I PUC CHEMISTRY SYLLABUS BLOW-UP

UNIT – I Some Basic Concepts of Chemistry 9 hrs

General introduction: Importance and scope of chemistry, nature of matter-classification, homogeneous and heterogeneous mixtures – examples, concept of elements, atoms, molecules and compounds.
Properties of matter and their measurement: seven basic physical quantities, their SI units and scientific notation (exponential notation).
Laws of chemical combination, with suitable examples.
Dalton's atomic theory – postulates.
Atomic and molecular masses: Atomic mass, amu (value of 1amu), average atomic mass with an example, molecular mass, examples, formula mass – NaCl as example. Mole concept and molar mass: Avogadro constant, mole and molar mass – examples.
Percentage composition, empirical formula and molecular formula- numerical problems. Stoichiometry relations –numerical problems to calculate amount of reactants/ products formed (in terms of mole and mass in grams) by giving balanced equations, limiting reagent –numerical problems.
Reactions in solutions: concentration terms – mass %, mole fraction, molality, molarity. Numerical problems.

UNIT – II Structure of Atom 10 hrs

Discovery of electron – name of the discoverer, characteristics of cathode rays, values of charge and mass.
Discovery of proton – characteristic of canal rays, values of charge and mass.
Discovery of neuron – name of the discoverer, value of charge and mass.
Atomic number, mass number, isotopes, isobars, problems.
Atomic models: Thomson atomic model and its limitations. Mention the observations and conclusions of α - ray scattering experiment. Rutherford atomic model and its limitations(based on Maxwell electromagnetic theory). Electromagnetic radiations – c, ν, $\bar{\nu}$, λ, their relationships, electromagnetic spectrum, particle nature of EMR($E = h\nu$), line spectrum of hydrogen, formula to calculate $\bar{\nu}$ of spectral lines in hydrogen – numerical problems.
Bohr's model-postulates and its limitations, concept of shells and subshells, dual nature of matter and light, de-Broglie relationship – numerical problems. Heisenberg uncertainity principle and its mathematical form. Concept of orbitals ,meaning of Ψ and Ψ^2, nodal surfaces or nodes.
Quantum numbers, shapes of s, p, d orbitals, rules for filling electrons in orbitals- (n + l) rule, Aufbau principle, Pauli exclusion principle, Hund's rule. Electronic configuration of atoms (1 to 36). Stability of half filled and completely filled orbitals.

UNIT – III Classification of Elements and Periodicity in Properties 5 hrs

Significance of classification, brief history of development of periodic table – law of triads with an example, law of octaves, Mendeleev periodic law – statement, Henry moseley observation based on X- ray spectra of elements, modern periodic law, long form of periodic table. Brief account of groups, periods, s, p, d and f blocks, Nomenclature of elements with atomic number greater than 100.

Periodic trends in properties of elements with reason: atomic radii, inert gas radii, ionic radii. compare radius of cation and anion with parent atom ,with reason, variation of radii of isoelectronic species, ionisation enthalpy, exception in first ionization enthalpy of N and O, with reason, electron gain enthalpy, compare $\Delta_{eg}H$ of F and Cl with reason. Electronegativity. valence – periodicity of valence or oxidation states (s and p block elements).

UNIT – IV Chemical Bonding and Molecular Structure 12 hrs

Chemical bond, valence electrons, Octet rule, Lewis symbols – significance, types of chemical bonds, Ionic bond (electrovalent bond) , example NaCl, Covalent bond- example Cl_2 (single bond formation), CO_2 (double bond formation), acetylene (triple bond formation), Lewis representation of some simple molecules (H_2, O_2, CO_3^{2-} as examples), formal charge – definition , calculation of formal charge on each oxygen atom in ozone, limitation of octet rule – with one example for each. Favourable conditions for the formation of ionic bond . Stability of ionic compound – lattice enthalpy. (details of lattice enthalpy to be dealt in thermodynamics).

Bond parameters: Bond length, covalent radius, Van der waals radius , bond angle, bond enthalpy and average bond enthalpy, bond order. Polarity of bonds- polar nature of covalent bond, dipole moment ,polarity in H_2O, BF_3, BeF_2, comparison of NH_3 and NF_3, Fajan's rule.

Geometry of molecules – VSEPR theory – postulates , shapes of molecules containing lone pair/s and bond pair/s, examples- $BeCl_2$, CH_4, H_2O, NH_3, SO_2. Resonance: concept, example- ozone. VBT: orbital overlap concept – s-s, s-p and p-p with examples, σ and π bonds.

Hybridisation concept-conditions for hybridization- types of hybridization, discuss sp^3 with CH_4, sp^2 with BCl_3, sp with C_2H_2, sp^3d with PCl_5, sp^3d^2 with SF_6, other examples to be mentioned.

MOT: Salient features, formation of molecular orbitals by LCAO method (qualitative approach),conditions for combination of atomic orbitals, formation of σ and π molecular orbitals, energy level diagrams for molecular orbitals for homonuclear diatomic molecules (H_2, He_2, C_2). Electronic configuration and molecular behaviour (bond order, nature of bond, bond length, magnetic nature, stability): H_2, He_2, Li_2, C_2, O_2.

Hydrogen bonding- types of hydrogen bonding, examples.

Unit V States of Matter- Gases and Liquids 9 hrs

Introduction-three states of matter, intermolecur forces- definition, types-dipole-dipole, dipole- induced dipole and London (dispersion) forces- a brief account with examples. Thermal energy- intermolecular forces *vs* thermal interactions .

Gaseous state: characteristics (mention), gas laws: Boyle's law and Charles' law - Statements, mathematical forms , graphs (P *vs*, V ,V *vs* T). Kelvin temperature scale, absolute zero-concept. Gay Lussac's Law (P, T relationship) - statement, mathematical form, graph. Avogadro law - Statement, mathematical form, Avogadro constant, STP conditions, molar volume.

Ideal gas: definition, ideal gas equation –derivation(from gas laws), gas constant R-value in SI units to be calculated, value of R in Latm $K^{-1}mol^{-1}$ to be mentioned. Relation between molar mass and density. Dalton's law of partial pressures - statement, mathematical form , aqueous tension and pressure of dry gas to be mentioned, relation between partial pressure of a gas and its mole fraction. Numerical problems on gas laws and ideal gas equation ,only.

Kinetic molecular theory of gases: assumptions, kinetic energy and molecular speeds (average, most probable, root mean square) - an elementary idea.

Behaviour of real gases- deviations from ideal behaviour, graph of PV *vs* P, causes for deviation and conditions for ideal behavior. van der Waals equation, compressibility factor (Z) - expression and its significance. Boyle temperature or Boyle point.

Liquifaction of gases - critical temperature, critical volume, critical pressure- meaning. (isotherms of CO_2 is not included) .

Liquid state: vapour pressure, normal and standard boiling points.

Surface tension and viscosity: definition and SI units (no mathematical derivations)

UNIT – VI Thermodynamics 11 hrs

Thermodynamic terms – concepts of system, surroundings, types of systems-examples, state of the system, state functions or state variables, energy- a state function, isothermal adiabatic, constant volume(isochoric)and pressure(isobaric) processes, reversible and irreversible processes, extensive and intensive properties.

Internal energy: as a state function .work and heat. Change in internal energy due to work and heat. First law of thermodynamics, mathematical form. Expression for ΔU under- adiabatic process ($\Delta U= -w$) and isothermal process ($\Delta U = q_v$). Expressions for work done during isothermal irreversible and reversible change. (derivation not included). Numerical problems.

Exothermic and endothermic reactions. Enthalpy: definition, change in enthalpy-sign convention, relationship between ΔH and ΔU,(derivation not included) examples. Numerical problems.

Heat capacity, specific heat, relationship between C_P and C_V for an ideal gas (derivation not included). Measurement of ΔU (bomb calorimeter) and of ΔH (calorimeter)-in brief.

Thermochemical equations- examples , enthalpy of a reaction – definition- example, factors affecting enthalpy of a reaction, standard state of a substance (specified temperature and 1 bar pressure). Standard enthalpy of a reaction: definition and examples of bond dissociation, phase transition, sublimation, formation, combustion, atomization, solution, dilution, ionization. Lattice enthalpy and Born – Haber cycle for NaCl. Hess's law of constant heat summation- statement-example. Numerical problems to calculate enthalpy of combustion and enthalpy of formation of CH_4, C_6H_6, CH_3OH.

Spontaneous and non spontaneous processes, examples, introduction of entropy as a state function, change in entropy of a system during a reversible process $\Delta S = \frac{q_{rev}}{T}$, entropy and spontaneity.Second law of thermodynamics,statement,Gibbs energy–definition(G = H–TS), Gibbs equation: $\Delta G = \Delta H - T\Delta S$, ΔG as a criterion for spontaneous and non spontaneous processes. Absolute entropy, third law of thermodynamics. Gibbs energy change and equilibrium, relationship between ΔG^0 and equilibrium constant (criteria for equilibrium), numerical problems.

UNIT – VII **Equilibrium** **13 hrs**

Introduction, equilibrium state of a system–equilibrium in physical processes-types-examples. Equilibrium involving dissolution of solid or gas in liquid- examples.

Equilibrium in chemical processes: meaning ($r_f = r_b$), dynamic nature, equilibrium equation (law of mass action).

Equilibrium constant (equilibrium law), $K^1 = \frac{1}{K}$ --- (a) , for reverse process,

K_p and K_c expressions for $aA + bB \rightleftharpoons cC + dD$ (to be assumed),

$K_p = K_c\,(RT)^{\Delta n}$ --- (b) (to be assumed) ,

examples for relation between K_p and K_c for reactions, $\Delta n = 0$, $\Delta n > 0$, $\Delta n < 0$ --- (c).
Numerical problems on (a), (b) and (c) and on K_P, K_C. (avoid quadratic equation).
Homogeneous and heterogeneous equilibria- examples.
Applications of equilibrium constant – predicting the extent of a reaction, direction of the reaction by reaction quotient Q, predicting the spontaneity of a forward or a reverse reaction based on ΔG of a reversible reaction.

Factors affecting equilibrium – Lechatelier's principle- effect of temperature, concentration, pressure, catalyst, addition of inert gas- in brief.
Effect of temperature : $2NO_2 \rightleftharpoons N_2O_4$; $\Delta H = -ve$
Effect of concentration: $Fe^{3+} + SCN^- \rightleftharpoons Fe(SCN)]^{2+}$, addition of Fe^{3+} and oxalate ion.
Effect of pressure: $CO + 3H_2 \rightleftharpoons CH_4 + H_2O$.
Ionic equilibrium – theories of acids and bases, with examples. Ionisation of acids and bases, degree of dissociation, strong and weak electrolytes, examples.
Ionic product of water: definition, expression, value at 298K, pH scale, pH- definition,
$pK_w = pH + pOH$(derivation).Numerical problems to calculate $[H^+]$, $[OH^-]$, from ionic product of water, pH, pOH.
Ionisation constant of weak acid and weak base: K_a and K_b, pK_a and pK_b and their relationship with K_w and pK_w. Ionisation of polybasic acid with an example. Factors affecting acid strength in brief (bond strength and electronegativity).Numerical problems (direct) on pK_a , pK_b
Common ion effect-definition, examples ($CH_3COOH + CH_3COONa$, $NH_4OH + NH_4Cl$), Buffer solutions-definition and examples (acetate and ammonia buffers), Henderson – Hesselbalch equation for acidic buffer to be derived, assume equation for basic buffer. Numerical problems.
Hydrolysis of salts, pH of their solutions (elementary idea), solubility product, solubility-examples, relationship between K_{sp} and S for AB, AB_2, A_2B type salts ($BaSO_4$, AgCl, Ag_2CrO_4, PbI_2) and a general expression for A_xB_y. Numerical problems taking $BaSO_4$, AgCl, Ag_2CrO_4, PbI_2 as examples.Condition for precipitation ($Q_{sp}>k_{sp}$). Common ion effect on solubility of ionic salts.

UNIT - VIII **Redox reactions** **5 hrs**

Concept of oxidation and reduction: classical idea–oxidation (addition of oxygen/ electronegative element or removal of hydrogen/ electropositive element, example for each), reduction - (removal of oxygen/electronegative element or addition of hydrogen/ electropositive element, example for each).

Redox reactions: in terms of electron transfer reactions with examples, oxidation & reduction - in terms of loss & gain of electrons, oxidising agent, reducing agent.

Oxidation number: definition, rules to calculate oxidation number, examples. Oxidation state, Stock notation – examples - FeO, Fe_2O_3, CuI, CuO, MnO and MnO_2.

Oxidation, reduction, oxidizing agent/oxidant, reducing agent/ reductant – in terms of oxidation number- examples.

Types of redox reactions:

1. Combination reactions : $C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2(g)}$
2. Decomposition reactions : $2KClO_{3\,(s)} \longrightarrow 2KCl_{\,(s)} + 3O_{2(g)}$
3. Displacement reactions (a) Metal displacement:
   $CuSO_{4(aq)} + Zn_{\,(s)} \longrightarrow Cu_{(s)} + ZnSO_{4\,(aq)}$
   b) Non-metal displacement: $2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$
   $2H_2O_{\,(l)} + 2F_{2\,(g)} \longrightarrow 4HF_{(aq)} + O_{2(g)}$
4. Disproportionation reactions: $2H_2O_{2\,(aq)} \longrightarrow 2H_2O_{(l)} + O_{2(g)}$
   $Cl_{2\,(g)} + 2\,OH^-_{\,(aq)} \longrightarrow ClO^-_{\,(aq)} + Cl^-_{\,(aq)} + H_2O_{\,(l)}$

Balancing of redox reactions :

a) Oxidation number method: $Cr_2O^{2-}_{7(aq)} + SO^{2-}_{3(aq)} \longrightarrow Cr^{3+}_{aq} + SO^{2-}_{4(aq)}$ (acidic medium)
   $MnO_4^-{}_{(aq)} + Br^-_{(aq)} \longrightarrow MnO_{2(s)} + BrO_3^-{}_{(aq)}$(acidic medium)

b) Half reaction method : $Fe^{2+}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + Cr^{3+}_{aq)}$ (acidic medium)
   $MnO_4^-{}_{(aq)} + I^-_{(aq)} \longrightarrow MnO_{2(s)} + I_{2(s)}$ (basic medium)

Applications: redox titrations, redox indicators - with examples. In electrode processes and cells (mention).

UNIT – IX **Hydrogen** **4 hrs**

Position of hydrogen in periodic table – similarities and differences with respect to alkali metals and halogens, occurrence, isotopes, preparation: laboratory method– Zn with acid, commercial – electrolysis of water, from methane and coal (as water gas). Properties: physical properties, chemical properties – reaction with halogens, dioxygen, dinitrogen, uses.

Hydrides – classification- one example for each type. Water – structure of the molecule, structure of ice, amphoteric nature (with NH_3, HCl), reaction with Na metal. Hard and soft water-differences, types of hard water-differences. H_2O_2 – preparation from BaO_2, volume strength of H_2O_2, structure, oxidizing property – with PbS, MnO^-_4 in acidic medium, reducing property – with I_2, storage of H_2O_2, uses. D_2O – uses. Dihydrogen as a fuel – meaning of hydrogen economy.

UNIT – X **s – Block Elements** **7 hrs**

Group – I, Group – II elements: general introduction, electronic configuration, occurrence, trends in ionization enthalpy, hydration enthalpy, atomic and ionic radii, trend in reactivity with oxygen (air), water, hydrogen , halogen. Uses.

Anomalous properties of lithium – reasons. Diagonal relationship with Mg – reasons, similarities in the properties of lithium with magnesium.

Preparation and properties of some compounds:

Sodium carbonate (washing soda):preparation by Solvay process (procedure and equations), properties-hydrolysis of CO_3^{2-} (Na_2CO_3),uses. Sodium chloride: sources, uses.

Sodium hydroxide: commercial process–using Castner - Kellner cell, properties–deliquescent, uses. Sodium bicarbonate (baking soda) – preparation from Na_2CO_3, uses.

Biological importance of sodium and potassium.

Anomalous behaviour of Beryllium- reasons, diagonal relationship with aluminium – reasons, similarities in properties of Beryllium with aluminium.

CaO: preparation, properties – reaction with water, CO_2, uses. $CaCO_3$: occurrence, preparation from slaked lime, uses, preparation of plaster of Paris form gypsum, uses. Biological importance of Ca, Mg.

UNIT – XI **Some p – Block Elements** **8 hrs**

General introduction to p– block elements-electronic configuration, oxidation states, inert pair effect, anomalous behavior of first member of each group.

Group 13 elements: General introduction, electronic configuration, occurrence , variation of atomic radii, ionization enthalpy, electronegativity, physical properties common oxidation states – considering inert pair effect, trend in chemical reactivity. Reaction of aluminium with air, acid, alkali (NaOH). Anomalous properties of boron.

Some important compounds of boron: Borax – reaction with water, action of heat, orthoboric acid – preparation from borax, properties – as a Lewis acid, action of heat, structure, diborane – preparation from BF_3 with $LiAlH_4$, physical properties- reaction with air, water, NH_3 – formation of inorganic benzene (borazine), structure. Uses of boron and aluminium.

Group – 14 elements: general introduction, electronic configuration, occurrence, variation of covalent radii, ionization enthalpy, electronegativity, oxidation states (inert pair effect) and trends in chemical reactivity towards oxygen and water.

Carbon: anomalous behaviour- reason, catenation, allotropic forms – graphite, diamond, fullerenes – their characteristics (structures not required). CO – preparation from HCOOH, carbon and air(producer gas), properties- reducing property- with Fe_2O_3, ZnO, poisonous nature, formation of metal carbonyls,uses. CO_2 – preparation from $CaCO_3$ (laboratory method), properties – weak dibasic acid, in photosynthesis, as dry ice,uses.

Important compounds of silicon: SiO_2 – structure, reaction with NaOH, HF, uses. Silicones –repeating unit $\left(R_2SiO \right)$, structure (partial) of the polymer, uses. Silicates – basic unit – SiO_4^{4-}, examples. zeolites – example, uses.

UNIT – XII Organic Chemistry – Some basic principles & Techniques 12 hrs

General introduction, mention urea as first organic compound synthesized by Wohler. Shapes of carbon compounds due to sp^3, sp^2 and sp to be mentioned. Structural representation – complete, condensed, and bond line formulas, wedge formula for CH_4.
Classification of organic compounds, functional groups, homologous series,IUPAC nomenclature of organic compounds (upto 6 carbons for aliphatic, 9 for aromatic),and bi-functional compounds.
Isomerism – structural – chain, position, functional, metamerism.
Fundamental concepts in organic reactions:mechanism – definition , fission of covalent bond – homolytic and heterolytic, carbanion, carbocation, alkyl free radicals, examples.Compare the stabilities of 1°, 2°, 3° carbocations and alkyl free radicals. Nucleophiles and electrophiles, examples.
Electron movement in organic reactions – Inductive effect – definition, example, electron withdrawing group(EWG,-I), electron donating groups (EDG,+I)-examples, resonance structures – concept to be recalled- resonance-definition, resonance energy, resonance effect, +R, –R effects with examples, electromeric effect , (+ E) and (–E) effects with examples, hyperconjugation (no bond resonance), examples $-C_2H_5^+$, $C_2H_5\bullet$, $CH_3CH = CH_2$ (orbital diagram not required).
Methods of purification of organic compounds: principle and examples - sublimation, crystallization, distillation, differential extraction.Chromatography:adsorption (column and TLC)and partition chromatography (all in brief). Diagrams for simple distillation, column and paper chromatography.
Qualitative analysis: detection of carbon and hydrogen, Lassigne's test: preparation of sodium fusion extract and tests to detect nitrogen, sulphur, halogens, and phosphorus (equations not expected).
Quantitative analysis: principle and calculations involved in the estimations of- carbon and hydrogen (labeled diagram), nitrogen by Duma's and KJeldahl's method(final equation only), halogens (Cl, Br, I) by carius method, sulphur by carius method and phosphorus. Numerical problems.

UNIT – XIII Hydrocarbons 12hrs

Classification of hydrocarbons.
Alkanes: nomenclature (upto 5 carbon atoms), isomerism,physical properties. Preparation by: hydrogenation of alkene and alkyne, examples (ethene, propene), from alkylhalide (reduction) and Wurtz reactions(methy and ethyl halides), Kolbe's electrolytic method for CH_3COONa (details of process not required).
Chemical properties: substitution reaction - halogenation - chlorination - mechanism, combustion (CH_4, C_4H_{10}), controlled oxidation (CH_4 to CH_3OH, H–CHO), aromatization (for hexane) pyrolysis.
Conformational isomerism: conformations- sawhorse and Newman projection formulae for eclipsed and staggered forms of ethane - compare stability and dihedral angle.
Alkenes: nomenclature (upto 5 carbon), structure of double bond (ethene, bond types and number).Geometrical isomerism – explain it as a type of stereoisomerism, cis and trans isomers,example – 2-butene. Physical properties.

Preparation: by hydrogenation of 2-butyne– by Lindlar's catalyst to get cis and Na/NH_3 to get trans isomers of 2-butene,dehydrohalogenation of alkyl halide, dehalogenation of vicinal halides- examples taking ethyl bromide and 2-cholopropane,1,2-dibromoethane, dehydration – ethene from alcohol.
Properties – chemical properties – addition reactions of ethene with H_2, Cl_2, Br_2 / CCl_4 (test for unsaturation). Markovnikoff's rule, addition of HBr to propene, mechanism, peroxide effect – for propene with HBr, addition of water to ethene and propene, oxidation (Baeyer's reagent) of ethene, ozonolysis (identification of products for ethene, propene, 2-butene), polymerization, uses.
Alkynes: nomenclature (up to 5 carbon), isomerism, structure of triple bond (ethyne-types of bonds and number). Preparation of ethyne – from calcium carbide,
1, 2-dibromoethane. Chemical properties for ethyne: acidic character – reaction with sodium metal, addition reactions with –H_2, Br_2, HBr, H_2O. Polymerization – example for linear polymer, ethyne to benzene.
Aromatic hydrocarbons: Introduction, IUPAC nomenclature, isomerism (position–o, p, m), structure of benzene – kekule structures, resonance and stability of benzene, aromaticity – characteristics for aromaticity (Huckel rule) – examples - benzene, cyclopentadienyl anion, naphthalene.
Chemical properties of benzene – electrophilic substitution reactions- halogenation, nitration, sulphonation, Friedel-carfts alkylation (R–X where R = CH_3, C_2H_5), acylation (CH_3COCl, $(CH_3CO)_2O$), benzene into hexachlorobenzene, addition reaction with H_2, Cl_2.
Mechanism of electrophilic substitution reaction – chlorination, nitration, alkylation (with CH_3Cl) acylation (CH_3COCl). Directive influence of a functional group in benzene – ortho and para directing groups (–OH, –OCH_3, –Cl, –CH_3) and meta directing groups (–NO_2, –CHO, –COOH) with examples.
Carcinogenicity and toxicity of benzene and polynuclear hydrocarbons to be mentioned.

Unit XIV Environmental chemistry 3 hrs

Environmental pollution:
Air pollution or troposphere pollution: gaseous air pollutants - oxides of sulphur, nitrogen, carbon, hydrocarbons - source and harmful effects to be mentioned. Global warming and greenhouse effect-brief note, acid rain - causes. Particulate pollutants- smoke, dust, mist and fumes, photochemical smog (composition)-source/formation and health problems-remedy.
Stratospheric Pollution: formation and breakdown of ozone (ozone hole), effects of depletion of the ozone layer. (Chemical reactions involved in the formation of smog and ozone depletion to be mentioned).
Water pollution: causes- organic wastes, pathogens, BOD and its significance, chemical pollutants and eutrophication.
Soil pollution: causes-pesticides, industrial wastes, biodegradable and non-biodegradable wastes. Strategies to control environmental pollution: waste management, collection and disposal.
Green chemistry: Introduction, green chemistry in day-to-day life, dry cleaning of clothes, bleaching of paper, synthesis of chemicals.

Guidelines for setting I PUC Chemistry question paper

1. The question paper has four parts: A, B, C and D. All the four parts are compulsory.
2. **Part A and B (I & II)**: Frame questions from all units as required.
   Part C (III): Frame questions from **Inorganic chemistry (Q.No.19 to 26)**. See the blueprint for spilt in the chapters)
   Part D (IV and V). Frame questions for part-IV from **Physical chemistry** (Q.No.27 to 34) and for part-V from **Organic chemistry** (Q.No.35 to 37).
3. **Blue print:** The question paper must be prepared based on the individual blue print which is based on the weightage of marks for each unit.
   - ❖ A variation of ±1 mark in the unit weightage is allowed.
   - ❖ **A blank blue print model is provided for reference.**
4. Answers to all the questions (except numerical problems) framed should be found in the syllabus provided by the Pre University Education Department.

Weightage to objectives:

Objective	Weightage	Marks
Knowledge	40%	43/105
Understanding	30%	31/105
Application	20%	21/105
Skill	10%	10/105

Weightage to level of difficulty

Objective	Weightage	Marks
Easy	40%	43/105
Average	40%	42/105
Difficult	20%	20/105

5. **Intermixing of questions** of different units is not allowed.
   5 marks question may be framed in (3+2) as for as possible.
   3 marks questions may be framed as 3 marks or (2 + 1) if inevitable.
6. **Questions based on numerical problems :** All the necessary data (i.e. like molecular mass, atomic mass, values of physical constants like **R, F, N_A etc.**,) should be given. **Final answer without appropriate unit carries zero mark.**
7. Each chapter (Unit) with marks weightage > 4 is split into two parts. Frame questions from part I for about half of the marks weightage and remaining from Part II. This is to give weightage to the entire unit and to avoid all questions from narrow range of the unit. For the approximate splitting of units, please refer the model blue print given.
8. **Numerical problems** worth of about **10 marks** should be given.
9. **Avoid questions from:**
   i) **Drawings** involving **3D diagrams**
   ii) **Boxed portions** of the units given in the text.
   iii) The **boxed materials with green colour** in the text book are to bring additional life to the topic and **are** non evaluative. (Please see the IV paragraph of the **preface** in the part I of the text book). Questions should not be framed on it

iv) Questions on numerical data given in the form of **appendix, numbered tables** containing **experimental data and life history of scientists** given in the chapters should be avoided.

10. One question on **mechanisms** (**3 marks**) in **organic chemistry** may be framed.
11. Prepare the **question paper** by strictly avoiding **½ mark** evaluation (or value points for ½ marks.)
12. Questions framed should not be vague and ambiguous.

I PUC CHEMISTRY (34)
Blue Print of Model Question paper - 1

GROUP	Unit	Title	Hrs	Marks	Part A 1x10 (VSA)**	Part B $2 \times \frac{5}{8}$ (SA)***	Part C $3 \times \frac{5}{8}$ (Inorganic)	Part D $5 \times \frac{7}{11}$ (Physical & Organic)	Total
Group-I Physical Hrs – 52 Marks= 47	I.	**Some basic concepts of Chemistry** Part 1= pg1-14; Part 2= pg 15-23;	9	8	✓(1)	✓(11)	-	✓(27)	08
	II.	**Structure of Atom** Part 1= pg 26-45; Part 2= pg 46-65;	10	9	-	-	-	✓✓ (28) (29)	10
	V.	**States of Matter: Gases and Liquids** Part 1= pg 132-143; Part 2= pg 143-152;	9	8	✓(2)	✓(12)	-	✓(30)	08
	VI.	**Thermodynamics** Part 1= pg154-164; Part 2= pg 164-180;	11	10	-	-	-	✓✓ (31) (32)	10
	VII	**Equilibrium** Part 1= pg185-205; Part 2= pg 205-222;	13	12	✓(3)	-	-	✓✓ (33) (34)	11
Group-II Inorganic Hrs – 41 Marks= 35	III.	**Classification of Elements and Periodicity in Properties**. Part 1= pg 70-82; Part 2= pg 82-92;	5	4	✓(4)	-	✓(19)	-	04
	IV.	**Chemical bonding and molecular structure** Part 1= pg 96-112; Part 2= pg 113-128;	12	11	-	✓(13)	✓✓✓ (20) (21) (22)	-	11
	VIII	**Redox Reactions** Part 1= pg255-266; Part 2= pg 266-272;	5	4	✓(5)	-	✓ (23)	-	04
	IX.	**Hydrogen.**	4	3	-	-	✓(24)	-	03
	X.	**S-Block Elements** Part 1= pg 291-298 (alkali); Part 2= pg 298-305 (Alkaline earth metals);	7	6	✓(6)	✓(14)	✓(25)	-	06
	XI.	**Some p-block Elements** Part 1= pg 307-314 (group- 13); Part 2= pg 314-322 (Group 14);	8	7	✓✓ (7) (8)	✓(15)	✓(26)		07
Group-III Organic Hrs – 27 Marks= 23	XII.	**Organic chemistry**: some basic principles and Techniques Part 1= pg326-341; Part 2= pg 341-360;	12	11	✓(9)	-	-	✓✓ (35) (36)	11
	XIII	**Hydrocarbons** Part 1= pg365-384; Part 2= pg -384-395;	12	10	✓(10)	✓✓ (16) (17)	-	✓ (37)	10
	XIV	**Environmental Chemistry**	3	2	-	✓(18)	-	-	02
		Total	**120**	**105**	**10**	**16**	**24**	**55**	**105**

Note : 1) The question paper must be prepared based on the individual blue print which is based on the Weightage of marks fixed for each unit/chapter.

Note : 2) In Chapters with marks wtg > 4, each chapter is split in 2 parts, about half of the total marks should be from part 1 and next half from part 2 of the chapter.

I PUC CHEMISTRY (34)
Blue Print of Model Question paper - 1

GROUP	Unit	Title	Hours	Marks	Part A 1x10 (VSA)**	Part B $2 \times \frac{5}{8}$ (SA)***	Part C $3 \times \frac{5}{8}$ (Inorganic)	Part D $5 \times \frac{7}{11}$ (Physical & Organic)	Total
Group-I Physical Hrs – 52 Marks= 47	I.	**Some basic concepts of Chemistry** Part 1= pg1-14; Part 2= pg 15-23;	9	8	(1)	(11)	-	(27)	08
	II.	**Structure of Atom** Part 1= pg 26-45; Part 2= pg 46-65;	10	9	-	-	-	(28) (29)	10
	V.	**States of Matter: Gases and Liquids** Part 1= pg 132-143; Part 2= pg 143-152;	9	8	(2)	(12)	-	(30)	08
	VI.	**Thermodynamics** Part 1= pg154-164; Part 2= pg 164-180;	11	10	-	-	-	(31) (32)	10
	VII	**Equilibrium** Part 1= pg185-205; Part 2= pg 205-222;	13	12	(3)	-	-	(33) (34)	11
Group-II Inorganic Hrs – 41 Marks= 35	III.	**Classification of Elements and Periodicity in Properties**. Part 1= pg 70-82; Part 2= pg 82-92;	5	4	(4)	-	(19)	-	04
	IV.	**Chemical bonding and molecular structure** Part 1= pg 96-112; Part 2= pg 113-128;	12	11	-	(13)	(20) (21) (22)	-	11
	VIII	**Redox Reactions** Part 1= pg255-266; Part 2= pg 266-272;	5	4	(5)	-	(23)	-	04
	IX.	**Hydrogen**.	4	3	-	-	(24)	-	03
	X.	**S-Block Elements** Part 1= pg 291-298 (Alkali); Part 2= pg 298-305 (Alkaline earth metals);	7	6	(6)	(14)	(25)	-	06
	XI.	**Some p-block Elements** Part 1= pg 307-314 (group- 13); Part 2= pg 314-322 (Group 14);	8	7	(7) (8)	(15)	(26)		07
Group-III Organic Hrs – 27 Marks= 23	XII.	**Organic chemistry: some basic principles and Techniques** Part 1= pg326-341; Part 2= pg 341-360;	12	11	(9)	-	-	(35) (36)	11
	XIII	**Hydrocarbons** Part 1= pg365-384; Part 2= pg -384-395;	12	10	(10)	(16) (17)	-	(37)	10
	XIV	**Environmental Chemistry**	3	2	-	(18)	-	-	02
		Total	**120**	**105**	**10**	**16**	**24**	**55**	**105**

Note : 1) The question paper must be prepared based on the individual blue print which is based on the Weightage of marks fixed for each unit/chapter.

Note : 2) In Chapters with marks wtg > 4, each chapter is split in 2 parts, about half of the total marks should be from part 1 and next half from part 2 of the chapter.

CHEMISTRY
I PUC MODEL QUESTION PAPER -1

Time: 3 Hours 15 min　　　　　　　　　　　　　　　　Max Marks: 70

INSTRUCTIONS:

i) The question paper has four parts A.B.C and D. All the parts are compulsory.
ii) Write balanced chemical equations and draw labeled diagrams wherever asked.
iii) Use log tables and simple calculators if necessary.
(Use of scientific calculators is not allowed)

PART-A

Answer all questions.　　　　　　　　　　　　　　　　**10x 1 =10**
(Answer each question in one word or in one sentence)

1. State 'law of definite proportions'.
2. Mention the type of intermolecular attractions that exists between non-polar molecules.
3. H^- is a Lewis base. Give reason.
4. Nitrogen has higher ionization enthalpy than that of oxygen. Give reason.
5. What is the oxidation state of Mn in MnO_4^- ?
6. Which alkali metal is the strongest reducing agent?
7. Give the composition of water gas?
8. Mention the type of hybridization of carbon in diamond.
9. Mention one use of chromatography.
10. Draw the staggered conformation of ethane.

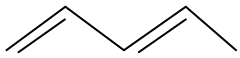
PART – B

Answer any FIVE questions (Each question carries two marks)　　　　**5x2=10**

11. a) Express 0.002568 in scientific notation.
    b) If the mass of one molecule of water is 18 u(amu), what is the mass of one mole of water molecules?
12. a) State Charles' law.
    b) Give the relationship between molecular mass and density of a gas.
13. Write the electronic configuration of H_2 molecule. What is its bond order?
14. Differentiate between the reactions of Li and Na on burning them in oxygen. Give equations.

15. What is the repeating unit in 'organo silicon polymer? Name the starting (raw) material used in the manufacture of organo silicon polymer.

16 Write the IUPAC names of the following hydrocarbons

i)  ii) $(CH_3)_3C - CH_3$

17. Give two tests to distinguish between alkanes and alkenes.

18. How is 'ozone layer' formed in the stratosphere? Name a chief chemical that causes its depletion.

PART – C

Answer any FIVE questions (Each question carries three marks) 5x3 = 15

19. a) Arrange the following in the decreasing order of their ionic radius:
N^{3-}, Mg^{2+}, Na^{+}, O^{2-}
b) State modern periodic law and assign IUPAC name to the element with atomic number 114. 1+2

20. a) Mention two conditions for the linear combination of atomic orbitals.
b) Draw the shapes of BMO and ABMO formed by the combination of 1s and 1s atomic orbitals 2+1

21. a) What are sigma and pi bonds?
b) Why is a sigma bond stronger than a pi bond? 2+1

22. a) Define dipole moment of a polar bond.
b) Show that BeF_2 molecule has zero dipole moment 1+2

23. Balance the Redox reaction using oxidation number method :
$MnO_4^-(aq) + Br^-(aq) \longrightarrow MnO_2(s) + BrO_3^-(aq)$ (in acidic medium)
3

24. Explain with equations the production of dihydrogen by coal gasification and water gas shift reaction. 3

25. a) Compare the hydration enthalpies and 2nd ionisation enthalpies of the alkali and alkaline earth metals.
b) Give the chemical formula of plaster of Paris. 2+1

26. a) Between boron and aluminium, boron cannot have covalency more than 4 but Al can have. Give reason.
b) Explain the reaction of diborane when it is exposed to air. 1+2

PART-D (IV & V)

IV Answer any FIVE questions. (Each question carries five marks) 5x5=25

27. a) Define i) Limiting Reagent ii) Molarity
b) $CaCO_3$ decomposes to give CO_2 gas according to the equation
$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$.
Calculate the mass of CaO(s) and CO_2 (g) produced on complete decomposition of 5.0 g of $CaCO_3$.
Given molar masses of CaO = 56 g, CO_2 = 44 g 2+3

28. a) The atomic number and atomic mass of Iron are 26 and 56 respectively. Find the number of protons and neutrons in its atom.

b) Calculate the wave number of the spectral line of shortest wavelength appearing in the Balmer series of H- spectrum. ($R = 1.09 \times 10^7\ m^{-1}$) 2+3

29. a) For the Element with atomic number 24

i) Write the Electronic configuration

ii) Write the value of **n** and ℓ for its electron in the valence shell

iii) How many unpaired electrons are present in it?

b) State Pauli's exclusion principle. 3+2

Is it possible to have the configuration $1s^3$.

30. a) Write any three postulates of kinetic theory of gases.

b) Two gases A & B have critical temperature as 250 K and 125 K respectively. Which one of these can be liquified first and why? 3+2

31. a) What is Intensive property of a system? Pick out the intensive property from mass, internal energy, density & volume.

b) 2 mol of an ideal gas undergoes a reversible and isothermal expansion from volume of 2.5 L to 10 L at 27^0 c. Calculate the work done by the gas in this expansion. Given R = 8.314 J/K/ mol 2+3

32. a) State Hess's law of constant heat summation.

b) Write Gibbs equation. Using ΔG, how do you decide whether a reaction at a given temperature is spontaneous or non spontaneous? 2+3

33. a) What is chemical equilibrium? What is meant by dynamic nature of chemical equilibrium?

b) Write the expression for equilibrium constant , K_c for the reaction

$aA + bB \rightleftharpoons cC + dD$. If the equilibrium constant for this reaction is 50, what is the equilibrium constant for it's reverse reaction

$cC + dD \rightleftharpoons aA + bB$? 2+3

34. a) Define acid and base by Bronsted - Lowry concept.

Identify a conjugate acid- base pair in the following.

$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$

b) What happens to the pH of water when NH_4Cl solid is dissolved in it and why? 3+2

V Answer any TWO questions. (Each question carries five marks) 2x5=10

35. For the compound $CH \equiv C - CH = CH - CH_3$ 5

i) Write its complete structure.

ii) Identify the number of sigma and pi bonds

iii) Identify the type of hybridisation of each carbon atom.

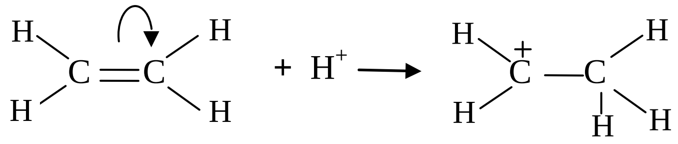
iv) Write the bond line formula of the compound.

v) Mention whether the compound in saturated or unsaturated

36. a) Identify the type of electron displacement effect in the following:

i) $\overset{\delta\delta+}{CH_3} \rightarrow \overset{\delta+}{CH_2} \rightarrow \overset{\delta-}{Cl}$

ii)

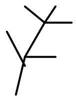
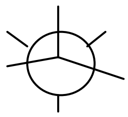


b) Give the principle and the formula involved in the estimation of sulphur by Carius method? 2+3

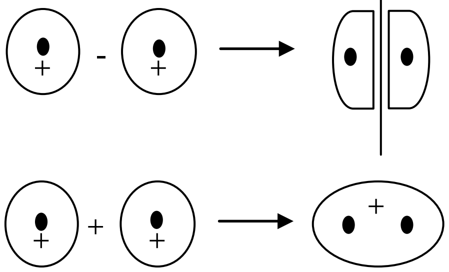
37. a) How is benzene prepared from ethyne?

b) Explain the mechanism of nitration of benzene. 2+3

I PUC CHEMISTRY
SCHEME OF VALUATION FOR MODEL PAPER-1

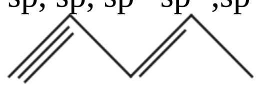
Q. No	PART - A	Mark
1.	Correct statement of the law	1
2.	Dispersive or London forces	1
3.	Because it donates an electron pair	1
4.	Nitrogen has more stable half filled orbitals (p^3) but Oxygen has less stable partially filled orbitals (p^4)	1
5.	+7	1
6.	Lithium or Li	1
7.	$CO + H_2$ or a mixture of carbon monoxide and hydrogen	1
8.	sp^3	1
9.	Separation of components in a mixture <u>or</u> to purify a compound or to test the purity of a compound. (Any one)	1
10.	Staggered conformation:  or 	1

	PART-B		
11.	a) b)	2.568×10^{-3} 18 g	1+1
12.	a) b)	Correct statement $d = \frac{PM}{RT}$	1+1
13.	$\sigma 1s^2\ \sigma^* 1s^0$ …. 1 Bond order = 1 …. 1		1+1
14.	Lithium on burning in oxygen gives its monoxide $4Li + O_2 \longrightarrow 2Li_2O$ Sodium on burning in oxygen gives its peroxide $4Li + O_2 \longrightarrow 2Li_2O$		1+1
15.	$\left(-R_2\,SiO-\right)$ or $\left(-\overset{R}{\underset{R}{Si}}-O-\right)$ Alkyl or Aryl substituted Silicon chloride		1+1
16.	a) b)	Pent– 1,3– diene 2,2 – dimethyl propane	1+1
17.	Two differences		1+1
18.	Due to action of UV radiations on oxygen ……………1 CFC or chlorine or chlorine containing compounds …………….1		1+1

		PART-C	
19.		a) $N^{-3} > O^{-2} > Na^{+} > Mg^{+2}$ b) Correct Statement ……………….. 1 IUPAC name of element 114 = Ununquadium ..………………..1	1+2
20.	a) b)	Two conditions…………2 	2+1
21.	a) b)	Covalent bond found by head-on/axial/end to end overlapping of bonding orbitals along the inter nuclear axis is called sigma bond. Covalent bond found by parallel/lateral overlapping of bonding orbitals perpendicular to the inter nuclear axis is called pi bond. Because in case of sigma bond the extent of overlapping of orbitals is more than that in a pi bond.	2+1
22.	a) b)	The product of the magnitude of the charge (q) and the distance between the centres of positive and negative charges (r) of a polar bond. The dipole moment of BeF_2 is zero because the two equal bond dipoles point in opposite directions hence cancel the effect of each other as shown below. F —— Be —— F $\longleftarrow\!\!+ \quad +\!\!\longrightarrow$	1+2
23.		$MnO_4^-\,(aq) + Br^-\,(aq) \longrightarrow MnO_2(s) + BrO_3^-\,(aq)$ +7 –1 +4 +5 $3e^- + MnO_4^- \longrightarrow MnO_2$ ….1 $Br^{-1} \longrightarrow BrO_3^- + 6e^-$ ..…1 Equalize the e^- lost to e^- gained $2MnO_4^- + Br^- \longrightarrow 2MnO_2 + BrO_3^-$ By inspection: $2MnO_4^- + Br^- + 2H^+ \longrightarrow 2MnO_2 + BrO_3^- + H_2O$ …..1	3
24.		Steam is passed over coal at 1000° C to get water gas. $C(s) + H_2O\,(g) \xrightarrow{1000°C} CO(g) + H_2(g)$ ….1 The mixture of steam and CO from water gas is passed over iron chromate at 400° C to get CO_2 and H_2 gas $CO(g) + H_2O\,(g) \xrightarrow{400°C/catalyst} CO_2\,(g) + H_2(g)$ ….2	3
25.	a) b)	Alkali metals have higher value for 2nd ionisation enthalpy ……1 Alkaline earth metal ions have higher value for hydration enthalpy ……1 $2\,(CaSO_4).\,H_2O$ …….1	2+1

26.	a) b)	Boron does not have 'd' sub shell in n = 2 level / it has only 4 valence orbitals / Al has d subshell in n = 3 level or has more valence orbitals to have Coordination Number – 6 ……1 Diborane catches fire spontaneously when exposed to air and burns in oxygen releasing large amount of energy. $B_2H_6 + 3O_2 \longrightarrow B_2O_3 + 3H_2O \quad \Delta_cH^\Theta = -Q \text{ kJ mol}^{-1}$ 2	1+2
		PART D	
IV Answer any Five			**5x5=25**
27.	a) b)	(i) & (ii)correct definitions ……(1+ 1) $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$. Molar masses 100 56 44 …(1 mark) 100 g of CaCO3 $\longrightarrow$ 56 g CaO 5 g of $CaCO_3 \longrightarrow$ …g of CaO $\quad 5 \text{ x } \frac{56}{100}$ = 2.8 g of CaO … (1 Mark) 100 g of $CaCO_3 \longrightarrow$ 44 g CO_2 5 g of $CaCO_3 \longrightarrow$ ……g of CO_2 $\quad 5 \text{ x } \frac{44}{100}$ = 2.2 g of CO_2 …(1 Mark)	2+3
28.	a) b)	Number of protons (Z)= 26; No. of Neutrons (A–Z)= 56–26 = 30; ....(1+1) $\bar{\nu} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ …..1 $\bar{\nu} = 1.09 \times 10^7 \left(\frac{1}{4} - 0\right)$ ….1 $\bar{\nu} = 2.725 \times 10^{-8} \text{ m}^{-1}$ …..1 (Answer without unit deduct 1 mark)	2+3
29	a) b)	i) $1s^2\ 2s^2\ 2P^6\ 3s^2\ 3P^6\ 3d^5\ 4s^1$ ….1 ii) n = 4, ℓ = 0 (½+ ½) mark each iii) Six or 6 ….1 Statement …1 No …1	3+2
30.	a) b)	Any three postulates ……… 1 x 3 Gas A ….. 1 Because on cooling the higher temperature is reached first. Hence the gas with higher critical temperature (T_c) will get liquefied first. ... 1	3+2
31.	a) b)	Definition of intensive property …1Mark Intensive property is density …1Mark $W = -2.203nRT \log \frac{V_2}{V_1}$ …1Mark Substitution: $W = -2.303 \times 2 \times 8.314 \times 300 \log \frac{10}{2.5}$ ………1Mark W = –3458.54 J ………1 mark	2+3
32.	a) b)	Correct statement of Hess's law ……..2 marks $\Delta G = \Delta H - T\Delta S$ ……………………………………………….1 mark ΔG = –ve means the reactive on is spontaneous …………1mark ΔG = +ve means the reaction is non spontaneous ………...1mark	2+3

33.	a) b)	Correct definition (meaning) of chemical equilibrium …………… 1 mark The state of a reversible reaction at which the rate forward reaction is equal to the rate of backward reaction………………1mark $K_c = \frac{C^c\ D^d}{A^a\ B^b}$ ………………..1 mark $K_C^{-1} = \frac{1}{K_C}$ ………………..1 mark $= \frac{1}{50} = 0.02$ ………………..1 mark	2+3
34	a) b)	Acid- species which donates a proton/ proton donor/ protogenic substance Base- species which accepts a proton/ proton acceptor/protophillic substance……………. 1+1 HNO_3 & NO_3^- or H_3O^+ & H_2O……………. 1 pH decreases……………. 1 Because NH_4Cl being a salt of strong acid and weak base undergoes hydrolysis to give acidic solution ……………. 1	2+3

V (Organic Chemistry) Answer any two		**2x5=10**	
35	a)	i) Write its complete structure ……………. 1 ii) Sigma = 10 and pi bond = 3 ……………. 1 iii) sp, sp, $sp^{2,}$ sp^2 ,sp^3 ……………. 1 iv)  ……………. 1 v) Unsaturated ……………. 1	5
36	a) b)	i) Inductive effect ……………. 1 ii) Electrometric effect. ……………. 1 Principle Procedure: organic compound + sodium peroxide/ fuming HNO_3 heat to get sulfuric acid. ……………. 1 Above mixture + excess of $BaCl_2$ solution to get white ppt of $BaSO_4$ filtered, washed, dried and weighted …………..1 Percentage of sulphur $= \frac{32 \times \text{Mass of } BaSO_4 \times 100}{233 \times \text{Mass of organic compound}}$ …………..1	2+3
37	a) b)	Ethyne on passing through a red hot Iron tube at 873 K undergoes cyclic polymerization to give benzene $3C_2H_2 \xrightarrow{\text{red hot iron tube}} C_6H_6$ or  …………..2 Nitration of benzene i) Equation for generation of electrophile ……..1 ii) Equation for formation of carbocation (arenium ion) ……..1 iii) Equation for removal of proton ……..1	2+3

Government of Karnataka

Commissionerate of Pre-University Education

I PUC Chemistry Practicals

EXPERIMENTS FOR CHEMISTRY PRACTICAL EXAMINATION

Time: 2 Hrs. Total Marks: 30

Q-I	**Salt analysis** Analyse the given simple inorganic salt systematically and report one **acid radical** and one **basic radical.**	10 marks
Q-II	**Titration (Volumetric Analysis)** Determination of the concentration (strength) of given **NaOH** solution by titrating against standard **Oxalic acid**. Or Determination of the concentration (strength) of given **HCl** solution by titrating against standard solution of **Sodium Carbonate**. (procedure of the titration should be given).	10 marks
Q-III	**Viva on the following experiments only (ask four simple questions each carrying one mark)** **Bunsen burner** **pH experiments** **Equilibrium experiment ($Fe^{3+} + SCN^{-} \rightleftharpoons [Fe(SCN)_3]^{2+}$** **Purification techniques**	4 marks
IV	Submission of the duly completed and **certified record**	6 marks
	TOTAL	**30 marks**

SCHEME OF VALUATION

Time: 2 Hrs. Total Marks: 30

Q-I	**Salt analysis (10 Marks)** **i) Preliminary tests** (any two correct) 1 mark **ii) Detection of Acid radical (4 Marks)** Group detection (correct group identification – 1 mark correct radical identification – 1 mark) 2 marks Confirmatory test 2 marks **iii) Detection of Basic radical (4 Marks)** Group detection (correct group identification – 1 mark correct radical identification – 1 mark) 2 marks Confirmatory test 2 marks For writing systematic procedure with absence of previous groups 1 mark	10

Q-II	**Titration (10 Marks)** i) For performing the experiment 3 marks For recording the readings in the tabular column 1 mark ii) For **accuracy** of the **Titre value** 3 marks up to ± 0.3 mL error 3 marks ± 0.4 mL error 2 marks ± 0.5 mL 1 mark ≥ 0.6 mL 0 mark * If a student reports abnormal error, the examiner may conduct the titration and assess the reading iii) **Calculations of Molarity (2 marks)** a. Formula 1 mark b. Substitution and answer (1+1) 2 mark	10 marks
Q-III	**Viva: four simple questions from the experiments mentioned above each carrying 1 mark ---- 1 x 4**	4 marks
IV	**Record** Submission of the duly completed and certified record	6 marks

Sl.No	% of experiments performed and recorded	Maximum marks to be awarded
1	> 90%	6
2	81% to 90%	5
3	71% to 80%	4
	41% to 70%	3
4	≤ 40%	0

	TOTAL	**30 marks**

Note:

a) The **following salts** are suggested to be given for analysis for practical examination: **NH_4Br, NH_4Cl, $Al_2(SO_4)_3$, $MgSO_4$, $CaCO_3$, $BaCl_2$, $MgSO_4$.**

b) **Inorganic salts** other than the mentioned above but given in the prescribed manual can be given to students in regular practical classes for practice.

c) All experiments as mentioned in the I PUC practical manual published by **Commissionerate of Pre-University Education** are to be conducted and recorded.