CLASS-XI

Work Book Cum Question Bank with Answers

CHEMISTRY

ଓଡ଼ିଶା ଶାସନ
ओड़िशा शासन

SCHEDULED CASTES & SCHEDULED TRIBES
RESEARCH & TRAINING INSTITUTE (SCSTRTI)
ST & SC DEVELOPMENT DEPARTMENT
BHUBANESWAR

WORK BOOK CUM QUESTION BANK WITH ANSWERS

CHEMISTRY

CLASS - XI

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SCHEDULED TRIBE & SCHEDULED CASTE
DEVELOPMENT DEPARTMENT
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Prepared by

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BHUBANESWAR – 751OO3

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CHEMISTRY

For Ist Year Science Syllabus

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Unit I:Some Basic Concepts of Chemistry

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 and equivalent mass of elements, acid, base, and salt,oxidants,reductants, and mole concept and molar mass, percentage composition, empirical and molecular formula, chemical reactions, stoichiometry and calculations based on stoichiometry.

Unit II: Structure of Atom

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

Unit III: Classification of Elements and Periodicity in Properties

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 and oxidation state.Nomenciature

of elements with atomic number greater than 100.

Unit IV : Chemical Bonding and Molecular Structure

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 molecules, molecular orbital theory of homonuclear diatomic molecules (qualitative idea only), hydrogen bond.

Unit V : States of Matter : Gases and Liquids

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

Unit VI : Chemical Thermodynamics

Concepts of System and surroundings and types of system, 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

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

Unit VIII : Redox Reaction

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

Unit IX : Hydrogen

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

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

Group 1 and Group 2 Elements

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

Preparation and Properties of Some Important Compounds :

Sodium Carbonate, Sodium Chloride, Sodium Hydroxide and Sodium Hydrogencarbonate, Biological importance of Sodium and Potassium.Calcium Oxide and Calcium Carbonate and their industrial uses, biological importance of Magnesium and Calcium.

Unit XI : Some p- Block Elements

General Introduction to p- Block Elements

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

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

Tetrachloride, Silicones, Silicates and Zeolites, their uses.

Unit XII : Organic Chemistry - Some Basic Principles and Technique

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 hyperconjugation. Homolytic and heterolytic fission of a covalent bond free radicals, carbocations, carbanions, electrophiles and nucleophiles, types of organic reactions.

Unit XIII : Hydrocarbons

Classification of Hydrocarbons

Aliphatic Hydrocarbons :

Alkanes - Nomenclature, isomerism, conformation (ethane only), physical properties methods of preparation 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 (Markownikoff'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.

Unit XIV : Environmental Chemistry

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

PERIODIC TABLE OF ELEMENTS

25 54.938
2,3,4,6,7
1962
1244
7.43
1.6
Mn
$[Ar]3d^5 4s^2$
Manganese

At. No. — At. Wt.
Ox. State
B.P. in °C
M.P. in °C
Density in g cm^{-3}
Electronegativity
(Allred, Rochow)
Symbol
Electronic Configuration
Name

NOTES

(1) Yellow-solid
Red-gas
Blue-liquid
Green-synthetically prepared.
(2) This table is scaled to the relative atomic mass, Carbon^{12} $A_r(^{12}C)=12$
(3) ( ) Indicates most stable or best known isotope.

←REPRESENTATIVE ELEMENTS→ s-Block (ns1–2)

← REPRESENTATIVE ELEMENTS → p-Block ($ns^2 np^{1-6}$)

d-Block [$(n-1)d^{1-10} ns^{0-2}$] or TRANSITION ELEMENTS

Group designation

Group →, PERIOD ↓

PERIOD	IA (1)		IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	(8)	VIII (9)	(10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (15)	VIIA (17)	INERT GASES (18)
1	1 1.0079 -1,1 -252.87 -259.2 0.0899 2.2 **H** $1s^1$ Hydrogen																	2 4.0026 0 -268.934 -272.2 0.178 **He** $1s^2$ Helium
2	3 6.941 1 1347 180.5 0.54 1.0 **Li** $[He]2s^1$ Lithium	4 9.0122 2 2970 1278 1.85 1.5 **Be** $[He]2s^2$ Beryllium											5 10.811 3 2550 2079 2.35 2.0 **B** $[He]2s^2 2p^1$ Boron	6 12.011 -4,2,4 4827 3367 2.26 2.5 **C** $[He]2s^2 2p^2$ Carbon	7 14.007 -3,2,3,4,5 -195.8 -209.86 0.81 3.1 **N** $[He]2s^2 2p^3$ Nitrogen	8 15.999 -2,-1 -182.96 -218.4 1.13 3.5 **O** $[He]2s^2 2p^4$ Oxygen	9 18.9984 -1 -188.14 -219.62 1.1 4.1 **F** $[He]2s^2 2p^5$ Fluorine	10 20.179 0 -246.05 -248.67 1.2 **Ne** $[He]2s^2 2p^6$ Neon
3	11 22.990 1 882.9 97.81 0.97 1 **Na** $[Ne]3s^1$ Sodium	12 24.305 2 1090 648.8 1.74 1.20 **Mg** $[Ne]3s^2$ Magnesium											13 26.982 3 2467 660 2.70 1.5 **Al** $[Ne]3s^2 3p^1$ Aluminium	14 28.0855 4 2355 1410 2.34 1.7 **Si** $[Ne]3s^2 3p^2$ Silicon	15 30.97376 -3,3,5 280w 44.2w 1.82.w 2.1 **P** $[Ne]3s^2 3p^3$ Phosphorus	16 32.066 -2,2,4,6 444.67 112.8 2.1 2.4 **S** $[Ne]3s^2 3p^4$ Sulphur	17 35.453 -1,1,3,5,7 -36.6 -100.98 1.1 2.8 **Cl** $[Ne]3s^2 3p^5$ Chlorine	18 39.948 0 -185.8 -189.2 -1.41 - **Ar** $[Ne]3s^2 3p^6$ Argon
4	19 39.098 1 774 63.65 0.86 0.90 **K** $[Ar]4s^1$ Potassium	20 40.078 2 1484 839 1.55 1 **Ca** $[Ar]4s^2$ Calcium	21 44.956 3 2831 1541 3.0 1.2 **Sc** $[Ar]3d^1 4s^2$ Scandium	22 47.88 3,4 3287 1660 4.5 1.3 **Ti** $[Ar]3d^2 4s^2$ Titanium	23 50.942 2,3,4,5 3380 1890 6.11 1.5 **V** $[Ar]3d^3 4s^2$ Vanadium	24 51.996 2,3,6 2672 1857 7.14 1.6 **Cr** $[Ar]3d^5 4s^1$ Chromium	25 54.938 2,3,4,6,7 1962 1244 7.43 1.6 **Mn** $[Ar]3d^5 4s^2$ Manganese	26 55.847 2,3,6 2750 1535 7.87 1.6 **Fe** $[Ar]3d^6 4s^2$ Iron	27 58.933 2,3 2870 1495 8.9 1.7 **Co** $[Ar]3d^7 4s^2$ Cobalt	28 58.69 2,3 2732 1453 8.91 1.8 **Ni** $[Ar]3d^8 4s^2$ Nickel	29 63.546 1,2 2567 1083 8.96 1.8 **Cu** $[Ar]3d^{10} 4s^1$ Copper	30 65.39 2 907 419.60 8.14 1.7 **Zn** $[Ar]3d^{10} 4s^2$ Zinc	31 69.723 3 2403 29.78 5.90 1.8 **Ga** $[Ar]3d^{10} 4s^2 4p^1$ Gallium	32 72.61 4 2830 937.4 5.32 2.0 **Ge** $[Ar]3d^{10} 4s^2 4p^2$ Germanium	33 74.922 -3,3,5 613* 817 5.17 2.2 **As** $[Ar]3d^{10} 4s^2 4p^3$ Arsenic	34 78.96 -2,4,6 684.9 217 4.19 2.5 **Se** $[Ar]3d^{10} 4s^2 4p^4$ Selenium	35 79.904 -1,1,3,5,7 58.78 -7.2 3.19 2.7 **Br** $[Ar]3d^{10} 4s^2 4p^5$ Bromine	36 83.80 2,4 -152.3 -156.30 2.16 **Kr** $[Ar]3d^{10} 4s^2 4p^6$ Krypton
5	37 85.468 1 688 38.89 1.5 0.9 **Rb** $[Kr]5s^1$ Rubidium	38 87.62 2 1384 769 2.63 1.0 **Sr** $[Kr]5s^2$ Strontium	39 88.906 3 3338 1522 4.5 1.1 **Y** $[Kr]4d^1 5s^2$ Yttrium	40 91.224 4 4377 1852 6.51 1.2 **Zr** $[Kr]4d^2 5s^2$ Zirconium	41 92.906 3.5 4742 2468 6.57 1.2 **Nb** $[Kr]4d^4 5s^1$ Niobium	42 95.94 2,3,4,5,6 4612 2617 10.28 1.3 **Mo** $[Kr]4d^5 5s^1$ Molybdenum	43 (98.906) 7 4878 2172 11.5 1.4 **Tc** $[Kr]4d^6 5s^2$ Technetium	44 101.07 3,4,8 3900 2310 12.41 1.4 **Ru** $[Kr]4d^7 5s^1$ Ruthenium	45 102.91 1,2,4 3527 1966 12.39 1.5 **Rh** $[Kr]4d^8 5s^1$ Rhodium	46 106.42 2,4 3140 1552 11.99 1.4 **Pd** $[Kr]4d^{10}$ Palladium	47 107.87 1,2 2212 961.9 10.49 1.4 **Ag** $[Kr]4d^{10} 5s^1$ Silver	48 112.41 2 765 320.90 8.65 1.5 **Cd** $[Kr]4d^{10} 5s^2$ Cadmium	49 114.82 3 2080 156.60 7.31 1.5 **In** $[Kr]4d^{10} 5s^2 5p^1$ Indium	50 118.71 2,4 2270 232 7.27 1.7 **Sn** $[Kr]4d^{10} 5s^2 5p^2$ Tin	51 121.75 -3,3,5 1750 630.70 100.70 1.8 **Sb** $[Kr]4d^{10} 5s^2 5p^3$ Antimony	52 127.60 -2,4,6 990 449.5 100.25 2 **Te** $[Kr]4d^{10} 5s^2 5p^4$ Tellurium	53 126.90 -1,1,3,5,7 184.40 113.5 4.94 2.2 **I** $[Kr]4d^{10} 5s^2 5p^5$ Iodine	54 131.29 2,4,6 -107.10 -111.9 3.5 **Xe** $[Kr]4d^{10} 5s^2 5p^6$ Xenon
6	55 132.91 1 678.4 28.40 1.9 0.90 **Cs** $[Xe]6s^1$ Cesium	56 137.33 2 1640 725 3.62 1.0 **Ba** $[Xe]6s^2$ Barium	57 138.91 3 3457 921 6.17 1.1 **La*** $[Xe]5d^1 6s^2$ Lanthanum	72 178.49 4 4602 2227 13.28 1.2 **Hf** $[Xe]4f^{14} 5d^2 6s^2$ Hafnium	73 180.95 5 5425 2995 16.65 1.3 **Ta** $[Xe]4f^{14} 5d^3 6s^2$ Tantalum	74 183.85 2,3,4,5,6 5660 3410 19.3 1.4 **W** $[Xe]4f^{14} 5d^4 6s^2$ Tungsten	75 186.2 2,4,7 5627 3160 21.0 1.5 **Re** $[Xe]4f^{14} 5d^5 6s^2$ Rhenium	76 190.2 2,3,4,5,6 5027 3045 22.57 1.5 **Os** $[Xe]4f^{14} 5d^6 6s^2$ Osmium	77 192.22 1,2,3,4,6 4130 2410 22.51 1.6 **Ir** $[Xe]4f^{14} 5d^7 6s^2$ Iridium	78 195.08 2,4 3827 1772 21.41 1.4 **Pt** $[Xe]4f^{14} 5d^9 6s^1$ Platinum	79 196.97 1,3 2807 1064 19.32 1.4 **Au** $[Xe]4f^{14} 5d^{10} 6s^1$ Gold	80 200.59 1,2 356.60 -38.84 13.53 1.5 **Hg** $[Xe]4f^{14} 5d^{10} 6s^2$ Mercury	81 204.38 1,3 1457 303.5 11.85 1.4 **Tl** $[Xe]4f^{14} 5d^{10} 6s^2 6p^1$ Thallium	82 207.2 2,4 1740 327.5 11.34 1.6 **Pb** $[Xe]4f^{14} 5d^{10} 6s^2 6p^2$ Lead	83 208.98 3,5 1560 271.3 9.81 1.7 **Bi** $[Xe]4f^{14} 5d^{10} 6s^2 6p^3$ Bismuth	84 (208.98) 2,4,6 962 254 9.14 1.8 **Po** $[Xe]4f^{14} 5d^{10} 6s^2 6p^4$ Polonium	85 (209.99) -1,1,3,5,7 337 302 - 2.0 **At** $[Xe]4f^{14} 5d^{10} 6s^2 6p^5$ Astatine	86 (222.02) 2 -61.8 -71 4.4 - **Rn** $[Xe]4f^{14} 5d^{10} 6s^2 6p^6$ Radon
7	87 (223.02) 1 677 27 - 0.90 **Fr** $[Rn]7s^1$ Francium	88 (226.03) 2 1140 700 5.5 1.0 **Ra** $[Rn]7s^2$ Radium	89 (227.03) 3 3200 1050 - 1 **Ac**** $[Rn]6d^1 7s^2$ Actinium	104 (261.11) - - - - - **Rf** $[Rn]5f^{14} 6d^2 7s^2$ Rutherfordium	105 (262.114) - - - - - **Db** $[Rn]5f^{14} 6d^3 7s^2$ Dubnium	106 (263) - - - - - **Sg** $[Rn]5f^{14} 6d^4 7s^2$ Seaborgium	107 (262) - - - - - **Bh** $[Rn]5f^{14} 6d^5 7s^2$ Bohrium	108 (265) - - - - **Hs** $[Rn]5f^{14} 6d^6 7s^2$ Hassium	109 (268) - - - - **Mt** $[Rn]5f^{14} 6d^7 7s^2$ Meitnerium	110 (269) - - - - **Ds** $[Rn]5f^{14} 6d^8 7s^2$ Darmstadtium	111 (272) - - - - **Rg** $[Rn]5f^{14} 6d^{10} 7s^1$ Rontgenium	112 (277) - - - - **Cn** $[Rn]5f^{14} 6d^{10} 7s^2$ Copernicium	113 **Uut** Un-Un Trium	114 (289) **Uuq** $[Rn]5f^{14} 6d^{10} 7s^2 7p^2$ Un-Un Quadium	115 **Uup** $[Rn]5f^{14} 6d^{10} 7s^2 7p^3$ Un-Un Pentium	116 (289) **Uuh** $[Rn]5f^{14} 6d^{10} 7s^2 7p^4$ Un-Un Hexium	117 **Uus** $[Rn]5f^{14} 6d^{10} 7s^2 7p^5$ Un-Un Septium	118 (293) **Uuo** $[Rn]5f^{14} 6d^{10} 7s^2 7p^6$ Un-Un Octium

METALS ← → NON-METALS

f–Block [$(n-2)f^{0-14}(n-1)d^{0-1} ns^2$] or INNER–TRANSITION ELEMENTS

* LANTHANOIDS ($4f^{1-14} 5d^{0-1} 6s^2$)	6	58 140.12 3,4 3426 799 5.9 1.1 **Ce** $[Xe]4f^2 6s^2$ Cerium	59 140.91 3,4 3512 931 6.5 1.1 **Pr** $[Xe]4f^3 6s^2$ Praseodymium	60 144.24 3 3068 1021 6.9 1.1 **Nd** $[Xe]4f^4 6s^2$ Neodymium	61 (146.92) 3 2460 1168 - 1.1 **Pm** $[Xe]4f^5 6s^2$ Promethium	62 150.36 2,3 1791 1077 7.54 1.1 **Sm** $[Xe]4f^6 6s^2$ Samarium	63 151.17 2,3 1597 822 5.1 1 **Eu** $[Xe]4f^7 6s^2$ Europium	64 157.25 3 3266 1313 8.0 1.1 **Gd** $[Xe]4f^7 5d^1 6s^2$ Gadolinium	65 158.93 3 3123 1356 8.27 1.1 **Tb** $[Xe]4f^9 6s^1$ Terbium	66 162.50 3 2562 1412 8.5 1.1 **Dy** $[Xe]4f^{10} 6s^2$ Dysprosium	67 164.93 3 2695 1474 8.8 1.1 **Ho** $[Xe]4f^{11} 6s^2$ Holmium	68 167.26 3 2900 1497 9.07 1.1 **Er** $[Xe]4f^{12} 6s^2$ Erbium	69 168.93 2,3 1947 1545 9.33 1.1 **Tm** $[Xe]4f^{13} 6s^2$ Thulium	70 173.04 2,3 1194 819 6.98 1.1 **Yb** $[Xe]4f^{14} 6s^2$ Ytterbium	71 174.97 3 3395 1663 9.84 1.1 **Lu** $[Xe]4f^{14} 5d^1 6s^2$ Lutetium
** ACTINOIDS ($5f^{1-14} 6d^{0-1} 7s^2$)	7	90 232.04 4 4790 1750 11.7 1.1 **Th** $[Rn]6d^2 7s^2$ Thorium	91 231.04 4,5 - 1600 15.4 1.1 **Pa** $[Rn]5f^2 6d^1 7s^2$ Protactinium	92 238.03 3,4,5,6 3818 1132 19.05 1.2 **U** $[Rn]5f^3 6d^1 7s^2$ Uranium	93 237.05 3,4,5,6, 3902 640 - 1.2 **Np** $[Rn]5f^4 6d^1 7s^2$ Neptunium	94 (244.06) 3,4,5,6 3232 641 - 1.2 **Pu** $[Rn]5f^6 7s^2$ Plutonium	95 (243.06) 3,4,5,6 2607 994 - 1.2 **Am** $[Rn]5f^7 7s^2$ Americium	96 (247.07) 3,4 - 1340 - 1.2 **Cm** $[Rn]5f^7 6d^1 7s^2$ Curium	97 (247.07) 3,4 - - - - **Bk** $[Rn]5f^9 7s^2$ Berkelium	98 (251.08) 3,4 - - - 1.2 **Cf** $[Rn]5f^{10} 7s^2$ Californium	99 (252.08) 3 - - - 1.2 **Es** $[Rn]5f^{11} 7s^2$ Einsteinium	100 (257.10) 3 - - - 1.2 **Fm** $[Rn]5f^{12} 7s^2$ Fermium	101 (258.10) 3 - - - 1.2 **Md** $[Rn]5f^{13} 7s^2$ Mendelevium	102 (259.10) 2,3 - - - - **No** $[Rn]5f^{14} 7s^2$ Nobelium	103 (260.11) 3 - - - - **Lr** $[Rn]5f^{14} 6d^1 7s^2$ Lawrencium

Compounds & their chemical formulae

Alabanite–MnS.
Alabaster–$CaSO_4$, $2H_2O$—Highly crystalline and natural Calcium sulphate.
Alcohol (Ethyl)–C_2H_5 (OH) grain alcohol || Alcohol (Methyl)–CH_3(OH), Wood alcohol
Alkhest–Universal Solvent of the Alchemists || Alkali Marine, Na_2CO_3 Soda.carb.
Alkali Vegetable—K_2CO_3 Potash carb, || Alkali Volatile $(NH_4)_2CO_3$. Alummon carb
Alum—K_2SO_4, Al_2 $(SO_4)_3$, $24H_2O$—Potassium Aluminium sulphate.
Alum– ammonium– $(NH_4)_2SO_4$, $Al_2(SO_4)_3$, 24 H_2O.
Alum –neutral—K_2SO_4, $Al_2(SO_4)_3$, $Al_2(OH)_6$.
Alim–Roman—It is a potash alum obtained from Alunite.
Alumina–Al_2O_3, Al-oxide || Alunite (alum-stone), K_2SO_4, $Al_2(SO_4)_3$, $4Al(OH)_3$ (n)
Amethyst (Violet)—yakut, Al_2O_3 (n) || Anthracite—C-Coal (n).
Antichlor–SO_2, Sulphur dioxide || Apatite (flour) $3Ca_3(PO_4)_2$, CaF_2(n).
Aragonite—$CaCO_3$—Rhombic variety of $CaCO_3$ (n).
Argenite—Ag_2S–Silver sulphide (n) || Argentum Vivum—Native Mercury.
Asbestos—$CaSiO_3$, 3 $MgSiO_3$—CaMg-Silicate (n).
Auricome—H_2O_2, Hydrogen peroxide peroxide || Austentite—a kind of hard steel.
Azurite—$2CuCO_3$ $Cu(OH)_2$, Basic carbonate of Cu (n) || Aquafortis—Nitric Acid.
Aqua regia—conc. HNO_3 + conc. HCl in the prop. of 1:. 3. (Royal water)
Baking powder—a mix. of $NaHCO_3$ & Tartaric acid || Baryta—BaO—Barium oxide
Baryta water—$Ba(OH)_2$ sol. Baking soda - $NaHCO_3$
Barytes—$BaSO_4$—Barium sulphate (n) || Bauxite—Hydrated Al-oxide (n). $Al_2O_32H_2O$
Benzene—C_6H_6—Hydrocarbon || Beryl— $Al_2(SiO_3)_3$ $3BeSiO_3$ – Al–Be silicate (n)
Bittern—Mother liquor from which common salt is crystallized out.
Black Ash—Mix of Na_2CO_3, CaS and Carbon.
Black Jack - ZnS–Zinc sulphide (n) || Black lead – C, Graphite.
Bog iron ore—Hydrated iron oxide (n) || Bolognian phosphorus—BaS
Bone-ash—Ca_3 $(PO_4)_2$—Calcium phosphate || Boracite, 2 $Mg_3B_3O_{15}$, $MgCl_2$ (n)
Borax—$Na_2B_4O_7$, $10H_2O$, Sodium pyroborate (n)
Borocalcite—CaB_4O_7, $6H_2O$ Calcium pyroborate || Bort—C—Impure diamond
Brimstone—S—Sulphur (n) || Brine—strong solution of common salt in water.
Britannia metal—alloy of Sb, Sn, Cu and Zn || Bronze—alloy of Cu and Sn.
Bronze (*aluminium*)—alloy of Cu and Al, looking like Gold.
Calamine—$ZnCO_3$. Zinc carbonate (n) || Calamine (electric). Zinc silicate.
Calcite—$CaCO_3$, Calcium carbonate (n) || Canton's phosphorus, CaS.
Calcspar—$CaCO_3$ (n), Caliche
Calomel—Hg_2Cl_2—Mercurous, chloride.
Carbonado—C—impure diamond || Carborundum–SiC, Silicon carbide.
Carnallite—$MgCl_2$, KCl, $6H_2O$—Double chloride of Mg and K (n)
Cassel yellow—$PbCl_2$, 7PbO—Basic load chloride.
Cassiterite—SnO_2—Tin Oxide (n) || Celestine—$SrSO_4$—Strontium sulphate (n)
Chile saltpetre – $NaNO_3$
Cerussite—$PbCO_3$—Lead carbonate (n) || Cinnabar—HgS—Mercuric sulphide (n).
Chalcopyrite—Cu_2S, Fe_2S_3, mixt. of Cu and Fe sulphides (n).
Coral—$CaCOa_3$, Calcium carbonate (n) || Corundum—Al_2O_3—Aluminium oxide (n)
Cryolite—AlF_3, 3NaF—Double fluoride of Al and Na (n) || Chalk—$CaCO_3$.
Chilli Saltpetre (Chili nitre)—$NaNO_3$— Sodium nitrate (n).
Corrosive sublimate—$HgCl_2$—Mercuric chloride,
Dephlogisticated air—Oxygen.

Diaspore—$Al_2O_3\ H_2O$ Hydrated Al-oxide (n) || Dihydrol—H_2O_2.
Dolomite–$CaCO_3$, $MgCO_3$—Double carb of Ca and Mg (n).
Dowflakc—Calcium chloride || Drier for paints—MnO_2, Mn-borate, etc.
Dutch liquid—$C_2H_4Cl_2$ — Ethylene chloride || Dutch metal—alloy of Cu and Zn
Eau de Javelle—Sol. of chlorine in cold dil. KOH sol.
Electrolytic gas—Mix. of H and O as obtd. by the electrolysis of H_2O.
Emerald—(green, Panna) —Al_2O_3 (n)
Epsom salt—$MgSO_4\ 7H_2O$—Magnesium sulphate.
Felspar—K_2O, Al_2O_3, $6SiO_2$—Potassium Al-silicate (n).
Fire-clay—Aluminium silicate || Formaldehyde—HCHO.
Franklinite—ZnO, Fe_2O_3, $6MnO_2$ — mix. of Zn, Mn and Fe-oxides (n)
Fluorspar—CaF_2—Calcium fluoride.
Glauber's salt—Na_2SO_4. $10H_2O$—Sodium sulphate.
Green vitriol—$FeSO_4$, $7H_2O$– Ferrous sulphate.
Galena—PbS—Lead sulphide (n) || Gangue—Impurities present in ores
Garnets—Mg, Fe or Ba-Al-silicates || Guano—Sea-bird's excreta.
Gun Powder—mix. of Charcoal, Sulphur and Nitre.
Haematite—Fe_2O_3—Ferric oxide (n).
Hausmannite—Mn_3O_4—Trimanganese tetroxide (n).
Heavy spar—$BaSO_4$—Barium sulphate (n).
Horn blende—Aluminium silicate (n) || Horn Silver—AgCl—Silver chloride (n)
Hydragillite—Al_2O_3, $3H_2O$—Hydrated Al.oxide.
Hydrogenite—Mix. of Silicon, NaOH and $Ca(OH)_2$
Hydrolith—CaH_2—Calcium hydride.
Inflammable air—H_2—Hydrogen.
Infusible white ppt—$HgCl(NH_3)$—Mercuric Am-chloride.
Ivory black—Animal charcoal.
Kainite—K_2SO_4, $MgCl_2$, $MgSO_4$, $6H_2O$.
Kelp—Ash obtd. by burning sea-weeds.
Kieselguhr—Silicious earth; remains of diatom (n). Diatoms are single celled aquati algae.
Kieserite—$MgSO_4$, H_2O—Magnesium sulphate (n).
Lapis Lazuli (ultramarine)—mixture of Sodium Al-silicate and Na_2S.
Lepidolite—Lithium mica || Leucite, K_2O, Al_2O_3, $4SiO_2$—Potash Al-Silicate (n).
Lignite—(brown coal) —A kind of coal.
Limonite—$2F_2O_3$, $3H_2O$ — Hydrated iron oxide (n).
Lithopone—white paint—Mix of $BaSO_4$ ZnS.
Liver of sulphur (Heper sulph)—obtd. by fusing S with K_2CO_3.
Loam–mixture of clay and sand || Luminous paint contains CaS, BaS, SrS or ZnS.
Lunar caustic—Silver nitrate || Laughing gas—N_2O —Nitrous oxide.
Magnesia—MgO—Magnesium oxide || Magnesite—$MgCO_3$—Magnesium carbonate (n)
Magnetite—Fe_3O_4—Ferroso-ferric oxide.
Malachite—$CuCO_3$, $Cu(OH)_2$—Basic carbonate of Cu (n).
Manganin—alloy of Cu, Mn and Ni || Marcasite—FeS_3—Iron diulphide.
Marl—mixture of clay and limestone.
Milk of lime—$Ca(OH)_2$—Suspension of $Ca(OH)_2$ in its solution.
Mosaic gold—SnS_2—Stannic sulphide.
Mohr's salt—Ferrous ammonium sulphate—$FeSO_4$, $(NH_4)_2SO_4$, $6H_2O$.
Nessler's reagent—Sol. of HgI_2 in KI sol. with excess of NaOH sol., $K_2[Hg\ I_4]$
Nitre—KNO_3—Potassium nitrate.
Nitroglycerine (Nobel's oil) — Glycerine trinitrate.

Olvine—Mg_2SiO_4 Magnesium silicate (n) || Opal—SiO_2—Silicon oxide (n).
Orpiment—As_2S_3—Arsenic trisulphide (n) || Orthoclase—Potash-Al-silicate (n).
Oxy-muriatic acid—Chlorine || Oil of vitriol—H_2SO_4—Sulphuric acid.
Pearl ash—K_2CO_3— Potash carb.
Perhydrol—strong sol. of H_2O_2 in water || Periclase – MgO—Magnesium oxide.
Permanent white—$BaSO_4$—Barium sulphate || Pewter— Alloy of Pb and Sb.
Pharaoh's Serpent—$Hg(CNS)_2$—Mercuric thiocyanate.
Phlogisticated air—N—Nitrogen
Phosphor-bronze — Alloy of Cu and Sn containing P.
Plaster of Paris— $(CaSO_4)_2, H_2O$—Calcium sulphate.
Platinized asbestos—Deposit of finely divided Pt on asbestos.
Potash—K_2CO_3—Potassium carbonate.
Prussian blue—$Fe_4[Fe(CN)_6]_3, 9H_2O$ —Ferric ferrocyanide.
Purple of Cassius—Purple ppt. obtd. by adding $SnCl_2$ and $SnCl_4$ to $AuCl_3$ sol.
Pyrolusite—MnO_2—Manganese dioxide (n).
Qnartz—SiO_2—Silicon dioxide (n) || Quicklime—CaO—Calcium oxide.
Quicksilver—Hg—Mercury.
Realger—As_2S_2—Arsenic disulphide (n).
Resin of Copper—Cu_2Cl_2—Fused Cuprous chloride.
Red Prussiate of Potash—$K_3Fe(CN)_6$—Potassium ferricyanide.
Rouge—Fe_2O_3—Ferric oxide || Sal Sedativum—H_2BO_3—Boric acid.
Sal. Volatile—$(NH_4)_2\ CO_3$ —Ammon. carb. || Sapphire (blue, neelum)—Al_2O_3.
Satinspar—$CaSO_4, 2H_2O$— Fibrous Gypsum (n).
Selenite—$CaSO_4, 2H_2O$— Transparent Gypsum (n).
Serpentine—$Mg_3Si_2O_7, 2H_2O$ — Mag. silicate hydrated (n).
Siderite—$FeCO_3$—Iron carbonate (n) || Soluble glass—Na_2SiO_3 or K_2SiO_3.
Spathic Iron ore—$FeCO_3$— Iron carbonate (n).
Spelter—Zn—Impure commercial Zinc.
Spiegel—Alloy of Fe, Mn containing Carbon.
Spinel—$MgAl_2O_4$—Mag aluminate (n).
Stibnite—Sb_2S_3—Antimony trisulphide (n).
Sylvine—KCl — Potassium chloride.
Talc—$3MgSiO_3, H_2O$ — Hydratcd Mag. silicate (n).
Tartar emetic — $2K(SbO)C_4H_4O_6, H_2O$ —Potash-antimony tartrate.
Tinstone — SnO_2— Tin oxide (n).
Topaz—(yellow pokraj or Pusparaj) — Al_2O_3 (n)
Tourmaline—Aluminium silicate (n).
Turnbull's blue—$Fe_3[Fe(CN_6)]_2$— Ferrous ferricanide.
Turquoise—Al-Phosphate (n) || Type metal—Alloy of Pb, Sb and Sn.
Ultramarine blue—Mix. of Sodi-Al-Silicate and Na_2S.
Venetian white — Mix. of white lead and $BaSO_4$.
Venetian red—Fe_2O_3, Ferric oxide || Verdigris—Basic acetate of Copper.
Vermilion—HgS—Mercuric sulphide || Vinegar—Dil. and impure Acetic acid.
Water glass—Thick solution of Sodium or Potassium silicate in water.
White lead—$2PbCO_3, Pb(OH)_2$—Basic carbonate of Pb.
Witherite — $BaCO_3$—Barium carbonate.
Wood's metal — Alloy of Bi, Pb, Sn and Cd.
Zinc-white—ZnO—Zinc oxide.

(n)—means that the substance occurs in nature,

Points to remember :

UNIT - I

→ Branches of chemistry, the science of matter and its transformations are Inorganic Chemistry, Organic Chemistry, Physical Chemistry, Analytical chemistry, Industrial Chemistry, Bio-chemistry, Nuclear-radiation Chemistry, Polymer Chemistry, Nano-Chemistry, Environmental Chemistry etc.

→ Units of measurement, CGS, FPS, MKS and SI systems of units.

→ There are seven fundamental units to express physical quantities.

e.g. Length (m), mass (kg), time (s), electric current (A), Temperature (K), Luminous Intensity (Cd - Candela), Amount of substance (mol)

→ **Derived units :** Pressure $= \frac{\text{Force}}{\text{area}} = {}^{\text{N}}\!/_{\text{m}^2} = \text{Kg m}^{-1}\,\text{s}^{-2} = \text{Pa}$

Electric potential (V) = $\text{Kg m}^2\,\text{s}^{-3}\,\text{A}^{-1}$

Electrical resistance (R) = $\text{Kg m}^2\,\text{s}^{-3}\,\text{A}^{-2} = \text{ohm} = \Omega$

→ **Temperature scales** (i) Centigrade scale, 0^0C to 100 ^{0}C = 100 dvns.

(ii) Fahrenheit scale 32^0F to 212^0F = 180 dvn

Relation, $F = \left(\frac{9}{5}C + 32\right)$

(iii) Absolute scale, T(K) = 273 + t^0C

→ **Matter :** Any thing that has definite mass and occupies space. These are of two categories :

(i) Homogeneous (one phase)

e.g. Water, sulfur, oxygen etc.

(ii) Heterogeneous (two or more phases) water + benzene

→ **Elements :** So far 118 elements have been discovered and are positioned in the periodic table. These are metals (Fe, Cu, Pb etc), Non-metals- (O, S, I, C etc) Metalloids with properties of metals and non-metals. Such as B, Si, Ge, As, Sb, Te, Po

→ **Compounds :** Chemical union of two or more elements H_2O, CH_3CHO.

(a) Stoichiometric compounds obeying laws of stoichiometry.

NiO, LaH_3, CH_4, CO_2, CO etc.

(b) Non-stoichiometric compounds. $Ni_{0.9999}O$, $LaH_{2.87}$

→ **Mixture :** (a) Homogeneous. liquid - liquid, water + alcohol

Solid - liquid salt solution

gas - gas air

solid - solid alloys. (steel, brass)

(b) Heterogeneous mixture. Gun power (nitre, sulfur, charcoal.

→ **States of matter :** Solid $\xrightarrow{\text{heat}}$ liquid $\xrightarrow{\text{heat}}$ Gas

$\xleftarrow{\text{cool}}$ $\xrightarrow{\text{cool}}$

Others are ionised plasma, Quark - Gluon plasma, Bose Einstein condensate, Fermionic condensate (super fluid state)

→ **Radicals :** Group of similar or different elements combined chemically forming a single unit with positive or negative charge, or neutral.

(a) Simple radicals, H^+, O_2^{2-}, Na^+, K^+

(b) Compound radicals CO_3^{2-}, SO_4^{2-}, $\overset{+}{N}H_4$

(c) Complex radicals $\left[Fe\left(CN\right)_6\right]^{3-}$, $\left[Cu\left(H_2O\right)_4\right]^{2+}$

(d) Neutral radicals are referred as free radicals formed by bond fission.

e.g. $(Cl\bullet)$

→ **Basic and acid radicals :** A salt consists of two radicals, basic (cation) derived from base and acid (anion) derived form acid.

e.g. $\overset{+}{Na}\,\overline{Cl} \longrightarrow \underset{\text{Base}}{\underline{N}aOH} + \underset{\text{acid}}{H\underline{Cl}}$

→ **Law of conservation of mass :** In a chemical reaction, mass of the reactants = mass of the products. Thus, it's necessary to balance a reaction.

→ **Law of definite proportion :** A given compound always contains the same elements combined in the same proportion by mass.

e.g. carbon dioxide consists of C and O in the ratio-

12 : 32 i.e., 3:8 by mass irrespective of nature of sources.

→ **Law of multiple proportion :** when two elements combine to form more than one compound the masses of the element that combine with fixed mass of other element are in a ratio of simple whole numbers.

$$(C+O) \rightarrow \left[\begin{array}{l} CO,\ 12:16 \\ CO_2\ \ 12:32 \end{array}\right.$$

Different masses of oxygen that combine with 12 gram of 'C' are in the ratio 1 : 2.

→ **Law of reciprocal or equivalent proportions :**

The masses of two or more different elements which separately combine with a definite mass of another element are either the same as or simple multiple of the masses of different elements when they combine amongst themselves.

H

$\boxed{H_2S}$ $\boxed{H_2O}$

S O

$\boxed{SO_2}$

In H_2S H : S = 2 : 32

H_2O H:O = 2 : 16

Hydrogen mass is definite (2 gram)

If S, O combine it will have same ratio

i.e., 32 : 16 (2:1) or ratio simple multiple to it.

In $SO_2 \Rightarrow S:O = 32:32 = 1:1$ (by mass)

Expected ratio, 2 : 1

Observed ratio 1 : 1

These two ratios are simple multiple of each other i.e, 2:1/ 1:1 = 2 : 1

If it is $SO_3 \Rightarrow S:O = 32:48 = 2:3$

Then, the simple multiple will be $\frac{2:1}{2:3} = 3:1$

→ **Gay-Lussac's law of gaseous combination :**

When gases combine they do so in simple ratio by volume to each other and to the gaseous products, all the volumes being measured at the similar conditions of temperature and pressure.

→ Atoms, the tiny particles of matter are indestructible and can neither be created nor destroyed by any chemical reaction, Dalton.

→ Atomic mass of an element = $\frac{\text{mass one atom of an element}}{\frac{1}{12}\text{th of the mass of } {}^{12}\text{C}}$

e.g. atomic of F = 9 u (amu = atomic mass unit = 1.67×10^{-24} gram)

i.e. 9 times heavier than $\frac{1}{12}$th mass of an atom of ^{12}C.

1 gram atom of F = 9 gram containing 6.023×10^{23} atoms

i.e. mass of an atom of F = $9 \times \frac{1}{6.023 \times 10^{23}}$ gram = $9 \times 1.67 \times 10^{-24}$ gram

→ Average atomic mass : Calculated from the number of isotopes and their abundance. e.g. Cl has two isotopes ^{35}Cl and ^{37}Cl with abundance 3:1. Average atomic mass of Cl = $\frac{(35 \times 3) + (37 \times 1)}{3 + 1} = 35.5$.

→ **Dulong-Petit law :**

Relative atomic mass × sp.heat = 6.4 (appx.) = atomic heat.

→ Molecular mass (Relative molecular mass) = $\frac{\text{mass of one molecule of an element / compound}}{\frac{1}{12}\text{mass of an atom of } {}^{12}\text{C}}$

e.g. molecular mass of nitrogen = 28 u.

1 gram molecule = 28 gram of nitrogen.

= 6.023×10^{23} molecule of nitrogen (N_2)

= $2 \times 6.023 \times 10^{23}$ atoms of nitrogen (Nitrogen is diatomic)

Actual mass of one molecule of nitrogen = $28 \times 1.67 \times 10^{-24}$ gram

= $28 \times 1.67 \times 10^{-27}$ kg

→ 1 mol = Molecular mass in gram = Avogadro's number of molecules

= $(6.023 \times 10^{23} \times \text{atomicity})$ no. of atoms

= 22.4 ℓ at NTP (if gaseous)

→ Equivalent mass : It is number by parts of mass that can combine or displace directly or indirectly 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of Cl.

→ Equivalent mass of an element / ion = $\frac{\text{atomic mass or ionic mass}}{\text{valency or charge}}$

e.g. eq. mass of $\text{Zn} = \frac{65}{2} = 32.5$, eq. mass of $SO_4^{2-} = \frac{96}{2} = 48$

→ Equivalent mass of acid or base $= \frac{\text{Mol. mass acid / base}}{\text{basicity / acidity}}$

basicity = no. of proton replaced.

e.g. Eq. mass of H_2SO_4 forming Na H SO_4 (salt)

i.e. one H^+ is replaced $= \frac{98}{1} = 98$

Eq. mass of H_2SO_4 forming Na_2SO_4 (salt)

i.e. two H^+ are replaced $= \frac{98}{2} = 49$. Thus, eq. mass is variable for polybasic acid.

→ Eqv. mass of salt = eqv. mass of cation + eqv. mass of anion.

e.g. eqv. mass of Na_2SO_4 = Eqv. mass of Na^+ + eqv. mass of SO_4^{2-}

$$= \frac{23}{1} + \frac{96}{2} = 71$$

→ Eq. mass of oxidant / reductant = mol. mass/change in 0.S.

$$H_2\overset{+6}{S}O_4 \longrightarrow H_2O + \overset{+4}{S}O_2 + O$$

Eq. mass of H_2SO_4 (oxidant) $= {}^{98}\!/_{2} = 49$

→ Empirical formula mass × n = Molecular mass

n = any integer (1, 2, 3, 4 ....).

→ Mol. mass = 2 × Vapour density

$$= \frac{\text{density of any gas}}{\text{density of hydrogen}} \text{ (T, P kept constant)}$$

$$= \frac{\text{wt of certain volume of any gas (vapours)}}{\text{wt of same volume of hydrogen}} \text{ (under same conditions of T and P)}$$

→ Diffusion method of determining molecular mass :

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

→ Eqv. mass : 1 F of electricity discharges 1 gram eqvt. of mass of stabstance.

In metal displacement method, $\frac{\text{wt of metal(1) added }(w_1)}{\text{wt. of metal (2) displaced }(w_2)} = \frac{E_1}{E_2}$

(E = eqvt. mass)

UNIT - II

→ Electron is a sub atomic particle, symbol $\bar{e}$ or β^- with charge $-1.6021\times10^{-19}\,C$ or -4.8032×10^{-10}esu, discovered by J.J. Thomson. $m_e = 9.10938356\times10^{-31}\,Kg$ $\left(\simeq \frac{1}{1837}\text{mass of proton}\right)$.

The name 'electron' was given by Irish physicist G.J. Stoney.

→ $e/m = 1.7599\times10^{8}\,C/gm$, Electron is the universal constituent of all matter.

→ Proton discovered by Goldstein with mass 1.672×10^{-27} Kg

→ Neutron (Chadwick) neutral particle with mass 1.675×10^{-27} Kg

→ Nucleus (Rutherford) with diameter in the order 10^{-13} cm where as atom in the order 10^{-8} cm.

→ Atomic number (Z) : The square root of the frequency of emitted X-rays is approximately proportional to atomic number (Z).

$\sqrt{\nu} = a(Z-b)$, a and b are constants.

→ Electromagnetic wave theory : the energy is emitted from any source continuously and consists of electric and magnetic fields. oscillating perpendicular to each other and to the direction of propagation.

→ Electromagnetic spectrum follows in the order of increasing wave length cosmic rays < γ -rays < X-rays < U.V < visible < IR < microwaves < Radio waves.

→ Rydberg's formula to calculate wave number and hence wave length of spectral lines of hydrogen spectra.

$$\bar{\gamma} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right], \quad R = \text{Rydberg constant} = 109677\ cm^{-1}$$

$= 1.097\times10^{7}\,m^{-1}$. For
Lyman series $n_1 = 1, n_2 = 2, 3, 4, 5$ .... ,
Balmer series, $n_1 = 2, n_2 = 3, 4, 5, 6$ ....,
Paschen series, $n_1 = 3, n_2 = 4, 5, 6$ ....,
Brackett series, $n_1 = 4, n_2 = 5, 6, 7$ ....,
Pfund series, $n_1 = 5, n_2 = 6, 7, 8$ ....,
Humphrey series $n_1 = 6, n_2 = 7, 8, 9$ ....,

$$\gamma = c\times\bar{\gamma} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right], \quad R = 3.29\times10^{15}\ \text{cycles/sec (or) } (m^{-1})$$

→ E_n = energy of the $\bar{e}$ in nth orbit

$$= -\frac{21.8\times10^{-19}}{n^2}J/atom$$

$$= -\frac{13.6}{n^2}ev/atom \qquad \because\ 1\ ev = 1.6022\times10^{-19}\ J$$

$$= -\frac{1312}{n^2}\ kJ\ mol^{-1}$$

$$= -\frac{2\pi^2 m z^2 e^4}{n^2 h^2} \quad \text{i.e., } E_n \propto \frac{z^2}{n^2}$$

→ radius of nth orbit, $r_n = \dfrac{0.529n^2}{z}$

→ Velocity of $\bar{e}$ in nth orbit, $V = \sqrt{\dfrac{Ze^2}{mr}}$ $\quad v \propto \dfrac{1}{\sqrt{r}}$

$$v_n = \frac{2.185\times10^6.Z}{n}m.s^{-1}$$

→ No. of spectral lines when $\bar{e}$ returns from n_2 to n_1 energy level

$$= \frac{(n_2-n_1)(n_2-n_1+1)}{2}$$

→ Photoelectric effect : $\frac{1}{2}mv^2 = h\nu - h\nu_0,\ h\nu_0 = \text{min imum}$ energy required to emit photoelectron = threshold energy or work function.

→ de-Broglie equation : $\lambda = \dfrac{h}{p} = \dfrac{h}{m\nu},\ \nu = \sqrt{\dfrac{2KE}{m}},$

$$2\pi r = n\lambda$$

→ Heisenberg's uncertainly principle : $\Delta x.\Delta p \geq \dfrac{h}{4\pi}$

$$(\Delta x.\Delta v) \geq \frac{h}{4\pi m}$$

→ Angular momentum of electron = $\sqrt{\ell(\ell+1)}\dfrac{h}{2\pi}$

→ No. of subshells in nth shell = n^2 and no. of e^- is $2n^2$.

UNIT - III

→ **Mandeleev's periodic law :** The properties of elements are periodic function of their atomic masses.

→ **Modern periodic law :** Atomic mass in the above law is replaced by atomic number.

→ **Ionic radius :** The distance between the centre of the nucleus upto which the electron cloud is extended.

→ $\Delta_{eg}H$, the electron gain enthalpy is the energy involved (released) when electron is added to the neutral gaseous atom.

If $\bar{e}$ affinity is 20 kJ/mols its $\Delta_{eg}H = -20kJ/mol$.

$$A_{(g)} + \bar{e} \rightarrow \bar{A}_{(g)} + 20\ kJ$$

→ Modern periodic table consists of

* 's'-block, group 1 and 2 with general electronic configuration, ns^1 and ns^2 n = 2 to 7 $(\text{except } H = 1s^1)$.

* ‘p’-block. group 13-18, configuration $ns^2 np^{1-6} = 2$ to 7 except He = $1s^2$.

* ‘d’-block, groups 3-12, period, 4th to 7th. Configuration $(n-1)s^2p^6d^{1-10}ns^{1or2}$ $n = 4$ to 7 (4 rows/series)

* f-block (Two rows at the bottom)
group-3, period 6th and 7th,

Lanthanoids $4s^2p^6d^{10}4f^{0\text{-}14}$ $5s^2p^6d^{0or1}6s^2$ or $[Xe]4f^{0-14}5d^{0(or)1}6s^2$ (4f series)

Actinoids. $5f^{1\text{-}14}6d^{0(or)1}$ $7s^2$ (or) $(Rn)5f^{1-14}6d^{0(or)1}7s^2$

→ ENC (Z^*) = $Z - \sigma$(screening constant)

→ Electronegativity in Pauling scale

$$\Rightarrow \quad X_A - X_B = 0.208\left[E_{A-B} - \left(E_{A-A}\times E_{B-B}\right)^{1/2}\right]^{1/2}$$

X_A and X_B electronegativity of atoms A & B

E_{A-A}, E_{B-B} and E_{AB} are bond energies of A–A, B–B and A–B. 0.208 arises due to a conversion factor from KCal to eV.

e.g. electronegativity values N > C > P > Si

Along the period the value increases and decreases down the group.

→ Reactivity of the element depends on its tendency to lose $\bar{e}$ or gain $\bar{e}$ e.g. Cs is the most reactive metal and F is most reactive nonmetal.

UNIT - IV

→ Chemical bond is the force which holds the atoms of a molecule together reducing the potential energy.

→ The strong electrostatic force of attraction between these oppositely changed ions is called ionic / electrovalent bond. Here the atoms are dissimilar with widely varying electronegativity values, (Δx almost greater than 1.7). Elecetrovalency of an atom is the number of electron (s) that an atom can lose or gain to have stable electronic configuration.

→ Covalent bond is formed by mutual sharing of electron so as to complete duplet or octet. The number of $\bar{e}$ (s) needed to satisfy octet / duplet on mutual sharing is called covalency of the atom e.g. covalency $Cl(=1)$, $O(=2)$, $N(=3)$, $C(=4)$.

→ The sharing pair of $\bar{e}$ determines the nature of single, double and triple bond. i.e., $:N\vdots\vdots N:$, $N\equiv N$.

→ As per orbital concept an orbital with one umpair electron can overlap axially or laterally to form σ or π bond respectively.

→ A sigma bond is always stronger than π bond due to greater extent of axial overlapping.

→ Bond length, bond angle, bond enthalpy, bond order, bond polarity have significant effect on properties of compounds.

→ Percentage of ionic character of a covalent bond

$$= \quad 16(X_A - X_B) + 3.5\,(X_A - X_B)^2 \quad \text{: Hannay - Smith equation}$$

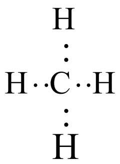
$$= \frac{\text{Actual (observed) dipolement}}{\text{Theoretical dipolement} = (\text{charge} \times \text{distance})} \times 100$$

→ Formal change (Fc) on atom in a Lewis structure =

[Total number of valence electrons in free atom] - [total number of lone pair $\overline{e}$ (s)]

$$-\frac{1}{2}[\text{total number of bonding / shared electrons}]$$

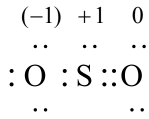
= No. of $\overline{e}$ in isolated state (g) – no. of $\overline{e}$ in combined state



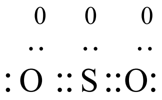
$$H \cdot\cdot \overset{\overset{H}{\vdots}}{\underset{\underset{H}{\vdots}}{C}} \cdot\cdot H$$

formal change on C = $4 - 0 - {}^{8}\!/_{2} = 0$

on H = $1 - 0 - {}^{2}\!/_{2} = 0$



$$\overset{(-1)}{:\ddot{\underset{..}{O}}}\ \overset{+1}{:\ddot{S}}::\overset{0}{\ddot{\underset{..}{O}}}$$



$$\overset{0}{:\ddot{O}}::\overset{0}{\ddot{S}}::\overset{0}{\ddot{O}}:$$

I II

	I	II
F.c.	$(O) = 6 - 6 - {}^{2}\!/_{2} = -1$	$0 = 6 - 4 - \frac{4}{2} = 0$
	$(S) = 6 - 2 - {}^{6}\!/_{2} = +1$	$S = 6 - 2 - \frac{8}{2} = 0$
	$(O) = 6 - 4 - \frac{4}{2} = 0$	$O = 6 - 4 - \frac{4}{2} = 0$

Generally the lowest energy structure is the one with the smallest formal charge on the atoms. Hence II is the structure SO_2.

→ Hybridisation is the intermixing of two or more atomic orbitals of comparable energy of same atom developing new hybrid orbitals having same energy, size and shape. The orienation is changed in such a manner that there occurs minimum repulsion among the hybrid orbitals. Hybrid orbital also contains maximum two $\overline{e}$(s) of opposite spin.

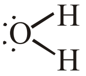
→ Hybrid orbital with unpair electron can combine with another hybrid or pure orbital forming stronger bond.

→ By means of hybridisation the shape, bond angle of different covalent molecules can be explained. e.g. sp^3 hybridisation provides three different types of molecules :

i) AB_4 e.g CH_4 (tetrahedral)

(ii) AB_3L (one hybrid orbital with pair of $\overline{e}$(s)) e.g. NH_3 (Pyramidal)

(iii) AB_3L_2 (Molecules with two lone pair of $\overline{e}(s)$.

e.g.  shape-Bent

→ Depending the mixing of orbitals hybridisations are of sp, sp^2 sp^3, dsp^2, sp^3d, sp^3d^2, sp^3d^3

→ **VSEPR theory :** The bond pairs and lone pairs present around central atom repel each other and hence move as far apart as possible to have minimum repulsion & more stability to the molecule / ion. The direction of the electron pairs gives a definite geometry due to electronic repulsion where $\ell-\ell$ repulsion > $\ell-b$ repulsion > b–b repulsion. Such repulsion is also governed by electronegativity of the central atom. e.g. Bond angle of $NH_3 > PH_3$.

→ According to the principle LCAO, bonding molecular orbital (BMO) is formed by addition of wave function of electron waves of atomic orbitals $(\psi_A+\psi_B)$, while antibonding m.o (AMO) is formed by substraction of wave function of the atomic orbitals $(\psi_A-\psi_B)$. $E_{BMO} < E_{AMO}$.

→ Only AO's of comparable energies and proper orientations combine to form MO's e.g. if Z is the molecular axis s-orbital can overlap with p_z orbital to form $\sigma-$bond so also $pz_{(A)}$ and $pz_{(B)}$ along z-axis.

→ The bonding MO's are represented as σ, π etc while the antibonding MO's aroas σ^*, π^* etc.

→ The shape of MO's formed depends on the type of the combining AO's.

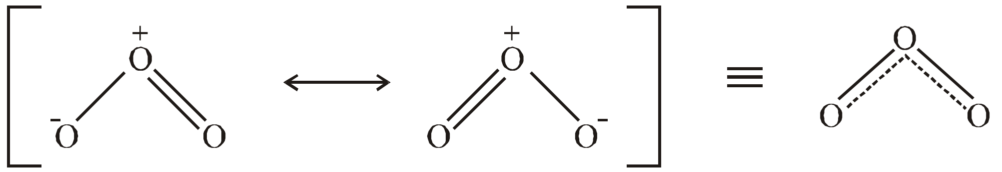
→ Filling of MOs follows the same rule as applicable to AO's.

→ The B.O = $\frac{1}{2}$ $(N_b - N_a)$, N_b = no. of $\bar{e}$(s) in BMO Higher the bond order smaller is the bond length and greater is the bond dissociation energy and greater is the stability.

→ No. of unpair $\bar{e}$ in the MOs determines the paramagnetic behaviour of the species.

→ **Resonance :** When a molecule is represented by two or more Lewis structures which differ in the position of electrons and not in the relative position of atoms and none of these structures can explain all the properties of the molecule, the actual structure of the molecule is said to be resonance hybrid of all these Lewis structures and these Lewis structures are called resonating / canonical / contributing structures.

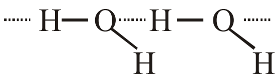
e.g.



Resonating structures **Resonance Hybrid**

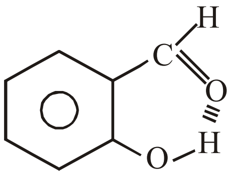
→ Hydrogen bond : Hydrogen of one molecule with most electronegative atom (F/O/N) of another molecule exhibiting dipole-dipole force.

i)



(Inter molecular H-bonding - molecules of same compound)

ii) Intramolecular (within the molecule itself)



(Salicylaldehyde)

UNIT - V

→ Pressure, volume, temperature, density etc. are the measurable properties of gases.

→ Diffusivity is a distinguished property of gas for which it's homogeneous. More than one gas always constitutes a single gaseous phase.

→ Boyle's law : $P_1V_1 = P_2V_2$ at constant temperature for a given mass of gas $\left(n = \frac{W}{M}\right)$.

→ Charles' law : $V_t = V_0\left(1+\frac{t}{273}\right)$ i.e., $V \propto T$ where $T = 273 + t^0 C$

Here P and 'n' are constants. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$.

→ Amonton's law : $P_t = P_0\left(1+ t/273\right)$ i.e., $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

→ Ideal gas equation : $PV = nRT$

R = Universal gas constant, = 0.0821 ℓ atom. $mol^{-1}\,K^{-1}$
= 1.987 cal $mol^{-1}\,K^{-1}$
= 8.314 J $mol^{-1}\,K^{-1}$
= 8.314 kP d_m^3 $mol^{-1}\,K^{-1}$ (P = Pascal)

→ $P = \frac{dRT}{M}$, $d = \frac{PM}{RT}$ (M=mol mass in gram mol^{-1})

$$= \frac{atm \times gm\,mol^{-1}}{L.atm.mol^{-1}\,K^{-1} \times K} = gm/L$$

→ $p_i = x_i P$ i.e, partial pressure of i^{th} gas in the mixture at given temperature = mole fraction of i^{th} component × total pressure

→ Real gas equation (vander Waals) : $\left(p + \frac{an^2}{V^2}\right)(V - nb) = n\,RT$ 'a', the constant is a measure of inter molecular attractive force within the gas independent of temperature and pressure. 'b' is the covalume i.e., volume of one mole of gas (excluded volume)

→ Compressibility factor (Z) = $\frac{PV}{n\,RT}$, $= V_{real}/V_{ideal}$.

Z = 1 for ideal gas.

→ T_B (Boyle temperature) The temperature at which a real gas obeys ideal gas laws over an appreciable range of pressure. $\left(T_B = \frac{a}{bR}\right)$, a & b are vander Walls gas constants.

→ Kinetic gas equation : $PV = \frac{1}{3} m N \bar{c}^2$, m = mass of a single molecule, N = total no. of molecules.

$$\overline{c} = \text{root mean square velocity} = \sqrt{\frac{n_1c_1^2 + n_2c_2^2 + ...}{n_1 + n_2}}$$

For 1 mole = mN = mN_A = M (molecular mass)

$$\overline{c} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3RT}{M}}$$

→ Average speed $(\nu) = \left(\frac{n_1\nu_1 + n_2\nu_2 + ...}{n_1 + n_2 + ...}\right) = \sqrt{\frac{8RT}{\pi M}}$

→ Most probable velocity $(\alpha) = \sqrt{\frac{2RT}{M}}$

$$\alpha : \nu : \overline{c} = 1 : 1.128 : 1.224$$

→ Critical temperature of a gas (T_C) is the temperature above which it cann't be liquefied how so ever high be the applied pressure

eg T_C (CO_2) = 31.1^{0}C = $\frac{8a}{27Rb}$

→ Critical volume at temperature T_C = 3b and $P_C = \frac{a}{27b^2}$

→ Compressibility factor at critical state, $z_C = \frac{P_C V_C}{RT_C} = \frac{3}{8}R.$

→ Surface tension of a liquid is defined as the force acting at right angle to the surface along unit length of the surface. Unit. dynes/cm or N/m (SI-unit)

→ Co-efficient of viscosity : It is the force of friction required to maintain a unit velocity (1cm/sec) difference between two parallel layers of unit distance (1-cm) having an area of unit crossection (1 cm^2)

$$f \propto A\frac{d\nu}{dx} \qquad \text{i.e.,} \left(f = \eta A\frac{d\nu}{dx}\right) \qquad (\nu = \text{velocity, } A = \text{area})$$

$$\eta = \frac{f\ dx}{A\ d\nu} = \frac{\text{dynes} \times \text{cm}}{\text{cm}^2 \times \text{cm/sec}} = \text{dynes cm}^{-2}\ \text{sec}.$$

1 poise = g cm^{-1} s^{-1} = $\frac{N.m}{m^2.m s^{-1}} = Nm^{-2}\ s$ f = g.cm sec^{-2}

= Pa.s (SI unit) = mass × acceleration

→ Surface tension and viscosity are due to strong inter molecular attractive forces which decrease with increase of temperature.

UNIT - VI

→ Thermodynamics deals with energy changes in chemical or physical process and enables us to study these charges qualitatively & to make useful predications.

→ Internal energy : The total energy stored in the substance by virtue of its chemical nature.

→ First law of thermodynamics. Energy can neither be created nor destroyed.

$$U_2 = U_1 + q + w$$

$$\Delta U = q + w, \text{ work of expansion} = -P\Delta V$$

$$q = \Delta U + P\Delta V$$

For an ideal gas undergoing an isothermal change, $\Delta U = 0$, have $q = -w$

i.e. heat absorbed by the system is work done by the system.

(work of expansion).

→ A part of the universe chosen for thermodynamic studies is called a system & the remaining part of the universe is called surrounding. system : (i) closed system, energy but not mass gets transferred. (ii) open system, mass and energy both transfer (iii) Isolated system : neither matter nor energy of the system changes with surrounding.

→ Properties of substance that depend on quantity are called extensive properties e.g. mass, volume, energy, heat capacity, enthalpy, free energy etc. while the properties like temperature, pressure, refractive index, density, surface tension, lelting point, boiling point depend on nature - not on quantity are calles intensive property.

→ A process which is carried out infinitesimally slowly is called reversible process.

→ $\Delta U = q_{(v)},\ \Delta H = q_{(p)},\ \Delta H = \Delta U + P\,\Delta\,V \Rightarrow \Delta U + \Delta n_g\,RT$

→ Work done in isothermal reversible expansion of an ideal ges,

$$W = -2.303\ n\ RT\ \log\frac{V_2}{V_1} = -2.303n\ RT\ \log\frac{p_1}{p_2}$$

→ Calorific value of a fuel or food is the amount of heat evolved in cal. when on complete combustion of 1 gram of food or fuel.

→ $\Delta_{neut}.H$ of strong acid - strong bose is always constant,

$$H^{H}_{(aq)} + OH^{-}(aq) \rightarrow H_2O + 57.1\,kJ/mol$$

→ $\Delta S = \frac{q_{rev.}}{T}\,J\,K^{-1}\,mol^{-1}$ e.g. entropy of fusion, $\Delta_{fus}S = \frac{\Delta_{fus}H}{m.p}$

→ $\Delta G = \Delta H - T\Delta S,\quad \Delta G^0 = -2.303\,\log K$

→ Second law of thermodynamics : It is impossible to convert heat completely to an equivalent amount of work without producing any change in the system or surrounding.

→ Third law of thermodynamics. At absolute zero temperature the enthalpy of every substance may become zero i.e. it's a perfectly crystalline solid i.e. at T = O, $\Delta G = \Delta H = 0$

$$\frac{\partial(\Delta H)}{\partial T} = -\Delta S \quad \text{and} \quad \frac{\partial\,\Delta H}{\partial T} = \Delta C_p$$

$$\underset{T\rightarrow 0}{\ell t}\quad \Delta S = 0 \quad \text{and} \quad \underset{T\rightarrow 0}{\ell t}\quad \Delta C_P = 0$$

UNIT - VII

→ A reaction (reversible) is said to be in an equilibrium state when two opposite reactors are in same rate.

→ Dynamic nature of equilibrium. Equilibrium is dynamic an even after equilibrium is attained the reactions (forward & backward) never stop.

→ At equilibrium the concentration of each of the reactants & products becomes constant.

→ Chemical equilibrium can be approached from either side.

(reactant side / product side)

→ Presence of catalyst affects the time of approaching the equilibrium, but does not alter the state equilibrium, i.e., the equilibrium constant.

→ $K_P = K_C(RT)^{\Delta n} = K_X(p)^{\Delta n}$

→ $Q = \text{reaction quotient} = \dfrac{[X]^x[Y]^y}{[A]^a[B]}$ for the equilibrium : $aA + bB \rightleftharpoons xX + yY$

i) $Q =$ K the reaction is in a state equilibrium.

$$\Delta G = \Delta G^0 + RT\ \ln Q,$$

at equilibrium, $\Delta G = 0$ i.e $-\Delta G^0 = RT \ln K$

ii) $Q > K$, $\Delta G =$ positive, the reaction proceeds in back ward direction.

iii) $Q < K$, $\Delta G = -ve$, the reaction proceeds in forward direction.

→ van't Hoff equation : $\dfrac{d \ln K_p}{dT} = \dfrac{\Delta H^0}{RT^2}$

$$\log \frac{K_{p(2)}}{K_{p(1)}} = \frac{\Delta H^0}{2.303R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$$

$K_{P(2)}$ = Equilibrium constant at temperature T_2.

→ K_p, K_C are dimension less quantity as expressed in the form activity where concentration / pressure are with respect to standard state i.e. 1 mole/L or standard state pressure of 1 bar.

e.g. pressure of 5 bar with respect to standard state pressure of 1 bar means 5 bar/1-bar = 5. (dimension less) similar in case of concentration.

→ Addition of inert gas at constant volume has no effect on the state of equilibrium where as at constant pressure, the equilibrium shifts in direction of increase of no. of moles.

→ Le-chatclier's principle : the favourable conditions for better yield of NH_3 by Haber's process–

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} + 92.5\,kJ$$

i) Optimum temperature of 500^0C.

ii) High pressure of 200 to 900 atm pressure.

iii) High concentration of pure and dry N_2 and H_2

iv) Finely divided iron as catalyst in presence of Mo as promoter.

→ Henderson equation - Acid buffers, $p^H = pKa + \log\dfrac{[\text{salt}]}{[\text{acid}]}$

Basic buffer $p^H = 14 - pK_b - \log\dfrac{[\text{salt}]}{[\text{base}]}$

→ For polyprotic acid $Ka = (K_a)_i \cdot (Ka)_2 .....$

→ A strong acid has weak conjugate base and viceversa.

→ Dissociation constant of water, $K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$

$[H_3O^+][OH^-] = Kw$ (ionic product of water)

Dissociation / ionisation constant of water, $K = \frac{K_w}{55.55 \text{ (molarity of water)}}$

→ Concentration of $[H^+]$ and $[OH^-]$ increases and hence Kw (increases with rise of temperature

e.g. $K_w(298K) = 1.0 \times 10^{-14}$, $K_w(323K) = 5.5 \times 10^{-14}$

→ $pH + pOH = p^{Kw} = 14$, Ka. Kb = Kw

$HF \underset{}{\overset{Ka}{\rightleftharpoons}} H^+ + F^-$, $F^- + H_2O \overset{Kb}{\rightleftharpoons} HF + OH^-$

→ Salt hydrolysis : (i) weak acid strong base, $K_h = \frac{Kw}{Ka}$

$$pH = 7 + \frac{1}{2}(pKa + \log C)$$

(ii) Salt of strong acid and weak base : $K_h = \frac{K_w}{K_b}$

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

(iii) Salt of weak acid and weak base : $K_h = \frac{K_w}{K_a.K_b}$

$$pH = 7 - \frac{1}{2}(\log K_a - \log K_b)$$

$$= 7 + \frac{1}{2}(pK_a - pK_b)$$

→ Solubility product of salt $AxBy = K_{sp} = [A^{y+}]^x.[B^{x-}]^y$

UNIT - VIII

→ Oxidation number is the residual charge left on an atom of a compound / ion where other atoms are removed as ion.

→ Redox reaction : The reaction where oxidation and reduction takes place simultaneously. Types :

i) Combination (Synthesis) reaction : $\overset{0}{H_2} + \overset{0}{Cl_2} \longrightarrow 2\overset{+1}{H}Cl^{-1}$

ii) Decomposition reaction : $2\overset{+}{K}\overset{+5}{Cl}\overset{-6}{O_3} \longrightarrow 2KCl^{-1} + 3\overset{0}{O_2}$

Oxidation : increase of O.N./ removal of $\overline{e}$

Reduction : Decreases of O.N. / addition of $\overline{e}$

iii) Metal displacement reaction : $\overset{+2}{Zn} + \overset{+2}{Cu}SO_4 \longrightarrow \overset{0}{Zn}SO_4 + \overset{0}{Cu}$

iv) Non-metal displacement reaction :

$$\overset{0}{Ca} + 2\overset{+2}{H_2}O^{-2} \longrightarrow \overset{+2}{Ca}(OH)_2 + \overset{0}{H_2}(g)$$

$$\overset{0}{Cl_2} + 2KI^{-1} \longrightarrow 2K\overset{+2}{Cl} + \overset{0}{I_2}$$

→ Equivalent mass of oxidant / reductant = $\dfrac{\text{Mol. mass}}{\text{Change in O.S. / No. of } \overline{e} \text{ involved}}$

$$\overset{+7}{M}nO_4^- + 8H^+ + 5\overline{e} \longrightarrow \overset{2+}{Mn} + 4H_2O$$

Equivalent mass of $KMnO_4$ is acidic medium $= \dfrac{\text{Mol. mass}}{5}$

→ Disproportion reaction : This is a redox reaction where same element undergoes oxidation and reduction.

$$2H_2S^{-2} + \overset{+4}{S}O_2 \longrightarrow 3\left(\overset{0}{S}\right) + 2H_2O$$

$$H_2O_2^{-2} + H_2O_2 \longrightarrow 2H_2\overset{-4}{O} + O_2^0$$

→ The maximum oxidation number of an element equal to no. of valence electrons(n) present in it and the minimum is (n–8).

e.g. maximum O.N. of 'S' in its compound is +6 (SO_3) and minimum, 6–8 = –2 (H_2S).

→ Spectactor ions : Ions those are present in solution but does not involve in the redox process. e.g.

$$Zn + 2H^+ + 2Cl^- \longrightarrow Zn^{2+} + 2Cl^- + H_2$$

$$\text{i.e. } Zn + 2H^+ \longrightarrow Zn^{2+} + H_2$$

The ion gets omitted in the net reaction, Here Cl^- ions are spectactor ions.

→ Redox reactions are basis of titration :

e.g. (i) Permanganometry (estimation of iron) / dichrometry

$$1 \text{ gram eqvt. of } MnO_4^- \text{ or } Cr_2O_7^{2-} \equiv 1 \text{ gram eqvt. of } Fe^{2+}$$

$$N_1V_1 = N_2V_2$$

(ii) Iodometry : Liberated iodine is titrated against thiosulfate solution where starch as indicator (adsorption indicator)

$$2Cu^{2+}_{(aq)} + 4I^-(aq) \longrightarrow Cu_2I_2(s) + I_2(Aq)$$

$$I_2(aq) + 2S_2O_3^{2-}(aq) \longrightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

→ Redox reactions are also involved in designing electrochemical cell (redox couple)

$$\mathrm{Zn}|\mathrm{ZnSO_4}||\mathrm{CuSO_4}|\mathrm{Cu}$$

$$\mathrm{Zn} + \mathrm{Cu}^{2+} \longrightarrow \mathrm{Zn}^{2+} + \mathrm{Cu}$$

Reduction potential of copper (Cathode) is greater than Zinc electrode (Anode).

UNIT - IX

→ $\mathrm{H} - e^- \longrightarrow \mathrm{H}^+$, resembles with elements with group 1

$\mathrm{H} + e^- \longrightarrow \mathrm{H}^-$, similarities with group 17.

→ Hydrogen exhibits in three isotopic forms ${}^1_1\mathrm{H}$ (Protium – 99.985%),

${}^2_1\mathrm{H}$ (Deuterium - 0.0156%), ${}^3_1\mathrm{H}$ (Tritium - radioactive 10^{-15}%)

$${}^{14}_{7}\mathrm{N} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{12}_{6}\mathrm{C} + {}^{3}_{1}\mathrm{H}$$

→ Isotopic effect : The difference in properties due to difference in atomic masses is called isotopic effect.

e.g. Reaction with water (H_2O) is much faster than that with D_2O

→ o–H_2 and p–H_2 : These are isomers of hydrogen where nuclear spin parallel in o–H_2, $\left(+\frac{1}{2}+\frac{1}{2}=1\right)$

and antiparallel in p–H_2 $\left[+\frac{1}{2}+\left(-\frac{1}{2}\right)\right]=0$

At room temperature, ordinary dihydrogen contains 75% of o.H_2 and 25% p.H_2.

→ The name hydrogen is referred to a isotopic mixture of H^+, D^+ and T^+ while H^+ (proton).

→ Depending upon water's behaviour towards soap, water that produces good lather with soap is called soft water & that does not give good lather with soap is hard water. Presence of Mg^{2+} and Ca^{2+} ions in water produces insoluble stearate or palmitate or oleate with soap inhibiting formation of good lather.

→ Presence of $Ca(HCO_3)_2$ or $Mg\ (HCO_3)_2$ produces temporary hardness as these can be removed simply by heating, whereas presence of $CaCl_2$, $CaSO_4$, $MgCl_2$ and $Mg\ SO_4$ causes permanent hardness, also called non-carbonate hardness.

→ Calgon (calcium gone), sodium hexametaphosphate, $Na_2[Na_4(PO_3)_6]$ is a water softner.

$$\underset{\text{From hard water}}{2\,\mathrm{CaCl_2}} + \mathrm{Na_2}\left[\mathrm{Na_4}\left(\mathrm{PO_3}\right)_6\right] \rightarrow \underset{\text{Complex salt (soluble)}}{\mathrm{Na_2}\left[\mathrm{Ca_2}\left(\mathrm{PO_3}\right)_6\right]} + 4\,\mathrm{NaCl}$$

→ Hydrogen acts as a reducing agent in many reaction including formation of vegetable ghee. It is used as rocket fuel. It has promising potential for use as non-polluting future fuel (Hydrogen economy).

→ In the formation of metallic hydride, hydrogen acts as an oxidising agent.

$$2\,\overset{0}{\mathrm{Na}} + \overset{0}{\mathrm{H_2}} \longrightarrow 2\,\overset{+}{\mathrm{Na}}\overset{-1}{\mathrm{H}}$$

→ Heavy water (D_2O) is manufactured by the electrolytic enrichment of normal water. It is used as moderator in nuclear reactor.

→ H_2O_2 is a non-planar, non-polar molecule with the properties of acidic oxidising agent, reducing agent and bleaching agent.

→ Volume strength of hydrogen peroxide : x vol. H_2O_2 means 1 ml of it gives 'x' ml of oxygen at NTP.

UNIT - X

→ s-block constitutes group 1 and 2 with general valence shell configuration ns^1 and ns^2 forming cations of M^{+1} and M^{+2} respectively.

→ There is a regular trend is physical and chemical properties with increasing atomic number, atomic / ionic radii, decrease of ΔiH values.

→ Among the elements of period 2 and period - 3 there present diagonal relationship due to comparable ionic potential.

e.g. Li is diagonally related to Mg.

→ Caustic soda is manufactured from sodium chloride by Castner-Kellner process and washing soda by Solvay process.

→ The partial covalent character in ionic compounds results through polarisation of the anion by cation as per Fajan's rule :

(i) Small cation - LiCl is more covalent than NaCl

(ii) Large anion - KI is more covalent than KCl

(iii) Large charge on cation or anion - $AlCl_3$ is more covalent than $MgCl_2$ and NaCl.

(iv) Pseudo inert configuration. AgCl is more covalent than KCl

→ Mg present in chlorophyl while calcium is present in bones and teeth. The calcium concentration in plasma is regulated at about 100mg/L. It is maintained by two hormones :

Calcitonin and parathyroid hormone. Bone is not unchanging, it is continuously being solubilised & redeposited to the extent of 400 mg per day in a man. All this calcium passes through the plasma.

→ Na and K function to mainain the normal osmotic presure of the different body fluids & thus protect the body against excessive loss of fluids. Gastric acid (HCl), base, pancreatic juice and bile are derived from blood sodium and potassium salts.

→ Oxone is a mixture of Na_2O_2 and dil HCl, used to bleach delicate fibres.

UNIT - XI

→ **p-block elements :** The valence shell electronic configuration is $ns^2\ np^{1-6}$ (n=2 to 7) 'He' ($1s^2$) is exceptional. Groups constitute 13 to 18. [Boron family to noble gases (aerogens)]

→ Inert pair effect : The reluctance of ns^2 electrons to take part in bonding, due to ineffective shielding of $\boxed{\uparrow\downarrow}$ electrons of valence shell by intervening 'd' and f- electrons. This effect increases down the group i.e, the bottom placed elements showing two oxidation states e.g. group-13, I and III; group-14, II and IV; group 15, III and V. Stability of lower oxidation increases towards the bottom. e.g. Tl is stable in I o.s. while B and Al are stable in III o.s..

→ Diborane (B_2H_6) BH_3 is unstable, an electron definient molecule exhibits two e^-(s)- three centre bridge (banana type) bond where B-undergoes sp^3 hybridisation.

→ Diamond, graphite and Fullerene are three crystalline forms of carbon. Fullerines C_{60}, C_{70}.....C_{350} (even) are only pure form of carbon, because they have smooth structure without having 'dangling' bonds.

→ Buck minster fullerene (C_{60}) is having 20-six membered rings and 12-five membered rings, each C- undergoing sp^2 hybridisation.

→ The remaining $\overline{e}$ at each C-atom is delocalised in molecular orbital providing aromatic character to the molecule containing single (143.5 pm) and double (138.3 pm) bonds.

→ CO is neutral wheress CO_2 is acidic. CO is a ligand due to presence of lone pair $\overline{e}(s)$ on C, forming a number of metal carbonyls (organo-metallic compounds). It is deadly poisonous due to higher stability of its hameoglobin complex in comparison to oxyhaemoglobin complex.

→ Increased content of CO_2 in atmosphere due to combustion of fossil fuels and decomposition of lime stone is feared to cause increase in 'green house effect', global warming.

→ Freons are mixed CFC, $CFCl_3$, CF_2Cl_2 , CF_3Cl' and used as refrigerator fluid.

→ Silicones are long chain organosilicon polymers. These are water repellants, good insulators.

→ Zeolites (silicates) have honeycomb like structure that absorbs or loses certain molecules and acts as molecular seives. Sodium zeolite used as cation exchanger for softening hard water.

→ Zeotite, ZSM-5 used to convert alcohol directly to gasoline.

→ Silicon carbide ($SiO_2 + 3C \longrightarrow SiC + 2CO$) used as abrassive material under the name carborundum.

UNIT - XII

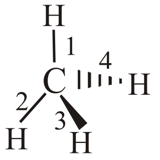
→ Organic compounds form inorganic substances -

(i) $NH_4Cl + KCNO \xrightarrow[-KCl]{} NH_4CNO \xrightarrow{\Delta} H_2NCONH_2\,(\text{Urea})$ - F. Wohler-1828

(ii) From constituting elements :

$$2C(s) + 2H_2(g) + O_2(g) \longrightarrow CH_3COOH(\ell), \qquad \text{Kolbe (1845)}$$

$$C(s) + 2H_2(g) \longrightarrow CH_4(g), \qquad \text{Berthelot (1856)}$$

→ 

1, 2 – bond in the plane of the paper

3 – Solid wedge (bond towards the observer)

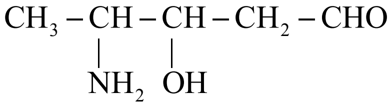
4 – Dashed wedge (bond away from the observer) lies below the plane of paper.

→ A functional group is an atom or group of atoms bonded together in a unique fashion and which determines the physical and chemical properties of the compounds.

→ During IUPAC naming the decreasing priority of some functional groups is ________ .

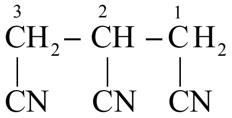
$-COOH$, $-SO_3H$, $-COOR$, $-COCl$, $-CONH_2$, $-CN$, $-CHO$, $>C=O$, $-OH$

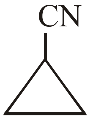
$-NH_2$, $>C=C<$, $-C\equiv C-$. Former one in the sequence is chosen as suffix & the following ones are as prefixes.

e.g. 
$CH_3-\underset{NH_2}{\underset{|}{CH}}-\underset{OH}{\underset{|}{CH}}-CH_2-CHO$ ⇒ 4–amino-3-hydroxypentanal.

→ –R, C_6H_5 –, halogens, $–NO_2$, – OR etc. are always prefix substituents.

→ The functional group having –C– is suffixed differently when present / repeated more than twice in a chain or bonded to cyclic / aromatic rings :


$\overset{3}{\underset{CN}{\underset{|}{CH_2}}}-\overset{2}{\underset{CN}{\underset{|}{CH}}}-\overset{1}{\underset{CN}{\underset{|}{CH_2}}}$ Propane -1, 2, 3– tricarbonitrile.


Cyclopropanecarbonitrile

Similarly, – CHO (–carbaldehyde), –COOH (Carboxylic acid),
– COCl (–Carbonylchloride), $–CONH_2$ (–carboxamide),
– COOR (alkyl – carboxylate)

→ Fission of covalent bonds leads to form free radicals (homolysis) and carbocation / Carbanion on heterolysis.

→ A reagent rich of $\bar{e}$ pair is nucleophile, Nu (:) (i.e., nucleus loving) and electron deficient / $\bar{e}$ -seeking species, electrophile $\left(E^+\right)$.

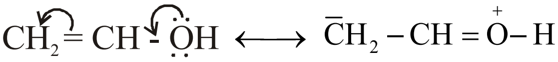
→ Shifting of $\bar{e}$ through covalent bond (σ) towards more electronegative atom/(EWG) causes –I effect.

$$\overset{\delta\delta\delta+}{CH_3} \longrightarrow \overset{\delta\delta+}{CH_2} \longrightarrow \overset{\delta+}{CH_2} \longrightarrow \overset{\delta-}{C\ell}$$

–C–Cl bond induces polarity which decreases with distance, almost negligible after three bonds.
Electron-releasing groups exhibits +I effect.

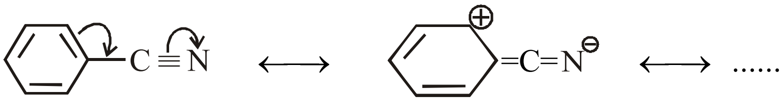
$$\left(CH_3\right)_2CH \longrightarrow, \left(CH_3\right)_3C \longrightarrow$$

→ Groups (i) which donate electrons to the double bond or a conjugated system are said to be +R or +M effect.
e.g. –OH, –OR, –SH, –SR, $–NH_2$, –NHR, $–NR_2$, –X etc.


$CH_2=CH-\ddot{O}H \longleftrightarrow \bar{C}H_2-CH=\overset{+}{O}-H$

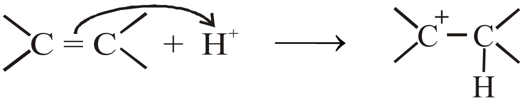
Groups (ii) which withdraw $\bar{e}$ (s) from the double bond or from a conjugated system are said to be –R or –M effect

e.g. >C=O, – CHO, –COOR, –CN, $–NO_2$, –NO etc.

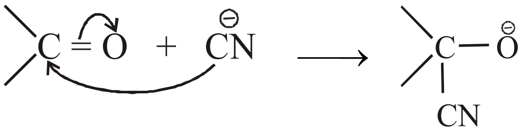
– Reffect : 

→ Electromeric effect (Temporary effect contrary to inductive effect)

+E effect : $\bar{e}$(s) of the double / triple bond are drawn towards the added reagent (electrophile).

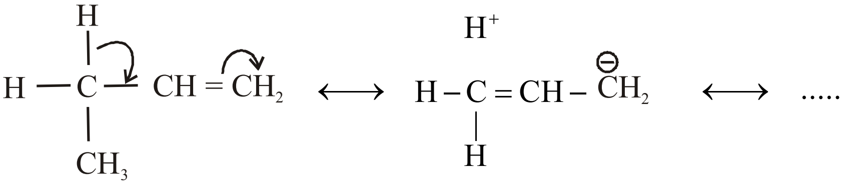


– E effect electrons draws from the reagent (nucleophile)



→ Hyper conjugation : (Baker - Nathan effect or no-bond resonance)

Delocalisation of σ – electrons through overlapping of p-orbitals of a double bond.



(other two hyper conjugatice structures.)

H^+ is found binding quite firmly to the π – cloud & hence not free to move. No. of α – H atoms determines the number of hyper-conjugative structures & greater stability.

e.g. i) $CH_3 - CH = CH - CH_3$ is more stable than $CH_3 - CH = CH_2$

Six – α . H atoms — Three- α –H atoms.

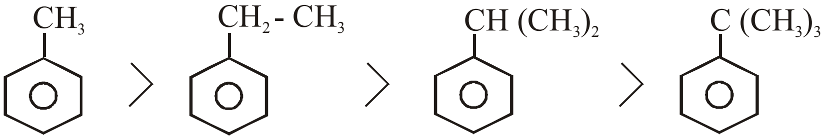
ii) Number of α -H atoms at $\overset{*}{C}$ decreases in the order

$$\overset{*}{C}H_3 - > CH_3 - \overset{*}{C}H_2 - > (CH_3)_2\overset{*}{C}H - > (CH_3)_3\overset{*}{C} -$$

(3) (2) (1) (0)

and decrease of hyper-conjugative structure

e.g.



iii) Stability of Carbocation -

$$(CH_3)_3\overset{+}{C} > (CH_3)_2\overset{+}{C}H > CH_3\overset{+}{C}H_2 > \overset{+}{C}H_3$$

In $(CH_3)_3\overset{+}{C}$ nine–CH bonds (9 pairs of σ -electrons) and carbocation has vacant p-orbital. $\overset{+}{C}H_3$ lacks hyper conjugative stability.

→ Different types of organic reactions -

(i) Substituation (ii) addition (iii) elimination (iv) rearrangement

→ The common techniques used for purification are :

(i) Sublimation (ii) Crystallisation (iii) Distillation (iv) Differential extraction

(v) chromatography.

→ Lassaigne's test is performed to detect 'N', 'S' and 'X' in organic compounds.

On sodium fusion : $Na + C + N \xrightarrow{\Delta} NaCN$

$$Na + S \xrightarrow{\Delta} Na_2S$$

$$Na + X \xrightarrow{\Delta} NaX \quad (X = Cl,\ Br,\ I)$$

Test for 'N' : $$6\,C\bar{N} + \underset{(FeSO_4)}{Fe^{2+}} \rightarrow \left[Fe(CN)_6\right]^{4-} \xrightarrow[(FeCl_3)]{Fe^{3+}} \underset{\text{Prusian blue}}{Fe_4\left[Fe(CN)_6\right]_3}$$

Test for 'S' $$S^{2-} + \underset{\text{nitropruside}}{\left[Fe(CN)_5\,NO\right]^{2-}} \longrightarrow \underset{\text{Violet}}{\left[Fe(CN)_5\,NOS\right]^{4-}}$$

Test for 'Cl' : The colour of precipitate with $AgNO_3$ determines the presence of Cl / Br / I.

AgCl (white) soluble in dil NH_4OH,

AgBr (Pale yellow) soluble in conc. NH_4OH,

AgI (yellow) insoluble in both dil NH_4OH and conc. NH_4OH.

Test for presence of 'P' in organic compound :

Compound with 'P' + Na_2O_2 (oxidant) → Phosphate $\xrightarrow[\text{(ii) }(NH_4)_2M_0O_4(Aq)]{\text{(i) }HNO_3\text{ boiled}}$ Yellow colour.

The yellow coloration on precipitate indicates formation of ammonium phosphomolybdate, $(NH_4)_3\ PO_4.\ 12\ MoO_3)$.

→ The % of different elements including C and H can be detected quantitatively.

e.g. (a) N can be detected by (i) Duma's method (ii) Kjeldahl's method.

(b) Halogen ans S by Carius method.

→ Carbocations carbanions, free radicals, carbens are the reaction intermediats by which mechanism of different organic reactions can be explained.

e.g. Chlorination of hydrocarbon is an example of free radical substitution reaction where

3^0 free radial > $2^0 > 1^0 > \cdot CH_3$ (order of stability)

→ Two or more compounds possessing the same molecular formula but different properties are known as isomers, the process is known as isomerism : such as structural isomerism and stereo-isomerism.

UNIT - XIII

→ Petroleum is the chief source of aliphatic hydrocarbons while coal is the main source of aromatic hydrocarbon.

→ Natural gas : Petroleum under the surface of earth is covered with a gaseous mixture called natural gas consisting of CH_4(90%) along with ethane, propane and butane and vapours of low boiling pentanes and hexanes.

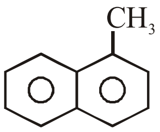
→ Liquified petrolium gas (LPG) is a mixture of propane and butane mainly used as fuel.

→ Octane number : % of isooctane (2,2,4–trimethyl pentane) by volume in mixture of iso-octane and n-heptane. Octane number - 80, means iso-octane 80% and n-heptane 20% by volume. n-nonane has octane number – 45, where 2,2,3-tri methyl butane (triptane) has an octane number of 124.

→ TEL, $(C_2H_5)_4Pb$ added to the extent 0.01% to gasoline, called gasoline or leaded gasoline, increases its octane number i.e., fuel quality.

→ Cetane number : Percentage of cetane by volume of cetane & α-methyl naphthalene which has same ignition properties as the diesel under examination in a test engine.

$CH_3 - (CH_2)_{14} - CH_3$
Hexa decane or cetane
Cetane no.100 ignites
spontaneously


α-methyl naphthalene,
Cetane no. Zero, ignites slowly,

→ Alkanes are saturated a cyclic hydrocarbons having the general formula $C_n H_{2n+2}$ where n = 1, 2, 3 .... In these C– is sp^3 - hybridised, non-polar, inert towards most of the reagents (called paraffins). However alkanes undergo free radical substitution reaction, isomerisation, cyclisation, etc with suitable reagents.

→ Marsh gas (CH_4) can be formed by decay of plants and animal tissues.

→ Alkenes and alkynes undergo electrophilic addition reaction while aromatic compounds (benzene and its derivative) undergo electrophilic substitution reaction.

→ Electron releasing groups (–X, –R, $-OCH_3$, $-NH_2$, –OH etc) are o/p directing while EWGs ($-NO_2$, – NO, –CHO, $-\underset{|}{C}O$, – COOH, –COX, $-CONH_2$ ....) are m-directing.

→ Saturated hydrocarbon containing C–C single bond undergoes free rotation exhibiting different conformers where the antiform/ staggered forms are more stable than the eclipsed form.

→ Alkene containing C=C can exhibit geometrical isomers (cis & trans) due to restricted rotation around the C–C double bond.

→ Some of the polynuclear hydrocarbons having fused benzene ring system have carcinogenic property. e.g. 1, 2 - Benzanthracene, 1, 2 - Benzpyrene, 1, 2, 5, 6–Dibenzanthracene etc.

UNIT - XIV

→ Environmental pollution is the effect of undesirable changes in the surrounding that have harmful effects on plants, animals and humanbeings.

→ Troposphere (upto ~10 Km) where usually the living organisms exist is polluted by oxides of sulfur, nitrogen, and carbon mainly responsible for acid rain.

→ Stratosphere (upto 50 km) above troposphere contain ozone layers which protect the earth from harmful U.V. rays coming from sun. There occurs depletion in ozone layers (Ozone holes) mainly due to CFCs. This increases the chances of skin cancer.

→ 75% of solar energy reaching earth is absorbed by earth surface and rest is radiated back to atmosphere. The polluting gases (oxides of S, N, C and halogen etc.) trap the heat resulting global warming.

→ The increase in the green house gases (CO_2, CFCs, methane, Nitrogenoxide etc) on rapid industrialisation raises the temperature of the earth's atmosphere which if not checked, may eventually result in melting polar ice caps and consequently may submerge the coastal land masses.

→ Rapid urbanisation, excessive industrialisation and uncontrolled use of pesticides, fertilisers in agriculture are the main source of air, water and soil pollution.

→ It is now high time to take stringent action to control environment pollution occuring from different angles. Measures to adopt preventing pollution : (i) waste management (ii) implementation of non-polluting green chemistry (iii) public awareness (iv) framing environment laws to control pollution.

◆◆◆

GROUP - A (1)
MULTIPLE CHOICE QUESTIONS

UNIT - I

1. After rounding off 1.235 and 1.225 one can write the answers respectively as –
(a) 1.23, 1.22 (b) 1.24, 1.123 (c) 1.23, 1.23 (d) 1.24, 1.22

2. Nk_g^{-1} is the unit of____
(a) momentum (b) velocity (c) pressure (d) acceleration

3. The most abundant element in the earth $\frac{1}{8}$ crust is——.
(a) iron (b) aluminium (c) oxygen (d) nitrogen

4. Which of the following not correctly matched ?
(a) $\mu = 10^{-6}$ (b) $M = 10^{6}$ (c) $f = 10^{-15}$ (d) $p=10^{-10}$

5. Two metallic oxides contain 27.6% and 30.0% oxygen respectively. If the formula of first oxide is M_3O_4 that of second will be –
(a) MO (b) MO_2 (c) M_2O_2 (d) M_2O_3

6. Which of the following statements is incorrect ?
(a) One gram atom of carbon N_A number of atoms
(b) One mole of oxygen contains Avogadro's no. of molecules
(c) One mole of hydrogen contains 6.023×10^{23} number of atoms
(d) Mass of one gram mole of nitric oxide is 30 gram.

7. Which of the following will have maximum number of atoms ?
(a) 1 gram of silver (b) 1 gram of sodium (c) 1 gram of chlorine (d) 1 gram of calcium

8. Normality of 2M sulfuric acid is______
(a) 2 N (b) 4 N (c) $\frac{N}{2}$ (d) $\frac{N}{4}$

9. Which pair of species have same percentage of carbon ?
(a) CH_3COOH & $C_6H_{12}O_6$ (b) CH_3COOH & C_2H_5OH
(c) $HCOOCH_3$ & $C_{12}H_{22}O_{11}$ (d) $C_6H_{12}O_6$ & $C_{12}H_{22}O_{11}$

10. Equivalent mass of $KMnO_4$ when it's converted to $MnSO_4$ is _______
(a) $\frac{M}{5}$ (b) $\frac{M}{6}$ (c) $\frac{M}{3}$ (d) $\frac{M}{2}$

11. 4 L of CO at NTP will require for complete oxidation —— volume of O_2 at same conditions.
(a) 4 L (b) 8 L (c) 2 L (d) 1 L

12. The sulfate of an element contains 42.2% element. Its equivalent mass is _______ .
(a) 17.0 (b) 35.0 (c) 51.0 (d) 68.0

13. What quantity of lime stone on heating will give 5.6 gram calcium oxide ?
(a) 10.0 gram (b) 5.6 gram (c) 4.4 gram (d) 1.0 gram

14. 100 ml of PH_3 when decomposed produces phosphorus and hydrogen, the change in volume is _____.
(a) 50 ml increases (b) 500 ml decreases (c) 900 ml decreases (d) nil

15. What mass of $CaCO_3$ is required to react completely with 25 ml of 0.75 M HCl ?

(a) 0.94 g (b) 0.93 g (c) 0.83 g (d) 1.0 gram

16. How many grams of NaOH should be dissolved to make 100 cm^3 of 0.15 M sodium hydroxide solution ?

(a) 0.6 g (b) 0.4 g (c) 0.015 mol (d) 1.6 g

17. In what medium the consumption of potassium permanganate is maximum in quantitative analysis ?

(a) acidic (b) alkaline

(c) neutral (d) identical in all media

18. The vapour density of a gas in 11.2. The volume occupied by 11.2 g of the gas at NTP will be _____.

(a) 11.2 L (b) 22.4 L (c) 1 L (d) 44.8 L

19. Two elements X (atomic mass = 75) and Y (atomic mass = 16) combined to give a compound having 75.8 % of X. The formula of the compound is

(a) XY (b) X_2Y (c) XY_2 (d) X_2Y_3

20. N_A is the Avogadro's number, then number of valence electrons is 4.2 g of nitride ions (N^{3-}) is _______.

(a) 2.4 N_A (b) 4.2 N_A (c) 3.0 N_A (d) 1.6 N_A

UNIT - II

21. Most elements have fractional atomic masses because

(a) they have isotopes.

(b) their isotopes have different masses and present in different ratio.

(c) their isotopes have non-integral masses.

(d) the constituent neutrons, protons and electrons combine to give fractional masses.

22. Isotopes having same neutrons are called ______

(a) isobars (b) isotones (c) isodiapher (d) isoelectronic

23. The limiting line in Paschen series corresponds to

(a) $n_1 = 2, n_2 = 3$ (b) $n_1 = 3, n_2 = 4$ (c) $n_1 = 3, n_2 = 10$ (d) $n_1 = 3, n_2 = \infty$

24. Which of the following electronic configuration is permissible ?

(a) $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1$ (b) $1s^2\ 2s^2\ 2p_x^2$

(c) $1s^2\ 2s^2\ 3s^2$ (d) $1s^2\ 2s^2\ 2p_y^2$

25. Electmagnetic radiation of maximum wavelength is :

(a) UV (b) Radio Waves (c) X-rays (d) Infrared

26. Which of the following sets of quantum numbers is not permissible ?

(a) $2, 1, 0, +\frac{1}{2}$ (b) $2, 2, -1, +\frac{1}{2}$ (c) $2, 1, +1, -\frac{1}{2}$ (d) $3, 2, 0, -\frac{1}{2}$

27. The wavelength of spectral line for an electronic transition is inversely related to –

(a) number of electrons undergoing transition

(b) nuclear charge of the atom

(c) Difference in energy of the energy levels involved in transition

(d) velocity of the electron undergoing transition.

28. Energy of the electron in an orbit is determined by :

(a) Principal quantum number (b) Azimuthal quantum number

(c) Spin quantum number (d) Magnetic quantum number

29. Number of electrons present in an element having 2K, 8L and 5M electrons is :
(a) 15 (b) 8 (c) 6 (d) 5

30. Bohr's model can explain
(a) spectrum of hydrogen atom only
(b) spectrum of an atom or ion having one electron only
(c) spectrum of hydrogen molecule (d) solar spectrum

31. The triad of nuclei that is isotonic is :
(a) ${}_6C^{14}$, ${}_7N^{15}$, ${}_9C^{17}$ (b) ${}_6C^{12}$, ${}_7N^{14}$, ${}_6C^{19}$
(c) ${}_6C^{14}$, ${}_7N^{14}$, ${}_9C^{17}$ (d) ${}_6C^{14}$, ${}_7N^{14}$, ${}_9C^{19}$

32. Nitrogen is having three unpaired electrons according to
(a) Hund's Rule . (b) Aufbau Principle
(c) Heisenberg's Principle (d) None of the above.

33. In an atom no two electrons can have the same value for all the quantum numbers. This was proposed by
(a) Hund (b) Pauli (c) Dalton (d) Avogadro.

34. Correct set of four quantum numbers for valence electrons of rubidium (Z = 37) is
(a) 5, 0, 0, + $\frac{1}{2}$ (b) 5, 1, 0, + $\frac{1}{2}$ (c) 5, 1, 1, + $\frac{1}{2}$ (d) 6, 0, 0, + $\frac{1}{2}$

35. For azimuthal quantum number 1=3, the maximum number of electrons will be
(a) 2 (b) 6 (c) 0 (d) 14

36. Krypton has the electronic configuration [Ar] $4s^2\ 3d^{10}\ 4p^6$, the 37th electron will go into which of the subshells ?
(a) 4f (b) 4d (c) 3p (d) 5s

37. If uncertainty in position of an electron is zero, the uncertainty in its momentum would be :
(a) Zero (b) $\frac{h}{2\pi}$ (c) $\frac{h}{\pi}$ (d) Infinite

38. The magnetic quantum number for valence electron of sodium is
(a) 3 (b) 2 (c) 1 (d) 0

39. Which of the following transition metal double charged cations has maximum unpaired electrons ?
(a) Mn (b) Ni (c) Co (d) Fe.

40. A particular element has configuration $1s^2\ 2s^2\ 2p^5$. In its chemical reaction, it is most likely to
(a) gain one electron (b) Lose one electron
(c) Lose three electrons (d) gain three electron.

41. Number of electrons in the nucleus of an element having atomic number 14 is
(a) 14 (b) 20 (c) 0 (d) 7

42. The nucleus of tritium consists of
(a) 1 Proton +1 neutron (b) 1 proton +3 neutrons
(c) 1 proton + no neutron (d) 1 proton + 2 neutrons

43. The species isoelectromic with CN^- are
(a) CO (b) O^+_2 (c) F_2 (d) NO

44. 2p orbital have :
(a) n = 1, 1 = 2 (b) n = 2, 1 = 1 (c) n = 1, 1 = 0 (d) n = 2, 1 = 0

45. The azimuthal quantum number of the 17th electron of chlorine atom is :
(a) one (b) two (c) three (d) zero

UNIT - III

46. Sn and Pb exhibit variable valency due to –
(a) intertpair effect (b) screening effect
(c) Lanthanide contraction (d) various $\Delta_i H$ values.

47. Long form of periodic table is based on
(a) Atomic mass (b) Atomic number (c) Atomic size (c) Electron affinity

48. Which pair of elements is chemically most similar ?
(a) Na, Al (b) Cu, S (c) Ti, Zr (d) Zr, Hf

49. Elements of group-1 give colour in Bunsen burner due to
(a) Low IP (b) Low MP
(c) Softness (d) one electron in the outermost shell.

50. Elements of same vertical group of the periodic table have
(a) Same atomic size (b) Same electronic configuration
(c) Same number of electrons in the outermost shell of their atoms.
(d) Same number of atoms.

51. The element with highest value of 1st ionisation potential is
(a) B (b) C (b) N (d) O

52. The ion having highest radius is
(a) Al^{3+} (b) N^{3-} (c) Na^{+} (d) F^{-}

53. Atomic radii of Fluorine and Neon in A° units are respectively given by
(a) 0.72, 1.60 (b) 1.60, 1.60 (c) 0.72, 0.72 (d) None of these

54. Which of the following electronic configuration represents a noble gas ?
(a) 2, 8, 1 (b) 2, 8, 8 (c) 2, 1, 7 (d) 2, 8, 2

55. The maximum tendency to form unipositive ion is for the element with the following electronic configuration,
(a) $1s^2 2s^2 2p^6 3s^1$ (b) $1s^2 2s^2 2p^6 3s^2 3p^1$
(c) $1s^2 2s^2 2p^6 3s^2 3p^2$ (d) $1s^2 2s^2 2p^6 3s^2 3p^3$

56. Characteristic of transition element is
(a) Incomplete d-orbital (b) Incomplete f- orbital
(c) Incomplete s- orbital (d) Incomplete p- orbital.

57. Which of the following has the largest ionic radius ?
(a) Be^{2+} (b) Mg^{2+} (c) Ca^{2+} (d) Sr^{2+}

58. The first ionisation energy of lithium is
(a) greater than Be (b) less than Be (c) equal to Na (d) equal to F

59. The electronegativity of the following elements increases in the order
(a) C, N, Si, P (b) N, Si, C, P (c) Si, P, C, N (d) P, Si, N, C

60. The 1st transition series contains elements having atomic numbers from
(a) 22 to 30 (b) 21 to 30 (c) 21 to 31 (d) 21 to 29

61. Which is the lightest metal in the periodic table ?
(a) H (b) Mg (c) Ca (d) Li

62. A newly discovered element ‘X’ is placed in the group 1 of the Periodic table because it forms
(a) an oxide which is acidic
(b) A volatile chloride having formula XCl
(c) An ionic chloride having formula XCl
(d) An insoluble $XHCO_3$

63. The correct order of metallic character of the elements B, Al, Mg and K
(a) B > Al > Mg > K (b) Al > Mg > B > K
(c) Mg > Al > K > B (d) K > Mg > Al > B

64. The correct order of the chemical reactivity of the elements F, Cl, O and N is
(a) F > O > Cl > N (b) F > Cl > O > N
(c) O > F > N > Cl (d) Cl > F > O > N

65. The correct order of non-metallic character of the elements B, C, N and F is
(a) F > N > C > B (b) C > B > N > F
(c) B > C > N > F (d) F > N > C > B

66. Size of the isoelecteonic species F^-, Ne and Na^+ is affected by
(a) Nuclear charge
(b) Valence principal quantum number
(c) Electron-electron interaction in the outer orbitals
(d) None of the above

67. In periodic table, on moving along a period, the ionisation potential
(a) Increases from left to right generally
(b) Remains unchanged mostly
(c) First increases then decreases
(d) Decreases from left to right occasionally

68. Which of the following atoms possesses the smallest volume?
(a) S (b) Si (c) P (d) He

69. Ionisation potential of an element does not depend upon
(a) Electrical neutrality (b) Nuclear charge
(c) Penetrating effect (d) Shielding effect

70. Zero group was introduced by
(a) Mendeleev (b) Ramsay (c) Lothar Meyer (d) Johnson

UNIT - IV

71. Which of the following pairs molecules is isostructural ?
(a) CO_2, C_2H_2 (b) $HgCl_2$, C_2H_4 (c) $HgCl_2$, $BeCl_2$ (d) $SnCl_2$, $BeCl_2$

72. Octet rule is not followed in
(a) CCl_4 (b) H_2O (c) O_2 (d) BCl_3

73. The species having one unpaired electron is –
(a) NO (b) CO (c) CN^- (d) O_2

74. Dipole moment of BeF_2 is
(a) very low (b) very high (c) zero (d) not definite.

75. Carbon in ethylene involves the hybridisation
(a) sp^3 (b) sp^2 (c) sp (d) none of these.

76. Hybridisation involves

(a) orbitals of same atom with slightly different energies

(b) orbitals of different atoms but with equal energies

(c) orbitals of same atom but with widely different energies

(d) orbitals of different atoms with different energies

77. Which of the following statements is wrong ?

(a) Covalent compounds are generally soluble in polar solvents

(b) Covalent compounds have low melting and boiling points

(c) Ionic solids do not conduct electricity

(d) Ionic compounds conduct electricity in the fused state.

78. Which of the following contains both covalent and ionic bonds ?

(a) CCl_4 (b) $CaCl_2$ (c) NH_4Cl (d) H_2O.

79. In which of the following the central atom does not use sp^3- hybrid orbitals in its bonding.

(a) BeF_3^- (b) OH_3^+ (c) NH_2^- (d) NH_3

80. The molecule which does not exhibit dipole moment is :

(a) NH_3 (b) $CHCl_3$ (c) H_2O (d) CCl_4.

81. What is the maximum number of hydrogen bonds in which a water molecule can participate ?

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

82. Which one of the following has zero dipoie moment ?

(a) ClF (b) PCl_3 (c) SiF_4 (d) $CFCl_3$

83. Two lone pairs of electrons and two bond pairs of electrons are present in

(a) NH_3 (b) BF_3 (c) H_2O (d) CO_2

84. NH_3 and BF_3 form an adduct readily because they form

(a) an ionic bond (b) a covalent bond

(c) a coordinate bond (d) a hydrogen bond

85. NH_3 has a much higher boiling point than PH_3 because

(a) NH_3 has larger molecular weight

(b) NH_3 undergoes umbrella inversion

(c) NH_3 forms hydrogen bond

(d) NH_3 contains ionic bonds whereas PH_3 contains covalent bonds.

86. Which of the following has net dipole moment ?

(a) CCl_4 (b) BF_3 (c) NH_3 (d) CO_2

87. Hydrogen bonding is maximum in

(a) Ethanol (b) Diethyl ether (c) Ethyl chloride (d) Triethyl amine.

88. Which of the following hydrogen halide is most volatile ?

(a) HF (b) HCl (c) HBr (d) HI

89. Arrange the following compounds in order of increasing dipole moment

Toluene (I), m-dichlorobenzene (II), o- dichlorobenzene (III), p-dichlorobenzene (IV)

(a) I < IV < II < III (b) IV < I < II < III (c) IV < I < III < II (d) IV < II < I < III.

90. KF combines with HF to form KHF_2. The compound contains the species

(a) K^+, F^- and H^+ (b) K^+, F^- and HF (c) K^+ and $[HF_2]^-$ (d) $[KHF]^+$ and F_2

91. Which one of the following compounds has sp^2 hybridisation ?

(a) CO_2 (b) SO_2 (c) N_2O (d) CO

92. The cylindrical shape of an alkyne is due to

(a) three sigma C—C bonds

(b) three π C—C bonds

(c) two sigma C—C bonds and one π C—C bond

(d) One sigma C—C bond and two π C—C bonds.

93. Which one of the following has highest dipole moment ?

(a) NH_3 (b) PH_3 (c) SbH_3 (d) AsH_3

94. In crystals of which of the following ionic compound would you expect maximum distance between the centres of cations and anions ?

(a) CsF (b) Cs I (c) Li I (d) Li F

95. Hydrogen bonding is not present in

(a) Glycerine (b) Water

(c) Hydrogen sulphide (d) Hydrogen fluoride

96. The correct order of increasing C—O bond length of CO, CO_3^{2-}, CO_2 is

(a) $CO_3^{2-} < CO_2 < CO$ (b) $CO_2 < CO_3^{2-} < CO$

(c) $CO < CO_3^{2-} < CO_2$ (d) $CO < CO_2 < CO_3^{2-}$

97. Covalent compounds have low melting point because

(a) Covalent molecules have definite shape

(b) Covalent bond is weaker than ionic bond

(c) Covalent bond is less exothermic

(d) Covalent molecules are held by weak van der Waals forces of attraction

98. Number of π-bonds in Naphthalene is

(a) 6 (b) 3 (c) 4 (d) 5

99. Shape of O_2F_2 is similar to that of

(a) C_2F_2 (b) H_2O_2 (c) H_2F_2 (d) C_2H_2

100. Which of the following is arranged in the increasing order of enthalpy of vaporisation ?

(a) NH_3, PH_3, AsH_3 (b) AsH_3, PH_3, NH_3

(c) NH_3, AsH_3, PH_3 (d) PH_3, AsH_3, NH_3

101. If a molecule MX_3 has zero dipole moment, the sigma bonding orbitals used by M are

(a) sp^3 d-hybrid (b) sp-hybrid (c) sp^3-hybrid (d) sp^2-hybrid

102. Which of the following is least volatile ?

(a) HF (b) HCl (c) HBr (d) HI

103. Which of the following is not electron deficient ?

(a) NH_3 (b) BF_3 (c) $AlCl_3$ (d) BH_3

UNIT - V

104. For one mole of gas the total Kinetic energy at temperature T is equal to

(a) RT (b) 3/2 RT, (c) 2/3 RT (d) (Cp - Cv) RT.

105. For a given mass of gas if pressure is reduced to half and the temperature is doubled, the volume will become

(a) 4 V (b) 2 V (c) $\frac{V}{4}$ (d) 8 V

106. The value of universal gas constant in litre atmosphere per degree per mole is

(a) 8.200 (b) 8.020 (c) 1.987 (d) 0.082

107. Volume of 20 grams of hydrogen gas at NTP is

(a) 224 litres (b) 22.4 litres (c) 2.24 litres (d) 112 litres

108. 1 cc at NTP of a hydrocarbon vapour is as heavy as 4cc of oxygen. What quantity of hydrocarbon will occupy a volume of 22.4 litres ?

(a) 6.4 (b) 32 (c) 128 (d) 64

109. The temperature at which a real gas obeys the ideal gas laws over a wide range of pressure is called the

(a) Critical temperature (b) Boyle temperature

(c) Inversion temperature (d) Reduced temperature

110. A gas at constant temperature 250 K is containted in a closed vessel. If it is heated through 1^0C, the percentage increase in its pressure is

(a) 0.4 % (b) 0.6 % (c) 0.8 % (d) 1.0 %

111. Equal weights of methane and oxygen are mixed in an empty container at 25°C. The fraction of total pressure exerted by oxygen is

(a) $\frac{1}{3}$ (b) $\frac{1}{2}$ (c) $\frac{2}{3}$ (d) $\frac{1}{3}\times\frac{273}{298}$

112. In a closed room of 1000 m^3 a perfume bottle is opened up. The whole room develops smell. This is due to which property of the gases ?

(a) Viscosity (b) Density (c) diffusion (d) None

113. Helium diffuses twice as fast as another gas B. If the vapour density of He is 2, the molecular weight of B is

(a) 4 (b) 8 (c) 16 (d) 24

114. 50ml of gas A diffuses through a membrane in the same time as taken for the diffusion of 40 ml of another gas B under identical pressure temperature conditions. If the molecular weight of A is 64 that of B would be

(a) 100 (b) 250 (c) 200 (d) 80

115. Which is not true incase of an ideal gas ?

(a) It cannot be converted into a liquid.

(b) There is no interaction between the molecules

(c) All molecules of the gas move with same speed.

(d) At a given temperature PV is proportional to the amount of the gas.

116. An ideal gas cannot be liquefied because

(a) Its critical temperature is always above 0°C.

(b) Its molecules are relatively smaller in size.

(c) It solidifies before becoming a liquid.

(d) Forces operative between its molecules are negligible.

117. The vander Waals equation explains the behaviour of

(a) Ideal gas (b) Real gases (c) Vapours (d) Non - real gases

118. Real gases approximate to ideal behaviour only

(a) If temperature is lower (b) If pressure is high

(c) If pressure is low and temperature is high

(d) If both temperature and pressure are low.

119. The deviation of behaviour of real gas from ideal gas will depend on

(a) Force acting between molecules

(b) Space occupied between molecules

(c) Mass of molecules only

(d) The force acting between molecules, space occupied by molecules and mass of molecules.

120. The pressure at which one mole of a gas at 0°C occupies a voume of one litre is :

(a) 2.24 atm (b) 4.48 atm (c) 11.2 atm (d) 22.4 atm

121. At what temperature in the celsius scale, volume of certain mass of gas at 27°C will be doubled keeping the pressure constant.

(a) 54^0 (b) 327°C (c) 427°C (d) 527^0C

122. Partial pressure of CO_2 in a mixture of CO_2 and N_2 is 1.25 atmosphere while the total pressure of the gas mixture is 5 atmosphere. Mole fraction of N_2 in the mixture is :

(a) 0.82 (b) 0.75 (c) 0.80 (d) 0.65

123. A comparision of relative rates of diffusion of H_2 and O_2 at a given temperature shows that.

(a) O_2 diffuses four times faster than H_2 (b) H_2 diffuses two times faster than O_2

(c) H_2 and O_2 diffuse at the same rate. (d) H_2 diffuses four times faster than oxygen

124. The root mean square velocity is expressed as

(a) $(\frac{3}{2}RT)^{1/2}$ (b) $(\frac{3RT}{M})^{1/2}$ (c) $(\frac{2RT}{M})^{1/2}$ (d) $(\frac{3RT}{M})$

125. If 3 litre of oxygen is heated from 27°C to 100°C keeping the pressure constant, the new volume will be

(a) 3.5 litres (b) 4.5 litres (c) 3.73 litres (d) 5.00 litres

126. In vanderwaals' equation of state for a non ideal gas, the term that accounts for intermolecular forces is

(a) (v - b) (b) RT (c) $\left(P+\frac{a}{V^2}\right)$ (d) $(RT)^{-1}$

127. A gas is initially at 1 atm pressure. To compress it to $\frac{1}{4}$ th its initial volume, pressure to be applied is :

(a) 1 atm (b) 2 atm (c) 4 atm (d) $\frac{1}{4}$ atm

128. At what temperature will the average speed of CH_4 molecules have the same value as O_2 has at 300 K

(a) 1200 K (b) 150 K (c) 600 K (d) 300 K

129. Viscosity of a liquid

(a) Increases with increase in temperature. (b) Decreases with decrease in temperature.

(c) Independent of temperature. (d) Decreases with decrease in pressure

130. Cleaning action of soap is due to.

(a) Viscosity of water. (b) Surface tension of water.

(c) Polarity of water (d) high boiling point of water.

131. The unit of viscosity is poise which means

(a) Dyne Cm^{-2} (b) dyne (c) dyne cm^{-2} sec^{-1} (d) dyne cm

132. The internal resistance to flow possessed by liquid is called its

(a) Surface tension (b) Fuidity (c) vapour pressure (d) Viscosity

133. A liquid in capillary tube rises due to

(a) Surface tension (b) Vapour pressure (c) Osmosis (d) Viscosity

134. At high altitudes the water boils at low temperature because

(a) water molecules are bound with strong hydrogen bonds.
(b) low atmospheric pressure
(c) high atmospheric pressure
(d) at low altitude the fuel does not give sufficient energy

135. When temperature increases, the surface tension of a liquid is

(a) increased (b) decreased
(c) neither increased nor decreased (d) depends on the nature of the liquid

UNIT - VI

136. A well stoppered thermosflask contains some ice Cubes. This is an example of a

(a) Closed system (b) open system
(c) isolated system (d) non-thermodynamic system

137. For the reaction

$N_2 + 3H_2 = 2NH_3, \Delta H = ?$

(a) $\Delta U = 2RT$ (b) $\Delta U - 2RT$ (c) $\Delta U + RT$ (d) $\Delta U + 2RT$

138. For an ideal gas, the relation between the enthalpy change and internal energy at constant temperature is given by

(a) $\Delta H = \Delta U + PV$ (b) $\Delta H = \Delta U + \Delta nRT$
(c) $\Delta H = \Delta U - P\Delta V$ (d) $\Delta H = \Delta G + T\Delta S$

139. If $C_{(s)} + O_{2_{(g)}} \rightarrow CO_{2_{(g)}}, \ \Delta H = X$ and

$CO_{(g)} + \frac{1}{2}O_{2_{(g)}} \rightarrow CO_{2_{(g)}}, \ \Delta H = Y$ then the heat of formation of CO is

(a) $X + Y$ (b) $X - Y$ (c) $Y - X$ (d) XY

140. Enthalpies of elements in their standard states are taken as zero. Hence the enthalpy of formation of a compound

(a) should always be negative. (b) should always be positive.
(c) will be equal to twice the energy of combustion.
(d) may be +ve or –ve

141. Which of the following values of heat of formation indicates that the product is least stable ?

(a) –94 kcal, (b) –231.6 kcals (c) +21.4 kcal (d) +64.8 kcal.

142. For an adiabatic process, which of the following is correct

(a) $P \Delta V = 0$, (b) $q = +w$, (c) $\Delta E = q$, (d) $q = 0$

143. Idenify the intensive quantity from the following.

(a) Enthalpy and temperature (b) Volume and temperature
(c) Enthalpy and volume (d) Temperature and refractive index

144. An exothermic reaction is one which

(a) takes place.on heating (b) is accompanied by flame

(c) is accompanied by absorption of heat (d) is accompanied by evolution of heat.

145. An endothermic reaction is one in which

(a) Heat is converted into electricity (b) Heat is absorbed

(c) Heat is evolved (d) Heat is converted to mechanical work

146. Which of the following reactions is not exothermic ?

(a) $CaCO_3 \rightarrow CaO + CO_2$ (b) $CCl_{4(g)} + 2H_2O \rightarrow CO_{2(g)} + 4\ HCl_{(g)}$

(c) $NaOH + HCl \rightarrow NaCl + H_2O$ (d) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

147. Which of the following is an endothermic reaction.

(a) $2H_2 + O_2 \rightarrow 2H_2O$ (b) $N_2 + O_2 \rightarrow 2NO$

(c) $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$ (d) $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

148. In exothermic reactions

(a) ΔE is zero (b) ΔH is –ve (c) ΔS is zero (d) ΔH is +ve.

149. Enthalpy of a compound is equal to its.

(a) Heat of combustion. (b) Heat of formation.

(c) Heat of reaction. (d) Heat of solution.

150. Heat of neutralisation of a strong acid and strong base is always.

(a) –13.7 kcal/mole. (b) –9.6 kcal/ mole (c) +6 kcal/mole, (d) +11.4 kcal/mole.

151. In which of the following neutralisation reactions, the heat of neutralisation is the highest.

(a) NH_4OH and H_2SO_4 (b) HCl and NaOH

(c) CH_3COOH and KOH (d) CH_3COOH and NH_4OH.

152. Hess's law of heat of summation includes

(a) Initial reactants only (b) Initial reactants and final products

(c) Final products only (d) Intermediates only.

153. For which reaction, S will be maximum ?

(a) $Ca_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CaO_{(s)}$ (b) $CaCO_{3_{(s)}} \rightarrow CaO_{(s)} + CO_{2_{(g)}}$

(c) $C_{(s)} + O_{2_{(g)}} \rightarrow CO_{2_{(g)}}$ (d) $N_{2_{(g)}} + O_{2_{(g)}} \rightarrow 2NO_{(g)}$

154. Energy required to dissociate 4 gms of gaseous hydrogen into free gaseous atoms is 208 kcals. at 25^0 C. The bond energy of H–H bond will be

(a) 104 kcals (b) 10.4 kcals (c) 1040 kcal (d) 104 cals

UNIT - VII

155. When rate of forward reaction is equal and opposite to the rate of backward reaction,the state is said to be :

(a) Reversible state (b) Equilibrium.

(c) Chemical equilibrium. (d) None of the above.

156. Which of the following reactions will be favoured by low pressure ?

(a) $H_2 + l_2 \rightleftharpoons 2HI$ (b) $N_2 + 3H_2 \rightleftharpoons 2NH_3$

(c) $PCl_5 \rightleftharpoons PCl_3 + Cl_2$ (d) $N_2 + O_2 \rightleftharpoons 2NO$

157. Which of the following factor will be useful in manufacture of ammonia by Haber's process ?
(a) High pressure. (b) Low pressure. (c) High temperature
(d) Increase in the concentration of ammonia.

158. The rate at which a substance reacts is proportional to its active mass. This statement is:
(a) Le-Chatelier's principle. (b) Faraday's Law.
(c) Law of multiple proportions. (d) Law of mass action.

159. When chemical equilibrium is reached the
(a) Reaction stops.
(b) Rate of forward reaction is equal to the rate of backward reaction
(c) Rate of forward reaction is more than that of backward reaction
(d) None of the above

160. In a reversible reaction if there is no change in total number of molecules, the reaction will be favoured by
(a) High pressure (b) Low pressure
(c) High temperature (d) Higher concentration of a reactant

161. Chemical equilibrium is
(a) Stationary (b) Dynamic (c) Inertness (d) State of rest

162. For the reaction, $H_2 + I_2 \rightleftharpoons 2HI$, the Kp and Kc are related as
(a) $K_p = K_C (RT)^2$ (b) $K_p = K_C (RT)^0$ (c) $K_p = K_c(RT)^{-2}$ (d) $K_p = K_c(RT)^{-1}$

163. The partial pressure of PCl_3, Cl_2 and PCl_5 are 0.1,0.2 and 0.008 atmosphere respectively for a reaction, $PCl_5 \rightleftharpoons PCl_3 + Cl_2$. The value of K_p is
(a) 2.5 (b) 5.0 (c) 0.25 (d) 25

164. For the reaction $PCl_5 \rightleftharpoons PCl_3 + Cl_2$, the forward reaction at constant temperature is favoured by:
(a) Introducing an inert gas at constant volume
(b) Introducing chlorine gas at constant volume
(c) Introducing an inert gas at constant pressure
(d) Increasing the volume of the container

165. The pH of solution containing 0.4 gm NaOH per litre is,
(a) 2.0 (b) 12.0 (c) 10.0 (d) 11.0

166. The pH of the solution is 3.0. If its pH is changed to 6.0, then the $[H^+]$ of the original solution has to be
(a) Doubled (b) Halved
(c) Increased 1000 times (d) Decreased 1000 times.

167. Metal ions like Ag^+, Cu^{2+} etc. act as
(a) Bronsted acid (b) Bronsted bases
(c) Lewis acids (d) Lewis bases.

168. With increase in temperature, ionic product of water
(a) increases (b) decreases
(c) remains unaffected (d) may increase or decrease

169. The pH of a 10^{-8} M solution of HCl in water is
(a) 8 (b) –8 (c) between 7 and 8 (d) between 6 and 7

170. At 90°C pure water has $[H_3O^+] = 10^{-6}$ M. What is the value of K_w at this temperature ?

(a) 10^{-6} (b) 10^{-12} (c) 10^{-13} (d) 10^{-14}

171. Given that the dissociation constant for H_2O is $K_w = 1 \times 10^{-14}\ \text{mol}^2\ \text{litre}^{-2}$, what is the pH of a 0.001 M KOH solution?

(a) 10^{-11} (b) 10^{-3} (c) 3 (d) 11

172. Among the following hydroxides, the one which has the lowest value of K_{sp} at ordinary temperature (about 25°C) is

(a) $Mg(OH)_2$ (b) $Ca(OH)_2$ (c) $Ba(OH)_2$ (d) $Be(OH)_2$

173. The following equilibrium is established when hydrogen chloride is dissolved in acetic acid.

$$HCl + CH_3COOH \rightleftharpoons Cl^- + CH_3COOH_2^+$$

The set that characterises the conjugate acid base pair is

(a) (HCl, CH_3COOH) and ($CH_3COOH_2^+$, Cl^-)

(b) (HCl, $CH_3COOH_2^+$) and (CH_3COOH, Cl^-)

(c) ($CH_3COOH_2^+$ HCl) and (Cl^-, CH_3COOH)

(d) (HCl, Cl^-) and ($CH_3COOH_2^+$, CH_3COOH)

174. The pH of a 10^{-10} M NaOH solution is nearest to

(a) 10 (b) 7 (c) 4 (d) –10

175. The precipitate of CaF_2 ($K_{sp} = 1.7 \times 10^{-10}$) is obtained when equal volumes of the following are mixed:

(a) 10^{-4} M Ca^{2+} + 10^{-4} M F^- (b) 10^{-2} M Ca^{2+} + 10^{-3} M F^-

(c) 10^{-5} M Ca^{2+} + 10^{-3} M F^- (d) 10^{-3} M Ca^{2+} + 10^{-5} M F^-

176. The strongest base of the following species is

(a) NH^{2-} (b) OH^- (c) O^{2-} (d) S^{2-}

177. Which of the following sulphides has the lowest solubility product ?

(a) FeS (b) MnS (c) PbS (d) ZnS

178. The pH of a solution obtained by mixing 50 ml of 0.4 N HCl and 50 ml of 0.2 N NaOH is

(a) $-\log 2$ (b) $-\log 0.2$ (c) 1.0 (d) 2.0

179. The solubility of $CaCO_3$ in water is 3.05×10^{-4} moles/litre. Its solubility product will be

(a) 6.1×10^{-4} (b) 9.3 (c) 3.05×10^{-4} (d) 9.3×10^{-8}

180. pH of water is 7. When a substance Y is dissolved, the pH becomes 13. The substance 'Y' is a salt of__________ .

(a) weak acid and weak base (b) strong acid and strong base

(c) weak acid and strong base (d) None of these

UNIT - VIII

181. Oxidation number of sodium in sodium amalgam is

(a) +1 (b) 0 (c) –1 (d) + 2

182. For the reaction, $M^{x+} + MnO_4^- \longrightarrow MO_3^- + Mn^{2+} + 1/2O_2$ if one mole of MnO_4^- oxidies 1.67 moles of M^{x+} to MO_3^-, then the value of x in the reaction is

(a) 5 (b) 3 (c) 2 (d) 1

183. The oxidation state of chromium in $Cr(CO)_6$ is

(a) 0 (b) + 2 (c) – 2 (d) + 6

184. Which of the following is not a redox reaction ?

(a) $CaCO_3 \rightarrow CaO + CO_2$ (b) $O_2 + 2H_2 \rightarrow 2H_2O$

(c) $Na + H_2O \rightarrow NaOH + 1/2\ H_2O$ (d) $MnCl_3 \rightarrow MnCl_2 + 1/2\ Cl_2$

185. When copper is treated with a certain concentration of nitric acid, nitric oxide and nitrogen dioxide are liberated in equal volumes according to the equation,

$x\ Cu + y\ HNO_3 \rightarrow Cu(NO_3)_2 + NO + NO_2 + H_2O$

The coefficients of x and y are

(a) 2 and 3 (b) 2 and 6 (c) 1 and 3 (d) 3 and 8

186. In the reaction, $2FeCl_3 + H_2S \rightarrow 2FeCl_2 + 2HCl + S$

(a) $FeCl_3$ acts as an oxidizing agent (b) Both H_2S and $FeCl_3$ are oxidized.

(c) $FeCl_3$ is oxidised while H_2S is reduced (d) H_2S acts as an oxidizing agent.

187. Number of moles of $KMnO_4$ required to oxidise one mole of $Fe(C_2O_4)$ in acidic medium is

(a) 0.6 (b) 1.67 (c) 0.2 (d) 0.4

188. A standard hydrogen electrode has zero electrode potential because

(a) hydrogen is easiest to oxidize (b) this electrode potential is assumed to be zero

(c) hydrogen atom has only one electron (d) hydrogen is the lighest element

189. Which of the following is a redox reaction ?

(a) H_2SO_4 with NaOH (b) In atmosphere ,O_3 from O_2 by lightning

(c) Nitrogen oxides from nitrogen and oxygen by lightning

(d) Evaporation of H_2O

190. Which among the following is the strongest reducing agent ?

Given, $Fe^{2+} + 2e^- \longrightarrow Fe(-0.44V)$; $Ni^{2+} + 2e^- \longrightarrow Ni(-0.25V)$;

$Sn^2 + 2e^- \longrightarrow Sn(-0.14V)$ and $Fe^{3+} + e^- \longrightarrow Fe^{2+}(-0.77V)$

(a) Fe (b) Fe^{2+} (c) Ni (d) Sn

191. The oxidation number of sulphur in S_8, S_2F_2, H_2S respectively, are

(a) 0, + 1 and –2 (b) + 2, +1 and –2 (c) 0, +1 and + 2 (d) –2, + 1 and -2.

192. The partial ionic equation,

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ suggests that the equivalent weight of $Cr_2O_7^{2-}$ will be equal to its formula weight divided by

(a) 3 (b) 6 (c) 1 (d) 14

193. Oxidation state of osmium (Os) in OsO_4 is

(a) + 7 (b) + 6 (c) + 4 (d) + 8

194. Fluorine is the best oxidising agent because it has

(a) highest electron affinity (b) highest E^0 reduction

(c) highest E^0 oxidation (d) lowest electron affinity

195. Equivalent mass of oxidising agent in the reaction, $SO_2 + 2H_2S \rightarrow 3S + 2H_2O$ is

(a) 32 (b) 64 (c) 16 (d) 8

196. The oxidation state of chromium in $[Cr(P\,Ph_3)_3\,(CO)_3]$

(a) + 3 (b) + 8 (c) zero (d) + 5

197. Which of the following is a redox reaction ?

(a) $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$

(b) $CaC_2O_4 + 2HCl \rightarrow CaCl_2 + H_2C_2O_4$

(c) $Mg(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + 2NH_4OH$

(d) $Zn + 2AgCN \rightarrow 2Ag + Zn(CN)_2$

198. A smuggler could not carry gold by depositing iron on the gold surface since

(a) Gold is denser (b) Iron rusts

(c) Gold has higher electrode potential than iron

(d) Gold has lower electrode potential than iron

199. For decolourization of 1 mole of $KMnO_4$, the moles of H_2O_2 required is

(a) 1/2 (b) 3/2 (c) 5/2 (d) 7/2

200. Excess of KI reacts with $CuSO_4$ solution and $Na_2S_2O_3$ solution is added to it. Which of the statement is incorrect for the reaction ?

(a) Evolved I_2 is reduced (b) CuI_2 is formed

(c) $Na_2S_2O_3$ is oxidised (d) $Cu_2\,I_2$ is formed

UNIT - IX

201. Which of the following statements is most applicable to hydrogen ? It can act

(a) as a reducing agent (b) as an oxidising agent

(c) both as oxidising and reducing agents (d) neither as an oxidising nor as a reducing agent.

202. Hydrogen combines with other elements by

(a) losing an electron (b) gaining an electron

(c) sharing an electron (d) losing, gaining and sharing of an electron.

203. The first ionization energy (kJ mol^{-1}) for H, Li, F, Na has one of the following values: 1681, 520, 1312, 495. Which of these values corresponds to that of hydrogen ?

(a) 1681 (b) 1312 (c) 520 (d) 495

204. The correct increasing order of the acidity of CO_2, H_2O and H_2O_2 is

(a) $CO_2 < H_2O_2 < H_2O$ (b) $H_2O < H_2O_2 < CO_2$

(c) $H_2O < H_2O_2 < CO_2$ (d) $H_2O_2 > CO_2 > H_2O$

205. The volume of 10 volume H_2O_2 solution that decolourises 200 ml of 2N $KMnO_4$ solution in acidic medium is:

(a) 112 ml (b) 336 ml (c) 200 ml (d) 224 ml.

206. Which of the following metals cannot be used for liberating dihydrogen from dilute hydrochloric acid?

(a) Zinc (b) Copper (c) Iron (d) Magnesium.

207. Which of the following is used as a moderator in nuclear reactors ?

(a) Hard water (b) Heavy water (c) Deionized water (d) Mineral water.

208. The temporary hardness of water due to calcium bicarbonate can be removed by adding

(a) $CaCO_3$ (b) $Ca(OH)_2$ (c) $CaCl_2$ (d) HCl

209. Hydrogen will not reduce

(a) heated cupic oxide (b) heated ferric oxide

(c) heated stannic oxide (d) heated aluminium oxide.

210. The oxidation states of the most electronegative element in the products of the reaction, BaO_2 with dil. H_2SO_4 are

(a) 0 and –1 (b) –1 and –2 (c) –2 and 0 (d) –2 and +1

211. 30 volumes H_2O_2 means

(a) 30 % H_2O_2 (b) 30 cm^3 of the solution contains 1g of H_2O_2

(c) 1 cm^3 of the solution liberates 30 cm^3 of O_2 at STP

(d) 30 cm^3 of the solution contain one mole of H_2O_2.

212. An oxide which gives H_2O_2 on treatment with dilute acid is

(a) PbO_2 (b) Na_2O_2 (c) MnO_2 (d) TiO_2

213. Hydrolysis of one mole of peroxodisulphuric acid produces

(a) two moles of sulphuric acid (b) two moles of peroxomonosulphuric acid

(c) one mole of sulphuric acid and one mole of peroxomonosulphuric acid

(d) one mole of sulphuric acid, one mole of peroxomonosulphuric acid and one mole of hydrogen peroxide.

214. Which contains both polar and non-polar bonds ?

(a) NH_4Cl (b) HCN (c) H_2O_2 (d) CH_4

215. When a substance A reacts with water, it produces a combustible gas B and a solution of substance C in water. When another substance D reacts with this solution of C, it produces the same gas B on warming but D can produce gas B on reaction with dilute sulphuric acid at room temperature. A imparts a deep golden yellow colour to a smokeless flame of Bunsen burner. A, B, C and D respectively are:

(a) Na, H_2, NaOH, Zn (b) K, H_2, KOH, Al

(c) Ca, H_2, $Ca(OH)_2$, Sn (d) CaC_2, C_2H_2, $Ca(OH)_2$, Fe

216. Which of the following pairs of substances on reaction will not evolve H_2 gas ?

(a) Fe and H_2SO_4 (aqueous) (b) Copper and HCl (aqueous)

(c) Sodium and ethyl alcohol (d) Iron and steam

217. Amongst H_2O, H_2S, H_2Se and H_2Te, the one with the highest boiling point is

(a) H_2O because of hydrogen bonding (b) H_2 Te because of higher molecular weight

(c) H_2S because of hydrogen bonding (d) H_2 Se because of lower molecular weight

218. Para and ortho hydrogen differ in

(a) atomic number (b) atomic mass

(c) spins of protons (d) number of neutrons

219. Heavy water is obtained by

(a) boiling water (b) fractional distillation of H_2O

(c) prolonged electrolysis of H_2O (d) heating H_2O_2

220. Which one of the following processes will produce permanent hard water ?

(a) Addition of Na_2SO_4 to water (b) Saturation of water with $CaCO_3$

(c) Saturation of water with $MgCO_3$ (d) Saturation of water with $CaSO_4$

221. The structure of H_2O_2 is

(a) planar (b) non-planar (c) spherical (d) linear

222. H_2O_2 acts as an oxidising agent in

(a) Acidic medium (b) Alkaline medium

(c) Alkaline and neutral medium (d) Acid and alkaline medium

UNIT - X

223. The raw materials used in the manufacture of sodium carbonate by Solvay's process are:

(a) Ammonia and carbon dioxide (b) Sodium chloride and carbon dioxide

(c) Sodium chloride, limestone and carbon dioxide

(d) Sodium chloride, limestone and ammonia.

224. Glauber's salt is

(a) $FeSO_4.7H_2O$ (b) $MgSO_4.7H_2O$

(c) $Na_2CO_3.10H_2O$ (d) $Na_2SO_4.10H_2O$

225. Plaster of Paris is

(a) $(CaSO_4)_2.H_2O$ (b) $CaSO_4.2H_2O$

(c) $CaSO_4.H_2O$ (d) $CaSO_4$

226. Which of the following has the lowest melting point?

(a) LiCl (b) NaCl (c) KCl (d) RbCl

227. Which of the following alkaline earth metal sulphates is least soluble in water ?

(a) $BaSO_4$ (b) $MgSO_4$ (c) $SrSO_4$ (d) $CaSO_4$

228. The basic character of the oxides, MgO, SrO, K_2O, NiO and Cs_2O increases in the order:

(a) $MgO > SrO > K_2O > NiO > Cs_2O$ (b) $Cs_2O < K_2O < MgO < SrO < NiO$

(c) $NiO < MgO < SrO < K_2O < Cs_2O$ (d) $K_2O < NiO < MgO < SrO < Cs_2O$

229. Which of the following are arranged in increasing order of solubilities ?

(a) $CaCO_3 < KHCO_3 < NaHCO_3$ (b) $NaHCO_3 < KHCO_3 < CaCO_3$

(c) $KHCO_3 < NaHCO_3 < CaCO_3$ (d) $CaCO_3 < NaHCO_3 < KHCO_3$

230. The thermal stability of alkaline earth metal carbonates decreases in the order :

(a) $BaCO_3 > SrCO_3 > CaCO_3 > MgCO_3$ (b) $BaCO_3 > SrCO_3 > MgCO_3 > CaCO_3$

(c) $CaCO_3 > SrCO_3 > MgCO_3 > BaCO_3$ (d) $MgCO_3 > CaCO_3 > SrCO_3 > BaCO_3$

231. A solution of sodium metal in liquid ammonia is strongly reducing due to the presence of

(a) sodium atoms (b) sodium hydride (c) sodium amide (d) solvated electrons.

232. Sodium reacts with water more vigorously than lithium because, it

(a) has higher atomic weight (b) is a metal

(c) is more electropositive (d) more electronegative

233. The hydration energy of Mg^{2+} is greater than that of

(a) Al^{3+} (b) Na^{+} (c) Be^{2+} (d) Mg^{3+}

234. The compound insoluble in acetic acid is

(a) calcium oxide (b) calcium carbonate

(c) calcium oxalate (d) calcium hydroxide.

235. The metallic lustre exhibited by sodium is explained by

(a) diffusion of sodium ions (b) oscillation of loose electrons

(c) excitation of free electrons (d) existence of body centred cubic lattice

236. Gypsum on heating to 390 K gives

(a) $CaSO_4.2H_2O$ (b) $CaSO_4$

(c) $CaSO_4. 1 / 2 H_2O$ (d) SO_3 and CaO

237. The by-product of Solvay ammonia process is

(a) Carbon dioxide (b) Ammonia

(c) Calcium chloride (d) Calcium carbonate.

238. The drying agent which absorbs carbon dioxide and reacts violently with water is

(a) Sodium carbonate (b) Alcohol

(c) Conc. H_2SO_4 (d) Calcium oxide

239. Property of alkaline earth metals that increases with their atomic number is

(a) ionization energy (b) solubility of their hydroxides

(c) solubility of their sulphates (d) electronegativity

240. The compounds of alkaline earth metals have the following magnetic nature

(a) Diamagnetic (b) Paramagnetic (c) Ferromagnetic (d) Antiferromagnetic

241. The stability of the following alkali metal chlorides follows the order:

(a) LiCl > KCl > NaCl > CsCl (b) CsCl > KCl > NaCl > LiCl

(c) NaCl > KCl > LiCl > CsCl (d) KCl > CsCl > NaCl > LiCl

242. Epsom salt is

(a) $MgSO_4.7 H_2O$ (b) $CaSO_4.H_2O$

(c) $MgSO_4.2 H_2O$ (d) $BaSO_4.2 H_2O$

243. Among the alkaline earth metals, the element forming predominantly covalent compound is

(a) Barium (b) Strontium (c) Calcium (d) Beryllium

244. KO_2 (potassium superoxide) is used in oxygen cylinders in space and submarines because it

(a) absorbs CO_2 and increases O_2 content (b) eliminates moisture

(c) absorbs CO_2 (d) produces ozone

UNIT - XI

245. Which of the following is a p-block element ?

(a) Magnesium (b) Aluminium (c) Sodium (d) Neon

246. Borax on heating with cobalt oxide forms a dark blue bead of

(a) $Co(BO_2)_2$ (b) $CoBO_2$ (c) $Co_3(BO_3)_2$ (d) No reaction

247. The product formed in the reaction, $BCl_3 + H_2O \longrightarrow$ is

(a) $H_3BO_3 + HCl$ (b) $BO_3 + HOCl$ (c) $B_2H_6 + HCl$ (d) No reaction

248. Which of the following is not an ionic trihalide?

(a) AlF_3 (b) BF_3 (c) InF_3 (d) GaF_3

249. In which of the following, +1 oxidation state is stabler than +3 ?

(a) Ga (b) Al (c) Tl (d) B

250. Boron carbide, B_4C is widely used in ————.

(a) Making plaster of pans (b) Makmg boric acid

(c) As a hardest substance after diamond (d) Making acetylene

251. One that marks the paper like lead is ________.

(a) Ga (b) B (c) Ti (d) Tl

252. The two type of bonds present in diborane are covalent bond and _____.
(a) Hydrogen bridge bond (b) Co-ordinate bond
(c) Ionic bond (d) None.

253. Borax bead test is responded by
(a) divalent metals (b) metals which form coloured metaborates
(c) light metals (d) heavy metals

254. Alumina is
(a) Acidic (b) Basic (c) Amphoteric (d) Neutral

255. Which metal is protected by layer of its own oxide ?
(a) Al (b) Ag (c) Au (d) Fe

256. Aluminium forms
(a) Electrovalent compounds only (b) Covalent compounds only
(c) Electrovalent and covalent compounds both
(d) Coordinate compounds only.

257. Which of the following metals bums in air at high temperature with the evolution of much heat ?
(a) Cu (b) Hg (c) Pb (d) Al

258. Aluminium deposited as vapour on glass forms a good mirror, essentially because :
(a) Coating is much smoother (b) It does not tarnish in air
(c) It has better shine than silver (d) It does not scratch

259. The most covalent aluminium halide is
(a) Aluminium fluoride (b) Aluminium chloride
(c) Aluminium bromide (d) Aluminium iodide

260. The blue coloured minerals 'Lapos Lazule' which is used as a semi precious stone is a mineral of the following class.
(a) Sodim aluminium silicate (b) Boron trioxide
(c) Zinc cobaltite (d) Basic copper carbonate

261. Beryllium and aluminium exhibit many properties which are similar. But the two elements differ in
(a) exhibiting maximum covalency in compound.
(b) exhibiting amphoteric nature in their oxides
(c) forming covalent halides which are Lewis acids.
(d) forming bridged chloride in vapour phase.

262. Which of the following occurs in free state ?.
(a) C (b) Si (c) Ge (d) Sn

263. Diamond and graphite are
(a) isomers (b) allotropes (c) isotopes (d) polymers

264. A gas which burns with blue flame is
(a) CO (b) CO_2 (c) O_2 (d) N_2

265. Poisonous gas present in the exhaust fumes of a car is
(a) CH_4 (b) C_2H_6 (c) CO_2 (d) CO

266. Sugar of lead is
(a) $PbSO_4$ (b) $(CH_3COO)_2Pb$ (c) $PbCl_2$ (d) $PbCO_3$

267. Which of the following can attack glass ?
(a) H_2SO_4 (b) HI (c) HCl (d) HF

268. Silicon is an important- constituent of

(a) Chlorophyll (b) Haemoglobin (c) Rocks (d) Amalgams

269. Which of the following compounds has peroxide linkage ?

(a) Pb_2O_3 (b) CO_2 (c) PbO_2 (d) SiO_2

270. Property of catenation is strongest in

(a) C (b) Si (c) N (d) O

271. Which of the following is not a crystalline form of silica ?

(a) Azurite (b) Quartz (c) Tridymite (d) Crystobalite

272. The chemical name of Phosgene is

(a) Phosphine (b) Canbonyl chloride

(c) Phosphorus oxychloride (d) Phosphorus trichloride

273. Artificial Carborundum is

(a) Silicon carbide (b) Boron nitride (c) quartz (d) Calcium carbide

274. Red lead is

(a) PbO_2 (b) PbO^{2+} (c) $PbO_2 2PbO$ (d) None

275. The softest form of carbon is

(a) Diamond (b) Graphite (c) Charcoal (d) Lamp black

276. Which of the following is not used as a refrigerant

(a) NH_3 (b) CO_2 (c) SO_2 (d) CO

277. Thermodynamically the most stable form of carbon is

(a) diamond (b) graphite (c) fullerence (d) coal

278. Carbon forms a large number of compounds because of

(a) catenation (b) high electron affinity

(c) high ionisation energy (d) variable valency.

279. The coccept of vital force theory was given by –

(a) Lavoisier (b) Wohler (c) Berzelius (d) Berthelot

280. A compound with molecular formula C_2H_2 must contain

(a) all sigma bonds (b) all π-bonds

(c) three sigma and two π-bonds (d) two sigma and three π-bonds.

281. Alicyclic compound is

(a) an aromatic compound (b) a carbocyclic compound

(c) an acyclic compound (d) a heterocyclic compound.

282. The compound $R - CO - NH_2$ is an example of

(a) an amide (b) an amine (c) a ketone (d) a cyanide.

283. The hybridisation of carbons C_1 and C_3 in the compound is

$^{1}CH_3 - {}^{2}CH = {}^{3}CH - {}^{4}CH_3$

(a) sp^2, sp^3 (b) sp^3, sp^2 (c) sp, sp^3 (d) sp^2, sp^2

284. IUPAC name of succinic acid is :

(a) Buten –1, 4 – dioic acid (b) 3 – carboxypropanoic acid

(c) Butane–1, 4–dioic acid (d) 1, 2 – Dicarboxyethane

285. Which of the following compounds has isopropyl group?

(a) 2,2,3,3–Tetramethylpentane (b) 2,2 - Dimethylpentane

(c) 2,2,3–Trimethylpentane (d) 2 – Methylpentane

286. Isomerism exhibited by acetic acid and methyl formate is

(a) functional (b) geometrical (c) chain (d) position

287. Electrophiles are

(a) electron loving species (b) electron hating species

(c) nucleus loving reagents (d) nucleus hating reagents.

288. Homolysis leads to form–

(a) electrophile (b) nucleophile (c) free radicals (d) ions

289. Nucleophiles are

(a) electron loving species (b) electron hating species

(c) nucleus loving species (d) nucleus hating species.

290. Heterolytic cleavage of a carbon - carbon covalent bond gives

(a) carbanion (b) carbocation

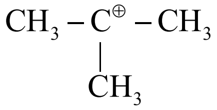
(c) free radical (d) both carbanion and carbocation

291. Chemically similar compounds differing by $-CH_2-$ group are

(a) isomers (b) allotropes (c) homologues (d) tautomers.

292. Which of the following is the most stable ion?

(a) $CH_3-CH_2-\overset{\oplus}{C}H-CH_3$ (b) $CH_3-CH_2-CH_2-\overset{\oplus}{C}H_2$

(c) 

$$CH_3-\underset{\displaystyle CH_3}{\underset{|}{C^{\oplus}}}-CH_3$$

(d) $\overset{\oplus}{C}H_3$

293. Which of the following would be optically active?

(a) tert. butanol (b) sec. butanol

(c) n - butanol (d) 4 - chloro butan–1–ol

294. Acetone and propen – 2 – ol are

(a) position isomers (b) keto-enol tautomers

(c) geometrical isomers (d) chain isomers

295. A pair of optically active compounds which are not mirror images are called

(a) mesomers (b) tautomers (c) enantiomers (d) diastereoisomers .

296. Which are isomers ?

(a) Ethanol and Ethoxyethane (b) Methanol and Methoxymethane

(c) Propionic acid and Ethylacetate (d) Propionaldehyde and Acetone

297. The hybridization of carbon atoms in C – C single bond in $HC\equiv C-CH=CH_2$ is

(a) sp^3-sp^3 (b) sp^2-sp^3 (c) $sp-sp^2$ (d) sp^3-sp

298. Which of the following orders regarding the electronagativity of hybrid orbitals of carbon is correct ?

(a) $sp > sp^2 > sp^3$ (b) $sp < sp^2 < sp^3$ (c) $sp > sp^2 < sp^3$ (c) $sp < sp^2 > sp^3$

299. A molecule is said to be chiral if it

(a) contains a centre of symmetry (b) contains a plane of symmetry

(c) can not be superimposed on its mirror image (d) exists as cis and trans forms

300. Which of the following is not a nucleophile ?

(a) CN^- (b) NH_3 (c) BF_3 (d) OH^-

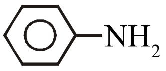
301. Which of the following in the most stable radical ?

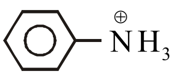
(a) $\dot{C}H_3$ (b) $R\dot{C}H_2$ (c) $R_2\dot{C}H$ (d) $R_3\dot{C}$

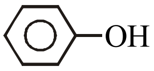
302. In Kjeldahl's method, the nitrogen present in organic compound is estimated in the form of_____

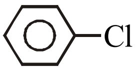
(a) NO (b) N_2O (c) N_2H_4 (d) NH_3

303. In which of the following molecules, the resonance effect is not present ?

(a) 

(b) 

(c) 

(d) 

304. An electrophilic reagent must have

(a) a vacant orbital
(b) an orbital containing one electron
(c) an orbital containing two electrons
(d) all completely filled orbitals

305. Methanol and acetone can be separated by

(a) fractional distillation
(b) distillation
(c) steam distillation
(d) vacuum distillation

306. Which of the following mixture would steam distillation be the most appropriate method of separation

(a) Diethyl ether and water
(b) Ethyl alcohol and water
(c) Aniline and sodium chloride
(d) none of these

307. In quantitiave estimation of hydrogen 0.1 g of an organic compound yielded 0.09g of water. The percentage of hydrogen in the compound is

(a) 8 (b) 10 (c) 12 (d) 15

308. In Duma's method for the estimation of nitrogen the sample is heated with copper oxide and the gas evolved is passed over

(a) Nickel (b) Copper oxide (c) Copper gauze (d) None of these.

309. In the Carius method for the estimation of halogens, the sample is heated with

(a) Conc. HNO_3
(b) Conc. H_2SO_4 and Ag
(c) Conc. HNO_3 and $AgNO_3$
(d) None of these

UNIT - XIII

310. The general formula of alkyne is

(a) C_nH_{2n} (b) C_nH_{2n-2} (c) C_nH_{2n+2} (d) C_nH_n

311. The C—C bond is shorter in

(a) Ethane (b) Ethylene (c) Acetylene (d) Propane.

312. When potassium maleate is electrolysed, the compound produced at the anode is

(a) Ethyne (b) Ethane (c) Ethene (d) Propyne

313. The product of the reaction $CH_3—C\equiv CH + 2HCl$ is

(a) $CH_3.C.Cl_2 - CH_3$
(b) $ClCH_2 - CHCl - CH_3$
(c) $ClCH_2 - CH_2 - CH_2 - Cl$
(d) $CH_3 - XH_2 - CHCl_2$

314. After ozonolysis of but-2-yne the product on hydrolysis is :

(a) Formic acid
(b) Acetic acid
(c) Formic acid and acetic acid
(d) Propionic acid

315. The alkyne $R—C\equiv C—R$ does not react with

(a) HOCl
(b) Br_2
(c) O_2
(d) Ammoniacal cuprous chloride

316. Ethene and ethyne may be distinguished by using

(a) Baeyer's reagent
(b) Tollen's reagent
(c) Bromine water
(d) Schiffs reagent

317. What is the number of possible alkynes with formula C_5H_8 ?
(a) 12 (b) 3 (c) 4 (d) 5

318. Acetylene reacts with sodamide in liquid ammonia and then reacts with CH_3I producing
(a) But-2-yne (b) 3-Methylbut-l-yne
(c) Propyne (d) Acetylene tetraiodide

319. Acidic hydrogen is present in
(a) Ethyne (b) Ethene (c) Benzene (d) Ethane

320. Lindlar's catalyst is
(a) Na in liquid NH_3 (b) Pt in ethanol
(c) Ni in ether (d) Pd with $BaSO_4$ and quinoline

321. When propyne is treated with aqueous H_2SO_4 in the presence of $HgSO_4$, the major product is
(a) Propanol (b) Propyl hydrogen sulphate
(c) Acetone (d) Propanol

322. When phenol is distilled with zinc dust the product is
(a) Toluene (b) Benzene (c) Xylene (d) None of the above

323. In Benzene the 'C' atoms are
(a) sp^3 hybridised (b) sp^2 hybridised (c) sp hybridised (d) Unhybridised

324. Coaltar is a main source of
(a) Aromatic compounds (b) Aliphatic compounds
(c) Cycloalkanes (d) Heterocylic compounds

325. The following products are formed by fractional distillation of Coal tar. Which one is basic ?
(a) Phenol (b) Toluene (c) Anthracene (d) Pyridine

326. All bonds is benzene are equal due to
(a) tautomerism (b) I-effect (c) Resonance (d) Isomerism

327. Phenol is less acidic than
(a) o-Cresol (b) p-methoxy phenol (c) p-nitrophenol (d) Ethanol

328. Nitro group in nitrobenzene is
(a) ortho director (b) meta director
(c) para director (d) ortho & para director

329. Nitration of benzene is
(a) Nucleophilic substitution (b) Electrophilic substitution
(c) Homolytic substitution (d) Electrophilic addition

330. The compound that is most reactive towards electrophilic nitration is
(a) Toluene (b) Benzene (c) Benzoicacid (d) Nitrobenzene

331. Benzene on oxidation in presence of V_2O_5 at 500^0C gives ——
(a) CO_2 and H_2O (b) Maleic anhydride (c) Maleic acid (d) Succinic acid

332. Anhydrous $AlCl_3$ is used in Friedel Craft reaction because it is
(a) electron rich (b) soluble in ether (c) insoluble (d) electron deficient

333. In which can the C—C bond length is same
(a) But-2-ene (b) But-1-ene (c) Benzene (d) Prop-1-yne

334. Which of the following pairs will give the same bond angle ?
(a) Ethane & ethyne (b) ethane & ethene (c) Ethene & benzene (d) Ethyne & benzene

335. The reaction of Toluene with Cl_2 in presence of $FeCl_3$ gives predominantly
(a) m-chlorotoluene (b) benzoyl chloride
(c) o - & p chlorotoluene (d) benzal chloride

336. Which of the following compound is likely to be carcinogemic ?
(a) Toluene (b) Aniline (c) naphthalane (d) 1, 2-benz pyrene

337. C — C bond lenght in benzene is
(a) $1.54A^0$ (b) $1.20A^0$ (c) $1.39\ A^0$ (d) $1.34\ A^0$

338. Benzoic acid when heated with sodalime gives
(a) benzene (b) sodium benzoate (c) benzaldehyde (d) benzyl alcohol

339. p-Nitrophenol is a stronger acid than phenol because nitro group is
(a) Electron donating (b) Electron withdrawing
(c) Basic (d) Acidic

340. Which is the correct statement ?
(a) Benzyl alcohol is more acidic than phenol (b) Ethanol is a powerful oxidising agent
(c) Phenol is more acidic than propanol. (d) Ethane has higher boiling point than ethanol

341. The direct iodination of benzene is not possible because
(a) Iodine is an oxidizing agent (b) Resulting C_6H_5 I is reduced to C_6H_6 by HI
(c) HI is unstable (d) The ring gets deactivated

342. The electrophile in aromatic sulphonation is
(a) $^{\oplus}SO_3H$ (b) SO_3 (c) HSO_3^- (d) H_2SO_4

343. Chlorination of toluene in the presence of $FeCl_3$ gives predominantly
(a) Benzoyl chloride (b) m-chlorotoluene
(c) benzyl chloride (d) o– and p–chlorotoluenes

UNIT - XIV

344. Which of the following is the uppermost region of the atmosphere ?
(a) Stratosphere (b) Troposphere (c) Exosphere (d) Ionosphere

345. Which of the following is the coldest region ?
(a) Troposphere (b) Mesosphere (c) Stratosphere (d) Thermosphere

346. Which of the following is a secondary pollutant ?
(a) NO (b) NO_2
(c) SO_2 (d) peroxyacyl nitrate (PAN)

347. Which of the following is not considered to be a pollutant ?
(a) NO_2 (b) CO_2 (c) O_3 (d) C_xH_y

348. Which of the following has greatest affinity for haemoglobin ?
(a) CO (b) NO (c) O_2 (d) CO_2

349. The most abundant hydrocarbon pollutant is
(a) Methane (b) Ethane (c) Propane (d) Butane

350. Which of the following is not involved in formation of photochemical smog ?
(a) NO (b) O_3 (c) C_xH_y (d) SO_2

351. The aromatic compounds present as particulates are
(a) Benzene (b) Toluene
(c) Nitrobenzene (d) Polycyclic hydrocarbons

352. Which of the following is present in maximum amount in acid rain ?

(a) HNO_3 (b) H_2SO_4 (c) HCl (d) H_2CO_3

353. London smog is found in

(a) Summer during day time (b) Summer during morning time

(c) Winter during morning time (d) Summer during day time

354. Photochemical smog is formed in

(a) Summer during morning time (b) Summer during day time

(c) Winter during morning time (d) Winter during day time

355. Which of the following statement is false ?

(a) London smog is oxidising in nature (b) London smog contains H_2SO_4 droplets

(c) London smog is formed in winter (d) London smog causes bronchitis.

356. Which of the following is true about photochemical smog ?

(a) It is reducing in nature (b) It is formed in winter

(c) It is a mixture of smoke and fog (d) It causes irritation in eyes.

357. 'White lung cancer' is caused by

(a) asbestos (b) silica (c) textiles (d) paper

358. Ozone layer is present in

(a) troposphere (b) stratosphere (c) mesosphere (d) exosphere

359. Depletion of ozone layer causes

(a) blood cancer (b) lung cancer (c) skin cancer (d) breast cancer

360. Which of the following is not a greenhouse gas ?

(a) CO_2 (b) CH_4

(c) chlorofluorocarbons (d) O_2

361. Ozone in the stratosphere is deleted by

(a) CF_2Cl_2 (b) C_7F_{16} (c) $C_6H_6Cl_6$ (d) C_6F_6

362. The smog is essentially caused by the presence of

(a) O_2 and O_3 (b) O_2 and N_2

(c) Oxides of sulphur and nitrogen (d) O_3 and N_2

◆◆◆

ANSWERS TO MULTIPLE CHOICE QUESTIONS

GROUP - A (1)

1.	d	2.	d	3.	c	4.	d	5.	d	6.	c	7.	b
8.	b	9.	a	10.	a	11.	c	12.	b	13.	a	14.	a
15.	a	16.	a	17.	b	18.	a	19.	d	20.	a	21.	b
22.	b	23.	b	24.	a	25.	b	26.	b	27.	c	28.	a
29.	a	30.	b	31.	a	32.	a	33.	b	34.	a	35.	d
36.	d	37.	d	38.	d	39.	a	40.	a	41.	c	42.	d
43.	a	44.	b	45.	a	46.	a	47.	b	48.	d	49.	a
50.	c	51.	b	52.	b	53.	a	54.	b	55.	a	56.	a
57.	d	58.	b	59.	c	60.	b	61.	d	62.	c	63.	d
64.	a	65.	d	66.	a	67.	a	68.	d	69.	a	70.	b
71.	a	72.	d	73.	a	74.	c	75.	b	76.	a	77.	a
78.	c	79.	a	80.	d	81.	d	82.	c	83.	c	84.	c
85.	c	86.	c	87.	a	88.	b	89.	b	90.	c	91.	b
92.	d	93.	a	94.	b	95.	c	96.	d	97.	d	98.	d
99.	b	100.	d	101.	d	102.	a	103.	a	104.	b	105.	a
106.	d	107.	a	108.	c	109.	b	110.	a	111.	a	112.	c
113.	c	114.	a	115.	c	116.	d	117.	b	118.	c	119.	d
120.	d	121.	b	122.	b	123.	d	124.	b	125.	c	126.	c
127.	c	128.	b	129.	d	130.	b	131.	c	132.	d	133.	a
134.	b	135.	b	136.	c	137.	b	138.	b	139.	b	140.	d
141.	d	142.	d	143.	d	144.	d	145.	b	146.	a	147.	b
148.	b	149.	b	150.	a	151.	b	152.	b	153.	b	154.	a
155.	c	156.	c	157.	a	158.	d	159.	b	160.	d	161.	b
162.	b	163.	a	164.	d	165.	b	166.	d	167.	c	168.	a
169.	d	170.	b	171.	d	172.	d	173.	d	174.	b	175.	b
176.	a	177.	c	178.	c	179.	d	180.	c	181.	b	182.	c
183.	a	184.	a	185.	b	186.	a	187.	a	188.	b	189.	c
190.	b	191.	a	192.	b	193.	d	194.	b	195.	c	196.	c
197.	d	198.	c	199.	c	200.	b	201.	c	202.	d	203.	b
204.	b	205.	d	206.	b	207.	b	208.	b	209.	d	210.	b
211.	c	212.	b	213.	c	214.	c	215.	a	216.	b	217.	a
218.	c	219.	c	220.	d	221.	b	222.	d	223.	d	224.	d
225.	a	226.	a	227.	a	228.	c	229.	d	230.	a	231.	d
232.	c	233.	b	234.	c	235.	b	236.	c	237.	c	238.	d
239.	b	240.	a	241.	d	242.	a	243.	d	244.	a	245.	b
246.	a	247.	a	248.	b	249.	c	250.	c	251.	d	252.	a
253.	b	254.	c	255.	a	256.	c	257.	d	258.	b	259.	d

260.	a	261.	a	262.	a	263.	b	264.	a	265.	d	266.	b
267.	d	268.	c	269.	c	270.	a	271.	a	272.	b	273.	a
274.	c	275.	d	276.	d	277.	b	278.	a	279.	c	280.	c
281.	b	282.	a	283.	b	284.	c	285.	d	286.	a	287.	a
288.	c	289.	c	290.	d	291.	c	292.	c	293.	b	294.	b
295.	d	296.	d	297.	c	298.	a	299.	c	300.	c	301.	d
302.	d	303.	b	304.	a	305.	a	306.	c	307.	b	308.	c
309.	c	310.	b	311.	c	312.	a	313.	a	314.	b	315.	d
316.	b	317.	b	318.	a	319.	a	320.	d	321.	c	322.	
323.	b	324.	a	325.	d	326.	c	327.	c	328.	b	329.	b
330.	a	331.	b	332.	d	333.	c	334.	c	335.	c	336.	d
337.	c	338.	a	339.	b	340.	c	341.	b	342.	b	343.	d
344.	c	345.	b	346.	d	347.	b	348.	b	349.	a	350.	d
351.	d	352.	b	353.	c	354.	b	355.	a	356.	d	357.	c
358.	b	359.	c	360.	d	361.	a	362.	c				

Hints to some multiple choice questions (Group - A-1)

4. P (pico) = 10^{-12}

5. M 72.4% $\frac{72.4}{M}$, $\frac{72.4}{M} \times \frac{6}{27.6} = \frac{3}{4} \Rightarrow M = \frac{72.4 \times 16 \times 4}{27.6 \times 3} = \frac{4633.6}{82.8} = 56$

O 27.6% $\frac{27.6}{16}$

M 70% $\frac{70}{56}$, $\frac{70}{56} \times \frac{16}{30} = 0.66 = \frac{2}{3} = M : O$

O $\frac{30}{16}$

The formula = M_2O_3

$\frac{\text{Percentage}}{\text{atomic mass}}$ = Relative number of atoms.

7. Less is the no. of gram atom, Less is the number of atoms.

9. % C in $CH_3COOH = \frac{24}{60} \times 100$

% C is $C_6H_{12}O_6 = \frac{72}{180} \times 100$

12. mass of $so_4^{2-} = 57.8$ $\frac{57.8}{48} = \frac{42.2}{E}$, $E = 35.0$

mass element = 42.2

14. $PH_3(g) \longrightarrow P_{(s)} + \frac{3}{2}H_2(g)$ increase of 50 ml

100 ml 150 ml,

15. 25 ml of 0.75 M $= \frac{25\times 0.75}{1000}$ mole

$CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2$

2 mole of HCl $\longrightarrow$ 1 mole of $CaCO_3$ = 100 gram

$$\frac{25\times 0.75}{1000}\text{mole} \longrightarrow \frac{1000\times 25\times 0.75}{2\times 1000}$$

19. $X = 75.8 \qquad \frac{75.8}{75} = 1.01$

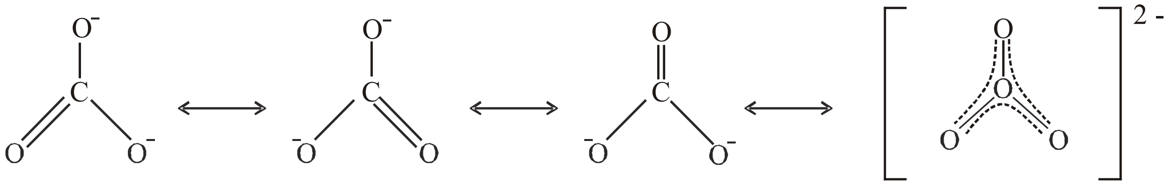
$Y = 24.2 \qquad \frac{24.2}{16} = 1.5 \qquad$ Ratio X : Y = 2 : 3 i.e., X_2Y_3

20. 4.2 gram $= \frac{4.2}{14} = 0.3$ gram atom

1 gram atom contains $8.N_A$ valence electrons

0.3 gram $\Rightarrow 8\times 0.3 = 2.4\,N_A$ valence electrons.

96. $\overline{C}\equiv\overset{+}{O},\quad \overline{O}-C\equiv\overset{+}{O} \longleftrightarrow O=C=O \longleftrightarrow \overset{+}{O}\equiv C-\overline{O}$



105. $$\frac{P_1\ V_1}{T_1} = \frac{{}^{P_1}\!/_{2}\ V'}{2\ T_1}$$

$$V' = \frac{P_1V_1\times 2\ T_1\times 2}{P_1\ T_1} = 4\ V_1$$

107. PV = nRT Molar volume at NTP = 22.4
Volume of 10 moles = 224 L

108. 1 cc. of hydrocarbon NTP = x gram = 4 cc of oxygen

22400 c.c. oxygen $= \frac{x}{4}\times 22400 = 32$

$x = \frac{32\times 4}{22400}$ = mass of 1 c.c Hydrocarbon.

Mass of 22400 cc. Hydrocarbon $= 32\times 4 = 128$

109. $$\frac{P_1}{T_1} = \frac{P_2}{T_2},\quad \frac{P_1}{250} = \frac{P_2}{251}$$

$P_2 = \frac{P_1.251}{250}$ Increase in pressure $= \frac{P_1.251}{250} - P_1 = \frac{P_1}{250}$

% of increase $= \frac{P_1\times 100}{250\times P_1} = 0.4$

113. $\frac{{}^{r}\mathrm{He}}{r_B} = 2 = \sqrt{\frac{d_B}{2}} \quad \Rightarrow \; d_B = 8$

Mol. mass = $2 \times d_B = 16$

114. $\frac{\frac{50}{t}}{\frac{40}{t}} = \sqrt{\frac{M}{64}} = \frac{5}{4} \quad \Rightarrow \quad \frac{M}{64} = \frac{25}{16} \quad \Rightarrow \quad M = 100$

120. $P = \frac{nRT}{V} = \frac{1 \times 0.082 \times 273}{1}$

128. $PV = \frac{1}{3} mN\bar{C}^2 = RT$

$\frac{16\bar{C}^2}{T} = \frac{32 \times \bar{C}^2}{300} \quad \Rightarrow T = 150$

165. $\left[OH^-\right] = \frac{0.4}{4.0} \mathrm{mol\,L^{-1}} = 10^{-2}$

pH = 14 –pOH = 12

166. [H^+] is reduced from 10^{-3} to 10^{-6} i.e, decreased 1000 times.

169. From acid [H^+] = 10^{-8}M

From water [H^+] = 10^{-7}, total [H^+] = $(10^{-8} + 10^{-7})$M = 11×10^{-8} M

pH = – $\log^{11}$ + 8 = – 1.04 + 8 = 6.96

178. 50 ml of 0.4 N HCl = 0.02 gram eqvt.

50 ml of 0.2 N NaOH = .01 gram eqvt.

Residual acid = 0.01 gram eqvt in 100 ml

= 0.1 gram eqvt in 1000 ml

[H^+] = 10^{-1}M, pH = 1.0

182. $M^{x+} + \underbrace{\overset{+7}{M}_n O_4^- \longrightarrow \overset{+5}{M}O_3^-}_{-5} + M_n^{2+} + \frac{1}{2}O_2$

1 mole of MnO_4^- oxidises 1.67 moles of M^{x+}

1 gram eqvt of MnO_4^- $\frac{1.67}{5}$ mole = 1gram eqvt of M^{x+}

$= \frac{1}{5-x}$ mole $\quad \left(\text{Eqvmass} = \frac{\text{mol. mass}}{\text{change in O.N}} \right)$

$\Rightarrow -1.67\,x + 8.35 = 5$

$x = \frac{3.35}{1.67} = 2$

187. $\left[MnO_4^- + 8H^+ + 5e^- \longrightarrow M_n^{2+} + 4H_2O\right] \times 3$

$$\begin{cases} Fe^{2+} \longrightarrow Fe^{3+} + e^- \\ C_2O_4^{2-} \longrightarrow 2CO_2 + 2e \end{cases}$$

$$\left[5.\ FeC_2\ O_4 \equiv 3 \text{ moles } MnO_4^-\right]$$

$$1 \text{ mole of } Fe\ C_2\ O_4 \equiv \frac{3}{5} = 0.6 \text{ mole of } MnO_4^-$$

199. $$\left.M_nO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O\ \right] \times 2$$

$$\left.H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-\right] \times 5$$

2 moles of MnO_4^- $\equiv$ 5 moles of H_2O_2

200. $2\ CuSO_4 + 5\ KI \longrightarrow Cu_2\ I_2 + I_2 + 2\ K_2SO_4$

$2\ Na_2\ S_2\ O_3 + I_2 \longrightarrow Na_2\ S_4O_6 + 2\ Na\ I$

Sod. tetrathionate

201. Hydrogen as oxidant, $\overset{0}{Na} + \frac{1}{2}\overset{0}{H}_2 \longrightarrow \overset{+}{N}a\,H^-$

as reductant, $CuO + H_2 \longrightarrow Cu + H_2O$

205. Normality of 10 V $H_2O_2 = \frac{10\times 68}{22.4\times 17}N$

$$V^{/} \times \frac{40}{22.4}N = 200\times 2N$$

$$V^{/} = \frac{400\times 22.4}{40} = 224 \text{ ml}$$

214. In H_2O_2, O–H bond polar, O–O bond polar.

215. $A(Na) + H_2O \longrightarrow \underset{(C)}{NaOH} + \frac{1}{2}H_2(B)$

Zn (D) + 2NaOH(C) $\longrightarrow$ $Na_2\ ZnO_2 + H_2$(B)

Zn (D) + $H_2SO_4 \longrightarrow ZnSO_4 + H_2$

Na imparts golden yellow flame.

237. $2NH_4Cl + Ca(OH)_2 \longrightarrow CaCl_2 + 2NH_3 + 2H_2O$

241. ΔH_f KCl (-428 kJ/mol) CsCl (-424), NaCl (-400), LiCl (-399) More negative value of the heat of formation more is the stability.

244. $4KO_2 + 2CO_2 \longrightarrow 2K_2CO_3 + 3O_2$

271. Azurite $Cu_3(CO_3)_2\ (OH)_2$, others are crystalline forms of silica.

307. 18 gram water contains 2 gram 'H'

0.09 gram contains $\frac{2}{18}\times 0.09$ gram

0.1 gram organic sample contains $\frac{0.18}{18}$ gram of H

So % of H $= \frac{0.18\times 100}{18\times 0.1} = 10\%$

318. $H-C\equiv C-H \xrightarrow[NH_3(\ell)]{NaNH_2} NaC\equiv CNa \xrightarrow{2CH_3I} CH_3-C\equiv C-CH_3$

But – 2 – yne

GROUP - A (2)
(Very short questions - One mark each)
Fill in the blanks :

UNIT - I

1. Mass of a single water molecule is ___________ gram.
2. Mass of 0.1 mole of water molecule is __________ kg.
3. The percentage of nitrogen in urea is ________ .
4. The mass of CO_2 formed on buring 2 moles of carbon in 16 gram of dioxygen in _______ gram.
5. Number of moles of hydrogen atoms in 3 moles of ethane is _______.
6. 200 ml of 0.1 N NaOH contains _____ moles of sodium hydroxide.
7. Equivalent mass of zinc is ______ gram.
8. Balance the equation. $P + HNO_3 \rightarrow H_3PO_4 + NO_2 + H_2O$ by partial equation method.
   The coefficient of NO_2 is _________.
9. Balance by oxidation number method. The coefficient of HNO_3 is _________.
10. Balance by ion-electron method. The coefficient of PO_4^{3-} is _________.
11. Iron pyrite on burning with oxygen provides _______ gas.
12. O.N. of chlorine is perchloric acid is ________ .
13. 25^0C = ——— F. = ——— ^0A.
14. Human body temperature in celsius scale is ________.
15. Unit of Luminous intensity is _________.

UNIT - II

16. Cathode rays are a source of _______ .
17. Cathode rays start from ________ move towards _________ .
18. Television picture tubes are nothing but _________.
19. Mass of an electrons is ________ kg.
20. The change carried by an electron is _______ C.
21. The particles emitted when Be is bombarded with $\alpha -$ rays are _______ .
22. _________ is the universal constituent of all matter.
23. The radius of the atom is about is about ______m while that of the nucleus is _______ m.
24. In the visible spectrum _____ zone has lower wave length.
25. The energy of 3rd orbit of Bohr atom is _____ ev.
26. An orbital with $n = 3 \quad \ell = 0 \quad s = -\frac{1}{2}$ is called ______ orbital.
27. The number of spherical nodes present in 4s orbital of hydrogen atom is _______.
28. Canal rays are discovered by ________ on perforation of _________ .
29. The number of unpair electron(s) nitrogen atom is _________.

30. In La(z = 57), the 57th $\overline{e}$ enters into ______ orbital

UNIT - III

31. Mendleef’s periodic table is based on ____________.

32. According to Newland’s octave the element identical to Be is ________ .

33. The first three elements in Dobereiner triad are ________ .

34. The atomic number of Eka-silicon is _______ .

35. According IUPAC carbon family belongs _________ group.

36. The IUPAC name of element with Z = 109 is _____ .

37. The present name of element Z = 112 is ______ and its symbol is _______.

38. Group - 13 contains ______ number of elements.

39. Sixth and 7th periods are longest periods each containing _____ number of elements.

40. In third period of elements elements are getting filled into ______ orbitals.

41. The first two coloumns constitute ______ block of elements.

42. The general electronic configuration of transition elements is _____.

43. _____ is the first transition metal.

44. _____ is the liquid non-metal in periodic table.

45. Nb and _____ have identical radii.

46. Group 17 elements are called halogens while group-16 elements are ______ .

47. The elements after uramium are called _______ .

48. _______ is the most electronegative element.

49. The abundant transition metal is ______ .

50. The metal with highest m.p. is ______ .

UNIT - IV

51. Shape of NH_3 is ______ while that of $\overset{+}{N}H_4$ is __________ .

52. In NH_3 and $\overset{+}{N}H_4$ ‘N’ undergoes _________ hybridisation.

53. The shape of C_2H_2 is __________ where ‘C’ undergoes _______ hybridisation.

54. In NaCl each Na^+ is surrounded by _____ Cl^- ions.

55. Ionic bond is formed by _________ of $\overline{e}$ (s).

56. Ionic bond is formed between two __________ atoms.

57. o–nitrophenol exhbits ________ type of H-bonding.

58. The number of unpair electrons present in O_2 molecule is ________ .

59. The angle between C–H bonds is methane is ________.

60. Magnetic behaviour of NO is ________ .

61. The angle between two adjacent sp^2 hybrid orbital is ________ .

62. NH_4^+ contains ________ and _______ type of bonds.

63. $(CN)_2C = C\ (CN)_2$ contains ______ σ bonds and ________ π bonds.

64. π-bonds can be formed by overlapping of ________ and _______ orbitals.

65. _______ orbital fails to form ________ bond but can from _______ bond.

66. Higher in the lattice energy _______ is the melting point.

67. During bond formation the distance between the atoms _______ and the potential energy _______.

68. The shape of PO_4^{3-} ion is _________ .

69. In ethane _______ orbital overlapping takes place, to form C–C sigma bond.

70. Element, A has 3 electrons is outermost shell & element, B has 6 $\bar{e}$(s) in the outermost shell. The formula of the compound is _________.

UNIT - V

71. There are _________ elements exhibit as gases.

72. Boyle's law in terms of density is expressed as _________.

73. Molar volume at NTP is _________.

74. Critical temperature of CO_2 is _________ ^{0}C.

75. For ideal gas the compresibility factor is __________.

76. The temperature at which real gas obeys ideal gas is _________.

77. Surface tension of a liquid decreases with _________ of temperature.

78. Greater is inter molecular force _________ is viscous force.

79. The flow in which there is regular gradation of velocity from one layer to other is called ______ flow.

80. The unit of η. in SI unit is ________.

81. When gas is allowed to pass through a nozzle the process is called __________ .

82. The process of diffusion is directly related to __________ .

83. The collision between ideal gas molecules is called __________.

84. The volume of a gas at 0^0C is 20 L at a given pressure, its volume at 10^0C will be _________.

85. The Boltzman constant (K) is represented as _________.

UNIT - VI

86. At constant volume q equals to ________.

87. At constant pressure q equals to ________.

88. Thermodynamically, adsorption is an ________ phenomenon.

89. Entropy of gas ________ than that of solid.

90. Entropy is a measure of ________ of the system.

91. $NH_{3(g)} + HCl_{(g)} \longrightarrow NH_4Cl(s)$, the ΔS value of this reaction is _________.

92. Mass is an example of _________ property.

93. Entropy of pure substance is zero at _________.

94. $\Delta H = \Delta G$ at ________ as per third law.

95. The sign of heat of neutralisation is _________.

96. The entropy of combustion is always ________.

97. When ice melts to water, its entropy _________.

98. ΔG value at equilibrium is _________.

99. ΔG is ________ when the process is spontaneous.

100. The unit entropy is _________ .

101. $\left(\Delta S_{system} + \Delta S_{surr}\right) > 0,$ for _______ process.

102. Natural pressures are _________ and irreversible.

103. Work obtained in __________ process is maximum.

104. When gas expands in vacuum, the work obtained in this case is __________.

105. For expansion, $\Delta V = +ve,$ the workdone is having _________ sign .

106. $C_{(s)} + O_2(g) \longrightarrow CO_2(g),$ Here heat of formation is equal to heat of _________.

107. $\left(\Delta_{fus}H + \Delta_{vap}.H\right)$ is termed as heat of _________.

108. For exothermic reaction H_p is always _______ than $H_{(r)}$.

109. Bomb calorimeter is generally used to measure _________.

110. $p.\left(\frac{\partial V}{\partial T}\right)_P$ = ______ for 1 mole of ideal gas.

UNIT - VII

111. The state of equilibrium is reached for __________ reaction.

112. The concentration reactant(s) and product(s) remains ________ at the equilibrium state.

113. For liquid $\rightleftharpoons$ vapour equilibrium, _________ is constant at a given temperature.

114. For a gaseous system equilibrium is possible only is a ________ system.

115. Even if concentration / partial pressure remains constant for both reactants and products the equilibrium is ________.

116. The rate of formation of NH_3 is ________ times the rate of disappearance of hydrogen.

117. For the equilibrium, $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$, the equilibrium constant, K_p = _______.

118. If $Q_C > K_C$, the reaction proceeds _______ direction.

119. The catalyst used in Haber's process to synthesise NH_3 is _________ .

120. The concentration of H^+ of solution having pH = 2 is _______.

121. Conjugate base of a stronger acid is _________.

122. The volume of (pH + pOH) at 298K is ________.

123. p K_b significance indicates ________.

124. The acidic character of HF > H_2O > NH_3 > CH_4 so its conjugate bases will be in the order ______.

125. The nature of aqueous ammonia chloride is _________ while that of sodium acetate (Aq) is ______.

UNIT - VIII

126. Increase in oxidation number of an element in a process is called _________.

127. Addition of hydrogen _______ the oxidation number of the element is the process.

128. On removal of $\overline{e}(s)$ the oxidation state of the element ________.

129. $H_2S + Cl_2 \rightarrow 2\,HCl + S$, in this process, the O.N. of S ______ while that of chloride _______.

130. The reaction where oxidation and reduction take place simutaneously is called ________ reaction.

131. ________ reaction takes place in an electrochemical cell.

132. ________ takes place at anode while ________ reaction takes place at cathode.

133. The redox reaction where same element / compound undergoes both oxidation and reduction is called ___________.

134. The O.N. of Cr in CrO_5 is ________.

135. The element that exhibits only negative O.S. is ___________.

136. $Fe^{2+} \rightarrow Fe^{3+} + \overline{e}$ Here Fe^{2+} acts as _________.

137. The O.N. of 'O' is potassium superoxide is __________ .

138. In metal carbonyls and nitrosyls metal exhibits _______ O.N.

139. Zn on action of dil. H_2SO_4 undergoes ________ liberating ________ gas.

140. Phosphorous reacts with dil. NaOH to liberate PH_3 where phosphorous undergoes ________.

141. Equivalent mass of $K_2Cr_2O_7$ is __________ times its molecules mass in acidic medium.

142. Equivalent mass of H_3PO_4 (Mol. mass = 98) is _________ time its molecular mass if it forms Na_2HPO_4 on treatment with NaOH.

143. 100 ml of 0.1 N HCl completely reacts with 0.12 gram of a metal. So equivalent mass of metal is ________.

144. ________ ml of 1M H_2SO_4 would require 100 ml of 1 N NaOH.

145. The algebraic sum of oxidation number of all atoms of neutral molecule must be _________.

UNIT - IX

146. Zinc with NaOH liberates _________ gas.

147. __________ gas is prepared from petrochemicals by water-gas shift reaction.

148. Calcium carbide with heavy water liberates _________ gas.

149. 1L of '10 vol. of H_2O_2 liberated _______ L of ______ gas at NTP.

150. Number of water molecule undergoing H-bonding in $CuSO_4 . 5H_2O$ is _________.

151. A water molecule can form _________ number of H-bonding simultaneously.

152. The reaction where water acts as a reactant is generally called ________.

153. Water is a polar molecule with bond angle _______.

154. The major source of water is nature is ________.

155. In metal hydride hydrogen has ________ O.N.

156. Metal hydride on action with water liberates _________ gas.

157. During laboratory method to remove impurities like PH_3 and AsH_3, the impure gas is passed through _________.

158. P_2O_5 or anhydrous $CaCl_2$ acts as _________ during purification hydrogen gas.

159. The metal that liberates $H_2(g)$ on heating dil HNO_3 is _________.

160. 'Inflamable air' is named as ________ by ________ .

161. Tritium contains _________ number of neutrons.

162. o–H_2 and p-H_2 are _________ of hydrogen.

163. $IO_4^- + H_2O_2 \longrightarrow IO_3^- + O_2 + H_2O$, Here hydrogen peroxide acts as _______ agent.

164. The conversion of atomic hydrogen to ordinary hydrogen is ________.

165. Tritium on radioactive decay gives ________ and turns into __________.

UNIT - X

166. The general electronic configuration of s-block of elements is ________.

167. The penultimate shell of alkalimetal contains _________ electron except _________ that contains only ________ electron.

168. Between Mg and Ca _________ is more electro positive.

169. Na^+, Mg^{2+}, Al^{3+} are _________ and the kernal contains _________no. of electrons.

170. $BeCl_2$ fumes in air, as it forms _______ on hydrolysis.

171. Lime water becomes turbid with CO_2/SO_2 due to formation _______ / _________.

172. On excess of gas the turbidity disappears due to formation of_______ / ________.

173. Be is diagonally related to __________ .

174. BeO is _______ is nature.

175. __________ gas is liberated when Mg. nitride is heated with water.

176. The formula of epsom salt is _________ and that of gypsum salt is ___________.

177. All enzymes, undergoing phosphate transfer using ATP, need _______ as co-factor.

178. Most of the body calcium present in bones and __________.

179. The green pigment of the plant contains _________ which makes photosynthesis possible.

180. In the laboratory _______ is used as drying agent for NH_3(g).

181. Calcium hydride is commonly known as _________ .

182. Washing soda is synthesized by ________ process.

183. The magnetic behaviour of potassium superoxide is ________.

184. The alkali used in the manufacture of soft soap is ________.

185. ________ metal is used in devising photo-electric cells.

186. The alkali metal with least melting point is ________.

187. Li fails to form _______ oxide.

188. The alkali metal chloride forming hydrate is _________.

189. Sodium imparts golden yellow while potassium imparts _______ colour to the flame.

190. Soda-lime is a mixture of __________ and _________ .

UNIT - XI

191. Ionic potential of boron is almost same as to that of _________.

192. Anhydride of ortho boric acid (H_3BO_3) is _______.

193. Diborane acts as Lewis _________.

194. Banana bonds present in _________.

195. Inorganic benzene is named as ________.

196. The glassy bead formed on strongly heating borax is _________.

197. Nickel salt on borox bead test looks _________.

198. Borax on heating with dil. HCl results _________.

199. Complete the equation : $Al_2O_3 + NaOH \longrightarrow$ ______ + H_2O .

200. Aluminium oxide is _________ in nature.

201. $AlCl_3$ forms a dimer but ________ does not form dimer in the group-13.

202. Aluminium becomes passive with ________.

203. Alums are used as _________ in dyeing.

204. Aluminium carbide generates __________ gas on adding water to it.

205. The formula of corundum is _________.

206. In the alumino-thermite process Al acts as ________.

207. Solid carbon dioxide is known as _________.

208. The gas common to both producer gas and water gas is _________.

209. Diamond is an __________ of carbon.

210. Water glass constitutes of _______.

211. Inorganic polymer of silicon is called _________.

212. Hydrides of silicon are called _________.

213. The element of group . 14 which is a metalloid is _________.

214. The most stable form of carbon is __________.

215. Commercial name of silicon carbide is _________.

216. Lead pencil contains ________.

217. Variety of carbon that used as lubricant is ______ as abrassive is _______.

218. The most common form of crystalline silica & a common mineral _______.

219. 'C' undergoes —— hybridisation in graphite and ——— in diamond.

220. The allotrope used as moderator in atomic reactor is ________.

221. Boron undergoes ________ hybridisation in diborane.

222. Bauxite ore is concentrated by the process called _________.

223. The basicity of boric acid is __________.

224. PbI_4 is unstable as iodide gets _________ by ________ .

225. Aqueous solution of borax is ________ in nature.

UNIT - XII

226. There are large number of carbon compounds which is due to its unique property called _________.

227. Chemistry of carbon compounds is called _________.

228. ________ is first organic compound synthesized by ________.

229. The number of σ and π bonds in propenenitrile is _________ and _________ respectively.

230. Neopentane contains ________ no. of primary C-atoms.

231. IUPAC name of tert-butyl radical is _______.

232. Na-metal is fused with organic compound in ________ test.

233. KOH –tower. absorbs ________ gas.

234. Carius method is carried out to find out quantitative amount of _______ in organic compound.

235. Reagent, $\overset{+}{K}\overline{OH}$ acts as a _________ in organic reaction.

236. Prussian blue colour during detection of nitrogen is due to formation of ________.

237. $CH\equiv CH + Na \longrightarrow A \xrightarrow{2CH_3Br} B$. The compound, B is _________.

238. $A \longrightarrow B$ is a ________ type of organic reaction.

239. The hybridisations occurs in cycloalkane and benzene are ______ and ________ respectively.

240. The IUPAC name of  $-CH(CH_3)_2$ is ________.

241. The maximum number of C-atoms arranged linearly in $CH_3-C\equiv C-CH=CH_2$ is _________.

242. Orbital interaction between the sigma bonds of a substituent group and the neighbouring π-bond (orbital) is known as _______.

243. The consecutive members of a homogeneous series differ by ______ u.

244. Alkane undergoes ______ substitution reaction where as arenes undergo ______ substitution reaction.

245. Alkylation of benzene in presence of anhydrous $AlCl_3$ is named is ________ reaction.

246. In the nitrating agent HNO_3 acts as _______.

247. CH_3Br on homolytic cleavage forms ______ and ________.

248. The N–O bonds in nitro-methane are identical due to _________.

249. Polarisation of σ-bond is reffered to as __________ effect.

250. Shape of methyl cation is _________ and that of $\overset{\odot}{C}H_3$ is ________.

UNIT - XIII

251. Sodium salt of _______ acid on heating with soda-lime gives propane.

252. Methane to butane all are _________.

253. Mixture of higher alkanes is called _________.

254. Alkanes with odd number of C-atoms have ______ m.p. thane those with even number of C–atoms.

255. Halogenation of alkanes follows _________ substitution reaction.

256. Iodination of alkane is carried out in presence of __________.

257. Isobutane on oxidation with $KMnO_4$ gives ________.

258. _________ isomers are formed on rotation of C–C single bond.

259. Propanone on reduction with HI in presence of _________ gives _________.

260. Aluminium carbide on hydrolysis liberates — gas while calcium carbide liberates _________.

261. Dehalogenation vicinal dihalide gives __________.

262. Acidic dehydration ethanol provides ________.

263. Addition of HBr to alkene is governed by _________ rule.

264. Propene on treatment with hydrogen bromide gives ________.

265. HBr adds to propene in presence of benzoyl peroxide producing _________.

266. $(CH_3)_3CBr \xrightarrow{LiAlH_4} A,\ (CH_3)_2CHBr \xrightarrow{LiAlH_4} B$, A & B are respectively _______ and ______.

267. In dehydrohalogenation the halides are followed in the order ___________.

268. Dehydro bromination of 1-bromobutane gives ________ but 2-Bromo butane gives ________ as major product.

269. On acidic hydrolysis of isobutylene __________ is formed.

270. On boiling isopropyl hydrogen sulfate with water ________ is formed.

271. Ethene on oxidation with hot $KMnO_4$ solution gives _________.

272. Cyclobutene on reductive ozonolysis gives _________.

273. Orlon is formed on polymersisation of _________.

274. Ring chain isomer of propyne is ________.

275. Dehydrobromination of 1, 2-dibromo propane is _______.

276. On heating with Zn dust 1, 2– di bromopropane gives ________.

277.. Alkane burns at a moderate temperature due to its high __________ and the process is ______ in nature.

278. _______ hydrocarbon (alkane) easily gets oxidised.

279. Nitration of benzene is an example of ________ reaction.

280. Phenol on chlorination gives ________ white nitrobenzene on chlorination gives ________.

UNIT - XIV

281. The lowest region of atmosphere is called __________.

282. Above troposphere between 10 km to 50 km from the sea level the region of atmospheric is called _________.

283. Ozone layers belongs __________ region of atmosphere.

284. Ozone layer prevents Sun's harmful _________ radiations.

285. The major source of CO polluting the environment is ________.

286. The increased amount of ________ in air is mainly responsible for global warming.

287. Glass is opaque to _________ radiations.

288. _______ traps heat and is transparent to sunlight.

289. The Taj Mahal in India has been affected by ________.

290. smog is constituted by _______ and _______.

291. London smog is formed in winter during ________.

292. PAH and PAN are abbreviations for ________ and _______ respectively.

293. Exposure of U.V. rays may cause ________.

294. Normally rain water has pH of about ________.

295. The gas that can remove oxygen from oxyhaemoglobin is _________.

296. The other name of Los Angeles smog is _________.

297. Between CO and CO_2 ______ is more poisonous.

298. _______ gas is used during photosynthesis by plants.

299. _______ make the enamel on teeth much harder forming _________.

300. Tetrachloro ethene was used as solvent for _________ which is now replaced by suitable detergent with _________.

◆◆◆

(Group - A-2)

Fill in the blanks (Answers)

1. $18\times1.67\times10^{-24}\,g = 2.99\times10^{-23}\,g$

2. masses 0.1 mole of water = 1.8 gram = 1.8×10^{-3} Kg

3. percentage of N in urea

$$\left(NH_2CONH_2\right) = \frac{28}{60}\times100 = 46.6\%$$

4. $C + O_2 \longrightarrow CO_2$

1 mole, 32 gram, 44 g

2 mole 64 gram of oxygen (16 gram given)

$\frac{1}{2}$ mole, 16 gram, 22 gram, Here oxygen = limiting agent

mass of CO_2 = 22 gram.

5. C_2H_6, 1 mole contains 6 moles of H-atoms

3 moles of ethane contain 18 moles of hydrogen atoms

6. $$\text{Molarity} = \frac{W}{M}\times\frac{1000}{V} = \text{no. of moles}\times\frac{1000}{V}$$

$$\text{no. of moles} = \frac{\text{molarity}\times V}{1000}$$

$$= \frac{0.1\times200}{1000}$$

$$= \underline{0.02\text{ mole}}$$

7. $$\text{Equivalent mass} = \frac{\text{atomic mass}}{\text{Valency}}$$

$$\text{of Zn} = \frac{65}{2} = 32.5\text{ gram}$$

8. $$\left[2HNO_3 \longrightarrow H_2O + 2NO_2 + O\right]\times5$$

$$\underline{2P + 5(O) + 3H_2O \longrightarrow 2H_3PO_4}$$

$$\cancel{2}P + \overset{5}{\cancel{10}}HNO_3 \longrightarrow \cancel{2}H_3PO_4 + \overset{5}{\cancel{10}}NO_2 + \cancel{2}H_2O$$

9. $$+5\times1$$

$$P + 5H\overset{+5}{N}O_3 \longrightarrow H_3\overset{+5}{P}O_4 + 5\overset{+4}{N}O_2 + H_2O$$

$$-1\times5$$

10. $$\left[NO_3^- + 2H^+ + e^- \longrightarrow NO_2 + H_2O\right]\times5$$

$$\underline{P + 4H_2O \longrightarrow PO_4^{3-} + 8H^+ + 5e^-}$$

$$5NO_3^- + 2H^+ \longrightarrow PO_4^{3-} + 5NO_2 + H_2O$$

11. SO_2

12. $HClO_4$, O.N. of Cl = +7

13. $^0F = \frac{9}{5}\times25 + 32 = 77$

K or $^0A = 273.1 + 25 = 298.1$

14. 37^0C

15. candela

16. electrons

17. cathode, anode

18. cathode rays (discharge) tube

19. 9.1094×10^{-31} Kg

20. 1.6022×10^{19} C

21. neutrons

22. Electron

23. 10^{-10}m, 10^{-15} m

24. violet

25. $E_n = -\frac{13.6}{n^2}\,ev/atom$,

$$E_{3rd\,orbit} = \frac{-13.6}{9}$$

$$= \underline{1.51\text{ ev/atom}}$$

26. 3s

27. no. of spherical nodes = $n - \ell - 1$,

for 4s = 4 – 0 – 1 = 3

28. Goldstein, cathode

29. Three $\left(2s^2\,2p_x^1\,2p_y^1\,2p_z^1\right)$

30. 5d

31. atomic mass

32. Mg

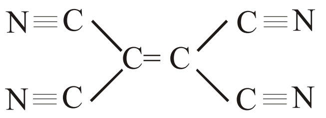
33. Triad (1) Fe (55.85) Co(58.93) Ni(58.71)

(2) Li (7.0) Na(23) K (39)

(3) Cl (35.5) Br (80) I (127)

(4) Ca (40) Sr (87.5) Ba (137)

34. 32 (Ge = Eka silicon)

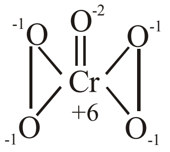
35. 14
36. 109 = Unnilennium
37. Copernicium, Cn
38. Six (B, Al, Ga, ln, Tl, Nh i.e. Nihonium)
39. 32
40. 3s and 3p
41. s-block
42. $3d^{1-10}\ 4^{0 \text{ or } 1}$
43. Sc (Z=21)
44. Br
45. Ta
46. Chalcogens
47. trans-uranium elements.
48. F (4.0)
49. Iron (Al. Second abundant)
50. tungsten (m.p. = 3422^{0}C)
51. pyramidal, tetrahedral
52. sp^3
53. Linear, sp
54. six
55. transference
56. dissimilar
57. Intra-molecular
58. two
59. $109^0\ 28^{/}$
60. paramagnetic
61. 120^0
62. covalent an co-ordinate
63.  , nine σ bonds nine π bonds
64. d, p $\left(d_\pi - d_\pi,\ p_\pi - p_\pi,\ d_\pi - p_\pi\right)$
65. s, π, σ
66. higher
67. decreases, lowers
68. tetrahedral
69. sp^3 hybrid
70. A_2B_3
71. ten (N, O, F, Cl, He, Ne, Ar, Kr, Xe, Rn)
72. $p \propto d$ at a given temperature.
73. 22.4 L
74. 31.1^{0}C
75. Z = 1
76. T_B (Boyle temperature)
77. rise
78. greater
79. laminar
80. N.s. m^{-2}
81. effusion
82. density of the gas
83. elastic
84. $V_t = v_0\left(1+\frac{t}{273}\right) = 20\left(1+\frac{10}{273}\right)L$
85. R/N_A
86. ΔU
87. ΔH
88. exothermic
89. greater
90. randomness / disorderedness
91. negative
92. extensive
93. 0 K
94. absolute zero
95. negative
96. positive
97. increases
98. Zero
99. negative
100. J K^{-1} mol^{-1}
101. spontaneous
102. spontaneous
103. reversible
104. zero
105. negative sign
106. reaction / combustion
107. sublimation
108. less
109. heat of combustion
110. R
111. reversible
112. unchanged

113. vapour pressure
114. closed
115. dynamic
116. $N_2 + 3H_2 \longrightarrow 2NH_3$

$$\frac{\frac{1}{2}d\left[NH_3\right]}{dt} = -\frac{1}{3}\frac{d\left[H_2\right]}{dt}$$

$$\frac{d\left[NH_3\right]}{dt} = -\left(\frac{2}{3}\right)\frac{d\left[H_2\right]}{dt}$$

117. p_{CO_2}
118. backward
119. finely divided iron
120. 10^{-2} M
121. weaker
122. 14
123. $-\log K_b$
124. $F^- < OH^- < \bar{N}H_2 < \bar{C}H_3$
125. acidic, alkaline
126. oxidation
127. decreases
128. increases $\left(Na - \bar{e} \longrightarrow \overset{+}{Na}\right)$
129. increases, decreases
130. redox
131. Redox
132. Oxidation $\left(\frac{1}{2}\right)$, reduction $\left(\frac{1}{2}\right)$
133. disproportionation
134. +6,



135. F
136. a reducing agent
137. $-\frac{1}{2}\left(\overset{+1}{K}O_2^{-1}\right)$
138. zero
139. Oxidation, hydrogen
140. disproportion

$$\overset{0}{P} + NaOH + H_2O \rightarrow \overset{-3}{P}H_3 + NaH_2\overset{+1}{P}O_2$$

141. $\frac{1}{6}$
142. $\frac{1}{2}$, Eq. mass of acid $= \frac{\text{mol. mass}}{\text{No. proton replaced}}$
143. 12
144. 50 ml
145. Zero
146. hydrogen
147. hydrogen,

$$CO(g) + H_2O(g) \xrightarrow[\text{Catalyst}]{673K} CO_{2(g)} + H_2(g)$$

148. $C_2\ D_2$
149. 10 L, di oxygen
150. One (other four by dative bonds)
151. Four
152. hydrolysis
153. 104.5^0
154. oceans (97.33%)
155. –1
156. hydrogen
157. silver nitrate
158. dehydrating agent
159. Mg or Mn
160. hydrogen, Lavoisier
161. two $\left({}^3_1H\right)$
162. isomers (nuclear)
163. reducing
164. exothermic
165. β–rays, ${}^3_2He$

$$\left({}^3_1H \rightarrow {}^0_{-1}e + {}^3_2He\right)$$

166. $ns^{1(or)2}\left(n = 2 \text{ to } 7\right)$

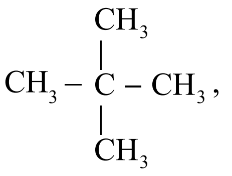
$$H = 1s^1$$

167. eight, hydrogen, one
168. Ca
169. isoelectronic, 10

170. HCl
171. $CaCO_3$ / $CaSO_3$
172. $CaHCO_3$ / $CaHSO_3$
173. Al
174. amphoteric
175. Ammonia
176. $MgSO_4.7H_2O$, $CaSO_4.\ 2H_2O$
177. magnesium
178. teeth
179. magnesium
180. calcium oxide
181. hydrolith
182. Solvay
183. Paramagnetic (n=1)
184. potassium hydroxide
185. Caesium (Cs)
186. Cs (302 K)
187. superoxide (K, Rb, Cs)
188. LiCl - $2H_2O$
189. violet
190. soda (NaOH) and lime (CaO)
191. silicon
192. $B_2\ O_3$
193. acid
194. diborane, $B_2\ H_6$
195. borazine
196. B_2O_3, $Na_2B_4O_7 \xrightarrow{\Delta} 2NaBO_2 + B_2O_3$
197. brown, $Ni\ (BO_2)_2$ nickel metaborate
198. boric acid (H_3BO_3)
199. $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$
200. amphoteric
201. BCl_3
202. conc. HNO_3
203. mordant
204. methane
205. $Al_2\ O_3$
206. reducing agent
207. dry ice
208. carbon monoxide
209. allotrope
210. sodium silicate
$Na_{2x}\ Si_y\ O_{2y+x}$ e.g. $Na_2\ SiO_3$
211. silicones
212. silanes
213. Germanium
214. Graphite
215. Carborundum
216. graphite
217. graphite, diamond
218. quartz
219. sp^2, sp^3
220. graphite
221. sp^3
222. Leaching
223. one
224. oxidised by Pb^{+4}

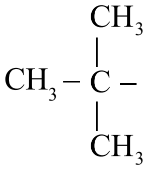
$$PbI_4 \longrightarrow PbI_2 + I_2$$

225. alkaline (salt of strong base and weak acid)
226. catenation
227. organic
228. Urea, Wohler (1828)
229. $CH_2 = CH - C \equiv N$, $6 - \sigma$ bonds
$3 - \pi$ bonds

230.


$$CH_3 - \overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}} - CH_3$$

, (4) - 1^0 -C atoms.

231.


$$CH_3 - \overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}} -$$

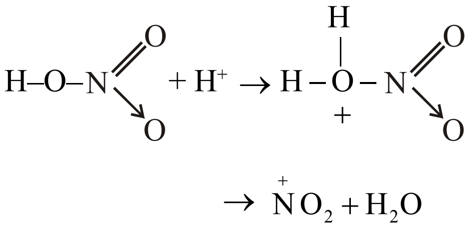
dimethyl ethyl –

232. Lassaigne (to detect N, S and X in organic compound)
233. CO_2
234. Halogen
235. nucleophile / alkali
236. Ferric ferrocyanide
237. A = $NaC \equiv CNa$

B = $CH_3 - C \equiv C - CH_3$

238. S_N (Nulceophilic substitution)

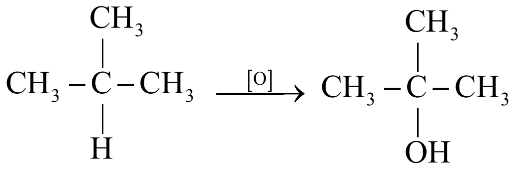
239. sp^3, sp^2
240. Methyethyl cyclopentane
241. Four (sp^2, sp - hybridised C)
242. hyperconjugation
243. 14
244. free radical , electrophilic
245. Friedel - Crafts
246. Lewis base



247. $CH_3^{\bullet}$, $\overset{\bullet}{Br}$
248. resonance
249. inductive
250. planar, pyramidal
251. butanoic acid

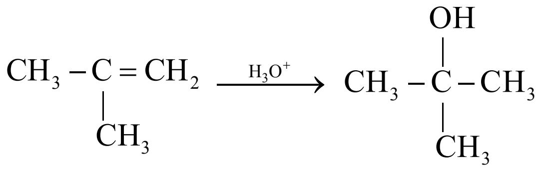
$$CH_3CH_2-CH_2COONa \xrightarrow[\Delta]{Sodalime} CH_3-CH_2-CH_3$$

252. gases
253. grease
254. lower
255. free radical
256. HNO_3 or HIO_3
257.



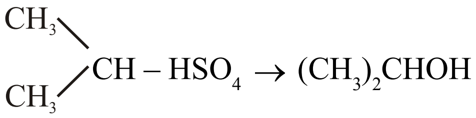
$$CH_3-\overset{CH_3}{\overset{|}{\underset{\underset{H}{|}}{C}}}-CH_3 \xrightarrow{[O]} CH_3-\overset{CH_3}{\overset{|}{\underset{\underset{OH}{|}}{C}}}-CH_3$$

258. Conformational
259. red P, propane
260. methane, acetylene
261. alkene
262. ethene
263. Markownikoff
264. 2-bromopropane
265. 1-bromopropane
266. $(CH_3)_2C = CH_2$, $CH_3-CH_2-CH_3$
267. RI > RBr > RCl
268. but-1-ene, but-2-ene
269.



$$CH_3-\underset{\underset{CH_3}{|}}{C}=CH_2 \xrightarrow{H_3O^+} CH_3-\overset{OH}{\overset{|}{\underset{\underset{CH_3}{|}}{C}}}-CH_3$$

270.



271. HCOOH
272. butanedial
273. $CF_2 = CF_2$ (tetra fluoroethene
274. Cyclopropene, 
275. $CH_3CHBr-CH_2Br$

$$\xrightarrow[-HBr]{KOH\ (alc)} CH_3-C\equiv CH$$

276. propene
277. activation energy, exothermic
278. tertiary
279. S_E
280. p-chlorophenol, m-chloronitrobenzene
281. troposphere
282. stratosphere
283. stratosphere
284. U.V.
285. automobile exhaust
286. CO_2
287. IR
288. CO_2 molecules
289. acid rain
290. smoke, fog
291. morning hours
292. polycyclic aromatic hydrocarbon, peroxy acetyl nitrate ($CH_3CO-O-O-NO_2$)
293. skin cancer
294. 5.6
295. CO or NO
296. Photochemical smog
297. CO
298. CO_2
299. Fluoride ion, 3 $Ca_3(PO_4)_2.CaF_2$
300. dry cleaning, $CO_2(l)$

◆◆◆

Group - A (3)

(Very short questions - one mark each)

UNIT - I

1. What is the masses of
   - (i) 0.02 mole of nitrogen gas.
   - (ii) 0.1 gram equivalent of sodium carbonate.
   - (iii) 6.023×10^{24} atoms of carbon.
   - (iv) 6.023×10^{22} molecules of nitrous oxide.
   - (v) Carbon in 0.1 mole of carbon dioxide.
2. Convert atomic mass unit to CGS unit.
3. Calculate the equivalent mass of potassium chlorate if it acts as an oxidant.
4. Calculate the molecular mass of a gas, 2.8 L of which weighs 4 grams at NTP.
5. Name the gases on heating copper turnings with
   (i) conc. H_2SO_4 (ii) Conc. HNO_3 (iii) dil HNO_3
6. Express the empirical and molecular formula of glucose.
7. Calculate the no. in grams of sodium present in 5.85 grams of NaCl.
8. Write down the formula of two oxides of iron.
9. Calculate the molecular mass of a trivalent metal chloride, if the equivalent mass of metal is 32.7.
10. Find out the number of water molecules in 1 dm^3 of water.

UNIT - II

11. Write down the electronic configuration of atom with Z = 24 and 29.
12. Find out the number of umpair e^-(s) in Fe^{2+}, Fe^{3+} ions.
13. Write quantum numbers of valence electron of K.
14. Draw the p-orbitals.
15. Arrange the orbital 4s, 4p, 3d, 3p in order of increasing energy.
16. Express de Broglie wave length of material wave.
17. The number of umpair electrons in dioxygen is 2. T/F.
18. Calculate the number of protons, neutrons and electrons is Li^+.
19. What are the $(n+l)$ values of 4f and 5d orbitals.
20. State Pauli's exclusion principle.

UNIT - III

21. Name the element with maximum $\Delta_{eg}H$ value.
22. Which element has highest density ?
23. Calculate number of d-block elements in the periodic table.
24. Assign the periodic position of element with Z = 118. Name it.
25. Choose the elements with highest and lowest $\Delta_i H$ values.
26. Arrange the elements of 2nd period in order of increasing $\Delta_i H$ values.

27. From among the elements choose the following :

Cl, Br, F, O, Al, C, Li, Cs and Xe

(a) element with greater electron effinity.

(b) the most electropositive element

(c) Identified as aerogen.

(d) the lightest metal

(e) the most reactive non-metal

(f) forming amphoteric oxide.

(g) forming diagonal relation with Mg.

(h) in liquid state under ordinary conditions.

(i) Forming largest number of compounds.

(j) Most abundant element on earth.

28. The electronic configuration of some of the elements A to E are given below.

(A) $1s^2\ 2s^2\ 2p^6\ 3s^2$ (B) $1s^2\ 2s^2\ 2p^6$ (c) $1s^2\ 2s^2\ 2p^3$

(D) $1s^2\ 2s^2\ 2p^6\ 3s^1$ (E) $1s^2\ 2s^2\ 2p^5$

Answer the following

(i) Name the elements

(ii) Which has low $\Delta_i H$ value

(iii) What is a halogen ?

(iv) Which is an alkaline earth metal ?

(v) What is a noble gas ?

UNIT - IV

29. How are ions formed ?

30. Define lattice energy ?

31. Show that $\overset{+}{N}H_4$ and CH_4 are of same geometry.

32. How many electrons are shared in N_2, O_2 molecules ?

33. Predict the hybridisation of central atom of the following :

PCl_5, SF_6, NH_3, H_2O, $BeCl_2$, BCl_3, SO_2, SO_3, SO_4^{2-}, CO_3^{2-}

34. Arrange the following in order of increasing bond angle :

BF_3, CF_4, CO_2, NH_3, H_2O, PH_3

35. Write the resonance structure of N_3^- ion. (azide)

36. Arrange :

(i) N_2, F_2 O_2, O_2^-, O_2^+ : is order of increasing bond order.

(ii) NO, $N\overset{+}{O}$, $N\overset{-}{O}$ L in increasing order of bond length

(iii) MgO, CaO, SrO, BaO : increasing order of lattice energy.

37. Is $SnCl_2$ a linear molecule ?

38. Write the Lithium halides in order of increasing covalent character.

39. Arrange the halides of Ag in order of increasing solubility in water.

40. The formula of a compound A_2B_5. Find out the number of $\overline{e}$ in the outer most orbit of A and B.
41. Identify an ion isoelectronic with carbon monoxide.
42. Draw the bonding pattern is nitrogen pentoxide.
43. Predict the nature of σ – bonds is ethane molecule.
44. Draw the resonating structures of CO_3^{2-} ion.
45. Show that PCl_5 is an unsymmetrical molecule
46. Identify molecule with (i) two electrons three centre bond
(ii) three electron two centre bond.
47. What is the hybridisation of Xe is XeF_2 / XeF_4 ?
48. How stability is related to bond order ?
49. Which orbitals generally involved in π -bond formation ?
50. Trans-but-2-ene is non-polar. (T/F).

UNIT - V

51. Express the Charle's law / combined gas law in terms of density.
52. Write down vander Waals gas equation for n moles of real gas.
53. Mention the units of 'a' and 'b' the vander Waals constants.
54. Draw the v-t and V-T plots as per Charles' law.
55. What do you mean by rms velocity ?
56. What is most probable velocity ?
57. Express Kinetic gas equation.
58. Define Dalton's law of partial pressure.
59. What is partial pressure ? How is it related to total pressure ?
60. Compare the diffusitivity of N_2, O_2, CO_2, He.
61. What is absolute zero temperature ?
62. What is the significance of 'R' ?
63. How kinetic energy is related to temperature ?
64. Under what conditions ideal gases behave ideal ?
65. Arrange the gas is order of ideality :
H_2, He, O_2, CO_2
66. Provide the relationship between probable, average and rms velocities.
67. What is the value of 'R' is Cal. unit ?
68. T_c of CO_2 is 31.1^0C and that of CH_4 – 81.9^0C. Which can have stronger intermolecular forces ?
69. Express 1 poise in SI unit.
70. Arrange the following liquids in order of increasing viscosity :
water, methanol, ethylene glycol, glycerol.

UNIT - VI

71. What is the value of Q for an isothermal process.
72. What is the work done for an aliabatic process ?
73. Can work be performed by the system without consuming heat energy ?
74. How Q_p and Q_V are related ?

75. Define internal energy ?

76. What is isochonic process ?

77. Mention (e) for extensive property and (i) for intensive property of the following :
mass, viscosity, volume, enthalpy, entropy, temperature, pressure, density, apecific heat, specific gravity, surface tension, ΔU.

78. Write the equation relating free energy, entropy and enthalpy.

79. How is ΔG related to equilibrium constant, K ?

80. Calculate the entropy vaporisation for liquid boiling at 300 K having enthalpy of vaporisation 30 kJ mol^{-1}.

81. If $N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$, $\Delta_r H^0 = -92.4\,kJ/mol$ what is standard enthalpy of formation of $NH_{3(g)}$?

82. $2Cl_{(g)} \rightarrow Cl_{2(g)}$, What is the sign of ΔH and ΔS.

83. Under what condition, the heat evolved or absorbed is equal to internal energy change ?

84. Under what condition an extansive prosperty can become intensite ?

85. We prefer ΔH to ΔU.

86. What are the conditions for expressing standard enthalpy change $\left(\Delta H^0\right)$?

87. What is the value of enthalpy of nutralisation of a strong acid by strong base ?

88. Why enthalpy of nutralisation of HF is greater than 57.1 kJ / mol ?

89. What is an isolated system ?

90. What is mechanical work ?

UNIT - VII

91. Predict the effect of
(i) addition of H_2 (ii) addition of CH_3OH
(iii) renoval of CO (iv) removal of CH_3OH
On the equilibrium of the reaction.

$$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$$

92. Find a relationship between K_p and K_C for the equilibrium

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

93. Which of the followings are Lewis acids ?

H_2O, BF_3, H^+ and NH_4^+

94. Find the conjugate acid / base of

HNO_2, CN^-, $HClO_4$, F^-, OH^-, CO_3^{2-} and S^{2-}

95. Calculate the pH of
(i) 0.005 M NaOH (ii) 0.002 M HBr

96. When does precipitation occur ?

97. What happens to the pH of the solution (NH_4OH) in presence of NH_4Cl (aq.) ?

98. Provide one example in each case of acidic and alkaline buffers.

99. How can you change the pH of solution from 3.0 to 6.0 ?

100. Describe the effect of acid and base on the $[S^{2-}]$ from H_2S (g).

101. What is acid / base according to Bronsted theory ?

102. What happens to solubility of calcium acetate with rise of temperture ?

103. What happens to solubility of AgCl(s) is pressence KCl (aq) ?

104. Arrange the following solution in water increasing pH.

$CaCl_2$, Na OAc, KCl

105. What is the group reagent required to precipitate group-I radicals ?

106. Define solubility product.

107. Compare the K_{sp} and solubility (s) of a sparingly soluble salt A_xB_y.

108. What is the simple buffer solution ?

109. Calculate the pH of a solution containing 0.98 grams of H_2SO_4 / litre of the solution.

110. Express is terms of Kp and Kx for the general reaction in equilibrium :

$$aA_{(g)} + bB_{(g)} \rightleftharpoons xX_{(g)} + yY_{(g)}$$

UNIT - VIII

111. Define oxidation and reduction is terms of electron with one example is each case.

112. $MnO_4^- + H^+ + C_2O_4^{2-} \longrightarrow$, Complete and balance.

113. $H_2S + HNO_3 \longrightarrow H_2O + NO + S$. Balance by oxidation number of method.

114. What are oxidation number of 'N' in hydrazoic acid & hydrazine ?

115. Identify reaction where

i) C. can exhibit o.s. –4 to +4

ii) N from –3 to +5

116. Justify the correctness of the following balanced equation.

$$6CO_2(g) + 6H_2O(\ell) \longrightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

117. $2AgBr + C_6H_6O_2(aq) \longrightarrow 2Ag(s) + 2HBr(aq) + C_6H_4O_2$

Identify the substance oxidised / reduced.

118. Arrange the following acids in order of increasing reducing agent :

HF, HCl, HBr and HI.

119. Write down formula of mercury (II) chloride, chromium (III) oxide.

120. $H_2O_{(s)} + F_2(g) \longrightarrow HF(g) + HOF(g)$

Show that it's a redox reaction.

121. Show that hydrogen acts as both oxidant and reductant.

122. Oxygen behaves as oxidant and reductant – explain.

123. Find out the oxidation number of Cl in HCl, HClO, ClO_4^-, ClO_3^-, ClO_2^-, Cl_2O .

124. What is mass of oxalic acid needed to decolorise 1 gram eqvt. of $KMnO_4$?

125. The equivalent mass of H_2O_2 is acidic and alkaline medium are identical — verify.

UNIT - IX

126. The positron of 'H' in periodic table is controversial. T/F
127. Name an element in each case that liberated hydrogen form (i) water (ii) alkali (iii) acid.
128. How is the impurity H_2S removed from 'H_2' gas prepared in the Lab ?
129. Which form of hydrogen is used for welding purposes ?
130. What is the use of Calgon ?
131. Name the linkage present in H_2SO_5 and $H_2S_2O_8$.
132. Why is H_2O_2 used as bleaching agent ?
133. What is Fenton's reagent ? Write its action with lactic acid.
134. How hydrogen peroxide acts as an antiseptic ?
135. How can you get water from hydrogen ?
136. Why D_2O has higher boiling point than H_2O ?
137. H_2O_2 acts as both oxidant and reductant — explain.
138. Name the chemical compound used in the preparation of H_2O_2 by auto-oxidation.
139. Draw the structure of hydrogen peroxide.
140. How can you distinguish H_2O from H_2O_2 ?

UNIT - X

141. Sodium sulfate. decahydrate is known as Glauber salt. T/F.
142. Arrange the alkali metal ions (M^+) in order of decreasing hydration enthalpy.
143. Identify the alkali metals easily forming superoxide.
144. Arrange the alkali metal halides in order of decreasing melting and b.p.
145. How can you get $NH_3(g)$ from NH_4Cl in the laboratory ?
146. Discuss the action of heat on washing soda.
147. What is common method to isolate NaCl from sea water ?
148. Compare the function of heat of $LiNO_3$ and $NaNO_3$.
149. Name the s-block elements failing to form solid bicarbonate.
150. What is dead burnt plaster ?
151. How is milk of lime formed ?
152. What happens to the solubility of group-2 metal sulfates and metal carbonates ?
153. Discuss the action of heat on hydroxides and carbonates of magnesium.
154. How can you prepare HCl (g) from common salt ?
155. Can you prepare 'K' from KCl ?
156. Arrange the group-2 metals in order of increasing reactivity with water / O_2 .
157. $BeO \xrightarrow[2175K]{C} A_{(s)} \xrightarrow{Water} B_{(g)}$

     Name A and B.
158. What is slaking ?
159. What are the conditions for making the salt soluble in water ?
160. Draw the structure $BeCl_2$ in solid form.

161. Arrange the group-2 metal sulphates in order of increasing stability.

162. Name the alkaline earth metal failing to form colour flame.

163. What are the compounds formed when chlorine is passed through Na(OH) cold/hot ?

164. Freshly cut sodium tarnishes on exposure – explain.

165. Write two uses of quick lime.

UNIT - XI

166. What is the general formula of alums ?

167. Arrange the boron halides in order of increasing acidic nature.

168. B–F bond lengths in BF_3 and BF_4^- are different, Why ?

169. How is sugar charcoal formed ?

170. How does Al metal react with methyl alcohol ?

171. What happens when CO is heated with ZnO ?

172. CO_2 is an acidic anhydride –explain.

173. Name the acid that affects glass.

174. What does it yield when trialkyl chloro silane, R_3SiCl is hydrolysed ?

175. Name the product formed on oxidation of graphite by HNO_3.

176. Write down the elements group-14 with atomic number.

177. Arrange in order of decreasing $\Delta_i H$ values of Si, Ge, Sn and Pb.

178. Why is carbon tetravalent ?

179. $AlCl_3$ is a Lewis acid – give reason.

180. Write two important ores of aluminium.

181. Draw the structure of diborane.

182. What are O.N. carbon in different carbides ?

UNIT - XII

183. Predict the hydridisation of carbon that it undergoes in ethane, ethene and ethyne.

184. What are homocyclic compounds ?

185. Name two heterocyclic compounds.

186. What are heteroatoms ?

187. Draw (i) the bond line formula of 3-bromo-but-1- ene.
(ii) Wedge - desh representation methane.

188. Arrange the following groups in order of sequence and name the suffixes and prefixes.
SO_3H, $-COOH$, $-CHO$, $-CN$

189. Name the compound CH_3SO_3H, $C_6H_5\ CONH_2$.

190. Arrange the following carbocations in order of increasing stability.

$$CH_2=CH-\overset{+}{C}H_2,\quad C_6H_5-\overset{+}{C}H_2,\quad CH_3-\overset{+}{C}H_2,\quad CH_3-CO-\overset{+}{C}H_2$$

191. What do you mean by nucleophile ? Give one example.

192. Name the isomers of the compound with formula $C_4H_{10}O$.

193. The compound,  is aromatic. True/False.

194. Draw the structure of anti-butane.

195. Arrange the following in order of decreasing stability :

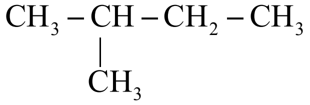
$$CH_3-\overset{+}{C}HCH_3,\ CH_3-\overset{+}{C}H-OCH_3,\ CH_3-\overset{+}{C}H\ COCH_3$$

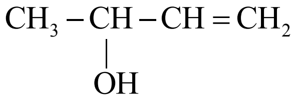
196. What is nitrating agent ?

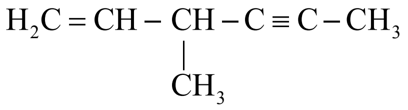
197. Write down the resonating structures of

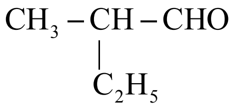
(i) C_6H_5-CHO (ii) C_6H_5 OH

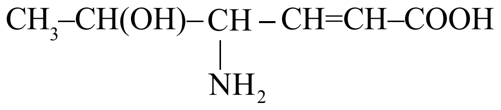
198. Provide the IUPAC names of

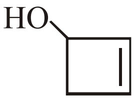
(i)

$CH_3-\underset{CH_3}{\underset{|}{CH}}-CH_2-CH_3$

(ii)

$CH_3-\underset{OH}{\underset{|}{CH}}-CH=CH_2$

(iii)

$H_2C=CH-\underset{CH_3}{\underset{|}{CH}}-C\equiv C-CH_3$

(iv)

$CH_3-\underset{C_2H_5}{\underset{|}{CH}}-CHO$

(v)

$CH_3-CH(OH)-\underset{NH_2}{\underset{|}{CH}}-CH=CH-COOH$

(vi)


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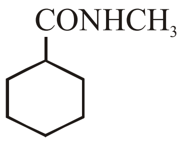
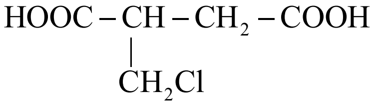
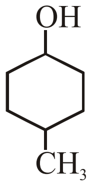
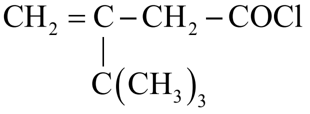


(viii)

$HOOC-\underset{CH_2Cl}{\underset{|}{CH}}-CH_2-COOH$

(ix)


(x) $CH_2=CH-COOC_2H_5$

(xi)

$CH_2=\underset{C(CH_3)_3}{\underset{|}{C}}-CH_2-COCl$

199. Write down the isomers of the following and provide IUPAC names.

(a) C_5H_{12} (b) C_5H_{10} (c) C_5H_8 (d) C_4H_8O

(e) $C_4\ H_6O$ (f) $C_5H_{10}O_2$ (g) $C_5H_{13}N$ (h) $C_6H_4Cl_2$

200. Name the Lewis acids used during ethylation of benzene.

UNIT-XIII

201. What is Sabatier - Senderen's reaction ?

202. Name the product formed when 2-chloropropane undergoes Wurtz reaction.

203. What are the gases formed on electrolysis of CH_3COONa (Aq.) ?

204. What is the rate of reaction of alkanes with halogens ?

205. Draw the Sawhorse projection of ethane.

206. Figure the Newman's projection of antibutane.

207. What is torsional strain ?

208. How can you get benzene from n-hexane ?

209. Provide the industrial synthesis of di hydrogen from methane.

210. What is marsh gas ?

211. Discuss the isomerism shown by alkene.

212. How does cis-isomer differ from trans-isomer ?

213. Identify the possible alkanes formed when a mixture of CH_3Br and C_2H_5Br undergoes Wurtz reaction.

214. How can you prepare isobutane from n-butane ?

215. What is β -elemination reaction ?

216. Zn –HCl reduces ethylbromide to ethane –explain.

217. Soda–lime is used to prepare methane from sodium acetate – give reason.

218. Sodium acetate on electrolysis forms ethane at anode – why ?

219. Cyclobutane is less stable than cyclopropane is it so ?

220. Write the structure of (Z)-but-2-ene, (E)-but-2-ene.

221. Identify E and Z.

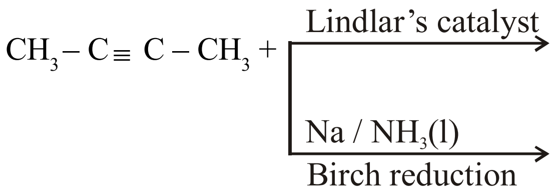
222. Arrange the following alkene is order of their increasing heat of hydrogenation :

trans-but-zene, cis-but-2-ele, but-1-ene, ethene

223. Name the product when Ethanol is heated with Al_2O_3.

224. What is the order of dehydration of different alcohols ?

225. Predict the product (s)



226. What is Lindlar's catalyst ?

227. Provide a test for unsaturation ?

228. What happens when ethene is treated with cold dil $KMnO_4$?

229. Name the alkene which on reductive ozonolysis gives
    - (i) acetone and acetaldehyde
    - (ii) acetaldehyde only
    - (iii) acetone only
    - (iv) acetone and formaldehyde
    - (v) acetaldehyde and formaldehyde
    - (vi) acetaldehyde, glyoxal and acetone
    - (vii) acetaldehyde, carbondioxide, acetone
    - (viii) Propanedial
    - (ix) 2 moles of ethanedial
    - (x) glyoxal + 2-oxopropanal

230. Name the potassium salt to prepare ethyne and propyne.

231. How can you prepare but-1-yne from acetylene ?

232. Provide a method to prepare

   (i) acetaldehyde from acetylene

   (ii) acetone form propyne

233. Write down the formula of westron and westrosol.

234. How can you prepare acetone from acetylene ?

235. Why alkanes are almost non-reactive towards different reagents ?

236. Cracked gasoline is preferred - why ?

237. Arrange the alkynes is order of increasing acidic nature : ethyne, propyne, but-2-yne, but-1-yne

238. What happen when propyne gets heated is a red hot iron tube ?

239. Name the salts to prepare toluene by decarboxylation method.

240. Distinguish between hexachloro benzene and benzene hexa chloride (BHC).

UNIT - XIV

241. Identify the major pollutants present in troposphere.

242. Name three major gaseous pollutants ?

243. Discuss the damages occur due to increase in concentration of NO_2.

244. What are harming effects of hydrocarbon in the atmosphere ?

245. What is Taj Trapezium ?

246. What is photo-chemical smog ?

247. Name the main compounds damaging the ozone layer.

248. What are acids generally present in acid rain ?

249. Give an example of non-biodegredable waste.

250. Unpolluted rain water coutained which acid ?

251. What is BOD ?

252. What is COD ?

253. Why photo chemical smog is so called ?

254. What is the nature of photochemical smog ?

255. Name two natural sources of air pollution.

256. Provide two examples herbicides.

257. What are polar stratopheric clouds ?

258. Identify certain (toxic) organic solvents.

259. What is chlorosis ?

260. Name the gas responsible for Bhopal tragedy.

◆◆◆

(Group - A-3)

Very short questions - One mark each (Answers)

1. (i) $28 \times 0.02 = 0.56\,\text{g}$

   (ii) $53 \times 0.1 = 5.3\,\text{g}$

   (iii) $10 \times 12 = 120\,\text{g}$

   (iv) $0.1 \times 44 = 4.4\,\text{g}$

   (v) 0.1 mole CO_2 contains 0.1 mole of C = 1.2 g

2. $1\text{u (amu)} = \frac{1}{12} \times$ mass of an atom of C-12

$$= \frac{1}{12} \times \frac{12}{6.023 \times 10^{23}} \text{ gram}$$

$$= 1.66 \times 10^{-24}\,\text{g} = 1.66 \times 10^{-27}\,\text{kg}$$

3. $KClO_3 \longrightarrow KCl + 3[O]$

122.5 parts 48 parts

$\frac{122.5}{6}$ 8 parts

Eq. mass = 20.42 gram/eqvt.

4. 2.8 L weighs 4 g at NTP

22.4 L weighs $\frac{4}{2.8} \times 22.4 = 32\,\text{g/mole}$

5. • $Cu + 2H_2SO_4 \longrightarrow CuSO_4 + 2H_2O + SO_2(g)$

• $2H^+ + NO_3^- + e^- \longrightarrow NO_2(g) + H_2O] \times 2$

$$Cu \longrightarrow Cu^{2+} + 2e^-$$

$$2NO_3^- + 4H^+ + Cu \longrightarrow 2NO_2(g) + Cu^{2+} + 2H_2O$$

• $NO_3^- + 4H^+ + 3e^- \longrightarrow NO + 2H_2O] \times 2$

$$Cu \longrightarrow Cu^{2+} + 2e^-] \times 3$$

$$2NO_3^- + 8H^+ + 3Cu \longrightarrow 3Cu^{2+} + 2NO(g) + 4H_2O$$

6. Empirical formula = CH_2O, molecular formula = $C_6H_{12}O_6$
7. Mole mass of NaCl = 58.5 = 1 mole

   5.85 g NaCl contains 2.3 g of Na.
8. Ferrous oxide, FeO; ferric oxide = Fe_2O_3
9. Formula of the metal chloride = MCl_3

   Molecular mass = $(3 \times 32.7) + (3 \times 35.5) = 204.6\,\text{g/mol.}$

10. 1 dm^3 of water = 1000 gram = $\frac{1000}{18}$ mole = 55.56 mole

No. of molecules = $55.56 \times 6.023 \times 10^{23}$

$= 334.6 \times 10^{23} = 3.346 \times 10^{25}$

11. Z = 24 (Cr) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$
= 29 (Cu) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

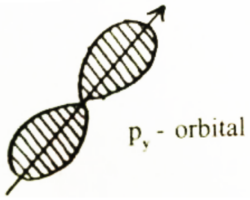
12. $_{26}Fe = 3d^6\ 4s^2$ Fe^{2+} $3d^6$ no. of unpair electrons = 4
Fe^{3+} $3d^5$ no. of unpair electrons = 5

13. $_{19}K = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

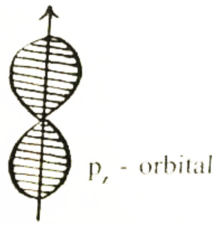
$n = 4,\ \ell = 0,\ m = 0\ \ s = +\frac{1}{2}$

14.

p_x - orbital



p_y - orbital



p_z - orbital

dumb-bell shape along X axis = p_x
along Y - axis = p_y
along Z-axis = p_z

15. $n + \ell$ values 4s, 4p, 3d, 3p
4, 5, 5, 4
3p < 4s < 3d < 4p

16. $\lambda = h/p$

17. T

18. proton = 3, neutron = 4, electron = 2

19. $(n + \ell)$ of 4f = 7, 5d = 7 (5+2)

20. No two electrons of an atom can have four quantum number values identical.

21. Chlorine

22. Osmium (/ Iridium)

23. Forty

24. Group - 18, period-7,
Name Ununoctium, Og (Oganesson)

25. Highest $\Delta_i H = He$, Lowest $\Delta_i H = C_S$

26. Li < B < Be < C < O < N < F < Ne

27. (a) Cl (b) Cs (c) Xe (d) Li (e) F (f) Al
(g) Li (h) Br (i) C (j) O

28. (i) A = Mg B = Ne C = N D = Na E = F
(ii) D (Na) (iii) E(F) (iv) D (Na) (v) B (Ne)

29. Ions are formed on gain of electron $A + e^- \longrightarrow A^-$ (anion)

i.e., by reduction and on loss of e^-, $M - e^- \rightarrow M^+$ (cation) i.e. oxidation.

30. It is the energy needed to break 1 mole of ionic compounds to constituting ions is gaseous state.

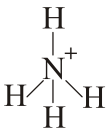
$$\overset{+}{N}aCl^{-}(s) + L.E \longrightarrow Na^{+}(g) + Cl^{-}(g)$$

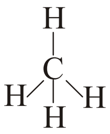
Lattice enthalpy is positive.

In formation of one mole compound same amount of energy is released.

$$\overset{+}{N}a(g) + Cl^{-}(g) \longrightarrow Na^{+}(s) + L.$$ Energy (lattice energy is exothermic.)

31. $\overset{+}{N}H_4$ and CH_4 both are isoelectronic $(10\,e^{-})$ and tetrahedral





32. $:N\vdots\vdots N:$ (sharing of $6e^{-}$), $.\ddot{\underset{..}{O}}{::}\ddot{\underset{..}{O}}$ (sharing of two pairs of e^{-})

33. PCl_5 (sp^3d), SF_6 (sp^3d^2), NH_3 (sp^3), H_2O (sp^3)
$BeCl_2$ (sp), BCl_3 (sp^2), SO_2 (sp^2), SO_3 (sp^2)
$SO_4^{2-}(sp^3)$, $CO_3^{2-}(sp^2)$

34. BF_3 (120^0), CF_4 ($109\ 28'$) CO_2 (180^0), NH_3 (107^0)
H_2O ($104^0\ 28'$), PH_3 (96.6^0)
$PH_3 < H_2O < NH_3 < CF_4 < BF_3 < CO_2$

35. N_3^{-} (azide ion) is iso-electonic with CO_2 so isostructural.

$$\overset{2-}{N}-\overset{+}{N}\equiv N \longleftrightarrow \overset{-}{N}=\overset{+}{N}=\bar{N} \longleftrightarrow N\equiv\overset{+}{N}-N^{2-}$$

36. (i) $N_2(14e^{-})$, $F_2(18\ e^{-})$, $O_2(16\ e^{-})$, $O_2^{-}(17\ e^{-})$, $O_2^{+}(15\ e^{-})$

B.O. 3 1 2 1.5 2.5

Order sequence : $F_2 < O_2^{-} < O_2 < O_2^{+} < N_2$

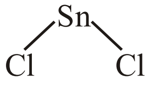
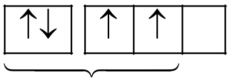
(ii) NO, NO^{+}, NO^{-}

$15e^{-}$ $14\ e^{-}$ $16\ e^{-}$

B.O. 2.5, 3, 2 Bond length order : $NO^{+} < NO < NO^{-}$

Higher is the bond order less is the bond length.

(iii) MgO > CaO > SrO > BaO

37.  bent  sp^2 –hybridisation.

AB_2L – type

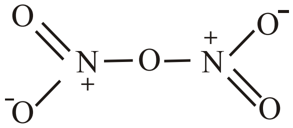
38. LiF < LiCl < LiBr < LiI

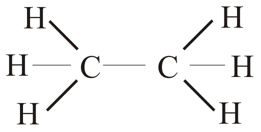
* Smaller cation has greater polarising power while larger anion has greater polarisability – Fajan rule.

39. AgI < AgBr < AgCl < AgF

40. 5 and 6

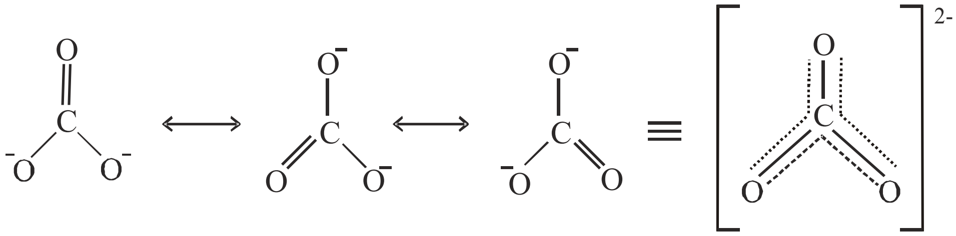
41. CN^{-} isoelectronic with CO ($14\ e^{-}$)

42. 

43. 

$C-H\left(sp^3-s\right)$, hybrid – pure

$C-C\left(sp^3-sp^3\right)$, hybrid – hybrid.

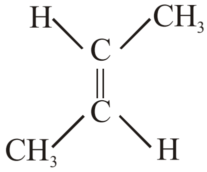
44. 

45. Here the bond angles are different (60^0 and 120^0) and bond lengths (equatorial and axial) are also not same. Equatorial P–Cl is smaller than axial, P–Cl.

46. B_2H_6 (diborane), $:\mathrm{N}\vdots\vdots\cdot\mathrm{O}:$ (11-valence $\bar{e}$s)

48. Greater is the bond order greater is the stability.

49. p and d - orbitals

50. 

non-polar (symmetrical) (T)

51. $d=\frac{m}{V}$ i.e, for a given mass $d\propto\frac{1}{V}$

$V\propto T$, so $d\propto\frac{1}{T}$ i.e, $d_1T_1=d_2T_2$

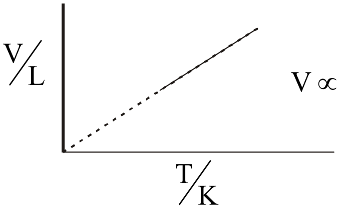
Combined gas equation, $\frac{P_1V_1}{T_1}=\frac{P_2V_2}{T_2}$

$d_1v_1=d_2v_2$ $\quad \frac{P_1\,d_2\,v_2}{T_1\,d_1}=\frac{P_2\,v_2}{T_2}\Rightarrow\frac{P_1}{d_1\,T_1}=\frac{P_2}{d_2\,T_2}$ i.e, $\frac{d_1\,T_1}{P_1}=\frac{d_2\,T_2}{P_2}$

52. $\left(P+\frac{an^2}{v^2}\right)\left(v-nb\right)=n.RT$

53. $p=\frac{an^2}{v^2}$ $\quad a=\frac{p^{/}\,v^2}{n^2}\,\text{atm.L}^2\,\text{mol}^{-2}$

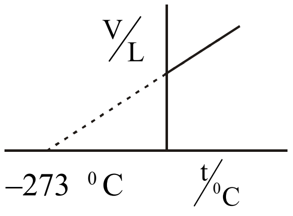
$V^{/}=nb$ $\quad b=\frac{V^{/}}{n}$ L mol^{-1}.

54. 

$V\propto T$

$V_t=V_0\left(1+\frac{t}{273}\right)$

V-t plot :



55. $\overline{c} = \sqrt{\dfrac{n_1 c_1^2 + n_2 c_2^2 + n_3 c_3^2 ....}{n_1 + n_2 + n_3 + ...}} = \sqrt{\dfrac{3RT}{M}}$

Square root of mean of square of all possible velocities.

56. It is the velocity shown by maximum fraction of molecules at a given temperature.

Mathematically most probable velocity = $\sqrt{\dfrac{2RT}{M}}$

57. $PV = \dfrac{1}{3}\ mN\overline{C}^2 = RT$ m = mass of a molecule

$= \dfrac{2}{3}KE$ N = total umber of molecules.

58-59. The total pressure of a mixture of gases at a given temperature is sum of their partial pressure that each gas experiences if present alone in the mixing container at that temperature.

$P = p_1 + p_2 + p_3 + ...$

$p_1 = X_1\ P$

* Partial pressure of a gas = mole fraction of that gas • total pressure.

60. $He > N_2 > O_2 > CO_2$

mol. mass ⇒ 4 28 32 44

rate of diffusion $\propto \dfrac{1}{\sqrt{\text{molecular mass}}}$

61. It is the temperature at which the volume of a given mass of gas theoretically reduces to zero at a given pressure i.e., – 273^0C. $\underset{-273}{V} = V_0\left(1 + \dfrac{-273}{273}\right) = 0$

In kelvin scale it is taken as O K

T (in absolute scale) = 273 +t (in celsius scale)

i.e. $5^0C = 273 + 5 = 278\ K = 278\ ^0A$

62. $R = \dfrac{PV}{nT}$ work $mol^{-1}\ K^{-1}$

R is reprsented as work done per mole per degree.

63. $PV = \dfrac{1}{3} mN\overline{C}^2 = RT \Rightarrow \dfrac{2}{3}E_K = RT$

$$E_K = \frac{3RT}{2}$$

Thus, E_K (kinetic energy) $\propto T$ (in absolute scale) kinetic energy of the particle increases with rise of temperature.

64. At high temperature and low pressure the real gas may behave ideal.

65. $He > H_2 > O_2 > CO_2$.

66. Most proabable velocity (C_P) : average velocity (C_a) : $\overline{C} = 1 : 1.128 : 1.224$

e.g. Average velocity = 0.9213. RMS velocity

Most probable velocity = 0.8164.RMS velocity.

67. $R = 1.987$ cal $mol^{-1} K^{-1}$

68. Higher is the value of T_C greater in the ease of liquefaction.

$$T_C = \frac{8a}{27Rb}$$, 'a' signifies the inter molecular force of attraction, $CO_2 > CH_4$.

69. 1 poise = 1 gram cm^{-1} s^{-1} = 10^{-1} Kg m^{-1} s^{-1}

70. Methanol < water < ethyleneglycol < glycerol

71. $\Delta T = 0$, $\Delta U = 0$, $Q = P\Delta V$ i.e. heat supplied in doing external work of expansion.

72. q = 0, $P\Delta V = -\Delta U$
work of expansion is due to decrease of internal energy.

73. No

74. $Q_P = \Delta H$, $Q = \Delta U$ and $\Delta H = \Delta U + \Delta n_{(g)}RT$

75. The total energy stored in the substance by virtue of its chemical nature is called internal energy.
$$U = U_r + U_v + U_t + U_e + U_n + U_{pe}$$
Ur = rotation energy U_v = vibrational, U_t = translational,
U_e = electronic energy, U_n = nuclear energy which arises due to binding nucleons within the nucleus.
U_{pe} = Potential energy which is possessed by the molecules by virtue of their position.

76. $\Delta V = 0$ i.e. at constant volume.

77. mass (e), viscosity (i), enthalpy (e), entropy (e) temperature (i), pressure (i), density (i), specific heat (i), specific gravity (i), surface tension (i), $\Delta U(e)$.

78. $\Delta G = \Delta H - T\Delta S$ i.e., $G = H - TS$

79. $-\Delta G^0 = RT\ln K = 2.303RT\log K$

80. $\Delta vS = \frac{30}{300} kJ\ K^{-1}\ mol^{-1} = 0.1 kJ\ K^{-1}\ mol^{-1}$

81. $\Delta_f H(NH_3) = \frac{-92.4}{2} = -46.2\ kJ/mol$

82. ΔH is negative, ΔS = –ve.

83. $\Delta V = 0$ i.e, no change is volume.

84. When unit amount of substance is taken e.g. mass is extensive but mass per unit unit volume is intensive.

85. Chemical reaction are generally performed at constant pressure. Heat change at constant pressure is turned as enthalpy.

86. Standard conditions i.e. 1 mole of substance at reference state i.e. temperature 298 K and pressure 1 bar. standard enthalpy change in written as ΔH^0 i.e., heat change for 1 mole of substance at 298 K and 1 bar.

87. $H^+ (aq.) + OH^- (aq) \longrightarrow H_2O(l) + 57.1\ kJ$
It is constant irrespective of nature of strong acid/base.

88. Due to greater hydration energy of F^- (smaller anion). This adds to heat of neutralisation.

89. No mass and energy get transferred to the surrounding.

90. Work obtained due to change in volume is called mechanical work i.e., = $P\Delta V$

$$= \frac{\text{Force}}{\text{area}} \times \text{area} \times \text{displacement}$$

$$= \text{Force} \times \text{displacement} = \text{work}$$

91. Equilibrium shift
(i) towards right
(ii) towards left
(iii) towards left reducing yield.
(iv) towards right yielding more

92. $\Delta n = -2$
$K_P = K_{C.}(RT)^{-2}$ i.e, $K_C = K_P \,.\, (RT)^2$

93. BF_3, H^+ Lewis acid as it's electron deficient.

$$NH_4^+ \longrightarrow NH_3 + H^+ \text{ (Bronsted acid)}$$

94. $HNO_2 \longrightarrow H^+ + NO_2^- \left(\text{conj. base}\right)$

CN^- (congujate base of HCN), F^- (conjugate base of HF)

ClO_4^- (conjugate base)

$$\underset{\text{Conj. acid}}{H_2O} \xleftarrow{H^+} \overline{OH} \xrightarrow{-H^+} \underset{\text{Conju. base}}{O^{2-}}$$

CO_3^{2-} conjugate base of HCO_3^- and H_2CO_3

S^{2-} conjugate base of HS^- and H_2S,

$$S^{2-} \xrightarrow{+H^+} HS^- \xrightarrow{+H^+} H_2S$$

Conjugate acids

95. (i) $OH^- = 5 \times 10^{-3}$ M

$-\log[OH^-] = -\log 5 + 3$

$\Rightarrow$ pOH $= -0.699 + 3 = 2.301$

pH $= 14 - 2.301 = 11.699$

(ii) $H^+ = 2 \times 10^{-3}$ M

$-\log\left(H^+\right) = -\log 2 + 3$

$= 2.699$

96. Precipitation occurs only when the ionic product value exceeds the K_{SP} value of the precipitate.

97. Due to common ion effect the dissociation of NH_4OH is suppressed reducing the $[OH^-]$ hence pH decreases. Further NH_4Cl becomes acidic on hydrolysis.

98. Acidic buffer : $CH_3COOH + CH_3COO\,Na$
Alkaline buffer : $NH_4OH + NH_4Cl$

99. pH = 3 i.e., $[H^+] = 10^{-3}$M
pH = 6 $[H^+] = 10^{-6}$ M
i.e., the solution concentration decreases 1000 times.

100. Dissociation/Ionisation of H_2S in acid Medium :

$$\xleftarrow[H^+]{H_2S \rightleftharpoons 2H^+ + S^{2-}}$$

In acidic medium due to common ion effect $[S^{2-}]$ decreases.
In alkaline medium –

$$H^+ + OH^- \longrightarrow H_2O$$

$$\xrightarrow[\text{alkaline medium}]{\substack{H_2S \rightleftharpoons 2H^+ + S^{2-} \\ OH^-}}$$

In alkaline medium the equilibrium shifts towards right on decrease of $[H^+]$. Thus $[S^{2-}]$ increases in alkaline medium.

101. Bronsted acid - proton donor e.g. HCl, NH_4^+, H_3O^+

base - proton acceptor Cl^- (all anions), $H_2O + H^+ \rightarrow H_3O^+$

102. Heat of solution of calcium acetate is exothermic, hence the solubility decreases with rise of temperature (Le Chateliers principle).

103. Solubility of AgCl is reduced in presence of KCl (Aq) where Cl^- acts a common ion.

$$\xleftarrow[Cl^-]{AgCl \rightleftharpoons Ag^+ + Cl^-}$$

104. $CaCl_2 < KCl <$ Na OAc

105. dil HCl.

106. Solubility product of a sparingly soluble salt at a given temperature is the mathematical product of the molar concentration of constituting ions raised to the power equal to the respective stoichiometric co-efficient.

$$K_{sp}(A_x\, B_y) = [A^{y+}]^x\ [B^{x-}]^y$$

e.g. $Ksp\ (Ag_2\, SO_4) = [Ag^+]^2 \left[SO_4^{2-}\right]$

107. Let 's' be the solubility of Ax By in mol/L

$$Ax\ By \rightleftharpoons x\ A^{y-} + y\ B^{x+}$$

$$\qquad\qquad x.s \qquad ys$$

$$K_{SP} = (xs)^x\ (ys)^y$$
$$= x^x\, y^y\ s^{x+y}$$

$$s = \left(\frac{K_{sp}}{x^x\, y^y}\right)^{1/x+y}$$

108. It is a salt solution of weak acid & weak base.

e.g. CH_3COONH_4.

109. 0.98 gram = 0.01 mol.

$$H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$$

$$\left[H^+\right] = 0.02\,mol/L$$

$$= 2\times10^{-2}\,M$$

$$-\log[H^+] = -\log 2 + 2$$

i.e., pH = 1.699

110. $K_p = \dfrac{p_X^x.p_Y^y}{p_A^a, P_B^b}$, $\qquad K_x = \dfrac{X_X^x.X_Y^y}{X_A^a, X_B^b}$,

p = partial pressure. X = mole fraction

111. Oxidation is de electronation e.g. $Na - e^- \rightarrow Na^+$

Reduction is electronation e.g. $Cl + e^- \rightarrow Cl^-$

112. $MnO_4^1 + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O]\times 2$

$$C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-]5$$

$$2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

113. $3\,\overset{(-2)}{H_2S} + 2\overset{(+5)}{HNO_3} \xrightarrow[\text{Decrease}3\times2]{\text{Increase }2\times3} 4H_2O + 2\overset{(+2)}{NO} + 3\,\overset{(O)}{S}$

114. HN_3, O.N. of $N = -\frac{1}{3}$; N_2H_4, $O.N = -2$

115. $\overset{-4}{C}H_4 + 4Cl_2 \longrightarrow \overset{+4}{C}Cl_4 + 4HCl$

$$\overset{-4}{C}H_4 \longrightarrow \overset{+4}{C}O_2$$

$$\overset{-3}{N}H_3 + 2O_2 \longrightarrow H\overset{+5}{N}O_3 + H_2O$$

116. Process of photosynthesis :

Step I $\quad 2H_2O \longrightarrow 2H_2 + O_2]\times 6$

Step II $\quad 6CO_2 + 12H_2 \longrightarrow C_6H_{12}O_6 + 6H_2O$

$$6CO_2 + 12H_2O \longrightarrow C_6H_{12}O_6 + 6O_2 + 6H_2O$$

Oxidation

117. $2\overset{+1}{Ag}Br + \overset{-2\ +6\ -4}{C_6H_6O_2}(aq) \longrightarrow 2\overset{0}{A}g(S) + 2HBr(aq) + \overset{0\ +4\ -4}{C_6H_4O_2}$

reduction

$C_6H_6O_2$ oxidised to $C_6H_4O_2$, hence it's a reductant

Ag^+ reduced to Ag, So AgBr is an oxidant

118. HI > HBr > HCl > HF

HI is a better reducing agent as it is easily oxidised.

$$2HI + O \rightarrow H_2O + I_2$$

119. $HgCl_2$, Cr_2O_3

120. $\overset{+2\ -2}{H_2O}(S) + \overset{0}{F_2}(g) \longrightarrow \overset{+1\,-1}{HF}(g) + \overset{+1\,-2\,+1}{HOF}$

O.N. of F increases from 0 to +1 is HOF and decreases from 0 to –1 (HF). Hence it's a redox reaction.

121. Hydrogen as oxidant, $Na + \frac{1}{2}\overset{0}{H_2} \longrightarrow Na\overset{-1}{H}$

as reductant, $\overset{0}{H_2} + Cl_2 \longrightarrow 2\overset{+1}{H}Cl$

122. Oxygen as oxidant, $Ca + \frac{1}{2}\overset{0}{O_2} \longrightarrow CaO^{-2}$

as reductant, $\frac{1}{2}\overset{0}{O_2} + F_2 \longrightarrow \overset{+2}{O}F_2$, oxygen reduces fluorine to fluoride.

123. O.N. Cl in HCl (–1), HClO (+1), $ClO_4^-(+7)$, $ClO_3^-(+5)$

$$ClO_2^-(+3),\ Cl_2O(+1)$$

124. 1 gram eqvt. of $KMnO_4 \equiv$ 1 gram eqvt. of $H_2C_2O_4 = 45$ gram.

125. Acidic–$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$
Alkaline – $H_2O_2 + 2e^- \rightarrow 2(OH)^-$

In both the cases $2e^-s$ are involved.

126. T (true)

127. $Na + H_2O \longrightarrow NaOH + \frac{1}{2}H_2$

$Zn + 2NaOH \longrightarrow Na_2ZnO_2 + H_2$

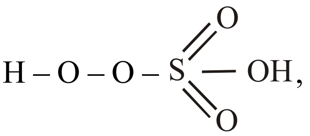
$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$

128. Passing through $(CH_3COO)_2Pb$ solution to remoke H_2S , $AgNO_3$ solution to remove PH_3 and AsH_3 , KOH solution removes CO_2, SO_2. Dried over P_2O_5 or anhydrous $CaCl_2$.

129. Atomic hydrogen. $H_2 \xrightleftharpoons{\text{electric arc}} 2H - 104.5 \text{ K cal}$

life period pf 'H' is only 0.3 sec. It is reversed to from H_2 which is highly exothermic & used for welding.

130. Calgon (sodim hexameta phosphate, $Na_6P_6O_{18}$).
It is a water softener to remove Mg^{2+} and Ca^{++} of hard water.

131. Peroxy linkage. 
$H-O-O-S(=O)(=O)-OH$, $HO_3S-O-O-SO_3H$

132. $H_2O_2 \longrightarrow H_2O + O$, pleaches on oxidation.

133. H_2O_2 + $FeSO_4$ (catalyst). With lactic acid $CH_3CH(OH)COOH$, it gives pyruvic acid $CH_3CO–COOH$.

134. Used as an antiseptic under the name 'perhydrol' as it is an oxidant.

135. $2H_2(g) + O_2(g) \xrightarrow[\text{electric discharge}]{\text{970K or}} 2H_2O$. The reaction is exothermic.

136. High molecular mass.

137. H_2O_2 as reductant. $H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^- \text{(Oxidation)}$

e.g. Ag_2O reduces to Ag

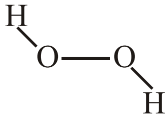
$$\begin{array}{c} H_2O_2 \longrightarrow O_2 + 2H \\ Ag_2O + 2H \longrightarrow 2Ag + H_2O \\ \hline H_2O_2 + Ag_2O \longrightarrow 2Ag + H_2O + O_2 \end{array}$$

As oxidant.

$$\begin{array}{c} H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O, \\ H_2S \longrightarrow S + 2H^+ + 2e^- \\ \hline H_2O_2 + H_2S \longrightarrow 2H_2O + S \end{array}$$

It oxidises H_2S leaving 'S'.

138. 2-ethyl anthraquinone.

139. 
,non-planar. The dihedral angle differs, in gaseous phase 111.5^0 and in solid phase (90.2^0).

So also the O–O and O–H bond lengths.

140. H_2O_2 when shakon with acidified $K_2Cr_2O_7$ with ether, the layer turns blue due to formation chromium peroxide (CrO_5). water does not respond this test.

141. True. $Na_2SO_4 . 10H_2O$ – Glauber salt.

142. $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

143. K, Rb, C_S

NaO_2 can be prepared with high oxygen pressure.

LiO_2 cann't be prepared in pure form.

144. $LiF > LiCl > Li\ Br > Li\ I$

$F^- > Cl^- > Br^- > I^-$

145. $2NH_4Cl + CaO \xrightarrow{\Delta} 2NH_3 \uparrow + CaCl_2 + H_2O$

146. $Na_2CO_3.10H_2O \xrightarrow[375\text{ K}]{\Delta} Na_2CO_3.H_2O \xrightarrow[7373\text{K}]{\Delta} Na_2\ CO_3$ (White powder) soda ash.

147. Solar evapouration

148. $2Li\ NO_3 \longrightarrow Li_2O + 2NO_2 + \frac{1}{2}O_2,\ \ Na\ NO_3(s) \xrightarrow{\Delta} NaNO_2(s) + \frac{1}{2}O_2(g)$

149. Li, Mg

150. $CaSO_4.2H_2O \xrightarrow{\Delta} 2\ CaSO_4.H_2O \xrightarrow{>393\text{ K}} CaSO_4$

151. A suspension of $Ca(OH)_2$ or hydrated lime in water

152. The solubility metal carbonate and metal sulfate decreases on moving down the group.

$BeSO_4 > MgSO_4 > CaSO_4 > SrSO_4 > BaSO_4$

Lattice energy almost remains constant $\left(SO_4^{2-}\text{ bigger ion}\right)$ but the hydration energy decreases from Be^{2+} to Ba^{2+}. Hence the solubility decreases.

153. $Mg(OH)_2 \xrightarrow{\Delta} MgO + H_2O$

$MgCO_3 \xrightarrow{\Delta} MgO + CO_2$

154. On heating sodium chloride with conc. H_2SO_4, HCl(g) is prepared.

$$2NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl \uparrow$$

155. Electrolysis of fused KCl separates K at cathode.

$$K^+ + e^- \longrightarrow K(s)$$

156. Reactivity : $Be < Mg < Ca < Sr < Ba$

157. $BeO \xrightarrow[2175\text{ K}]{C} Be_2C \xrightarrow{\text{Water}} CH_4$

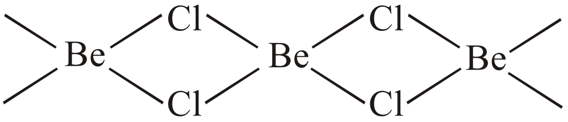
158. Active metal oxide (CaO, BaO, SrO) with water forming metal hydroxide is called slaking.

$$MO + H_2O \longrightarrow M(OH)_2$$

159. The hydration energy should exceed the lattice energy of the salt.

e.g. $MgSO_4$ is more soluble in water than $CaSO_4$. The hydration energy depends on charge (higher) on the ion and ionic size (smaller).

160. $BeCl_2$ (s)



Polymeric structure with Cl bridges.

161. $BeSO_4 < MgSO_4 < CaSO_4 < SrSO_4 < BaSO_4$
773 K 1168 K 1422 K 1647K
Decompositing temperature increases showing their stability.

162. Be and Mg (due to their higher $\Delta_i H$ values)

163.
$$\begin{array}{l} Cl_2 + H_2O \longrightarrow HCl + HOCl \\ NaOH + HCl \longrightarrow NaCl + H_2O \\ NaOH + HOCl \longrightarrow NaOCl + H_2O \\ \hline [2NaOH + Cl_2 \xrightarrow{Cold} NaOCl + NaCl + H_2O] \times 3 \end{array}$$

On heating $NaClO_3$ is formed.

$$\begin{array}{l} 3NaOCl \longrightarrow NaClO_3 \quad + 2NaCl \\ \hline 6\,NaOH + 3Cl_2 \longrightarrow 5\ NaCl + NaClO_3 + 3H_2O \end{array}$$

164. Sodium is an active metal. It loses its lusture due to surface oxidation and subsequent formation of carbonate (Sodium + oxygen + moisture + CO_2 in air)

$$\begin{array}{l} 4Na + O_2 \rightarrow 2Na_2O \\ [Na_2O + H_2O(\text{moisture}) \rightarrow 2NaOH] \times 2 \\ [2NaOH + CO_2(\text{air}) \rightarrow Na_2CO_3 + H_2O] \times 2 \\ \hline 4Na + 2CO_2 + O_2 + 2H_2O \rightarrow 2Na_2CO_3 + 2\,H_2O \end{array}$$

This is the reason why sodium is kept preserved in kerosene. Lighter Lithium is kept in paraffin.

165. Uses of CaO.
(i) Drying agent
(ii) As a flux during metal (Fe) extraction
(iii) Cement industry

166. $M_2SO_4, T_2(SO_4)_3 . 24\ H_2O$
M = monovalent cation like K^+, NH_4^+
T = Trivalent cation like Fe^{3+}, Al^{3+}, Cr^{3+}
e.g. Potash alum $K_2SO_4, Al_2(SO_4).\ 24\ H_2O$

167. $BF_3 < BCl_3 < B\,Br_3 < BI_3$

168. BF_3 (sp^2 hybridises), BF_4^- (sp^3 - hybridised)
Increased p-character increases bond length.

169. Sugar with conc. H_2SO_4 form sugar charcoal as conc. H_2SO_4 takes away water.

$$C_{12}\ H_{22}\ O_{11} \xrightarrow{\text{Conc. } H_2SO_4} 12C + 11\ H_2O$$

170. $6CH_3OH + 2Al \longrightarrow 2(CH_3O)_3\,Al + 3H_2 \uparrow$

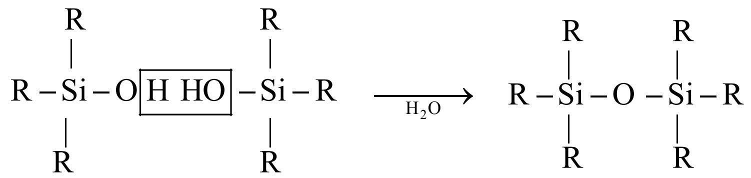
171. $CO + ZnO \longrightarrow Zn + CO_2$ (CO is a reducing agent)

172. $CO_2 + H_2O \longrightarrow H_2CO_3$ (carbonic acid)

173. HF (Hydrofluoric acid.)

174. Silicone is formed. (hexaalkyl siloxane is formed.

$$R_3SiCl \xrightarrow{H_2O} R_3Si\,OH$$



$$R_3Si{-}O\boxed{H\ HO}{-}SiR_3 \xrightarrow[H_2O]{} R_3Si{-}O{-}SiR_3$$

175. Graphite gets oxidised by con. HNO_3 to form insoluble yellowish green graphitic acid, $C_{11}H_4O_5$

176. $_6C$, $_{14}Si$, $_{32}Ge$ $_{50}Sn$, $_{82}Pb$, $_{114}Fl$ (Flerovium)

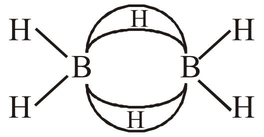
177. C > Si > Ge > Sn < Pb

178. $_6C = 1s^2\ 2s^2\ 2p^2$, $\underset{2s}{\boxed{\uparrow}}\ \underset{2p}{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}$

Due to presence of four unpair $\bar{e}$(s) as energy of 2s is comparable to E_{2p}.

179. $AlCl_3$ is in sextet state and short of two $\bar{e}$(s) to satisfy octet, $\bar{e}$-deficient. Species having tendency to accept ($\bar{e}$) are called Lewis acids.

180. Bauxite $Al_2O_3 \cdot 2H_2O$, Cryolite - Na_3AlF_6

181. 

Two B – H – B. two $\bar{e}$ three centre bridge / banana bonds

Four ter min al regular B – H bond (two $\bar{e}$ two centre bond)

182. C^{4-} (carbide), C_2^{2-} (acetylide) ________ ionic carbide.

C_3^{4-} (allylenide) e.g. $Li_4C_3 \xrightarrow{H_2O} C_3H_4$ i.e $CH_3-C\equiv CH$

There are also covalent carbide (B_4C, SiC) and intersitial carbide (T_iC)

183. sp^3, sp^2 and sp

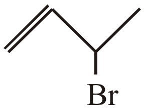
184. Ring compounds containing-C-C- chain are of two types

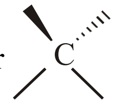
(i)	alicyclic e.g., cyclopropane	(ii)	aromatic e.g. benzene.

185.

(i)	pyridine	(ii)	Furan

186. S, O, N

187. (i) 

(ii) ▶ Towards the viewer  ┆ (iii) away from the viewer

- solid lines represented the bonds in the plane of the surface.

188.

	Prefix	**Suffix**
– COOH	Carboxy-	– Oic acid / - carboxylic acid as the case may be
– SO_3H	Sulfo -	– sulfonic acid
– CN	Cyano-	– nitrile / -Carbonitrile
– CHO	Fomyl-	– al / –carbaldehyde

189. CH_3SO_3H Methanesulfonic acid

$C_6H_5CONH_2$ Benzenecarboxamide

190. $CH_3CO\overset{+}{C}H_2 < CH_3-\overset{+}{C}H_2- < CH_2=CH-\overset{+}{C}H_2 < C_6H_5-\overset{+}{C}H_2$

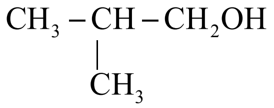
191. Electron rich species having tendency to donate $\bar{e}$ (s) are called nucleophile i.e. nucleus loving species.

e.g. $\bar{O}H$.

192. Isomers of $C_4H_{10}O$.

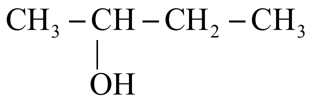
Alcohol : $CH_3-CH_2-CH_2-CH_2-OH$ Butan-1-ol

(1^0)



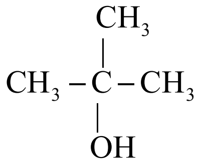
$CH_3-\underset{\displaystyle CH_3}{\underset{|}{C}H}-CH_2OH$ 2-methylpropan-1-ol

(2^0)



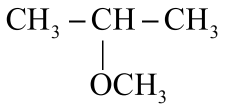
$CH_3-\underset{\displaystyle OH}{\underset{|}{C}H}-CH_2-CH_3$ Butan-2-ol

(3^0)



$CH_3-\overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle OH}{\underset{|}{C}}}}-CH_3$ 2-methylpraopan-2-ol

Ether. $CH_3-CH_2-CH_2-OCH_3$ 1-methoxypropane

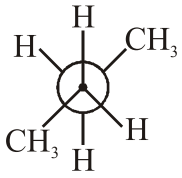


$CH_3-\underset{\displaystyle OCH_3}{\underset{|}{C}H}-CH_3$ 2-Methoxypropane

$CH_3-CH_2-O-CH_2-CH_3$ Ethoxyethane

193. False.

194.

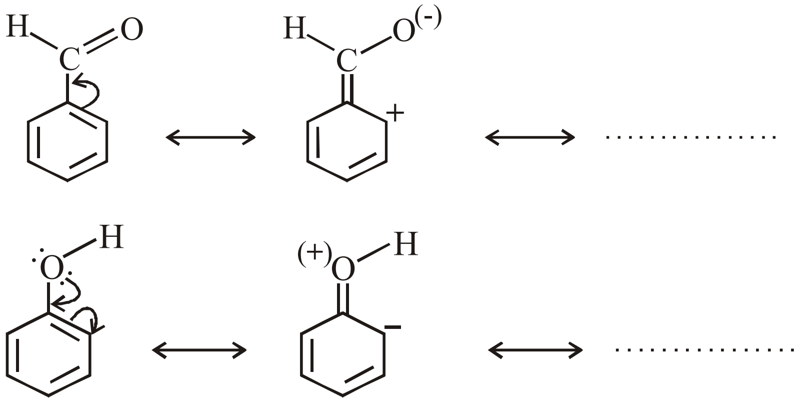


Antibutane

195. $CH_3-\overset{+}{C}HOCH_3$ $CH_3-\overset{+}{C}HCH_3$ $CH_3-\overset{+}{C}HCOCH_3$

196. $HNO_3+H_2SO_4 \longrightarrow \overset{+}{N}O_2+HSO_4^-+H_2O$

197.



198. (i) 2-Methyl butane

(ii) But - 3 - en - 2 - ol

(iii) 3-Methyl hex -1-en-4-yne

(iv) 2-Methyl butanal

(v) 4-amino-5-hydroxy hex–2–enoic acid.

(vi) Cyclobut–2–en–1–ol

(vii) N-methylcyclohexanecarboxamide

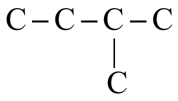
(viii) 2–(chloromethyl) butane–1, 4–dioic acid.

(ix) 4-methyl cyclohexan-1-ol

(x) Ethylpropenoate

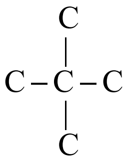
(xi) 3–(1,1-dimethylethyl-) but-3-enoylchloride

199. (a) $CH_3-CH_2-CH_2-CH_2-CH_3$ pentane



$$\begin{array}{c}C-C-C-C\\ \quad\;\; | \\ \quad\;\; C\end{array}$$

2-methyl butane,

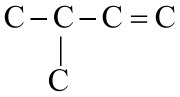


$$\begin{array}{c}C\\ |\\ C-C-C\\ |\\ C\end{array}$$

2, 2-dimethyl propane

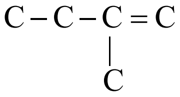
(b) $C_5\ H_{10}$ $CH_2-CH_2-CH_2-CH=CH_2$ Pent-1-ene

$C-C-C=C-C$ Pent-2-ene



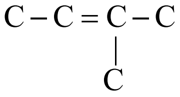
$$\begin{array}{l}C-C-C=C\\ \quad\;\; |\\ \quad\;\; C\end{array}$$

3-methyl but-l-ene



$$\begin{array}{l}C-C-C=C\\ \qquad\quad |\\ \qquad\quad C\end{array}$$

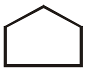
2-methyl but-1-ene



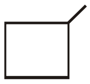
$$\begin{array}{l}C-C=C-C\\ \qquad\quad |\\ \qquad\quad C\end{array}$$

2-methyl but-2-ene

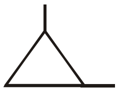
Ring isomers :



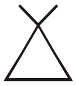
Cyclopentane



methylcyclobutane,

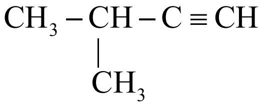


1,2-dimethylcyclopropane



1, 1– dimethylcyclopropane.

(c) C_5H_8 : $CH_3-CH_2-CH_2-C\equiv CH$ Pent-1-yne



$$\begin{array}{l}CH_3-CH-C\equiv CH\\ \qquad\quad |\\ \qquad\quad CH_3\end{array}$$

3 Methylbut-1-yne

$C-C\equiv C-C-C$ Pent-2-yne

(no chain isomer)

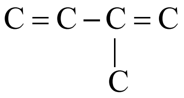
$C=C-C=C-C$ Penta-1,3–diene

(Conjugative diene)

$C=C=C-C-C$ Penta-1,2 - diene

(Consecutive diene)

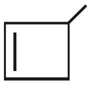
$C=C-C-C=C$ Penta-1, 4-diene



$$\begin{array}{l}C=C-C=C\\ \qquad\quad |\\ \qquad\quad C\end{array}$$

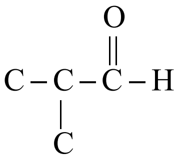
2-Methyl buta-1, 3–diene and others

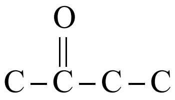
Cyclopentene.



3-methylcyclobut-1-ene and others

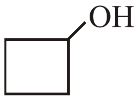
(d) C_4H_8O. $CH_2–CH_2–CH_2–CHO$ butanal

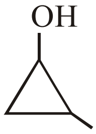
Aldehyde–

$C-\underset{\underset{C}{|}}{C}-\overset{\overset{O}{\|}}{C}-H$
2-methyl propanal

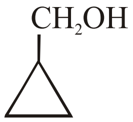
Ketone –

$C-\overset{\overset{O}{\|}}{C}-C-C$
butanone

– enol $CH_2 = CH - CH_2–OH$ but–3–en–1–ol

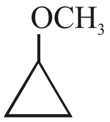
Alkoxyalkene (ether) $CH_2 = CH - CH_2 - OCH_3$ 1-methoxyprop-2-ene

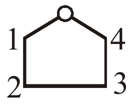
Ring isomer –

Cyclobutanol


2-methylcyclopropanol


cyclopropyl methanol

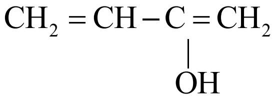
(Write other possible isomers)

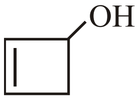

Methoxycyclopropane


(Tetrahydro furan) / Oxolane

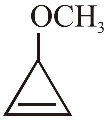
1, 4 – epoxy butane

(e) $CH_3-C\equiv C-CH_2-OH$ (C_4H_6O) But-2-yn-1-ol

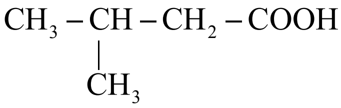

$CH_2=CH-\underset{\underset{OH}{|}}{C}=CH_2$
Buta-1, 3-dien-2-ol

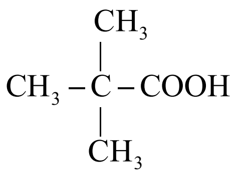

Cyclo but – 2 – en – 1 – ol and others (try yourself)

$H-C\equiv C-CH_2OCH_3$ Methoxypropyne


Methoxycycloprop-2-ene

(f) $C_5H_{10}O_2$ Acids $CH_3 - CH_2 - CH_2 - CH_2 - COOH$ Pentanoic acid


$CH_3-\underset{\underset{CH_3}{|}}{CH}-CH_2-COOH$
3-Methyl butanoic acid



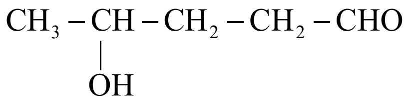
$CH_3-C(CH_3)_2-COOH$

2, 2–dimethylpropanoic acid etc.

Ester. $CH_3-CH_2-CH_2-COOCH_3$ Methylbutanoate

$CH_3-CH_2-COOC_2H_5$ Ethylpropanoate etc.

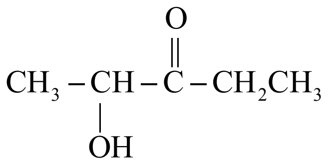
Hydroxyalkanal-



$CH_3-CH(OH)-CH_2-CH_2-CHO$

4-hydroxypentanal

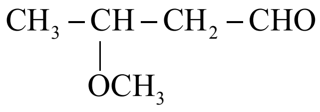
Hydroxyalkanone-



$CH_3-CH(OH)-\overset{O}{\overset{\|}{C}}-CH_2CH_3$

2-hydroxypentan-3-one

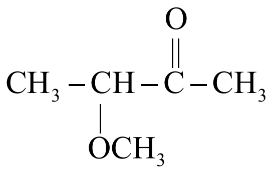
Alkoxyalkanal-



$CH_3-CH(OCH_3)-CH_2-CHO$

3-Methoxy butanal

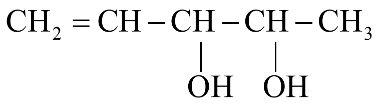
Alkoxyalkanone-



$CH_3-CH(OCH_3)-\overset{O}{\overset{\|}{C}}-CH_3$

3-methoxybutanone

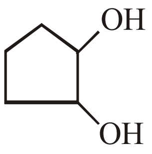
Unsaturated diol



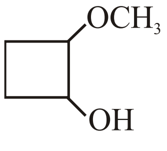
$CH_2=CH-CH(OH)-CH(OH)-CH_3$

$pent-4-ene-2,3-diol$

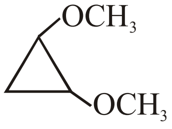
Cyclic diol



cyclopentane - 1,2 - diol

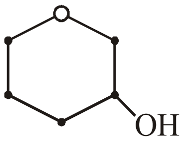


2-methyoxycyclobutan-1-ol



1,2–dimethoxy cyclopropane

Hydroxy cyclicether



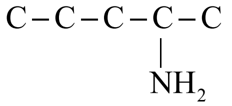
2-hydroxy-1,5-epoxy pentane

(Others try yourself)

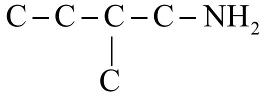
(g) $C_5H_{13}N$

$CH_3-CH_2-CH_2-CH_2-CH_2-NH_2$ Pentan-1-amine

1^0 amine :

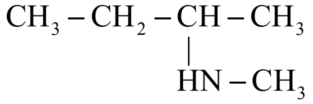


$C-C-C-\underset{\underset{NH_2}{|}}{C}-C$ Pentan-2-amine



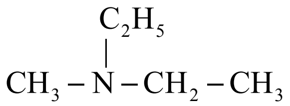
$C-C-\underset{\underset{C}{|}}{C}-C-NH_2$ 2-methylbutan-1-amine

2^0 amine :



$CH_3-CH_2-\underset{\underset{HN-CH_3}{|}}{CH}-CH_3$ N-methylbutan-2-amine

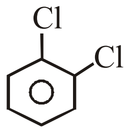
3^0 amine:



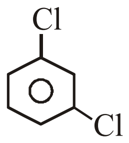
$CH_3-\overset{\overset{C_2H_5}{|}}{N}-CH_2-CH_3$ N-ethyl-N-methylethanamine

(Try others)

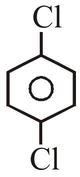
(h) $C_6\,H_4\,Cl_2$



1, 2-dichlorobenzene



1,3-dichlorobenzene



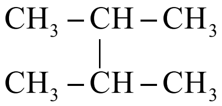
1,4-dichlorobenzene

200. $AlCl_3$ or $FeCl_3$. $AlCl_3+C_2H_5Cl \longrightarrow \underset{\text{(electrophile)}}{C_2^+H_5}+AlCl_4^-$

201. Catalytic hydrogenation of alkene and alkyne to alkane is known as Sabatier – Senderen's reactions.

$CH_2=CH_2+H_2 \xrightarrow[523-573K]{Ni} CH_3-CH_3$

202. $CH_3-CHCl-CH_3 \xrightarrow[\Delta]{Na/dry\ ether}$



$\begin{array}{c} CH_3-CH-CH_3 \\ | \\ CH_3-CH-CH_3 \end{array}$

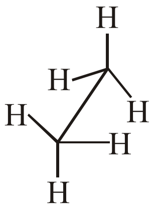
2, 3- dimethyl butane

203. At cathode H_2(g), At anode CH_3–CH_3 and CO_2

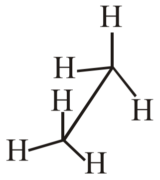
204. Fluorination > Chlorination ≃ bromination > iodination

Incase of alkane $3^0 > 2^0 > 1^0 > CH_3-$

205.

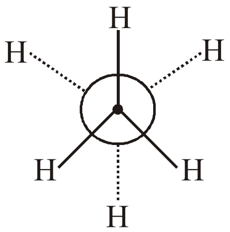


Staggered ethane

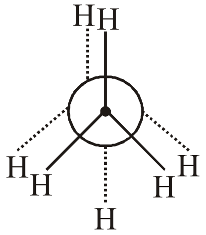


eclipsed ethane

206.



Staggered 60^0



eclipsed 120^0

207. The repulsive interaction between the electron clouds which affect the stability of conformation is called torsional strain e.g. staggered form has the least torsional strain while eclipsed form has maximum. The energy difference between these two extreme forms is only 12.5 kJ / mole which can be achieved by ethane molecule at ordinary temperature. So it's not possible to separate or isolate conformational isomers.

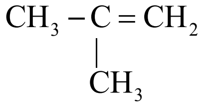
208. Aromatisation : n-hexane at 773 K under 10–20 atom in presence of Cr_2O_3 or V_2O_5 or Mo_2O_3 gives benzene.

209. $CH_4 + H_2O\ (\text{steam}) \xrightarrow[1273K]{Ni} CO + 3H_2$

210. A gaseous natural product produced from bacterial decomposition of organic materials (mostly cellulose) in marshes. It primarily consists of methane.

211. Isomerism of alkene (C_nH_{2n})

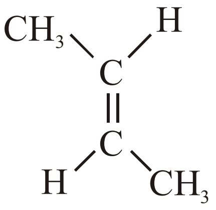
(i) Chain isomerism $CH_2 = CH - CH_2 - CH_3$ but-1-ene



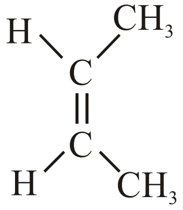
$CH_3 - \underset{\displaystyle CH_3}{\underset{|}{C}} = CH_2$ methylpropane

(ii) Position isomerism $CH_3 - CH = CH - CH_3$ but-2-ene

(iii) Geometrical isomerism

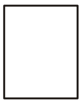


trans - but-2-ene

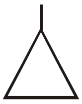


cis-but-2-ene

(iv) Ring isomerism.



cyclobutane



Methylcyclopropane

212. (i) They have same formula but different configuration.

(ii) Cis-isomer has greater dipolemoment than trans.

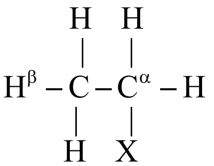
(iii) Cis - isomer has greater b.p. than trans. due to dipole-dipole force.

(iv) Due to symmetry, trans-isomer has greater m.p., less dis-ordered in solid state.

213. $CH_3Br + C_2H_5Br \xrightarrow[\text{reflux}]{\text{Na / dry ether}} \underset{\text{ethane}}{CH_3 - CH_3} + \underset{\text{butane}}{C_2H_5 - C_2H_5} + \underset{\text{propane}}{CH_3 - C_2H_5}$

214. $\underset{\text{n-butane}}{CH_3 - CH_2 - CH_2 - CH_3} \xrightarrow[HCl(g)]{\text{Anhy. } AlCl_3} \underset{\text{isobutane}}{CH_3 - CH(CH_3) - CH_3}$

215. Elimination of β -H atom in dehydration of alcohol or dehydrohologenation of alkyl halide is called β - elimination.



$$H^{\beta} - \underset{\displaystyle H}{\underset{|}{\overset{\displaystyle H}{\overset{|}{C}}}} - \underset{\displaystyle X}{\underset{|}{\overset{\displaystyle H}{\overset{|}{C^{\alpha}}}}} - H \xrightarrow[\Delta]{\text{alc. KOH}} CH_2 = CH_2 + HX, \qquad X = \text{halogen.}$$

216. Zn-HCl is a reducing agent. It reduces alkyhalide except RF to alkane.

$$RX + H_2 \xrightarrow{Zn+H^+} RH + HX$$

$$\left(Zn + 2H^+ \longrightarrow Zn^{2+} + H_2\right)$$

$$C_2H_5Br + H_2 \longrightarrow C_2H_6 + HBr$$

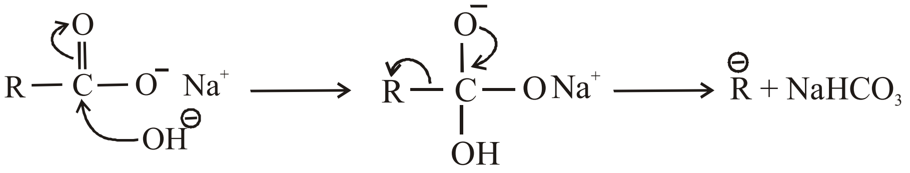
217. Soda-lime = NaOH + CaO in the ratio 3 : 1

NaOH alone can initiale the decarboxylation but the presence of CaO –

(i) permits the reaction to be carried out at relatively higher temperature (630 K) ensuring decarboxylation.

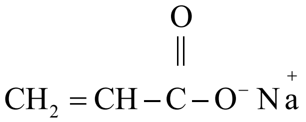
(ii) NaOH is kept dry in presence of CaO

$$RCOOH + NaOH \longrightarrow RCO\bar{O}\ Na^+ + H_2O$$



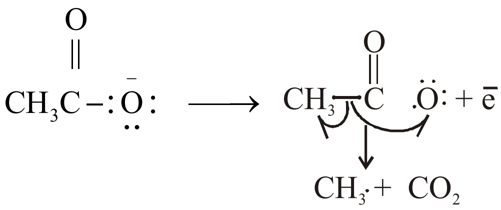
$$\underset{\text{(intermediate)}}{R^-} + NaHCO_3 \xrightarrow{\Delta} RH + Na^+\ CO_3^{2-}$$

Ease of Decarboxylation depends on stability of $\bar{R}$.



$CH_2 = CH - \overset{\overset{\displaystyle O}{\|}}{C} - O^-\ N\overset{+}{a}$ easily undergoes decarboxylation forming $CH_2{=}CH_2$ as $CH_2{=}\overset{\ominus}{C}H$ is more stable than $CH_3 - \bar{C}H_2$.

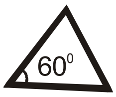
218. At anode -



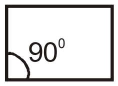
$$CH_3\overset{\overset{\displaystyle O}{\|}}{C} - :\overset{-}{\underset{\cdot\cdot}{O}}: \longrightarrow CH_3 - \overset{\overset{\displaystyle O}{\|}}{C}\ \ .\overset{\cdot\cdot}{O}: + \bar{e} \longrightarrow CH_3\cdot + CO_2$$

$$CH_3\cdot + \cdot CH_3 \longrightarrow CH_3 - CH_3 \text{(ethane)}$$

219. No,. cyclobutane is more stable than cyclopropane where bond is more strained due bond angle 60^0.

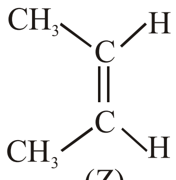


more reactive

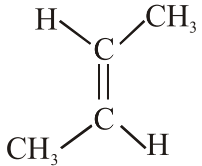


less reactive

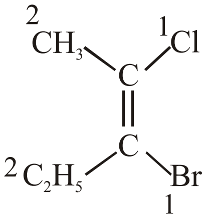
220.



(Z)

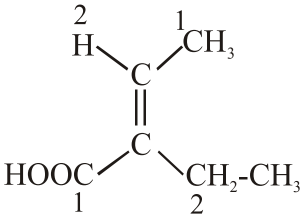


(E)

221. 

Z-3-bromo-2-chloropent-2-ene

Z = Zusammem



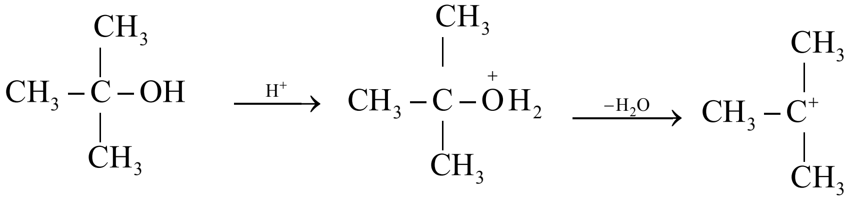
E-2-ethyl but-2-enoic acid

E = Entgegen

222. trans but-2-ene < cis-but-2-ene < but -1-ene < ethene, more is substituted alkene more is the stability & less is the heat of hydrogenation.

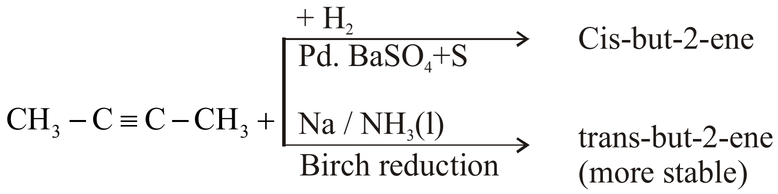
223. Ethene is formed.

224. $3^0 > 2^0 > 1^0$



$$CH_3-\underset{CH_3}{\overset{CH_3}{C}}-OH \xrightarrow{H^+} CH_3-\underset{CH_3}{\overset{CH_3}{C}}-\overset{+}{O}H_2 \xrightarrow{-H_2O} CH_3-\underset{CH_3}{\overset{CH_3}{C^+}}$$

(stable)

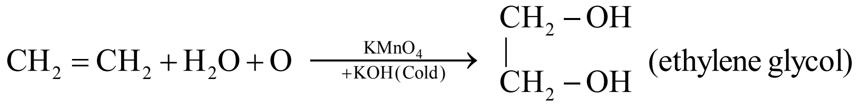
225. 

$$CH_3-C\equiv C-CH_3 + \begin{cases} \xrightarrow[\text{Pd. } BaSO_4+S]{+H_2} \text{Cis-but-2-ene} \\ \xrightarrow[\text{Birch reduction}]{Na / NH_3(l)} \text{trans-but-2-ene (more stable)} \end{cases}$$

226. It is heterogeneous catalyst i.e. Pd supported over $CaCO_3$ or $BaSO_4$ and partially poisoned by additron of sulfur or quinoline.

227. Br_2(Aq) Orangered (or) alkaline $KMnO_4$ (purple) loses colour on contact with $C=C$ or $C\equiv C$ (unsaturation) Alkaline $KMnO_4$ is called Baeyer's reagent.

Benzene does not affect the colour of Baeyers reagent.

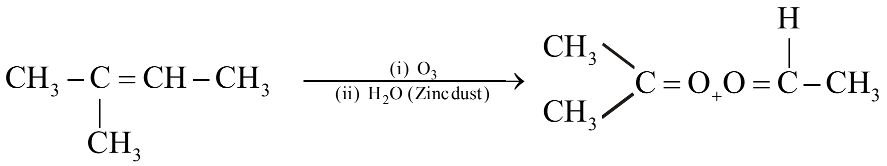
228. 

$$CH_2=CH_2+H_2O+O \xrightarrow[+KOH(Cold)]{KMnO_4} \begin{array}{l} CH_2-OH \\ | \\ CH_2-OH \end{array} \text{(ethylene glycol)}$$

$$CH_2=CH_2+4(O) \xrightarrow[\Delta(heat)]{KMnO_4,\ KOH} HCOOH \xrightarrow{(O)} CO_2+H_2O$$

$$2KMnO_4+H_2O \longrightarrow 2KOH+2MnO_2+3[O]$$

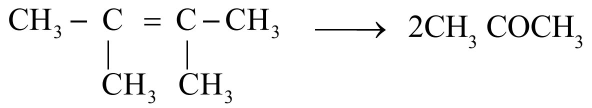
$$2KMnO_4+2\,KOH \longrightarrow 2\ K_2MnO_4+H_2O+O$$

229. (i) 

2-methyl but-2-ene

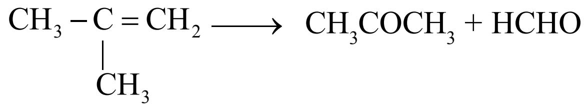
(ii) $CH_3-CH=CH-CH_3 \longrightarrow 2CH_3CHO$

But-2-ene

(iii) 

$CH_3-\underset{CH_3}{\underset{|}{C}}=\underset{CH_3}{\underset{|}{C}}-CH_3 \longrightarrow 2CH_3\,COCH_3$

2,3 - dimethyl but-2-ene

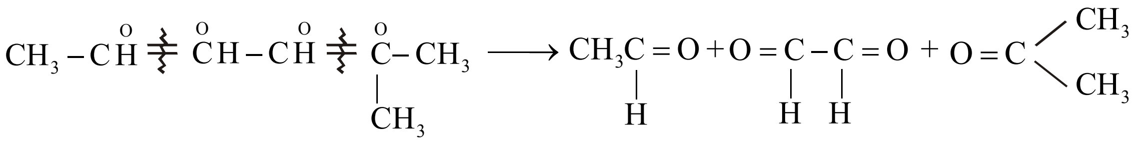
(iv) 

$CH_3-\underset{CH_3}{\underset{|}{C}}=CH_2 \longrightarrow CH_3COCH_3 + HCHO$

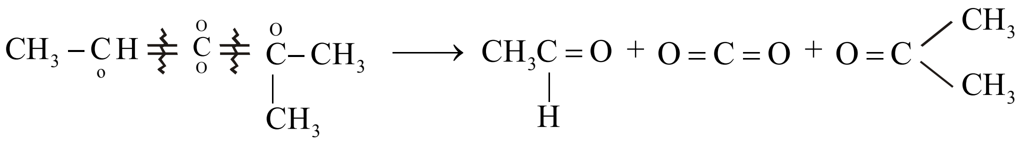
(2-Methylpropene)

(v) $CH_3-CH=CH_2 \longrightarrow CH_3CHO + HCHO$

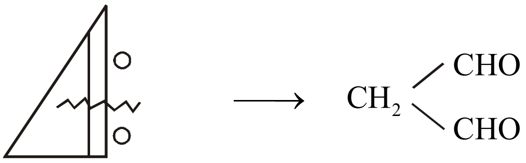
Propene

(vi) 

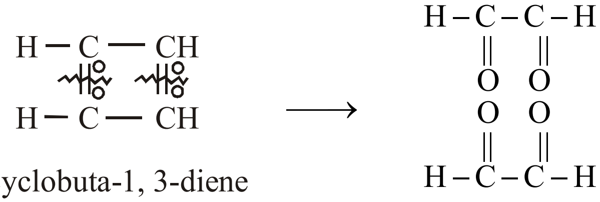
2-Methyl hexa-2, 4-diene glyoxal

(vii) 

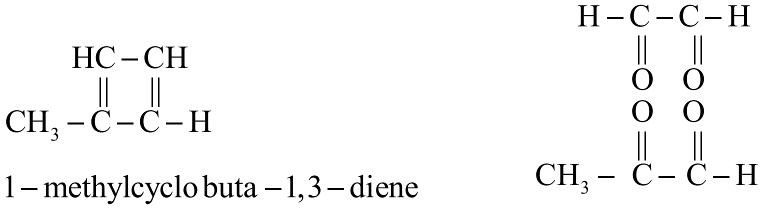
2-Methyl penta-2, 3-diene

(viii) 

Cyclopropene

(ix) 

Cyclobuta-1, 3-diene

(x) 

1 – methylcyclo buta – 1,3 – diene

230. 

$\underset{CH\ COOK}{\overset{CH\,COOK}{||}}$,

Pot. maleate / fumerate on electrolysis, produces ethyne at anode.

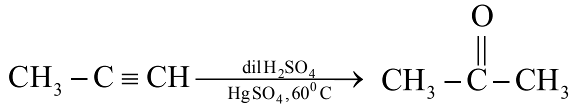
$CH_3-\underset{CH\,COOK}{\underset{||}{C}}-COOK$

(Potassium salt of methyl substituted maleic acid).

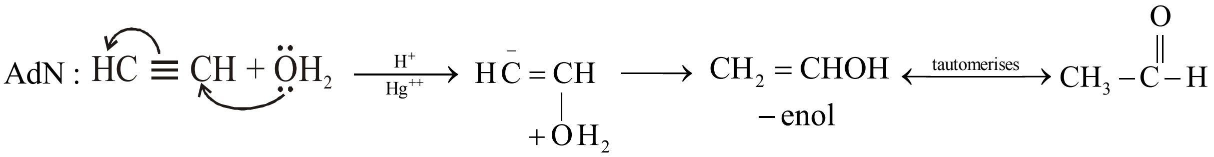
231. $HC\equiv CH \xrightarrow{Na/NH_3} \underset{\text{Monodium acetylide}}{HC\equiv C\ Na} \xrightarrow[-2NaBr]{C_2H_5Br} \underset{\text{But}-1-\text{yne}}{HC\equiv C-CH_2-CH_3}$

232. (i) $CH\equiv CH \xrightarrow[Hg\,SO_4,\,60^0\,C]{dil\ H_2SO_4} CH_3CHO$

(ii)

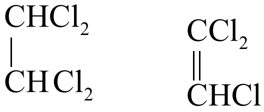


$$CH_3-C\equiv CH \xrightarrow[HgSO_4, 60^0\,C]{dil\,H_2SO_4} CH_3-\overset{\overset{\displaystyle O}{||}}{C}-CH_3$$



$$AdN: HC\equiv CH+\ddot{O}H_2 \xrightarrow[Hg^{++}]{H^+} H\overset{-}{C}=\underset{\underset{+OH_2}{|}}{CH} \longrightarrow \underset{-\,enol}{CH_2=CHOH} \xleftrightarrow{tautomerises} CH_3-\overset{\overset{\displaystyle O}{||}}{C}-H$$

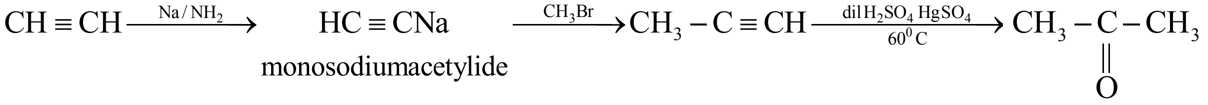
233.



$$\begin{matrix} CHCl_2 \\ | \\ CHCl_2 \end{matrix} \quad , \quad \begin{matrix} CCl_2 \\ || \\ CHCl \end{matrix}$$

Westron , Westrosol

234.



$$CH\equiv CH \xrightarrow{Na/NH_2} \underset{monosodiumacetylide}{HC\equiv CNa} \xrightarrow{CH_3Br} CH_3-C\equiv CH \xrightarrow[60^0\,C]{dil\,H_2SO_4\ HgSO_4} CH_3-\underset{\underset{\displaystyle O}{||}}{C}-CH_3$$

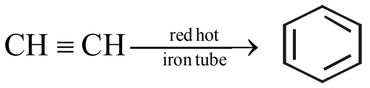
235. Alkanes are saturated hydro carbon non polar, sp^3 hybridised (C–C stronger bonds) don't react with common chemical (acid, alkali, oxidant, reductant) reagents i.e. why these are called paraffins.

236. On cracking lower alkanes / hydrocarbons are formed which have greater fuel efficiency.

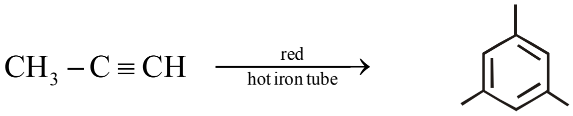
e.g.. $\underset{Dodecane}{C_{12}H_{26}} \xrightarrow[973K]{Ni} \underset{Lower\ b.p.}{C_7H_{16}+C_6H_{10}} + \text{other products}$

237. $HC\equiv CH > CH_3-C\equiv CH > CH_3-CH_2-C\equiv CH > CH_3-C\equiv C-CH_3$

238.



$CH\equiv CH \xrightarrow[iron\ tube]{red\ hot}$

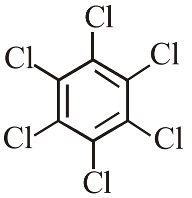


$CH_3-C\equiv CH \xrightarrow[hot\ iron\ tube]{red}$

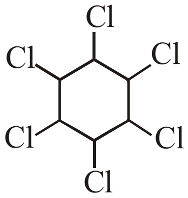
Propyne 1, 3, 5 - trimethyl benzene (mesitylene)

239. $\underset{Sodium\ phenyl\ ethanoate}{C_6H_5CH_2COONa} \xrightarrow[\Delta]{Soda-lime} \underset{Toluene}{C_6H_5-CH_3}$

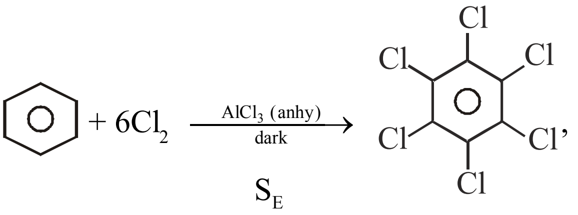
240.

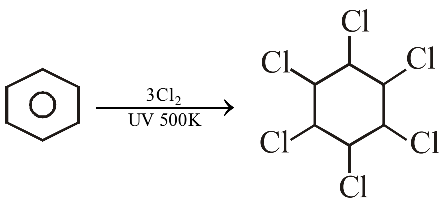


C_6Cl_6
Hexa chlorobenzene ,



$C_6H_6Cl_6$
Benzene hexachloride (gammaxene)





(BHC) Free radical addition reaction

241. Major air pollutants CO, NO_x ans SO_x, particluates, hydro-carbon ($C_x H_y$), PAN (peroxyacylnitrate).

242. Gaseous pollutants, CO, NO_x and SO_x

243. Increased conc. NO_2 leads to formation of HNO_3

$$4NO_2 + 2H_2O + O_2 \longrightarrow 4HNO_3$$

Presence of O_3 also initiates nitric acid formation. It comes down to the surface of the earth in the form of acid rain.

244. i) If present in excess (500–1000 ppm), hydrocarbon causes carcinogenic effect on the lungs.

ii) Due to their photochemical reactions with O_2 and NO_x they form plotochemical oxidants and photo chemical smong damaging living kingdom.

245. TTZ (Taj Trapezium zone) is a defined area of 10, 400 Sq. Km around Tajmahal to protect the monument from environmental pollution. World heritage sites include Taj Mahal, Agra Fort and Fatehpur Sikri.

246. It is a mixture of irritative compounds like NO_2, O_3, PAN (peroxyacylnitrate), aldehydes, ketones, hydrocarbons and CO. It is formed in presence U.V rays from sun and appears as a brown haze during morning and afternoon.

247. NO and freons.

248. H_2SO_4, HNO_3 $\left(Cl_2 + H_2 \xrightarrow{hv} 2Cl\bullet + 2H\bullet \longrightarrow 2HCl\right)$ and HCl

249. Materials cann't be broken down by natural organisms and acts as a source of pollution.

e.g. glass, plastic etc.

250. Un polluted water / clean water has pH of 5.6 counting dissolved CO_2 i.e., H_2CO_3 (carbonic acid).

251. BOD : The amount of oxygen consumed by micro organism in decomposing waste in a sample of sewage water is called BOD (Biochemical oxygen demand).

252. COD (chemical oxygen demand) COD is an indicative measure of amount of oxygen mg/L that can be consumed by reactions in a measured water sample. It is measured by treating the given sample of water with an oxidising agent, $K_2Cr_2O_7$ in presence of dil. H_2SO_4.

COD is preferred to BOD as BOD required at least 5 days to measure.

253. It is formed as a result of photochemical reaction between oxides of nitrogen and hydrocarbon.

254. It is oxidising in nature.

255. Volcanic eruption, forest fire, pollen grains of flowers.

256. Used to destroy unwanted vegetation. e.g. altrazine simazine. [$NaClO_3$, Na_3AsO_3]

257. PSC is a special type of clouds present over Antarctica in winter.

258. Benzene, carbon tetrachloride, trichloroethylene

259. Slowing down the formation of chlorophyll in plants due to presence of SO_2 pollutant is called chlorosis.

260. MIC – methyl isocyanate. (Linked to Bhopal gas tragedy)

◆◆◆

GROUP - B

(SHORT QUESTIONS)

(Two or Three marks each)

Unit - I

1. **Match the following sections.**

	Section - A		**Section - B**
(a)	Sodium hydroxide	(i)	Brittle
(b)	Calcium oxide	(ii)	deliquescent
(c)	gypsum	(iii)	efflorescent
(d)	Gold	(iv)	hygroscopic
(e)	Iodine	(v)	malleable
(f)	Bromine	(vi)	Elastic
		(vii)	Liquid

2. In an experiment, 2.4 g of iron oxide on reduction with hydrogen yield 1.68 g of iron. In another experiment 2.9 g of iron oxide give 2.03 g of iron on reduction with hydrogen. Show that above data illustrate the law of constant proportion.
3. Calculate the mass % of different elements present in sodium sulfate.
4. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1 % oxygen by mass.
5. The vapour density of a mixture containing NO_2 and N_2O_4 is 38.3. Calculate the moles of NO_2 is 100 gram of the mixture.
6. 1 gram of metal, M which has specific heat 0.06 combines with oxygen to form 1.08 g of oxide what is the atomic mass of M ?
7. How many moles of methane are required to produce 22 gram of $CO_2(g)$ after combustion ? Calculate the molar volume of O_2 for such combustion.
8. 50.0 kg of N_2 (g) and 10.0 kg of $H_2(g)$ are mixed to produce $NH_3(g)$. Calculate the $NH_3(g)$ formed. Which one is the limiting reagent ?
9. The density of 3 M solution of NaCl is 1.25 g / ml. Calculate the molality and mole fraction of the solute.
10. $A + B_2 \longrightarrow AB_2$

    Identify the limiting reagent, if any, in the following reaction mixtures.

    (i) 300 atoms of A + 200 molecules of B

    (ii) 2 moles of A + 3 mol of B

    (iii) 100 atoms of A + 100 molecules of B

    (iv) 5 mole A + 2.5 mol B

    (v) 2.5 mol A + 5 mol B

Unit - II

11. A microscope using suitable photons is employed to locate an electron is an atom within a distance of 0.1^0A. Calculate the uncertainly in measuring the velocity.
12. Electro-magnetic radiation of wavelength 242 nm is just sufficient to ionise sodium atom. Calculation I.E. of sodium in kJ / mole.
13. The wavelength of the first line is Balmer series is 656 nm calculate the wavelength of the second line and the limiting line in Balmer series.

14. In hydrogen atom, an electron jumps from 3rd orbit to the 2nd orbit. Calculate the wavelength (λ) of radiation $(h = 6.63 \times 10^{-34}\text{ J.s.})$

15. How much energy is required to ionise a hydrogen atom if electron occupies fifth orbit ? Compare it with its ionisation energy removing $\overline{e}$ from first orbit.

16. The mass of an electron is 9.1×10^{-31} kg . If its KE is 3.0×10^{-25} J , calculate its wavelength.

17. What is the number of photons of light with wavelength 4000 pm that provide 1 J of energy ?

18. 2×10^{8} atoms of carbon are arranged side by side. Calculate the radius of carbon atom. If the length of this arrangement is 2.4 cm.

19. A 25 watt bulb emits monochromatic yellow light of wavelength $0.57\mu m$. Calculate rate of emission of quanta / sec.

20. Calculate the energy associated with the first orbit of He^{+} and calculate its radius.

21. What are the drawbacks of Rutherford's model ?

22. Hydrogen atom contains one electron, but atomic spectrum consists of large number of lines – explain.

23. How does Bohr's atomic model explain the line spectrum of hydrogen ?

24. Electronic energy assigned negative – explain.

25. How does orbit differs from orbital ?

26. Define Borh's orbit in the light of de-Brogli material wave.

27. For macroscopic object wave concept as well as uncertainty principle is insignificant.

28. What is Aufbau principle ?

29. Discuss Pauli's exclusion principle.

30. State Hund's rule of maximum multiplicity.

31. Cr(Z=24) and Cu (Z=29) have exceptional electronic configuration. Explain.

32. Discuss the limitation of Bohr's atomic model.

33. What do you mean by shells, sub-shells and orbitals ?

34. Explain the terms with examples
isotopes, isobars, isotones, isodiaphers.

35. Derive an expression to calculate the radius of H-like atoms.

36. Calculate the number of revolutions of the electron is an orbit per sec & the time taken for one revolution.

37. How many nodes are present in 3-p orbital ?

38. To which orbit the electron in the hydrogen atom will jump after absorbing 1.94×10^{-18} J of energy ?

39. What is photoelectric effect ?

40. A bulb emits light of wave length 4500^{0}A. The bulb is rated as 150 watt and 8% of energy is emitted as light. How may photons are emitted by the bulb per second ?

Unit - III

41. Explain (i) The second $\Delta_i H$ of Na is greater than that of Mg.
(ii) $\Delta_i H_1$ of Mg is more than that of Al.
(iii) C and Pb belong to the same group.

42. Discuss the merits of long form periodic table over short form.

43. Write down the general characteristics of s / p/ d / f block elements.

44. Give reason :

(i) $\Delta_i H_1$ of 'N' is greater than that of oxygen but the $\Delta_i H_2$ is reversed.

(ii) The second ionisation enthalpy of 'Al' is more than expected.

(iii) $Na^+_{(g)}$ is isoelectronic with $Ne_{(g)}$ but they vary in their ionisation enthalpy values.

(iv) $\Delta_i H_1$ of Ga > $\Delta_i H_1$ (Al) but that of Tl > In in the same group.

45. What are metalloids ?

46. What are the defects of long form periodic table ?

47. Discuss the factors affecting $\Delta_i H$ values.

48. How does $\Delta_i H$ values change in the periodic table ?

49. The first and second electron gain enthalpies of Oxygen vary in sign. Give reson.

50. Electron gain enthalpy of F is less than that of Cl, explain.

51. The electron gain enthalpy of noble gases is positive.

52. The $\Delta_{eg}H$ of 'O' is highly negative while that of 'N' is slightly positive.

53. Would you regard $_{30}Zn$, $_{48}Cd$ and $_{80}Hg$ as s or d-block elements ? Give reason.

Unit - IV

54. Why do atoms combine ?

55. How does ionic bond differ from covalent bond ?

56. How do ionic compounds differ from covalent compounds ?

57. Differentiate atoms from ions.

58. Draw the Lewis structures of acetylene, carbonate ion, SO_3.

59. Discuss the factors effective (i) ionic bond formation (ii) covalent bond formation.

60. Explain different bond parameters.

61. Discuss the polar character of covalent bonds.

62. What is dipole moment ? Write down its significance.

63. The inter nuclear distance of HF is 0.92 ^{0}A. Calculate its % of ionic character if the observed dipolemoment is 1.98 D.

64. Discuss the covalent character of ionic bond.

65. How does σ-bond differ from π-bond ?

66. What is resonance ?

67. Write down the significance of resonance energy. Draw the resonating structures of N_2 and CO_3^{2-}.

68. Explain the following .

(i) CO_2 is linear but SO_2 is bent.

(ii) $BeCl_2$ is linear but $SnCl_2$ is V. shaped.

(iii) CH_4 and $\overset{+}{N}H_4$ are isostructural.

(iv) Bond angle of NH_3 differs from H_2O though N and O undergo sp^3 hybridisation.

69. How does pure orbital differ from hybrid orbital ?

70. Show that BCl_3 and PCl_3 have different shape.

71. Privide one example in each case of sp^3d, dsp^2, $d\,sp^3$, $sp^3\,d^2$, $d^2\,sp^3$ hybridisation.

72. How molecular orbitals are formed ?

73. Distinguish between bonding and antibonding molecular orbitals.

74. O_2 is paramagnetic while N_2 is diamagnetic. Explain.

75. Account for the following :
   - (i) Ice floates over water.
   - (ii) Methanol has higher b.p. than methanamine.
   - (iii) p-nitrophenol has higher m.p. than o-nitrophenol.
   - (iv) Water (H_2O) is liquid but H_2S is a gas.
   - (v) Among the halogen hydra acid HF is liquid.

Unit - V

76. Discuss the measurable properties of gases.

77. Write down the postulates of kinetic molecular theory of gases.

78. Find a relationship betweem average kinetic energy and absolute temperature.

79. Write down the kinetic gas equation. Deduce Boyle's law, Charle's law. Avogadro's law, Dalton's law, Graham's law.

80. What is compresibility factor ? Discuss the effect of pressure on deviation / compresibility factor.

81. Distinguish between real gas and ideal gas.

82. Derive equation of state for an ideal gas.

83. What is vapour pressure ? How does it control the boiling and melting point ?

Unit - VI

84. State and derive mathematically the first law if thermodynamics.

85. What are heat capacity and specific heat ? Show that $C_p - C_r = R$.

86. Explain standard enthalpy change of reaction, Find a relationship between the enthalpy change at constant pressure and cnstant volume.

87. The standard heat of formation for $CCl_4(g)$, $H_2O_{(g)}$, $CO_{2(g)}$ and HCl(g) are –25.5, –57.8, –94.1 and –22.1 Kcal / mole respectively calculate $\Delta H(298)$ is kJ and ΔU value for this reaction

$$CCl_4(g) + 2H_2O(g) \longrightarrow CO_{2(g)} + 4\,HCl(g)$$

88. What is heat of neutralisation ? Calculate the amount of heat released when 200 ml of 0.2 M HCl is mixed with 100 ml 0.1 M KOH solution ? 1s the solution acidic ?

89. Calculate the reasonance energy of benzene. Given that ${}^{BE}C-C \Rightarrow 347.3$, C=C $\Rightarrow$ 615 and C–H $\Rightarrow$ 416.2 kJ/mol. Experimentally $\Delta_f\,C_6\,H_6 = -5535.1$ kJ / mol.

90. Explain 3rd law of thermodynamics.

91. (a) What is the basic difference between enthalpy of reaction and enthalpy of formation ?
   (b) Heat of dissociation of CH_3 COOH is 1.9 kJ/mol. How much heat will be evolved when 1 mole of acetic acid is neutralised by NaOH (Aq) ?

92. Discuss enthalpy of atomisation with example.

93. How does ΔG change with temperature ?

Unit - VII

94. State law of mass action. Write down the expression for K_C, K_p and K_x.
95. Derive relationship between K_p, K_C and K_X, when there three forms of equilibrium constants are equal.
96. Mention the suitable conditions of better yield of NH_3 by Haber's process.
97. State law of chemical equilibrium. Relate K_p and K_C for the heterogeneous equilibrium $NH_4HS_{(s)} \rightleftharpoons NH_3(g) + H_2S\ (g)$.
98. What are the characteristics of equilibrium constant ?
99. What are features of chemical equilibrium ?
100. Calculate the equilibrium constant at 1395 K for the following :

$$CO_2 + H_{2(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)},\ \text{if}$$

$$2H_2O_{(g)} \rightleftharpoons 2H_{2(g)} + O_{2(g)}, \qquad K_1 = 2.1 \times 10^{-13}$$

$$2CO_{2(g)} \rightleftharpoons 2CO_{(g)} + O_{2(g)} \qquad K_2 = 1.4 \times 10^{-13}$$

101. Show that the decomposition of PCl_5 is reduced with increase of external pressure at a given temperature.
102. Find out the K_p value, if 1 mole of NH_4HS is subjected to attain equilibrium at 1-atom pressure and decomposition temperature.
103. Explain proton transfer theory what are its merits over Arrherrius theory.
104. Write down Lewis theory of acids and bases.
105. How does Lewis theory differ from Bronsted concepts of acid - base ?
106. Discuss the factors affecting acid strength.
107. What is a buffer solution ? Show that solution of ammonium acetate behaves as buffer solution.
108. How can you determine the pH a buffer solution ?
109. Classify the salts into different categories.
110. Show that ammonium chloride solution is acidic while sodium acetate (Aq) alkaline.
111. pH of ammonia acetate solution depends on K_a and K_b but not on the concentration of the solution.
112. Classify the cations into groups depending on their K_{sp} values of respective sparingly soluble salts.
113. 100 ml of 0.01 M HCl is added to 300 ml of 0.1 M NaOH. The resulting solution is not unto 1L. Calculate the pH of the solution.
114. What is reaction quotient ? How does it affect an equilibrium process / a reversible reaction ?
115. Discuss the effect of temperature on equilibrium constant.
116. What is common ion effect ? Discuss its applications.
117. $H_2S_{(g)}$ is passed to precipitate Cu^{++} in acidic medium while to precipitate Zn^{2+} in alkaline medium, explain.

Unit - VIII

118. Distinguish between oxidation number and valency.
119. Balance the following by O.N. and ion-electron methods.

i) $Mg + H^+ + NO_3^- \longrightarrow Ms^{2+} + N_2O_{(g)} + H_2O$

ii) $Cr_2O_7^{2-} + H^+ + Fe^{2+} \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$

iii) $MnO_4^- + B\overline{r} + H_2O \longrightarrow MnO_2 + BrO_3^- + OH^-$

120. Depict the Galvanic cell in which the reaction :

$$Zn(s) + 2AgNO_3 \longrightarrow Zn(NO_3)_2 + 2Ag(g) \text{ takes place,}$$

Further show :

(i) Which of the electrode is negative

(ii) The carrier of current in the cell

(iii) Individual reaction at each electrode.

(iv) Oxidising and reducing agents in it.

121. Predict the products on electrolysis –

(i) $AgNO_3$ solution with Ag electrodes.

(ii) Same solution using Pt-electrodes

(iii) A dilute solution of H_2SO_4 with Pt - electrodes.

(iv) An aqueous solution of $CuCl_2$ with Pt electrodes.

(v) With Cu electrodes.

122. Answer the following :

(i) Excess of chlorine during chlorination is removed by treating with $SO_2(g)$.

(ii) Write reactions where 'Cl', 'S' and 'P' undergo disproportionation.

(iii) What is the maximum weight of nitric oxide that can be obtained starting only with 10 gram of NH_3 and 20 gram of oxygen.

(iv) Select three metals that can show disproportionation reaction.

(v) $C_6H_5CHO(l) + 2Cu^{2+}(Aq) + 5OH^-$ does not react but C_6H_5CHO responds Tollen's reagent, $[Ag(NH_3)_2]^+$ forming Ag.

(vi) Toluence is oxidised to benzoic acid preferably with alcoholic $KMnO_4$ but not with acidic / alkaline $KMnO_4$.

123. It is economical to take $KMnO_4$ is acidic medium for quantitative analysis - Explain.

124. Discuss the factors governing electrode potential.

Unit - IX

125. How does hydrogen resemble with alkali metals, group 1 ?

126. Hydrogen can be placed along with halogens, explain.

127. How does 'H' differ from alkali metals and halogens ?

128. Can you justify the position of H above carbon ?

129. What is isotopic effect ?

130. How nascent 'H' is prepared ? Mention two of its reducing properties.

131. How is atomic hydrogen prepared ? Is it stable ? Give its action with carbon monoxide.

132. Explain the isomers of di hydrogen.

133. Classify hydrides with examples.

134. How does heavy water differ from ordinary water ?

135. How can you concentrate hydrgen peroxide ? Mention the way it is preserved.

136. Define the volume strength of H_2O_2. Find out the normality of '10 V H_2O_2'.

137. Complete and balance the following reactions .

$$NaHCO_3 + H_2O_2 \longrightarrow$$

$$Mn^{2+} + H_2O_2 + OH^- \longrightarrow$$

$$Fe(CN)_6^{4-} + H_2O_2 + H^+ \longrightarrow$$

138. 'Hydrogen economy is an alternative' ——— explain

139. Hydrogen peroxide can both act as oxidant and reductant, explain with one example in each case.

140. Account for the following :
   (i) H_2O_2 is tested with acidified $K_2Cr_2O_7$.
   (ii) Phosphorous fails to form PH_5.
   (iii) H_2O_2 is non - linear.
   (iv) Water is the commonly prefered solvent.
   (v) Hydrolysis differs from hydration.

Unit - X

141. (a) Why does first element of a group differ from other elements of the group ?
   (b) How Lithium differs from other elements of the group ?

142. Show that Lithium is diagonally related to Mg.

143. Lithium has higher $\Delta_i H$ value but it's a powerful reducing agent, why ?

144. Lithium is stored in wax while Na and K inside kerosine - explain.

145. Alkali metals show characteristic flame colouration - give reason.

146. Why do alkali metals show maximum photoelectric effect ?

147. Give reason :
   (i) Lithium forms monoxide, sodium peroxide while potasium forms superoxide.
   (ii) Basic character of hydroxides increases in moving down the group.
   (iii) LiF is insoluble in water.
   (iv) Alkali metals dissolve in $NH_3(\ell)$ to give conducting blue solution.
   (v) Conducting power of alkali metal chloride (Aq) increases on moving down the group
   (vi) Be and Mg fail to form flame colouration.

148. Discuss the principle of formatioin of caustic soda.

149. How is baking soda prepared ? What is the composition of baking powder ?

150. Discuss the biological functioning of Na and K.

151. Complete the following reactions –

$$Na + NH_3 \longrightarrow$$

$$K + CH_3OH \longrightarrow$$

$$NaHCO_3 \xrightarrow{\Delta}$$

$$Cl_2 + NaOH(hot) \longrightarrow$$

152. Sodium loses its lusture on exposure. Why ?

(or)

On exposure to air solid NaOH becomes a liquid and after sometime it changes to a white powder, is it so ?

153. Write down the principle of manufacture of washing soda, explain why potassium carbonate can't be synthesized by this process ?

154. Group 2 - metals form stable dipositive cation (M^{2+}), give reason.

155. Discuss the nature of carbonates and sulfates of the group-2 ?

156. Explain the anomalous behaviour of Be.

157. Provide equations of the following phenomena :

(i) On exhaling through a tube containing line water milkiness appears.
(ii) The above milkiness (turbidity) dis-appears on continuous exhaling.
(iii) The turbidity re appears on heating the above solution.
(iv) Anhydrous $CaSO_4$ acts as a drying agent.
(v) Magnesium metal burns in air to give white ash which on hydrolysis leaves a pungent odour.

158. 'Be' diagonally related to Al. Explain.

159. Write the biological importance of Mg and Ca.

Unit - XI

160. How does 'B' differ from Al ?

161. Boron is diagonally related to Si. Explain.

162. Give reason :

(i) BCl_3 exists as monomer but Al, Ga and In trihalide exist as dimer.
(ii) Non. traditional bonds are found in diborane.
(iii) Tl exhibits variable oxidation state.
(iv) Cobalt salt responds boraxbead test.
(v) Orthoboric acid (H_3BO_3) is a Lewis monobasic acid.
(vi) BF_3 is a weaker Lewis acid.
(vii) Aluminium becomes passive with conc. HNO_3.
(viii) $AlCl_3$ is not ionic.
(ix) CO_2 is a gas but SiO_2 is a solid.
(x) Carbon compounds are huge in number.

163. Diamond is used as abrassive but graphite as lubricant. Explain.

164. Discuss anomalous character of carbon.

165. How does diamond differ from graphite ?

166. What are fullerens ?

167. **Match the following :**

	I		II
(a)	Dry ice	(i)	Aluminium oxide
(b)	Cryolite	(ii)	Silicon carbide
(c)	Boron nitride	(iii)	Tincal
(d)	Carborundum	(iv)	3 NaCl. $AlCl_3$
(e)	Borax	(v)	Isostructural with graphite
(f)	Corundum	(vi)	Solid carbon dioxide

168. An element (A) is heated with lime in an electric furnance to give another compound (B). 'B' with water leaves a gas (C) which gives white precipitate (D) with ammoniacal $AgNO_3$ (Aq). Identify A to D and provide equation in each step.

169. What are silicones ?

Unit - XII

170. Discuss the formation shape and stability of (i) carbocation (ii) Carbanion (iii) free radicals.

171 What is inductive effect ? How does it affect acid and basic nature of different organic compounds.

172. Explain +R and –R effect with one example in each case.

173. What is hyper conjugation ? Explain the stability of alkenes.

174. How can you detect the presence N in aniline ?

175. Discuss various types of organic reactions with example.

176. What are different type s of attacking reagents in organic reactions ?

177. A sample of 0.5 gram of an organic compound was treated according to Kjeldahal's method. The ammonia evolves was absorbed in 50 ml of 0.5 M H_2SO_4. The residual acid required 60 ml of 0.5 M solution of NaOH for neutralisation. Fid the percentage composition of nitrogen in the compound.

Unit – XIII

178. How does halogenation of methane take place ?

179. Starting from suitable alkane discuss the preparation of
(i) CH_3OH (ii) HCHO (iii) t-butyl alcohol

180. How does HBr add to propene ?

181. Discuss the acidic nature of alkyne.

182. How can you prepare (i) acetaldehyde from acetylene
(ii) acetylene from ethylene
(iii) acetylene from benzene

183. What is Fridel Craft's reaction ?

184. Starting from benzene prepare (i) o-Chloro nitrobenzene
(ii) o-Nitrochlorobenzene

185. What are activating and deactivating goups ?

186. Show that benzoic acid on nitration forms m-nitro benzoic acid.

Unit - XIV

187. Why CO is more poisonous / dangerous than CO_2 ?

188. What are the different sources of air pollution ?

189. Write down the chief sources of water pollution.

190. Mention six detrimental effects of soil pollution.

191. Explain : (i) Ozone hole is being observed only in Antarctica.
(ii) Taj Mohal is losing its lustre day by day.

192. What is CFC ? How does it destroy ozone ?

193. What would have happened, if green house gases were totally missing in the earth's atmosphere ?

194. What is green chemistry ? How will it help to decrease environmental pollution ?

195. What do you mean by incineration ? How is it helpful for waste management ?

◆◆◆

GROUP - B (ANSWERS)

(Two or Three marks each)

Unit - I

1.

(a)	Sodium hydroxide	(ii)	deliquescent
(b)	Calcium oxide	(iv)	hygroscopic
(c)	gypsum	(iii)	efflorescent
(d)	Gold	(v)	malleable
(e)	Iodine	(i)	Brittle
(f)	Bromine	(vii)	Liquid

2. Sample I mass of iron 1.68 g

mass of oxygen = 2.4 – 1.68 = 0.72 Fe : 0 = 2 : 1

Sample II mass of iron = 2.03

mass of oxygen = 2.90 – 2.03 = 0.87 Fe : 0 = 2 : 1

In the fixed ratio.

3. Na_2SO_4 = 142 gm/mole

$$\% \text{ Na} = \frac{46}{142} \times 100 = 32.4\%$$

$$\% \text{ S} = \frac{32}{142} \times 100 = 22.5\%$$

$$\% \text{ O} = \frac{64}{142} \times 100 = 45.1\%$$

4. Fe % 69.9 Relative no. of atoms $= \frac{69.9}{56} = 1.25$ simplest at. ratio. $\frac{1.25}{1.25} = 1$

0 % 30.1 Relative no. of atoms $= \frac{30.1}{16} = 1.88$ $\frac{1.88}{1.25} = 1.5$

Fe : 0 = 2 : 3 Empirical formula = Fe_2O_3

5. Mol. mass of mixture $= 2 \times \text{V.D} = 2 \times 38.3 = 76.6$

moles of NO_2 + moles N_2O_4 = moles of the mixture

$$\frac{x}{46} + \frac{100 - x}{92} = \frac{100}{76.6} \Rightarrow x = 20.1\text{g}$$

$$\therefore \text{ moles of } NO_2 = \frac{20.1}{46} = 0.437$$

6. Dulong-Petits law, atomic mass $\times$ Sp. heat $\simeq$ 6.4

$$\text{Appx. atomic mass} = \frac{6.4}{0.06} = 106.6$$

0.08 gram oxy. combine with 1.0 gram metal, so equivalent mass of metal $= \frac{1}{0.08} \times 8 = 100$

$$\text{Valency} = \frac{\text{atomic mass}}{\text{equivalent mass}} = \frac{106.6}{100} = 1$$

Exact atomic mass $= 100 \times 1 = 100$

7. $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

1 mole 1 mole, 22 gram CO_2 = 0.5 mole

moles of methane required = 0.5 mole

1 mole requires 2×22.4 L oxygen at NTP

0.5 mole requires $2\times22.4\times0.5 = 22.4$ L (oxygen).

8. $\underset{28\,kg}{N_2} + \underset{6\,kg}{3H_2} \longrightarrow \underset{34\,kg}{2NH_3}$ (From stoichiometry)

50 kg $\frac{6}{28}\times50 = 10.7$ kg . H_2 given 10 kg H_2 (limiting reagent)

Amount of NH_3 formed

6 kg give 34 kg

10 kg give $\frac{34}{6}\times10 = 56.6$ kg

9. 3M solution $\Rightarrow$ 3 moles of solute in 1000 ml of solution

weight of the solution = $1000\,ml\times1.25\,g/ml = 1250\,g$

wt of solute = 3 mol of NaCl = $3\times58.5 = 175.5\,g$

wt of solvent = 1250 – 175.5 = 1074.5

$$\text{molality} = 3\times\frac{1000}{1074.5} = 2.79\,m$$

$$\text{mole fraction of the solute} = \frac{3}{3+(1074.5/18)} = \frac{3}{3+59.69} = 0.048.$$

10. $A + B_2 \longrightarrow AB_2$

1 mole atom A + 1 mole of B_2 i.e. 2 moles of atoms B $\longrightarrow$ 1 mole of molecule AB_2.

i) 300 atoms of A + 600 atoms of B (= 300 molecules B).
given 200 moleculs of B i.e B = limiting reagent.

ii) 2 moles A + 2 moles B, but given 3 moles
i.e., A will completely react = limiting reagent.

iii) 100 atoms A + 100 molecule (200 atoms)
In stoichiometric ratio. (No limiting agent)

v) 2.5 moles A + 2.5 moles B (given 5 moles)
B is in excess, A = limiting agent.

Unit - II

11. $\Delta x.\ \Delta p = \frac{h}{4\pi} \Rightarrow \Delta x.\ \Delta V = \frac{h}{4\pi m}$

$$\text{i..e.,}\quad \Delta V = \frac{h}{4\pi\Delta x\,m} = \frac{6.626\times10^{-34}\,J.s\left[\left(\cancel{Kg}\ m^{\cancel{2}}\ s^{-2}\right).s\right]}{4\times3.14\times0.1\times10^{-10}\,\cancel{m}\times9.11\times10^{-31}\,\cancel{Kg}}$$

$1J = 1\,Kg\,m^2\,s^{-2}$

$= 0.579\times10^7\,m/s$

$\therefore$ Velocity $= 5.79\times10^6\,m/s$

12. $E = N h \nu = N h \frac{C}{\lambda}$

$$= \frac{\left(6.02\times10^{23}\ \text{mol}^{-1}\right)\left(6.626\times10^{-34}\ \text{J}\cancel{\text{s}}\right)\times3\times10^{8}\ \cancel{\text{m}}\ \cancel{\text{sec}^{-1}}}{242\times10^{-9}\ \cancel{\text{m}}}$$

$$= 0.4945\times10^{6}\ \text{J mol}^{-1}$$

$$= 4.945\times10^{5}\ \text{J mol}^{-1}$$

13. According to Rydberg's formula –

$$\bar{\gamma} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

For Balmer first line, $n_1 = 2$ and $n_2 = 3$

$$\frac{1}{656} = R\left[\frac{1}{2^2} - \frac{1}{3^2}\right] = \frac{R.5}{36} \qquad \text{............. (i)}$$

For second line $\Rightarrow n_1 = 2$ and $n_2 = 4$

$$\frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{4^2}\right] = \frac{3R}{16} \qquad \text{............ (ii)}$$

Dividing (i) by (ii), we get

$$\frac{\lambda}{656} = \frac{5}{36}\times\frac{16}{3} \qquad \lambda = 485.9\ \text{nm}$$

For the limiting line, $n_1 = 2$ and $n_2 = \infty$

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{\infty}\right) = {}^{R}\!/_{4} \qquad \text{............ (iii)}$$

Dividing (i) by (iii)

$$\frac{\lambda}{656} = \frac{5\cancel{R}}{36}\times\frac{4}{\cancel{R}} \quad \Rightarrow \quad \lambda = 364.4\,\text{nm}$$

14. $E_n = -\frac{21.8\times10^{-19}}{n^2}\ \text{J/atom}$

$$\Delta E = E_3 - E_2 = 21.8\times10^{-19}\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$= 21.8\times10^{-19}\left(\frac{5}{36}\right)$$

$$= 3.03\times10^{-19}\ \text{J}$$

$$\Delta E = h\frac{c}{\lambda} = 3.03\times10^{-19}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{6.63\times10^{-34}\,\text{Js}\times3\times10^{8}\ \text{ms}^{-1}}{3.03\times10^{-19}\ \text{J}}$$

$\therefore \quad \lambda = 6.56\times10^{-7}\ \text{m}$

15. $n_1 = 5$ & $n_2 = \infty$

$$\Delta E = E_\infty - E_5 = -21.8\times10^{-19}\left(\frac{1}{\infty}-\frac{1}{25}\right)$$

$$= 0.872\times10^{-19} = 8.72\times10^{-20}\text{ J}$$

For $\Delta_i H \Rightarrow n_1 = 1 \;\&\; n_2 = \infty$

$$\Delta E = 21.8\times10^{-19}\text{ J} = 2.18\times10^{-18}\text{ J}$$

16. Since $KE = \frac{1}{2}mv^2$

$$\nu = \sqrt{\frac{2KE}{m}} = \left(\frac{2\times3\times10^{-25}\text{ kg m}^2\text{ s}^{-2}}{9.1\times10^{-31}\text{ kg}}\right)^{\frac{1}{2}} = 812\text{ m/s}$$

$$\lambda = \frac{h}{m.v} = \frac{6.626\times10^{-34}\left(\text{J.s} = \cancel{\text{Kg}}\text{ m}^2\ \cancel{\text{s}^{-1}}\right)}{9.1\times10^{-31}\ \cancel{\text{kg}}\times812\text{ m}/\cancel{\text{sec}}}$$

$$= 8967\times10^{-10}\text{ m} = 896.7\text{ nm}$$

17. $E = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\text{ J.s.}\times3\times10^{8}\text{ ms}^{-1}}{4000\times10^{-12}\text{ m}} = x\text{J}$ (1-photon) (or) $E = \frac{Nhc}{\lambda}$

x J = 1 photon

$N\text{ (no. of photons)} = \frac{E.\lambda}{hc}$

$$1\text{ J} = \frac{1}{x}\text{photon} = \frac{4000\times10^{-12}}{6.626\times3\times10^{-26}}$$

$$= 201.2\times10^{14} = 2.012\times10^{16}\text{ photons.}$$

18. oooo........

$$2\times10^{8}\times2r = 2.4\times10^{-2}\text{ m}$$

$$r = \frac{2.4\times10^{-2}}{4\times10^{8}}\text{ m} = 0.6\times10^{-10} = 6\times10^{-11}\text{ m}$$

19. 25 watt = 25 J/sec. $\lambda = 0.57\times10^{-6}\text{ m}$

$$E = N\frac{hc}{\lambda} \quad N = \frac{E\times\lambda}{h\times c} = \frac{25\text{ J}\times0.57\times10^{-6}\text{ m}}{6.63\times10^{-34}\text{ J.s}\times3\times10^{8}\text{ m s}^{-1}}$$

$$= 0.718\times10^{20}\text{ quanta/sec}$$

$$\text{rate} = 7.18\times10^{19}\text{ quanta/sec of emission.}$$

20. $E_n = -\frac{21.8\times10^{-19}z^2}{n^2}\text{ J}$

$$= -\frac{21.8\times10^{-19}\times4}{1^2}\text{ J} = -87.2\times10^{-19}\text{ J}$$

$$r_n = \frac{0.529\times n^2}{Z} = \frac{0.529\times1^2}{2} = 0.2645\ {}^{0}\text{A}$$

21. Draw books of Rutherford model :

(i) It fails to explain line spectrum of hydrogen.

(ii) According to Maxwell, whenever an electric charged body is subjected to acceleration, it emits radiation and loses energy. Electron is a charged particle and moves around the nucleus (+vely charged) it must have to lose / radiate energy. On subsequent loss of energy it will fall into nucleus. But it never happens.

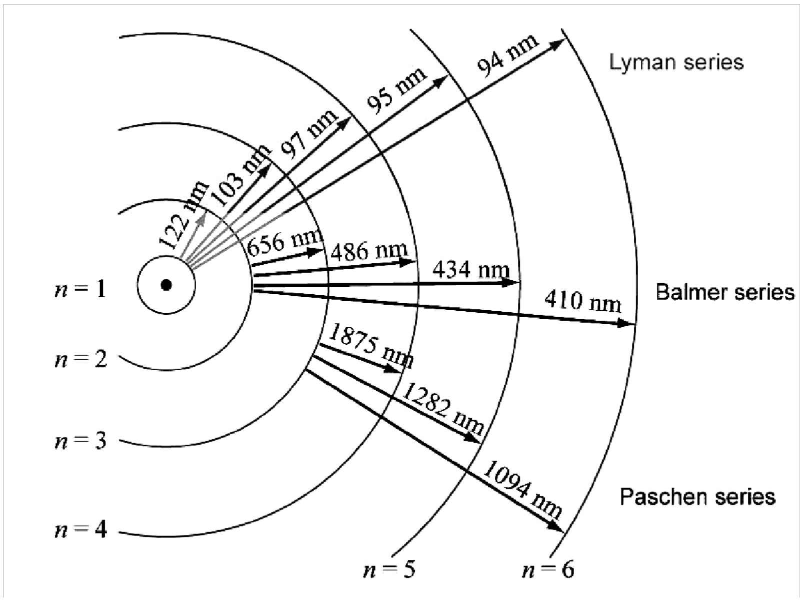
22. Although an hydrogen atom contains one electron, yet we observe large number of spectral lines as the sample of hydrogen gas contains a large number of atooms (e.g. 1 mole i.e. 2 gram hydrogen consist of 2 N no. of atoms).In the discharge tube under influence of high potential of electricity, the electrons in different hydrogen atom absorb different energy and excite to different energy levels which on retum emits different energy corresponding to different lines.

e.g. electrons in some atoms are excited to second energy level and other may be prompted to third energy level (orbit) and so on.

23. **Spectra of 'H' like atom :**

When electron jumps from higher energy level (n_2) to lower energy level (n_1), it emits energy and the spectra so obtained in known as emission spectra. The emission spectra for 'H' like atom (one electron system is as follows)

Name of the series	$\mathbf{n_1}$	$\mathbf{n_2}$	**Region**
Lyman series	1	2,3,4......	U.V
Balmer series	2	3,4,5	visible
Paschen series	3	4,5,6.......	Near IR
Bracket series	4	5,,6,7......	I. R.
P fund sries	5	6, 7..........	Far I.R



24. $E_n = -\frac{1312}{n^2}.Z^2 \text{ kJ/mol},$ when electron is at ∞ distance from the nucleus i.e., $n = \infty$, there is no force of attraction, where its energy is taken is zero. As it jumps to orbits nearer to nucleus it experiences a force of attraction as a result some energy is given out i.e., conventionaly assigned negative. More –ve is the energy value more is the attraction between nucleus and the rotating electron i.e., when $\overline{e}$ moves from $n = \infty$ to $n = 1$ (K. shel l) there is the maximum release of energy i.e., $-1312\,\text{kJ}$ per mole of $e^{-}(s)$ for mole of Hydrogen.

i.e., $-21.8\times10^{-19}\ \text{J}/\text{H}-\text{atom}$

Further as the energy of $n=\infty$, is zero, so the levels below 'n' will be less than zero. i.e., –ve. The Energy required to move the $\overline{e}$ from $n=1$ to $n=\infty$ is called ionisation energy which is positive.

$\Delta_i H$ of hydrogen is + 1312 kJ/mol.

25. **Difference between Orbit and Orbital**

Orbit	**Orbital**
1. It is two dimensional	1. It is three dimensional
2. It is signified by principal Q.N	2. It is signified magnetic Q.N.
3. It indicates the principal shell where electrons exist.	3. It indicates the shape of the region where there is maximum probabality of finding the electrons.

26. $mvr=\dfrac{nh}{2\pi}\ (\text{Bohr})$

$$\Rightarrow\quad 2\pi r=\frac{nh}{mv}=\frac{nh}{p}$$

$\Rightarrow\ \boxed{2\pi r=n\lambda}$, Thus, Bohr's orbits are those where its circumference is whole number (n=1, 2, 3......) muptiple of the wave length (λ) i.e. all material waves must be in phase during motion.

27. de-Broglie wave equation $\Rightarrow \lambda=\dfrac{h}{mv}$, i.e., for microscopic particle mass is much less, so, λ is significant i.e wave length measurable e.g. for an electron of $m=9.1\times10^{-31}\ \text{kg}$

$$\lambda=\frac{6.626\times10^{-34}\,(\text{J.s})}{9.1\times10^{-31}\ \text{Kg}\times(3\times10^{8})\,\text{m/s}}=\frac{6.626\times10^{-34}\,\text{Kg}\,\text{m}^2\,\text{s}^{-2}.\text{s}}{9.1\times3\times10^{-23}\ \text{kgm}\,\text{s}^{-1}}$$

$$=0.2427\times10^{-11}\text{m}\ =\ 2.427\,\text{pm}(\text{U.V.Zone})$$

But an object of mass 0.1 kg moving with a velocity of 10m/s, the λ will be =

$$\frac{6.626\times10^{-34}\,\text{Js}}{0.1\text{kg}\times10\text{m}\sec^{-1}}=6.626\times10^{-34}\ \text{m}\ \text{(insignificant)}$$

28. The electrons are arrange in different subshells of an atom in order of increasing energy. Such building up of electrons in an atom is based on Aufbau principle. i.e., e^- first enters to subshell of lower energy than to next higher one. The arrangement of sub shells in order of energy is given by (n+*l*) rule :

i) subshell with lower $(n+\ell)$ value possess lower energy.

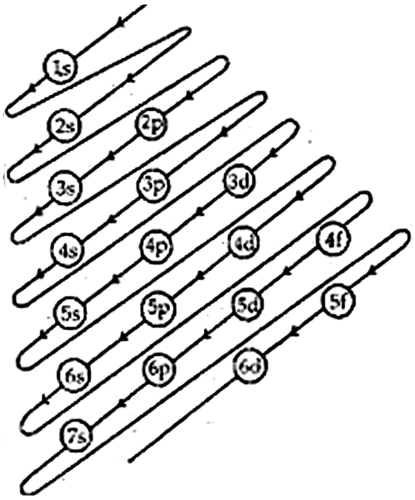
$$\underset{n+\ell=1+0}{E_{1s}}\quad<\quad\underset{n+\ell=2+0}{E_{2s}}$$

ii) orbitals / subshells with identical $(n+\ell)$ value possess lower energy for lower 'n' value.

e.g. 2p, 3s

$n+\ell$ 2 + 1 3+0, Thus $(E_{2p}<E_{3s})$, electrons enter to 3s after filling the 2p (i.e, 2p6).

The order of filling orbitals : in ground state



Limitation

(i) $E_{4f} < E_{5d}$, but the 57th $\overline{e}$ of La (Z = 57) enter to 5d instead of 4f. Similarly $E_{5f} < E_{6d}$ but 89th $\overline{e}$ of Ac (Z=89) enters to 6d instead of 5f (with lower energy)

ii) Cr (Z = 24), the configuration is $4s^1\ 3d^5$ instead of expected configuration $4s^2\ 3d^4$

Cu (Z = 29), the configuration is $4s^1\ 3d^{10}$ against expected arrangement of $4s^2\ 3d^9$

29. Pauli's exclusion principle.

According to Pauli "No two electrons of an atom will have four identical sets of quantum numbers".

Explanation. **When we write $1s^2$,**

First $\overline{e}$ n = 1 Second $\overline{e}$ n = 1

$\ell = 0$ $\ell = 0$

m = 0 m = 0

$s = +\frac{1}{2}$ $s = -\frac{1}{2}$ n, ℓ, m values are identical, but 's' value differs.

When we represent in box form,

$\boxed{\downarrow\uparrow}$ $\boxed{\uparrow\uparrow}$ or $\boxed{\downarrow\downarrow}$

Correct incorrect

30. **Hund's Rule**

According to Hund, "the pairing of electrons starts only when all the orbitals are being filled with electron of same spin or in other words there will be a maximum number of unpair $\overline{e}$ with same spin. For example, when we represent the electronic configuration of nitrogen in boxwise, we have two probabilities, they are :

1 s 2s $2p_x$ 2py 2pz

N (z = 7) $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$

or

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

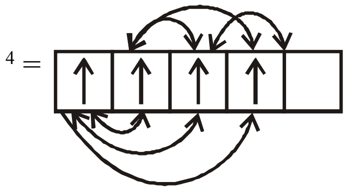
The second one is correct according to Hund's Rule.

31. The exceptional configuration of ${}_{24}Cr$ i.e., d^5 instead of d^4 and Cu(29) i.e., d^{10} instead of d^9 can be explained considering –

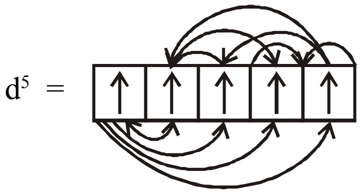
(i) Symmetrical distribution electrons: It is well known that symmetry leads to stability. The half filled, d^5 and completely filled, d^{10} subshells have symmetrical distribution of electrons in them & are therefre more stable. By such their shielding of one another is relatively small & the electrons are more strongly attracted by the nucleus.

(ii) Exchange energy : The $\bar{e}$ with same spin exchange their positions among the degenerate d-orbitals releasing energy (–ve) called excharge energy.

Further pairing energy is the energy expensed (+ve) in pairing the $\bar{e}$ in the orbitals.



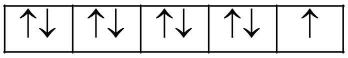
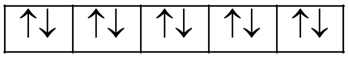
There are 6 possible exchanges,
lowering the energy 6E (say)
Exchange energy = –6E



10 possible exchanges, lowering the
energy 10 E (say)
Exchange energy = – 10 E

Greater is the lowering of energy greater is the stability.

Hence Cr (Z = 24), $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

d^9	d^{10}
	
Exchange of five $\bar{e}$ with $+\frac{1}{2}$ spin	Exchange of five $\bar{e}$ with $+\frac{1}{2}$ spin
= – 10 E	= – 10 E
Exchange of four $\bar{e}$ with $-\frac{1}{2}$ spin	Exchange of five $\bar{e}$ with $-\frac{1}{2}$ spin
= – 6 E	= – 10 E
Total exchange energy = – 16 E	Total exchange energy = – 20 E
Due to formation of four pairing effect,	Pairing energy = + 5 P
the pairing energy = + 4p	(five pairs of $\bar{e}$ s)
Total stabilisation energy = –16 E + 4P	Total stabilisation energy = (–20E + 5P)

Seen that (–20E + 5P) is more negative than (–16 E + 4P). So d^{10} configuration is more stable than d^9.

Thus, Cu(29) $\Rightarrow$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

32. Limitations of Bohr's model :

Draw backs :

1. Bohr's model is only applicable to hydrogen like atm i..e, it could explain the spectral behaviour of only one electron system like H, He^+, Li^{2+} etc.
2. The quantisation of angular momentum is purely arbitary one.
3. It could not explain "Zeeman's effect", "The splitting of a spectral line in to closely spaced cluster of spectral lines under the influence of an external magnetic field is known as Zeeman's effect".
4. It could not explain the "Stark effect".
   " The splitting of a single spectral line into closely spaced cluster of spectral lines under the influence of an external electric field is called Stark effect".

33. **Shell** represents the main energy level occupied by the $\overline{e}$ given by principal quantum number, 'n'. e.g n = 1, K-shell (1st orbit), n = 2, L-shell (2nd orbit) and so on.

Subshell represents the sub energy level occupied by the $\overline{e}$ given by azimuthal quantum number (ℓ).

For a given 'n' value, the subshells are with 'ℓ' values 0 to (n–1)

i.e., n = 1 (K – shell) $\ell = 0$ (s – sub shell)

n = 2 (L–shell) $\ell = 0, 1$ (s and p - sub shell)

Thus, sub shells corresponding to $\ell = 0, 1, 2$ and 3 are represented by s, p, d and f - respectively.

orbital represents the orientation when a subshell takes place under the influence of an external magnetic field. It is given by the magnetic quantum number (m) with values, $-\ell$ to $+\ell$

i.e., for $\ell = 0$, m = 0 ; $\ell = 1$, m = –1, 0, +1 (three orbitals p_x, p_y & p_z)

$\ell = 2$ (d), m = –2, –1, 0, +1, +2 (five orbitals d_{xy}, d_{yz} d_{zx} $d_{x^2-y^2}$ and d_{z^2}). Similarly there are seven f-orbitals (–3 to +3) $f_{y(3x^2-y^2)}$, $f_{2(x^2-y^2)}$, f_{yz^2}, f_{z^3}, f_{xz^2}, f_{xyz}, $f_{x(x^2-3y^2)}$

34. **Isotopes :** Atomic species with same atomic number but different mass numbers (A) i.e., there are same atom with different number of neutrons in the atom.

e.g.	^1_1H (protium),	^2_1H (deutterium),	^3_1H (Tritium)
no. of proton	1	1	1
no. of neutron	0	1	2

Among there three Isotopes of Hydrogen ^3_1H (Tritium) is radioactive and emanates β – rays , other two isotropes are non-radioactive and stable.

Isobars : Species with different atomic number (no. of protons) i.e., different elements with different number of neutrons but having same mass number.
(no of protons + no. of neutrons)

e.g.	$^{14}_7N$ and	$^{14}_6C^*$	
Z	= 7	6	
no. of neutrons	= 7	8	(Here C^{14} is radioactive)

Isotones : Atomic species with same number of neutrons in the nucleus.

e.g. $^{16}_8O$, $^{14}_6C$

n = 16 – 8 = 8, n = 14 – 6 = 8

Isodiapher : These are atomic species with identical value of (n–z) i.e, neutron and proton difference. This is obtained on disintegratio of 'α' rays from a radioactive species.

e.g. $^{226}_{88}Ra \longrightarrow ^{222}_{86}Rn + ^4_2He$

Here $^{226}_{88}Ra$ and $^{222}_{86}Rn$ are isodiaphers.

Z = 88	Z = 86
n = 138	n = 136

Difference of neutron and proton or (A-2Z) value

50	50

35. Radius of the orbit (r).

We know that force of attraction on the electron by nucleus = Centripetal force acting on the electron

i.e., $\frac{Ze^2}{r^{\cancel{2}}} = \frac{mv^2}{\cancel{r}}$ $\quad$ $Z = 1$ for H, $= 2$ for He^+ and $= 3$, Li^{2+}

Bohr's postulate : $mvr = \frac{nh}{2\pi}$

$$v = \frac{nh}{2\pi\, mr}$$

Substituting in the first equation

$$\frac{Ze^2}{\cancel{r}} = \cancel{m} \times \frac{n^2h^2}{4\pi^2 m^{\cancel{2}} r^{\cancel{2}}}$$

$$\text{Now } r = \frac{n^2h^2}{4\pi^2\, m\, Ze^2}$$

1st Bohr's radius $r_0 = 0.529\ {}^0A$

$$r_n\left(H - \text{like atom}\right) = \frac{0.529\, n^2}{Z}\ {}^0A$$

36. Number of revolutions / sec $= \frac{\text{velocity of the electron}}{\text{circumference of the orbit}}$

$$\begin{aligned} mvr = \frac{nh}{2\pi} \qquad & = \frac{v}{2\pi r} \\ v = \frac{nh}{2\pi mr} \qquad & = \frac{nh}{2\pi mr} \times \frac{1}{2\pi r} \\ & = \frac{nh}{4\pi^2 mr^2} \end{aligned}$$

$\frac{nh}{4\pi^2 mr^2}$ revolution in one second

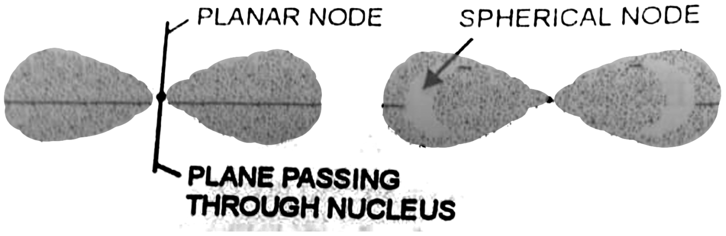
Time taken for one revolution $= \frac{4\pi^2 mr^2}{nh}$

37. Number of planar nodes = 1

Here $n = 3$, $\ell = 1$. No of radial / spherical nodes

$= n - \ell - 1 = 3 - 1 - 1 = 1$

Total number of nodes in 3p - orbital = $(n-1) = 3-1 = 2$



38. $E_n = \dfrac{-21.8\times10^{-19}}{n^2}\,J/atom$

for $n=1, \quad E_1 = -21.8\times10^{-19}\,J$ (existing energy)

After absorbing energy, $\overline{e}$ possesses an energy of

$$= \quad -21.8\times10^{-19}+1.94\times10^{-18}$$

$$= \quad -21.8\times10^{-19}+19.4\times10^{-19}$$

$$= \quad -2.4\times10^{-19}$$

Thus, $-\dfrac{21.8\times10^{-19}}{n^2} = -2.4\times10^{-19}$

$n^2 = \dfrac{21.8}{2.4} = 9$ i..e., $n = 3$

$\therefore$ The electron will jump to the third orbit.

39. Photoelectric effect

(i) The electrons are ejected from the metal surface as soon as the beam of light strikes the surfae.

(ii) The number of $\overline{e}(s)$ ejected is proportional to the intensity or brightness of the light.

(iii) The minimum frequency needed to eject the electron is called threshold frequency (ν_0) below which no photoelectric effect.

(iv) The K.E. of the $(\overline{e})$ increases with the frequency of the light used.

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

40. 150 watt = 150 Joul per sec.

Light emitted is $150\times\dfrac{80}{100} = 12.0\,J$

$E = nh\nu$

$$n = \frac{E\times\lambda}{h\times c} = \frac{12\cancel{J}\times4500\times10^{-10}\,\cancel{m}}{\left(6.626\times10^{-34}\,\cancel{Js}\right)\left(3\times10^{8}\,\cancel{m}\,\cancel{s^{-1}}\right)}$$

$= 2.717\times10^{19}$ number of photons.

Unit - III

41. (i) Na^+ $1s^2\,2s^2\,2p^6$. Stable (octet) configuration, expulsion of $\overline{e}$ involves very high energy while from Mg^+ ($1s^2\,2s^2\,2p^6\,3s^1$), expulron of $\overline{e}$ will be easier & it occupies stable configuration $\left(Mg^{2+} - 1s^2\,2s^2\,2p^6\right)$ on removed of $\overline{e}$.

(ii) ${}_{12}Mg$, $1s^2\,2s^2\,2p^6\,3s^2$ ${}_{13}Al$, $1s^2\,2p^6\,3s^2\,3p^1$

Δ_iH_1 is less for Al as it involves removal of 3p $\overline{e}$, but is Δ_iH_1 of Mg involves removal of $\overline{e}$ from much closer / penetrated 3s electron, requiring more energy against the attractive force between nucleus & 3s- electron.

(iii) $_{6}C = 1s^2\,2s^2\,2p^6$

$_{82}Pb = 1s^2\,2s^2\,2p^6\;\;3s^2\,3p^6\,3d^{10}\;\;4s^2 4p^6\;\;4d^{10}\;\;4f^{14}\;\;5s^2\,5p^6\,5d^{10}\;\;6s^2\,6p^2$. The general valence shell configuration of group - 14 is $ns^2\,np^2$. Pb belongs to 6th period (n=6) of this group.

42. Merits of long form periodic table :

(i) **Atomic number :** The elements of long form periodic table are arranged in order of increasing atomic number. This eliminates the anomaly of atomic masses of e.g. $_{18}Ar$ (39.9)$_{19}K$ (39.1) ; $_{27}Co$ (58.9)$_{28}Ni$ (58.7) ; $_{52}Te$ (127.6) $_{53}I$ (126.9).

(ii) **Placement of Isotopes :** Isotopes of an elements need no separate place as they have same atomic number.

e.g. $^{1}_{1}H$, $^{2}_{1}H$ and $^{3}_{1}H$. mass differs but placed in one place (Z=1)

(iii) **Demarcation** of metals and non metals.

The elements left to the zigzag line are metals, right to this line are non-metals, nearer to the line are metalloids.

(iv) **Reproduceability :** As the arrangement of elements based on electronic configuration (Aufbau principle), the properties of s, p, d and f block element can be specified and easily reproduced.

43. (a) General characteristics of s-block elements

(alkali metals - group -1, alkaline earth metals group - 2)

(i) These are soft metals with low m.p. and b.p.

(ii) They have low $\Delta_i H$ value and hence more electro positive mostly forming electrovalent compounds with + 1 (alkalimetals) and +2 (alkaline earth metals) oxidation states.

(iii) These are highly reactive metals and the reactivity increases on moving down the group.

(iv) Most of the elements impart characteristic colours to the flame.

(v) These have strong reducing character & they have more –ve reduction potatial value. Among these Lithium behaves as a powerful reducing agent.

(b) General characteristics of 'p' block elements.

(i) General electronic configuration $ns^2\,np^{1-6}$ i.e., elements group 13 to 18.

(ii) p-block consists of metals, non-metals and metalloids. The metallic character increases with decrease of $\Delta_i H$ values.

(iii) These elements mostly form covalent compounds.

(iv) As the elements are more electro negative can act as better oxidising agent. The oxidising character increases from left to right in a period & reducing character increases from top to bottom in a group.

(c) General characteristics of d-block elements.

(i) The general electronic configuration of d-block elements is $(n-1)\,s^2\,p^6\,d^{1-10}\,ns^{1\text{ or }2}$ where n = 4, 5, 6 or 7.

(ii) These are hard malleable and ductile metals with high melting and boiling points.

(iii) They are good conductors of heat and electricity.

(iv) They transi their properties between more electropositive, s-block and more electronegative p-block elements.

(v) They exhibit variable oxidation states.

(vi) Their compounds are generally coloured and paramagnetic.

(vii) Most of them (V, Cr, Mn, Fe, Pd, Ni, Pt.....) used as catalysts.

(viii) They form a number of complexes, alloys.

(ix) Zn, Cd, Hg with $(n-1)\ d^{10}\ ns^2$ configuration, do not show most of the properties of transition elements.

(d) General characteristics of f-block elements.

(i) General outer shell configuration : $(n-2)\ f^{0-14}\ (n-1)\ d^{0 \text{ or } 1}\ ns^2$

(ii) They are heavy metals.

(iii) They have generally high melting and boiling points.

(iv) They show variable oxidation states.

(v) High tendency to form complexes

(vi) Most of the actinoids are radioactive.

44. (i) $_7N = 1s^2\ 2s^2\ 2p^3$ $\qquad$ $_8O = 1s^2\ 2s^2\ 2p^4$

$\overset{+}{N} = 1s^2\ 2s^2\ 2p^2$ $\qquad$ $\overset{+}{O} = 1s^2\ 2s^2\ 2p^3$

Due stable half filled coniguration of 'N' its Δ_iH_1 is very high exceptionally. $\overset{+}{O}$ is isoelectronic with 'N' having half filled configuration (p^3). That is way Δ_iH_2 i.e., removal of $\overline{e}$ from $\overset{+}{O}(g)$ needs higher energy, further there is increase of proton number.

(ii) $_{13}Al : 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Δ_iH_1 involves removal of $\overline{e}$ from less penetrated 3p, but the 2nd $\overline{e}$ Δ_iH value involves removal of $\overline{e}$ from much penetrated 3s orbital. So extraction of 3s $\overline{e}$ ivolves exceptionally high energy.

(iii) $_{11}Na^+_{(g)}$ $1s^2\ 2s^2\ 2p^6 = {}_{10}Ne$ (isoelectronic) , $s^2\ p^6$ is much stable configuration, further Na-nucleus contains 11 proton. Thus, due to increase of nuclear change, the Δ_iH value increases. Expulsion of –vely changed $\overline{e}$ from cation (Na^+) is difficult i.e., involves higher energy.

So. Δ_iH_2 of (Na) > Δ_iH_1 (Ne)

(iv) $_{13}Al\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

$_{31}Ga\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$

Ga has higher Δ_iH value than Al due to presence of 3d-electrons which have less shielding effect. This increases the effective nuclear change (Z^*) increasing the greater nuclear effect on the removing $\overline{e}$ increasing the Δ_iH value.

$_{49}In\ \ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^1$

$_{81}Tl\ \ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 4f^{14}\ 5s^2\ 5p^6\ 5d^{10}\ 6s^2\ 6p^1$ d and f - electrons have least shielding effect increasing Z^* ($Z^* = Z - \sigma$). Higher the Z^* greater is the Δ_iH value.

$Z^*_{Tl} > Z^*$ In it. $\Delta_iH(T\ell) > \Delta_iH(ln)$ inspite of increase of shell number.

45. Elements closed to zig-zag line of the long form periodic table are metalloids with both metallic and non-metallic behaviour e.g. Si, Ge, As, Sb ad Te. (Elements bordering the line and running diagonally across the periodic table.)

Properties :

(i) Metalloids have a solid state of matter

(ii) These have metallic lusture with low elasticity and brittleness.

(iii) These are semiconducted elements with average transmission of heat.

46. Defects of long form periodic table :

(i) Position of hydrogen is still controversial as some of its property demands its placement in group-1 and it has also similarities with halogens (group - 17).

(ii) Lanthanides and actinides have not been accomodated in the main body of the periodic table.

(iii) In short form periodic table the vacant place indicates discovery of new element, the vacant space between s and p-block never encourage placement of new element.

47. Factors affecting $\Delta_i H$ values :

Ionisation enthalpy $(\Delta_i H)$ depends on the following factors :

(a) nuclear charge : $\Delta_i H$ value increases with increase of nuclear charge.

(b) atomic size : Greater is the atomic size smaller is the $\Delta_i H$ value.

(c) Penetrating effect. s > p > d > f. e.g. the ionisation enthalpy will be more to knock out a s-electron than p-electron of the same shell, which will be in turn, will be more than that required to remove a 'd' - electron and so on.

(d) Shielding or screening effect of the inner shell electrons : Imperfect shielding increrases the effective nuclear change (ENC) and hence the ionisation enthalpy ($\Delta_i H$ values)

Zeff = Total nuclear change (Z) – screening constant (σ). The $\sigma_s > \sigma_p > \sigma_d > \sigma_f$. An increase in the number of electrons in the inner shells tends to decrease the ionisation enthalpies.

(e) Electronic configuration : More stable is the electronic configuration (half filled, completely filled or noble gas configuration, ns^2 or $ns^2\ np^6$) more is the energy required to expell the $\overline{e}$ and greater is the $\Delta_i H$ value.

$$\Delta_i H_1 \left(Be - 1s^2\, 2s^2\right) > \Delta_i H_1 \left(B - 1s^2\, 2s^2\, 2p^1\right)$$

48. The $\Delta_i H$ values increases along a period from left to right with increase of nuclear change and decrease of atomic size. However, the increase is not regular / sequential due to stable electronic configuration. Atoms with stable electronic configuration have higher $\Delta_i H$ value.

e.g. $\Delta_i H_1$ of $_7N$ > $\Delta_i H_1$ of $_8O$ due to half filled stability of N ($2s^2\ 2p^3$), similar is the case is Be ($2s^2$) i.e, $\Delta_i H_1$ of $_4Be$ > $\Delta_i H_1$ of $_5B$ inspite of less nuclear charger.

In a group : The $\Delta_i H$ value decreases on moving down the group with increases of shell number i.e. atomic size. However, the sequence becomes irregular on introduction of d and f - electron which influence the Z^* values. e.g. In group . 14 $\Delta_i H_1\ ({}_{82}Pb) > \Delta_i H_1({}_{50}Sn)$

49. The electron gain enthalpy $(\Delta_{eg} H)$ is the energy released when an $\overline{e}$ adds to an isolated gaseous atom, more is the energy released more is the electron affinity. On approach of an $\overline{e}$ to electron rich anion (less $P/_e$ ratio), it experiences repulsion and becomes endoergic i.e., energy needed to surpass the repulsive force between the anion (g) and incoming / added $\overline{e}$.

$$\Delta_{eg} H_1 \Rightarrow O_{(g)} + \overline{e} \rightarrow \overline{O}_{(g)} + 141\,kJ/mol$$

$$\Delta_{eg}H_2 \Rightarrow \bar{O}_{(g)} + \bar{e} + 780\ kJ/mol = O^{2-}{}_{(g)}$$

i..e, $\Delta_{eg}H_1 = -141\ kJ/mol$ (attraction)

$\Delta_{eg}H_2 = +780\ kJ/mol$ endoergic i.e., repulsion

The overall reaction $O_{(g)} + 2\bar{e} \longrightarrow O^{2}{}_{(g)}$ is endoergic.

50. $Cl_{(g)} + \bar{e} \xrightarrow{\text{more attraction}} Cl^{-}_{(g)} + 349\,kJ/mol$

$F_{(g)} + \bar{e} \xrightarrow{\text{less attraction}} F^{-}{}_{(g)} + 328\,kJ/mol$

The data show, F the most electronegative atom has lower $\Delta_{eg}H$ value than that of Cl. This is due to small size of 'F' ($2s^2\,2p^5$). The electron–electron repulsions in the relatively compact '2p' subshell are comparatively large & hence the incoming $(\bar{e})$ is not accepted favourably as in the case with Cl ($3p^2\,3p^5$).

51. Noble gases have completely filled subshells ($s^2\ p^6$ except He (s^2) – Octet and duplet). There is no room in their valence orbitals $\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$ & the added $\bar{e}$ has to be placed in an orbital of next higher shell. As a result energy is needed to to add that additional $\bar{e}$ (endoergic i.e., $\Delta_{eg}H = \text{positive}$)

e.g. $Ne(g) + \bar{e} + 116\ kJ/mole = N\bar{e}(g)$

$Rn(g) + \bar{e} + 68\ kJ/mol = Rn^{-}(g)$

The $\Delta_{eg}H$ values decreases from Ne or Rn as the size increases.

$\Delta_{eg}H$ of He is however lowest of all $(+48\,kJ/mol)$. This unexcepted behaviour is that due to its smallest size, it has much higher tendency to accept an additional $\bar{e}$ than any other noble gas i.e., experiencing less repulsion showing less addition of external energy.

Due to stable electronic configuration these elements have zero electron affinity.

52. $O_{(g)} + \bar{e} \rightarrow \bar{O} + 141\ kJ/mol,\ \ \Delta_{eg}H = -141\,kJ/mol$

$N_{(g)} + \bar{e} + 31\ kJ \rightarrow \bar{N}\qquad ,\ \ \Delta_{eg}H = +31\ kJ/mol$

On addition of $\bar{e}$ to an isolated 'O' atom energy is released while $\bar{e}$ addition of $N_{(g)}$ needs 31 kJ/mol. i.e., endoergic. This is due to repulsion between N atom and the incoming $\bar{e}$ arising out of stable (half filled) configuration of N ($1s^2\,2s^2\ 2p^3$)

53. In these elements there penultimate shell is filled with 18 (pseudonoble gas configuration) electrons and ns^2 electron in the outer shell i.e., Zn ($4s^2$) Cd ($5s^2$) and Hg ($6s^2$). Inspite of $\bar{e}$ in 's', these elements resemble more the d-block rather than s-block elements as listed below :

(i) Zn, Cd and Hg like other d-block elements form complexes but s-block elements usually don't.

(ii) Zn, Cd, Hg form covalent compound while 's' block elements form only ionic compounds.

(iii) Like other d-block elements these are having higher $\Delta_i H$ values than s-block elements in the same period.

(iv) Like other d-block elements the electrode potential of Zn, Cd, Hg are much less negative than those of s-block elements of the same period.

Unit - IV

54. **Causes of combination of atoms :**

(i) Electronic structure - To attain octet / duplet i.e. stable electronic configuration.

(ii) Net attractive forces between atoms :

(a) Attractive forces between ${}^{n}A - {}^{e}B$ and ${}^{n}B - {}^{e}A$

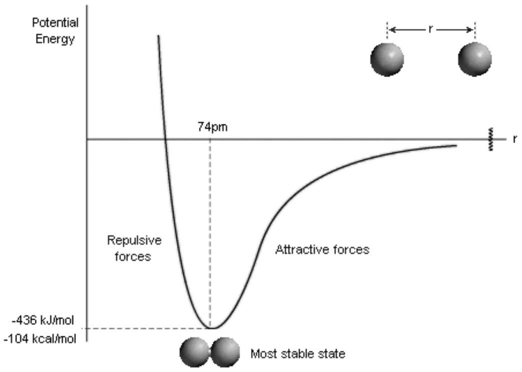
(b) Repulsive forces between ${}^{n}A - {}^{n}B$ & ${}^{e}A - {}^{e}B$

${}^{n}A$ = nucleus of atom 'A'	${}^{e}A$ = electron of atom 'A'
${}^{n}B$ = nucleus of atom 'B'	${}^{e}B$ = electron of atom 'B'

When resulting force is attractive, the two atoms combine.

(iii) Lowering of energy : At the time of bond formation the energy gets lowered to a minimum internuclear distance. If the distance is reduced further there will be repulsion $\left({}^{n}A - {}^{n}B \text{ and } {}^{e}A - {}^{e}B\right)$ & no-bond formation. The depth of the potential pool is called lowering of energy (or) bond energy & the inter nuclear distance of two atoms at which there is maximum decrease of potential energy (PE) is called bond length (r_0).

When r_0 is infinity the PE is zero. It decreases where the attractive force arises as the atoms come closer.



(iv) **Octet rule :** Noble gas has the stable configuration with 8 electrons in its valence shall (except He = 2) Thus atoms with less than 8-electron (or 2 - electrons) enter into chemical reaction (bond formation) in order to attain stable configuration (Octet or duplet).

Such a stable state can occur by mutual sharing or transfer of one or more electrons.

55.

Ionic / electrovalent bond	**Covalent bond**
(i) Formed by transference of electron undergoing redox reaction. One loses e^- & other gains it.	(i) Formed by mutual sharing as a result both occupy stable configuration (duplet/octet)
(ii) It occurs between two different atoms of widely varying electrongativity i.e., $\Delta x \geq 1.7$	(ii) It occurs betwen to similar or dissimilar atoms of less electronegativity difference e.g. H–Cl, Cl–Cl.
(iii) Does not form isomerism. Some compound show different geometric arrangement i.e. polymorphic.	(iii) Due to directional nature of covalent bond it shows isomerism.
(iv) Electrovalent bond is designated as + – . e.g $Na^+ Cl^-$	(iv) Represented by solid stick, A – B.

(v) Charge on ion is multiple of e-charge i.e., 1.6×10^{-19}C. e.g. chargeon Na^+ = Charge on chloride ion = 1.6×10^{-19}C., charge on Mg^{2+} = Charge on SO_4^{2-} = $2\times1.6\times10^{-19}$ C.	(v) Due to polarity, charge on an atom is partial i..e. δ^+ & δ^-.

56.

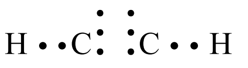
Ionic compounds	**Covalent compounds :-**
(a) Due to transfer of $\overline{e}(s)$ it consisting of ions (cation and anion)	(a) As formed by sharing $\overline{e}(s)$ / overlapping of atomic orbitals, these exhibit in molecular form
(b) Bonding occurs between more electropositive metal & less electro psitive or more electronegative non-metals. e.g. NaCl.	(b) Bonding between non-metals, e.g. Cl–Cl
(c) These compounds are soluble polar solvents due to ion-solvent attraction i.e, hydration energy surpasses the lattice energy.	(c) Covalent compounds are soluble in non-polar solvents. e.g. Cl_2 is less soluble in water but more soluble in CCl_4.
(d) Ionic compounds due to lattice force exhibit in solid state providing different crystalline. structure.	(d) Covalent compounds generally exhibit in liquid or gaseous state exhibiting no structure, however the molecule provides different structure / shape. e.g. CH_4 is tetrahedral atomic arrangement but in gaseous state molecules held by vander Waals force.
(e) Ionic compounds are good electrolytes i..e, conducting electricity in molten state or in solution.	(e) Covalent compounds are bad conductors.
(f) These possess high m.p. and b.p.	(f) These are with low mp (even in solid e.g. iodine) and b.p.
(g) These compounds undergo ionic reactions (fast) in solution. $\overset{+}{K}C\overline{l} + A\overset{+}{g}NO_3^- \longrightarrow AgCl\downarrow + \overset{+}{K}NO_3^-$	(g) Undergo molecular reactions which are very slow as bond breaking (old bond) and bond making (new bond) consumes some time.

57.

Atoms	**Ions**
i. Atoms are neutral particles	i. Ions are charged particles +ve charged particles are (cations) formed on loss of $\overline{e}$ while –vely changed particles (anions) are formed on gain of $\overline{e}(s)$.
ii. The outer (valence) shell configuration is as per Auf bau principle. All atoms not having stable configuration.	ii. The outer shell configuration is 0, 8, 18 $\overline{e}(s)$ (stable configuration) e.g. H^+ $1s^0$ Na^+ $2s^2\ 2p^6$ Zn^{2+} $3s^2\ 3p^6\ 3d^{10}$ (Pseudonoble gas configuration)

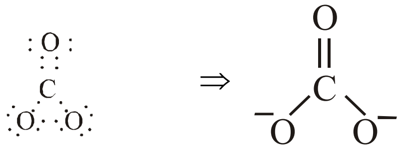
iii. Atom is having definite covalent radius.	iii. Cation is having smaller & anino having bigger than corresponding atom depending on $^{p}/_{e}$ ratio. $Na - e^{-} \longrightarrow Na^{+}$, smaller, $\frac{p}{e} >$ $Cl + \overline{e} \longrightarrow C\overline{l}$, bigger $\frac{p}{e} <$

58. C_2H_2 ,

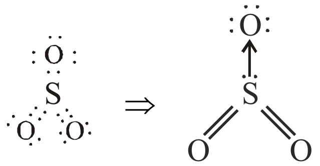


Triple bond

, CO_3^{2-}



SO_3



59. (i) Factors affecting ionic bond formation :

(a) Low heat of sublimation of metal

(b) Low ionisation enthalpy

(c) Low bond dissociation enthalpy of non-metal

(d) high electron gain enthalpy (electron affinity)

(e) high lattice energy (greater electrostatic force of attraction between oppositely charged ions, $F = \frac{q_1 \cdot q_2}{r^2}$. Smaller is the inter ionic distance greater will be inter ionic force of attraction and greater is the strength of ionic bond.

(ii) Factors influencing covalent bond formation :

(a) High $\Delta_i H$ value atoms with high I.E. favour covalent bond.

(b) Equal electron affinity.

(c) Electronegativity. Δx i.e., the difference of elctronegativity must be low.

(d) According to Fajan's rule smaller cation (high polarising power) and larger anion (high polarisability) favour formation of covalent bond.

60. **Bond parameters**

(a) Bond length : The equilibrium distance between the nuclei of two bonded atoms (inter nuclear distance) is known as bond length.

Significance (i) Covalent radius = $\frac{d_{A-A}}{2}$ d_{A-A} = inter nuclear distance

(ii) Greater is the bond length weaker is the bond strength and more instable is the molecule. Stability decreases in the order HI < HBr < HCl < HF

(b) **Bond order** / bond multiplicity : Number of bonds between two atoms in a covalent compound.

e.g.	Single bond H –H	(bond order = 1)
	Doubl bond $O = O$	(bond order = 2)
	Triple bond $N \equiv N$	(bond order = 3)

Higher is the bond order smaller is the bond length and greater is the stability.

Bond order $\propto \frac{1}{\text{bond length}} \propto$ bond energy $\propto$ stability.

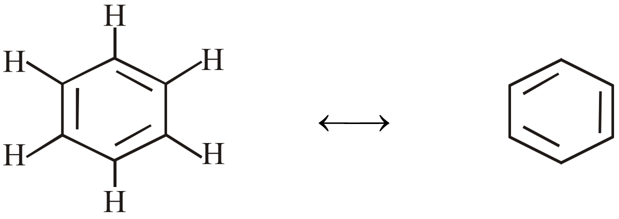
Calculation : (i) $\frac{1}{2}$ (No of electrons shared)

(ii) As per MOT bond order = $\frac{1}{2}$ (No. of $\overline{e}s$ is bonding – No. of $\overline{e}s$ is antibonding orbitals)

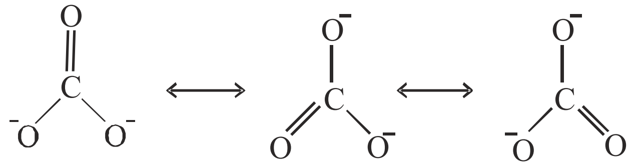
(iiii) $\frac{\text{No. of bonds between two given atoms}}{\text{Number of resonating structures}}$ = bond order

e.g. bond order of 'C' in benzene = No. of bonds between C and C

$$= \frac{1+2}{2} = 1.5$$



Two resonating structure



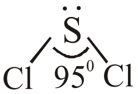
bond order of C – O bond = $\frac{2+1+1}{3} = 1.3$

(c) Bond angle. The angle between the lines representing the direction bond i.e., orbitals containing the bonding electrons is called bond angle (This supports the directional nature of covalent bond).

e.g. H – C –H bond angle in methane is $109^0\ 28'$. The bond angle signifies the shape of the molecule. The magnitude of bond angles is affected by –

(i) Number of lone pair on the central atom as as per VSEPR theory $\ell - \ell$ repulsion > $\ell - b$ repulsion > b-brepulsion

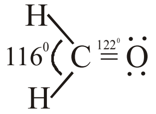
e.g. bond angle of CH_4 ($109^0\ 28'$) , : NH_3 (107^0), $\ddot{O}$ H_2O ($104^0 . 28'$)



(ii) Electronegativity and size of the central atom.
e.g. bond angle of NH_3 > that of PH_3 (93^0) (Thought of not undergoing hybridisation)
Electronegativity of N > P .
However Bond angle of $PF_3 > PH_3$, Here F is more electronegative than H causing greater electronic repulsion.

(iii) Presence of multiple bond. Due presence of π (electrons), repulsion increases the bond angle.

e.g.,



(iv) Nature of hybridisation.

$$\underset{(109^0\,28')}{sp^3} < \underset{(120^0)}{sp^2} < \underset{(180^0)}{sp}$$

e.g. H C H / H–C–C bond angle in ethane < ethylene < acetylene

(d) **Bond energy :** The amount of energy required to break one mole of particular type of bond so as to separate them into gaseous atoms is called bond dissociation energy or bond energy. Unit : kJ / mol

Factors affecting B . E :

(i) Bond length - Higher the bond length less is the BE.

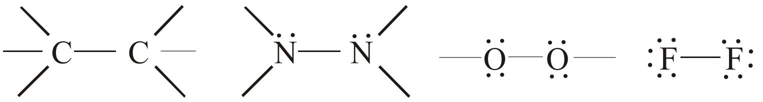
(ii) Multiplicity of the bond - Higher the BO greater the B.E.

e.g. $\overset{B-E}{N \equiv N} > O = O > F - F$

(iii) Number of lone pair of electrons present –

Greater the lone pairs of $\overline{e}$, greater is the electronic repulsion less is the BE.

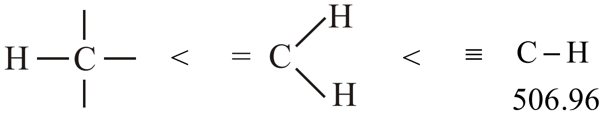
e.g.



No. of lone pairs/atom	0	1	2	3
BE (kJ/mol)	348	163	146	139

(iv) Nature of hybridisation. Greater is the s-character smaller is the bond length & greater is the BE.

C – H BE :



411 444.14 506.96 J / mol

61. **Polar character of covalent bonds :**

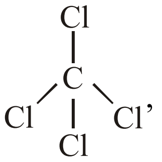
A covalent bond is formed by sharing of valence shell electrons of consisting atoms in a molecule. But electronegativity of an atom is the tendency to attract the shared pair of electrons towards itself. Due to varying electronegativity of atom the shared pair is unequally distributed between the two atoms. e.g. ln H–F, $F^{(4.0)}$ is more electroneative than H(2.1). Thus the shared pair is more towards F acquiring partial negative charge $(\delta-)$ and as shared $\overline{e}$ (s) are away from 'H', it acquired parital +ve change $(\delta+)$. As a result a dipole results $\left(\overset{\delta+}{H}-\overset{\delta-}{F}\right)$, the bond gets polarised or polar.

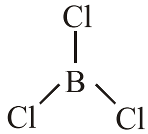
Such unequal / unsymmetrical distribution of shared electrons between two atoms makes the bond polar.

When two idential atoms or atoms with almost equal electronegativity value form a covalent bond, there occurs symmetrical distribution of $\overline{e}$ between two atoms i.e., the shared $\overline{e}$ pair remains midway between two nuclei & no charge separation takes place.

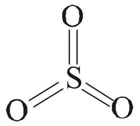
e.g. H_2, Cl_2, F_2, O_2, N_2, CH_4 etc. (non-polar molecules)

In certain cases the bonds are polar due to greator difference of electronegativity values, but the molecul behaves non-polar e.g. CCl_4, BCl_3, CO_2, SO_3 etc. In such case the structure / shape of the molecule is symmetrical. Symmetry leads to non-polarity.





, O = O = O,



62. Dipole moment is a quantitative way to express the degree of polarity of a covalent bond. It is a vector quantity, mathematically expressed as –

Dipole moment (μ) = change $\times$ distance = e.d.

e = magnitude of net +ve or -ve charge in the order of 10^{-10} esu. or in coulomb.

d = distance between the charges i.e., bond length is order 10^{-8}cm or 10^{-10} m (i.e., in SI unit.

Thus it is in Debye unit in the order 10^{-18} esu. cm

e.g. $\mu_{HCl} = 1.03\text{ D}$ i.e., 1.03×10^{-18} esu cm.

1 D = 3.336×10^{-30} C.M.

Dipolemoment is represented as $+\!\!\!\!\longrightarrow$ (crossed arrow) heading towards more electronegative atom e.g. $\overset{+\!\!\!\!\longrightarrow}{H - F}$

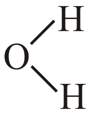
Significance :

(i) Polarity of the bond / molecule : More is the dipole moment more polar is the moelcule.

$$\underset{(1.98\text{D})}{\mu_{H-F}} > \underset{(1.03\text{D})}{\mu_{H-Cl}} \quad \text{i.e., HF is more polar than HCl.}$$

(ii) Shape of the molecule. Zero value of dipole moment indicates symmetrical molecule. CO_2 and H_2O both are triatomic, $\mu_{CO_2} = 0.0\text{D}$ while that of H_2O is 1.84 D indicating that CO_2 is linear while H_2O bent.

O = C = O ,



(iii) Distinction between polar and non polar molecule.

$\mu = 0$, indicates the molecule is non-polar.

e.g. BF_3, CH_4, CCl_4, H_2, N_2, O_2 etc.

(iv) Calculation of % of ionic character.

$$\text{i.e.,} = \frac{\text{Actual dipolemoment of the bond}}{\text{Dipole moment of pure ionic bond}} \frac{(\text{experimental})}{(\text{Theoretical})} \times 100$$

63. μ_{HF} = charge $\times$ distance

$= 4.8 \times 10^{-10}\text{ esu} \times 0.92 \times 10^{-8}\text{ cm}$

$= 4.42\text{ D}$

% of ionic character of H–F bond

$$= \frac{\text{ActualDM}}{\text{calculated}} \times 100 = \frac{1.98}{4.42} \times 100 = 44.8 \simeq 45$$

64. **Covalent character of ionic bond**

Ionic bonds posses some covalent character due to polarisation of ions. Anions are symmetrical when cation comes closer the symmetry is lost, the tendency of an anion to get polarised / deformed by cation is called its polarisability. The deforming power of the cation is called polarising power. The polarising power of cation and polarisability of anion is governed by Fajan's rule leading to partial covalent character of ionic bond.

Rules -1: Higher positive charge on cation

2 : Smaller cationic size

3 : Cation with $ns^2\ np^6\ nd^{10}$ (18) configuration possess higher polarising power than cation having $ns^2\ np^6$ configuration. Polarising power of $\underset{(2,8,18)}{Zn^{2+}} > \underset{(2,8,8)}{Ca^{2+}}$

(This is due to less shielding effect of d-electrons.)

4. Greater negative change on anion.

5. Larger size of the anion.

e.g. covalent character of Lithium halides follows in the order : $LiF < LiCl < LiBr < LiI$

I^- has greater polarisability while being smaller cation Li^+ has greater polarising power.

65. Difference between σ bond and π bond :

Sigma bond	**Pi bond**
(i) It is formed by overlaping of half filled atomic orbital along their inter nuclear axis.	(i) A - Pi - bond is formed by side wise / lateral overlapping p-p or d-p orbitals.
(ii) Due to greater extent of overlaping of the charge cloud it's stronger.	(ii) π-bond is weaker
(iii) It is having a single $\bar{e}$-cloud, symmetrical along intermolecular axis.	(iii) Electron cloud distributes above and below the inter nuclear axis (two $\bar{e}$ clouds)
(iv) About σ-bond free rotation is possible. e.g. conformers of alkane.	(iv) π-bond restricts free rotation. e.g. cis- trans - isomerism

66. **Resonance :**

It is seen that certain molecules and ions can be represented by none than one valence bond / Lewis structure but none of these structures can correctly account for the properties like bond length, bond energy etc. of the molecule or ions. The alternative structures written for a given molecule / ion without change of position of nuclei, bonding and non-bonding pairs of $\bar{e}(s)$ are called resonating / canonical structures. The actual structure is said to be a resonance or mesomeric hybrid of all these alternative forms or canonical structures. The process of representing different Lewis structures and writing the hybrid structure of a given molecule or ion is called resonance.

Example : CO_2

$$\underset{1.22\,A^0}{O=C=O} \longleftrightarrow \underset{1.10\ ^0A}{\overset{+}{O}\equiv C-\bar{O}} \longleftrightarrow \bar{O}-C\equiv\overset{+}{O} \qquad \equiv \qquad \underset{1.15\,A^0}{O \underline{\cdots} C \underline{\cdots} O}$$

When the actual bond length / bond energy differs from the expected bond length and energy, the species must have under gone resonance. This resonance affects bond energy, bond length, stability

$$\text{Bond order of C and O in } CO_2 = \frac{\text{No. of bonds between C \& O}}{\text{No. of resonating structures}}$$

$$= \frac{4}{3} = 1.3$$

Resonance is represented by a double headed arrow. $\left(\longleftrightarrow\right)$

Features :

(i) The canonical forms have no real existence and cann't depict the properties of the molecule.

(ii) All these structures contribute to actual structure of the molecule which is resonance hybrid of all these contributing structures.

(iii) In the canonical from similar charge should not reside on adjacent atoms and also the positive charge should not rest on more electronegative atom.

(iv) Greater is the resonating structure on delocalisation of $\overline{e}(s)$ greater is the stability of that structure or species.

(v) There is no such equilibrium between canonical forms as we have between tautomeric forms / tautomers, $\left(\text{Keto form} \rightleftharpoons \text{enol form}\right)$

(vi) Resonance affects the bond energy and bond length e.g. C–O bonds in CO_3^{2-} are identical.

67. **Resonance energy :** It is defined as the difference in energy between the actual structure of the entity and most stable of hypothtical structures (canonical structures)

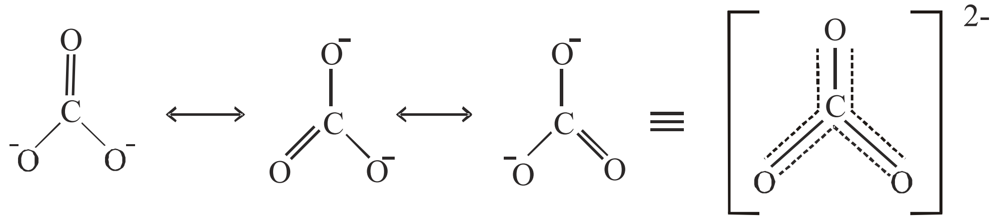
resonance energy = Experimental – theoretical = negative.

for Benzene = 208 kJ–360 kJ = – 152 kJ/mol. More negative is the values more is the stability i.e., resonance lowers the energy of the moelcule.

Resonance structure of Nitrogen molecule :

$$:N \equiv N: \quad \longleftrightarrow \quad {}^{(-)}:\ddot{N} =: \ddot{N}\,{}^{(+)} \quad \longleftrightarrow \quad \overset{+}{\ddot{N}}::\overset{-}{\ddot{N}}: \quad \equiv \quad N \equiv N$$

(10 valence electrons)



All bonds of C–O are identical due to resonance.

$$\text{Bond order of C–O bond} = \frac{4}{3}$$

68. (i) $\overset{+4}{C}\overset{-4}{O_2}\,(2\sigma \text{ and } 2\pi \text{ bonds})$

$C^*\ \underbrace{\boxed{\uparrow}\ \boxed{\uparrow}}\boxed{\uparrow}\boxed{\uparrow}$

Two $p-\overline{e}(s)$ left for 2π bonds

sp-hybridisation hence linear.

O=C=O $\quad C \Rightarrow$ sp hybridised (AB)

$O \Rightarrow sp^2$ hybridised (ABL_2)

$\underbrace{\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\boxed{\uparrow}}\boxed{\uparrow}$ π – bond (pure p)

$\overset{+4}{S}\overset{-4}{O_2}$

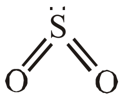
$S^*\ \underbrace{\boxed{\uparrow\downarrow}\ \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}\ \boxed{\uparrow}\boxed{\ }\boxed{\ }\boxed{\ }$

S undergoes sp^2 hybridisation

i.e. AB_2L type

i.e. Bent due to presence of a lone pair of $\overline{e}$ in sp^2 hybrid orbital.

So it's bent.



(ii) Be $1s^2\ 2s^2\ 2p^0$

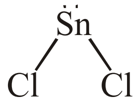
$Be^*\ \underbrace{\boxed{\uparrow}\ \boxed{\uparrow}}_{2s\ \ 2p}\boxed{\ }\boxed{\ }$

sp-hybridisation

Shape : Linear

$SnCl_2$

Sn $5s^2\ 5p^2$



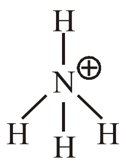
$\underbrace{\boxed{\uparrow\downarrow}\ \boxed{\uparrow}\boxed{\uparrow}}_{5s\ \ \ 5p}\boxed{\ }$ sp^2– hybridisation.

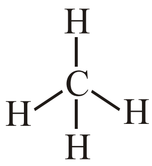
Shape : Bent (AB_2L - type)

(iii) CH_4 - No. valence electrons = 4 + 4 = 8

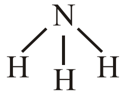
$\overset{+}{N}H_4$ - No. valence electrons - (5-1) + 4 = 8

i.e. these two species are isoelectronic. Hence both are isostrutual i.e. tetrahedral.

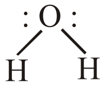




(iv)



sp^3



sp^3

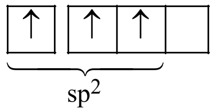
AB_3L type $\qquad$ AB_2L_2 type with ℓ - ℓ, ℓ -b and b-b repulsions.

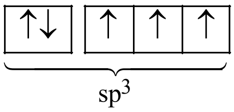
As per VESPR theory, l-l repulsion in water is greater than the b–b. repulsion, hence bond angle contracts much from the tetrahedral angle ($109^0\ 8^{/}$)

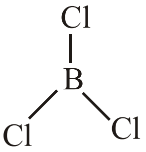
69.

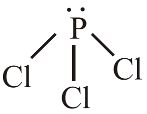
	Pure orbital		Hybrid orbital
(i)	Determined by (ℓ and) 'm' quantum number values $\ell = 1(p)$, $m = -1, 0, -1$ p-subshel has three possible orientation. i.e., having three (p_x, p_y, p_z) orbitals.	(i)	On mixing the orientiation changes. e.g. three sp^2 orbitals are oriented at an angle of 120^0.
(ii)	Lobes identical. e.g. p-orbital dumb-bell shaped. 	(ii)	On mixing the $\bar{e}$ distribution changes and the mixed or hybrid orbital lobes become unsymetrical,  More is the 'p' character more is the elongation.
(iii)	There are atomic orbitals like s, p, d and f.	(iii)	Molecular orbitals formed by mixing of atomic orbitals. e.g. sp, sp^2, sp^3, sp^3d, sp^3d^2, sp^3d^3 etc.
(iv)	Each orbital can accmodate two $\bar{e}$ s maximum.	(iv)	Hybrid orbitals can have also 2 $\bar{e}$s maximum.

70. ${}_5B$ $1s^2$ $2s^2$ $2p^1$ PCl_3 ${}_{15}P$ $3s^2$ $3p^3$ $3d^0$

B^* 



Shape : Trigonal 



Type : AB_3L (one lone pair)

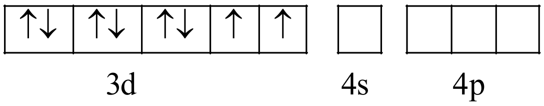
Angle = 120^0 AB_3 type

Shape : Pyramidal

Angle- < $109^0 28'$ (100^0)

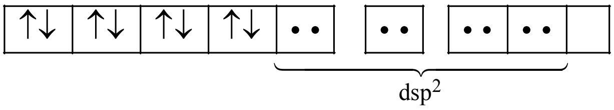
71. sp^3d - PCl_5 Trigonal bipyramidal

dsp^2 $[Ni(CN)_4]^{2-}$ ${}_{25}Ni$ (II)



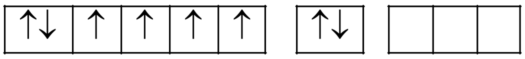
$[Ni(CN)_4]^{2-}$

Square planar

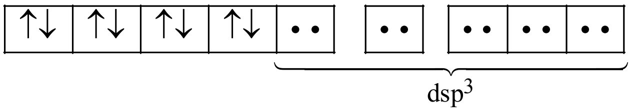


•• ($\bar{e}$ pair of ligand , $C\bar{N}$)

dsp^3 Fe $(CO)_5$ Fe (0) $3d^6$ $4s^2$ $4p^0$



$[Fe(CO)_5]$

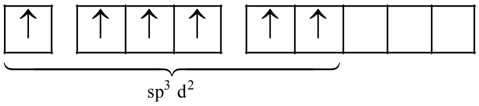


(•• from ligand CO.)

Shape - Trigonal bipyramidal

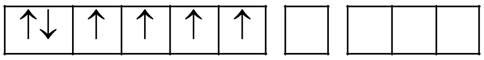
sp^3d^2 SF_6 S(0) $3s^2$ $3p^4$ $3d^0$

S(VI)

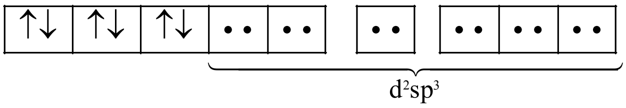


Shape Octahedral

d^2sp^3 $[Fe(CN)_6]^{4-}$ Fe (II) $3d^6$ $4s^0$ $4p^0$



$[Fe(CN)_6]^4$



(•• from ligand $C\bar{N}$)

72. **Formation MOs (LCAO)**

The formation of molecular orbital in the light of wave functions is described by an approximate method known as linear combination of atomic orbitals (LCAO).

For two electron molecular orbitals system (i.e. H_2) may be described by LCAO that can take place by addition and by substraction of wave functions of individual atomic orbitals.

$\psi_{mo} = \psi_A \pm \psi_B$ ψ_A = wave function of atom A

ψ_B = wave function of atom B

Thus, two MOs can be written as

$\sigma = \psi_A + \psi_B$ (Bonding MO) - constructive interference of electron waves (additive)

$\sigma^* = \psi_A - \psi_B$ (AMO) – destructive interference of electron waves.(substractive)

In a bonding molecular orbital $\bar{e}$ density concentrates between the two nuclei while is AMO $\bar{e}$ density is located away from the space between the nuclei. In fact there is a nodal plane between nuclei.

Hence the repulsion between the nuclei is high. Electrons present in BMO tend to hold the nuclei together & fovours bond formation. Electrons in AMO destabilise the molecule. The total energy of two MOs however remains same as that of two original atomic orbitals.

73.

	Bonding M.O. (BMO)		**Antibonding M.O. (AMO)**
a.	It is formed by linear combination of two atomic orbitals when their wave functions are added i.e., $\psi_A + \psi_B$. (additive)	a.	AMOs are formed by linear combination of two atomic orbitals when their wave functions are substracted, $\psi_A - \psi_B$ (substractive)
b.	Its energy is less than the combining atomic orbitals.	b.	$E_{AMO} > E_{A.O.(s)}$
c.	It increases the electron density between the nuclei stabilising the molecule.	c.	It decreases the electron density between the molecule, therefore destabilises the molecule.
d.	It has no nodal plane.	d.	It has nodal plane where electron density is zero.
e.	It is symmetrical about inter nuclear axis.	e.	It is symmetrical about the intermolecular axis and about a line perpendicular to it.

74. Electronic configuration of N_2 (14 es)

$\sigma_{1s^2}\ \overset{*}{\sigma}_{1s^2}\ \sigma_{2s^2}\ \overset{*}{\sigma}_{2s^2}\ \pi_{2p_x^2}\ \pi_{2p_y^2}\ \sigma_{2p_z^2}$ (No unpair $\overline{e}$, diamagnetic)

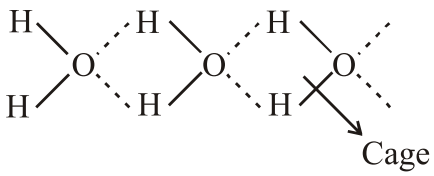
Electronic configuration of $O_2\ (16\,e^-)$

$\sigma_{1s^2}\ \overset{*}{\sigma}_{1s^2}\ \sigma_{2s^2}\ \overset{*}{\sigma}_{2s^2}\ \sigma_{2p_z^2}\ \pi_{2p_x^2}\ \pi_{2p_y^2}\ \pi^*_{2p_x^1}\ \pi^*_{2p_y^1}$

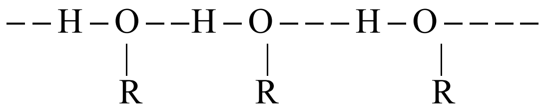
Two unpair $\overline{e}$ in antibonding π orbitals, hence paramagnetc.

75. (i) In ice the polar water molecules are so arranged that due to H-bonding cages (vacants paces) are formed increasing the volume and decreasing the density.

Hence floats over water.

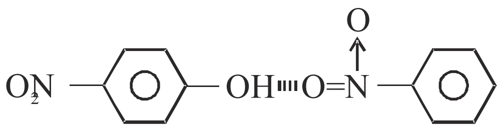


(ii) Strength of H - bond : $O\cdots\cdots H$ bond > $N\cdots\cdots H$ bond as oxygen more electronegative (3.5) than N (3.0). This is the reason why alcohols are much associated with higher b-p than amines.

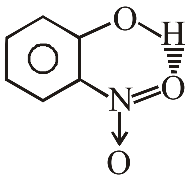


$$--\mathrm{H}-\underset{\substack{|\\ \mathrm{R}}}{\mathrm{O}}--\mathrm{H}-\underset{\substack{|\\ \mathrm{R}}}{\mathrm{O}}---\mathrm{H}-\underset{\substack{|\\ \mathrm{R}}}{\mathrm{O}}----$$

(iii) p-nitrophenol is much associated due to intermolecular H-bond & hence is having higher m.p while o-nitrophenol undergoes intramolecular H-bonding & the molecules are lean associated by vander Waals force with less m.p.



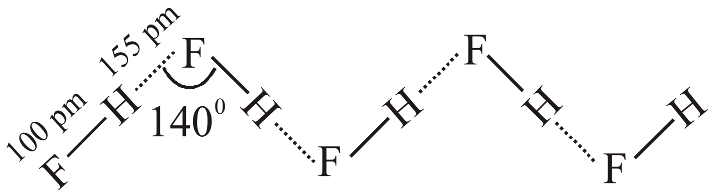
P-nitrophenol
(inter molecular H-bond)



0-nitro phenol (intramolecular-H-bond)

(iv) The much polar H_2O molecules are held stronger by intermolecular H-bonding & henc e exhibitted in the form of liquid with higher b.p.. Sulfur is less electronegative and bigger in size, fail to form H-bond. Thus H_2S molecules are held by weaker vander Waals force and exhibits in gaseous state.

(v) F (4.0) is the most electronegative atom and forms strong inter molecular H-bond. Due to stronger association of much polar H-F it exhibits as liquid . Others are held by weaker vander Waals force, hence as gases.



Unit - V

76. These are four measurable properties of a gas. These are (i) mass (ii) volume (iii) pressure (iv) Temperature, the state variable of gases.

(i) **Mass.** It is expressed in gram or k.g. It is also related to number of moles (n)

$$n = \frac{W}{M} = \frac{\text{wt. of the gas}}{\text{Mole mass.}}$$

i.e, Mass (m) = no. of moles × mol. mass.

e.g. 0.2 moles N_2 weighs $28 \times 0.2 = 5.6\,\text{gram}$

(ii) **Volume.** Volume of the gas = volume of the container at given temperature and pressure. It is expressure as $1\ell = 1\,\text{dm}^3 = 10^3\ \text{cm}^3 = 1000\,\text{ml}$.

(iii) **Pressure**. The randomly moving molecules in a vessel not only collide among themselves but also against the inner wall of the container. The colliding force per unit area of the container is called pressure i.e., Pressure = Force / area

Units : 1 atm = 76 cm of Hg = 760 mm Hg = 760 torr

SI unit = Pascal i.e, N/m^2

1 atm = 10^2 K Pa = 10^5 Pa

P (1 atm) = h ρ g= $76\,\text{cm} \times 13.6\ \text{gm/cm}^3 \times 981\,\text{cm/sec}^2$

$$= 1.012 \times 10^6 \frac{\text{gm.cm}}{\text{cm}^2\ \text{sec}^2} = 1.012 \times 10^6\ \text{dyne/cm}^2$$

(iv) Temperature - It indicates the intensity of heat possessed by the molecules. With increase of temperature the kinetic energy increases. It is represented in ^{0}C or Kelvin temperature (K) where $K = {}^0C + 273$ e.g. boiling point water = 373 K i.e, 100^0C

77. Postulates of kinetic molecular theory :

(a) Every gas consists of very minute particles called molecules which are of similar in shape, size and mass and different from other gas molecules physically & chemically.

(b) The actual volume of the molecule is negligible in comparison to the volume of the container.

(c) The molecules are always in a state of constant rapid, zig-zag motion in all possible direction with different velocities. They travel in a straight line. The direction of motion is changed either on collision with other molecules or with the walls of the container. There is no molecule with zero velocity.

(d) The period during which the collision occurs is negligible compared to the time taken to traverse the distance between two successive collisions.

(e) The molecules are supposed to be incompressible spheres and unergo elastic collisions i.e., no loss of energy during collisions with one another.

(f) Un-equal bombardments of molecules on the walls of the container give rise to pressure. i.e, pressure is the average force per unit area of the wall which the moelcules exert during their impacts on walls per unit time. More is the no. of collisions per unit area per unit time, more will be the pressure.

(g) The molecules move independently. The inter molecular force of attraction is negligible.

(h) Molecules are not affected by gravitational force.

(i) Average K.E. of the gas molecules is directly proportional to absolute temperature.

78. From kinetic gas equation, it's known that

$$PV = \frac{1}{3} m N \bar{c}^2$$

m = mass of a single molecule

N = total number of molecules

$\bar{c}$ = root mean square velocity

$$= \sqrt{\frac{n_1 c_1^2 + n_2 c_2^2 + n_3 c_3^2 + ...}{n_1 + n_2 + n_3 + ... = N}}$$

If it contains 1 mole of gas, $N = N_A$ (Avogadro's number) and mN_A = molecular mass of the gas.

$$PV = RT = \frac{2}{3} \times \frac{1}{2} m N_A \bar{c}^2$$

$$= \frac{2}{3} \text{ K.E. (K.E. of 1 mole of molecules} = \frac{1}{2} m N_A \bar{c}^2$$

$$= \frac{2}{3} . N_A . KE$$

Per molecule $\boxed{K.E = \frac{3RT}{2N_A} = \frac{3}{2} K.T}$ $K = \frac{R}{N}$ = Boltzmann constant. i.e., Average KE $\propto$ T

It shows :

(i) Average K.E of all gases is same at same temperature.

(ii) At absolute zero, the kinetic energy is zero
i.e., molecules cease to move at 'O' K, tending to behave as a perfect crystal (no randomness).

79. Kinetic gas equation : $PV = \frac{1}{3} m N \bar{c}^2$

$$PV = \frac{2}{3} . K.E.$$

(i) As KE is constant at a given temperature, PV = constant (at a given temperature)

i.e., $P_1V_1 = P_2V_2 = ...$ (Boyle's law).

(ii) Derivation of Charles law:

$$PV = \frac{2}{3} . KE$$

$$KE \propto T$$

i.e., $PV \propto T$, at constant pressure $V \propto T$, i.e., $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ .... Charles' law for a given mole of gas at a given pressure.

(iii) **Avogadro's law** Under similar conditions of temperature & pressure equal volume of all gases contain equal number of molecules. i.e, $V \propto N$ at a given T and P.

For gas I $P_1V_1 = m_1 N_1 \bar{C}_1^2$

For gas II $P_2 V_2 = m_2 N_2 \bar{C}_2^2$

At a given temperature, $T = \frac{1}{2} m_1 c_1^2 = \frac{1}{2} m_2 \bar{c}_2^2$ i.e., the gases have same average K.E.

i.e., $m_1 \bar{c}_1^2 = m_2 \bar{c}_1^2$

At given P and V, $P_1V_1 = P_2V_2$ (for both the gases I and II)

$m_1 N_1 \bar{c}_1^2 = m_2 N_2 \bar{C}_2^2$

As $m_1\bar{C}_1^2 = m_2 \bar{c}_2^2$, $N_1 = N_2$

i.e., both container with volume $V_1 = V_2$, contain equal number f molecules, thus $N_1 = N_2$ (Proved).

(iv) **Dalton's law of partial pressure.** Total pressure of a mixture of gases is equal to sum of their partial pressures i.e., each gas would exert if present alone in the mixing container at a given temperature.

Kinetic gas equation : $PV = \frac{1}{3} m\ N\ \bar{c}^2$

$$= \frac{2}{3}.\ K.E$$

$$KE = \frac{3}{2}\ PV$$

If a number of gases is under consideration then the total K.E. of the gaseous mixture must be equal to K.E. of the individual gases.

$(KE)_{Total} = (KE)_1 + (KE)_2 + (KE)_3 + ...$ 1, 2, 3............ are individual gases.

or $\frac{3}{2}\ PV = \frac{3}{2}\ P_1V + \frac{3}{2}\ P_3V + ...$

V = volume of the mixing container containing gases 1, 2, 3..........

P_1 = pressure exerted by gas 1 if present alone in the volume 'V'

= partial pressure of gas - 1

Thus $\boxed{P = P_1 + P_2 + P_3 + ...}$ proved.

(v) Graham's law of diffusion : At a given temperature & pressure the rate of diffusion gases is inversely proportional to the square root of their densities and molecular masses.

i..e, $r \propto \frac{1}{\sqrt{d}} \propto \frac{1}{\sqrt{M}}$

Kinetic gas equation : $PV = \frac{1}{3} m N \bar{c}^2$

For one mole of gas $PV = \frac{1}{3}\ m.N_A\ \bar{c}^2 = \frac{1}{3}\ M\ \bar{c}^2$

M = Molecular mass of the gas

$$\bar{C}^2 = \frac{3PV}{M} = \frac{3P}{M/V} = \frac{3P}{d}$$

$$\overline{C}^2 \propto \frac{1}{d} \text{ at a given pressure.}$$

Greater velocity of the gases, greater is the rate of diffusion. $r \propto \overline{C}$

$$r \text{ (or) } \overline{C} \propto \frac{1}{\sqrt{d}}. \text{ Proved.}$$

80. **Compressibility factor .**

Compressibility factor (Z) is defined as $Z = \frac{PV}{(PV)\text{ideal}} = \frac{PV}{n\,RT} = \frac{P\,V/n}{RT} = \frac{P\,V_m}{RT}$

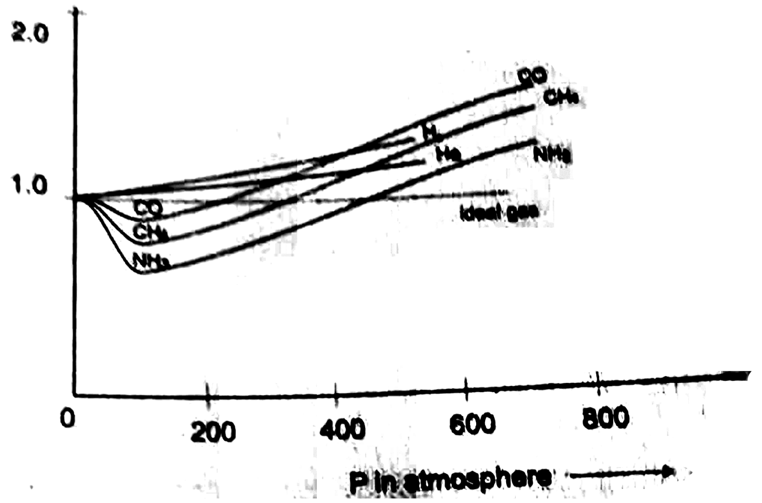
Where V_m is the molar volume i.e. volume occupied by 1 mole of the gas $\left(\frac{V}{n}\right)$

For ideal gas, Z = 1 as PV = n RT

The deviation of Z from unity is thus measure of imperfection of the gas under consideration and deviation from ideal behaviour.

Effect of change of pressure on deviation :

The compressibility factors are determined for a number of gases over a wide range of pressures at a constant temperature i.e. 0^0C.



It is seen from the graph ———

(i) At extremely low pressure ‘Z’ is very close to unity indicating that all gases at very low pressure behave almost ideally.

(ii) At very high pressure ‘Z’ is greater than unity indicating that the gases are less compresible than ideal gas, due to dominance of molecular repulsive forces.

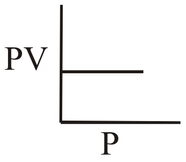
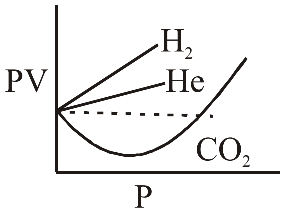
(ii) At moderately low pressure the gases are more compressible than ideal gas. PV is less than PV_{ideal} & Z< 1. This is due to dominance of long range attractive forces at low pressure which favour compression.

The value of Z initially goes on decreasing with increase in pressure, passes through minimum & then begins to incease with further increase of pressure.

(iv) For H_2 and He, $Z > 1$. These gases seem to be less compressible than ideal gas at all pressure and at 0^0C. However, if temperature is sufficiently low i.e. below – 165^0C for H_2 and below -240^0 for He, than Z–P plot are shown by both these gases as for other gases.

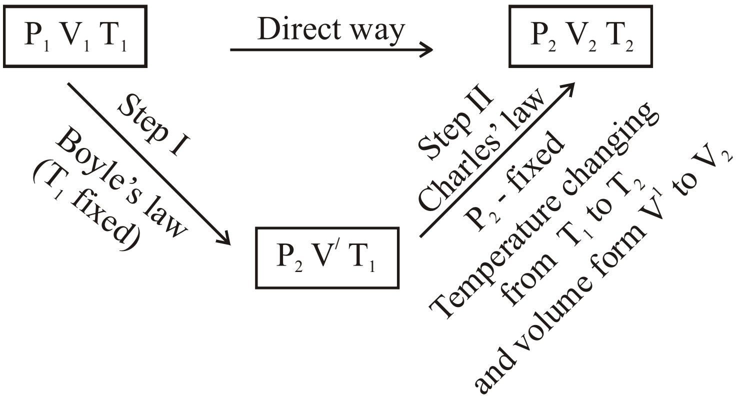
On the other hand, if the temperature is sufficiently high, then other gases show Z–P plots similar to those given by H_2 and He.

81. Distinction between ideal gas and real gas

	Ideal gas		**Real gas**
i)	It obeys gas laws under all conditions of temperature & pressure.	i)	It obeys gas laws at low pressure & high temperature.
ii)	Ideal gases are theoretical gases.	ii)	All gases are real.
iii)	Volume of the gas molecules is negligible in comparison to the total volume occupied by the gas.	iii)	Volume of the gas molecules can't be neglected.
iv)	The forces of attraction among the molecules the gas are negligible.	iv)	Forces can't be neglected at high pressure and low temperature.
v)	It obeys ideal gas equation $PV = nRT$	v)	It obeys vander Waals equation. $\left(P + \frac{an^2}{v^2}\right)(v - nb) = nRT$
vi)	 PV – P plot is palallel to pressure axis.	vi)	 PV–P deviates from ideal plot.
vii)	Compresibility factor. Z = 1 (ideal gas)	vii)	Z is greater or less than one depending on condition of pressure, temperature and nature of the gas.

82. The equation which gives the simultaneous effect of pressure and temperature on the volume of a gas is known as ideal gas equation or equation of state for an ideal gas.

Derivation :



Step I - As per Boyle's law –

$$P_1 V_1 = P_2 V^1$$

$$V^1 = \frac{P_1 V_1}{P_2}$$

Step II - As per Charles' law, $V \propto T$

$$V^{/}\big/_{T_1} = V_2\big/_{T_2}$$

Substituting $V^{/}$ value

$$\frac{P_1 V_1}{P_2 \times T_1} = V_2\big/_{T_2} \Rightarrow \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Ideal gas equation in terms of density :

$$d = \frac{m}{v}, \text{ for a given mass of gas } d_1 v_1 = d_2 v_2 \quad (\text{i.e, } dv = \text{constant})$$

$V \propto T$ Charles law

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

$$T \propto \frac{1}{d} \qquad d_1\ T_1 = d_2\ T_2$$

Boyles law $V \propto \frac{1}{P}$, $V \propto \frac{1}{d}$ i.e. $d \propto P$ and from above $d \propto \frac{1}{T}$

$$d \propto \frac{P}{T} \quad \text{i..e,} \quad \frac{dT}{P} \text{ is constant}$$

Thus, $\boxed{\frac{d_1\ T_1}{P_1} = \frac{d_2\ T_2}{P_2}}$

83. Vapour pressure of a liquid at any given temperature is the pressure exerted by the vapour of a liquid when liquid & vapour are is a state of equilibrium.

$$\text{Liquid} \rightleftharpoons \text{Vapour}.$$

Boiling point of a liquid is the temperature when vapour pressure becomes equal to atmospheric pressure. Higher is the V.P. lower is the boiling point.

e.g. (i) dieyhyl ether has lower b.p. (308 K) than water (373 K)

(ii) Boiling temperature increases in a pressure cooker as the ease of vaporisation decreases under higher pressure, less is the v.p. higher is the boiling point.

Melting point. It is the temperature at which both solid and liquid have identical vapour pressure.

$$\text{solid} \rightleftharpoons \text{liquid}$$

Unit - VI

84. First law of thermodynamics (conservation of energy)

– Meyer and Helmholtz

Statement -

(i) Energy can neither be created nor destroyed but can be transformed from one form to another.

(ii) It is impossible to construct or perpetual motion machine which could produce work without consumption of energy.

(iii) The charge in inter elergy in cyclic process is zero. Mathematical formulation of the law :

$$\boxed{\begin{matrix} U_1 = \text{internal energy} \\ V_1 = \text{volume of the gas} \\ n_1 = \text{no of moles} \end{matrix}} \xrightarrow[\text{at pressure (P)}]{+Q} \boxed{\begin{matrix} U_2 \\ V_2 \\ n_2 \end{matrix}}$$

Initial state A   Final State – B

On absorption of Q amount of heat from the surrounding, state changs for A to B with change in internal energy, $U_1 \longrightarrow U_2$, $v_1 \longrightarrow v_2$ and $n_1 \longrightarrow n_2$ at a given pressure 'P'.

The amount of heat absorbed is used in two ways :

(i) It changes the internal energy of the system from U_1 to U_2, change of internal energy $\Delta U = U_2 - U_1$.

(ii) It does some work on the system by consuming rest amount of heat energy. Hence Q = change in inter energy + work done.

Work done is work of compression i..e, $P(v_2 - v_1) = P\ \Delta V = \Delta n\, RT$

Δn is the change of no. of moles of gases.

Thus, $Q = \Delta U + \Delta n\, RT$

For infinitesimal amount of heat absorbed by the system.

dq = du + dw, the differential form of first law of thermodynamics.

$$\boxed{\begin{matrix} \text{Intital} & \\ & U_1 \\ \text{state} & \end{matrix}} \xrightarrow[\text{Heat absorbed}]{+q} \boxed{U_1 + q} \xrightarrow[\text{work done on the system}]{+w} \boxed{\begin{matrix} U_1 + q + w \\ = U_2 \end{matrix}} \qquad \boxed{\Delta U = q + w}$$

Final state $\Delta U = q + P\Delta V$

If work done is negative i.e., work of expansion, work done by the system.

$$w = -P(\Delta V)$$

$$\Delta U = q - P\Delta V$$

$$\boxed{q = \Delta U + P\Delta V}$$

Thus, (i) neither 'q' nor 'w' is a state function, yet quantity $q + w = \Delta U$, is a state function.

(ii) For an ideal gas undergoing isothermal change, $\Delta U = 0$, hence $q = -w$

heat absorbed by the system is work done by the system.

(iii) For an adiabatic system $q = 0$,

$$Q = \Delta U + \Delta nRT = 0$$

$\Delta U = -\Delta n\ RT(w)$ work done is equal to decrease in internal energy.

(iv) For cyclic system, $\Delta U = 0$, $q = P\Delta V$ (work of expansion)

Heat absorbed by the system is used only in doing work of expansion.

(v) On no change in volume, no mechanical work, $q = \Delta U$, Heat absorbed by the system is used only in raising the internal energy. This is also applicable to the case of solid or liquid state where the change in volume is negligible $(\Delta V = 0)$.

Significance of 1st law :

(i) It helps to establish relationship between the heat absorbed by the system and work done.

(ii) Work can not be performed by the system without consuming energy.

(iii) Work and energy can't be produced from nothing.

85. **Heat capacity :** Heat capacity is the capacity of a system to absorb heat and preserve the same within: The increase of temperature is proportional to the heat transferred.

$$q = \text{coefficient} \bullet \Delta T$$

The magnitude of the coefficient depends on size, composition and nature of the system. It can also be written as $q = C\Delta T$. The coefficient 'C' is called heat capacity.

When 'C' is large, a given amount of heat results only a small temperature rise. Water has large heat capacity (C) i..e, a lot of energy is needed to raise its temperature (say by 1^0)

Molar heat capacity (C_m) : The quantity of heat needed to raise the temperature of one mole of substance by one degree celsius (or Kelvin).

Specific heat / Specific heat capacity is the heat required by unit mass of a substance by one degree rise of temperature.

Thus, for finding out the heat (q) required to raise the temperature of a sample, we multiply the specific heat of the substance, (C) by mass (m) and change in temperature, ΔT

i.e., $\boxed{q = m.c.\Delta T}$

Types of heat capacity :

(i) Molar heat capacity at constant volume (C_V)

By definition, $C = \frac{dq}{dT}$, from first law, $dq = dU + Pdv$

$$C = \frac{dU + Pdv}{dT}$$

at constant volume, dv = 0

$$\text{So } C_V = \frac{dU}{dT} = \left(\frac{\partial u}{\partial T}\right)_V$$

C_V is defined as the rate of change of internal energy with temperature.

(ii) Molar heat capacity at constant pressure (C_P)

$$C = \frac{dU + Pdv}{dT}$$

We know $dH = dU + Pdv$ $\qquad (\because \; H = U + Pv)$

$$\text{Thus } C_P = \frac{dH}{\partial T} = \left(\frac{\partial H}{\partial T}\right)_P$$

Relationship between C_P and C_V

Molar heat capacity at constant pressure (C_P) is always greater than that at constant volume (C_V). This is because where gas is heated at constant volume the heat absorbed by the gas is used to increase the internal energy only without doing any external work (Pdv = 0), But heating the gas at constant pressure results in doing external work and also raising the internal energy of the system.

The difference between C_P and C_V, $(C_P - C_V)$ gives the work done by 1 mole of the gas during expansion when heated through 1^0C at constant pressure.

$$C_P = \frac{\Delta H}{dt}, \quad \Delta H = C_P\, dT$$

$$C_V = \frac{\Delta U}{dt}, \quad \Delta U = C_V\, dT \qquad \therefore \qquad H = U + PV$$

for 1 mole of ideal gas, $\Delta H = \Delta U + \Delta(PV)$

$$\Delta H = \Delta U + \Delta(RT)$$

$$\Delta H = \Delta U + R\,\Delta T$$

Substituting ΔH and ΔU values,

$$C_P\, dT = C_v dT + R\, dT$$

$$\Rightarrow \quad \boxed{(C_P - C_V) = R}$$

R is nothing but the work done per mole of ideal gas per degree rise of fall of temperature.

86. Heat of reaction $(\Delta_r H)$: It is the quantity of heat evolved or absorbed when the number of gram molecules of the substances taking part in the chemical reaction as indicated in the balanced thermochemical reaction has completely reacted giving the products. i.e.,

$\Delta_r H$ = (Heat content of products) – (Heat content of the reactants)

$$= \sum H_P - \sum H_r$$

e.g. $C_{(S)} + O_2(g) \longrightarrow CO_2(g) + 393.5\,kJ$ (exothermic)

$\Delta_r H = (H_{CO_2}) - [H_{(C)} + H_{(O_2)}]$, heat content of the elements taken as zero.

$$\Delta_r H = \Delta_f\ H_{(CO_2)} = -393.5\ kJ$$

When the process is carried out at 298 K and 1 atm. pressure the enthalpy change ΔH of a reaction is called standard enthalpy change and represented as ΔH^0.

Factors affecting enthalpy f reaction, $(\Delta_r H)$.

(i) Amount of reactants : More is the amount of reactant (s) more is the heat evolved, hence it's an extensive property.

(ii) Physical state of the reactants & products : The change of physical state is also accompanied by heat change.

(iii) The value of ΔH for a reaction changes with temperature. (Kirchoff equation)

* **Enthalpy change of reaction at constant volume, Q_v**

Let us consider a chemical reaction taking place at constant temperature and constant volume.

In such case, W = O

As $Q = \Delta U + W$, So $\Delta U = Q_v$

* Enthalpy change of reaction at constant pressure (Q_P)

we know $\Delta H = \Delta U + P\Delta V$

If Q_P be the amount of heat exchanged in above reaction i.e., enthalpy change of reaction at constant pressure,

$$\text{Then},\ \Delta H = \boxed{Q_P = Q_V + P\Delta V}$$

$$Q_P = Q_v + \Delta n_g\ RT$$

Δn_g is the difference in number of moles of the gaseous products and reactants.

e.g. Dissociation of NH_3 (g)

$$2\,NH_3(g) \longrightarrow N_2(g) + 3H_2(g) \qquad \Delta n_g = (1+3) - 2 = 2$$

So $Q_P = Q_V + 2RT$

87. $\Delta_r H = \left(\Delta_f H_{CO_2} + 4.\Delta f\ H_{HCl}\right) - \left(\Delta_f H_{CCl_4} + 2\,\Delta_f H_{H_2O}\right)$

$$= (-94.1 + 4\times -22.1) - (-25.5 + 2x - 57.8)$$

$$= (-94.1 - 88.4) - (-25.8 - 115.6)$$

$$= -182.5 + 141.4$$

$$= -41.1\ \text{K Cal} \qquad 1\ \text{K cal} = 4.184\ \text{kJ}$$

$$= -41.1 \times 4.184 = -172.\ \text{kJ}$$

$$CCl_4(g) + 2H_2O(g) \longrightarrow CO_2(g) + 4HCl(g) \quad \Delta n = (1+4) - (1+2) = 2$$

$$\Delta H = \Delta U + \Delta n_g\ RT \qquad R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$$

$$-172.\ \text{kJ} = \Delta U + (2 \times 0.008314 \times 298)$$

$$= \Delta U + 4.955\ \text{kJ}$$

$$\therefore \quad \Delta U = -(172. + 4.955)\,\text{kJ} \equiv 176.955\ \text{kJ}$$

88. **Enthalpy of neutralisation .** It is the enthalpy change occurs when one gram equivalent of an acid is added to 1 gram equivalent of a base or viceversa in a fairly dilute solution.

strong acid + strong base $\longrightarrow$ salt + water + 57.1 kJ

i.e., $H^+(aq) + OH^-(Aq) \longrightarrow H_2O + 57.1$ kJ

i.e., enthalpy of formation of water from H^+ and H^- ions. The heat released is reduced depending on the strength of the acid, base as the degree of ionisation is reduced.

e.g., NH_4OH (weak base) + HCl (strong acid) $\longrightarrow$ NH_4Cl + 51.9 kJ

NH_4OH (weak base) + CH_3COOH (weak acid) $\longrightarrow$ CH_3COONH_4 + 12.kJ.

Problem : 200 ml of 0.2 M HCl $= \dfrac{200 \times 0.2}{1000} = 0.04$ mole of HCl

100 ml of 0.1 M KOH $= \dfrac{100 \times 0.1}{1000} = 0.01$ mole of KOH

0.01 mole of KOH $\equiv$ 0.01 mole of HCl

Residual HCl = 0.03 mole, hence the solution is acidic.

The heat released $= -57.1 \times 0.01 = 0.571\,\text{kJ}$

$= 571\ \text{J}$

89. Theoretical energy from bond energy data.

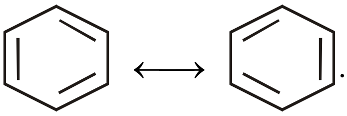
Benzene has (i) 3 C–C single bonds
(ii) 3 C=C double bonds
(iii) 6 C–H bonds

$$\Delta_d H = (3\times 347.3) + (3\times 615) + (6\times 416.2)\,\text{kJ}$$

$= 5384.1\ \text{kJ/mol}$ (Bond dissociation energy, theoretical)

Experimental dissociation energy $= +5535.1\ \text{kJ/mol}$

Reosnance energy $= (5384.1 - 5535.1) = 151.\text{kJ/mol}$

i.e., benzene is stabilised by 151 kJ/mole due to resonance i.e, .

Benzene molecule is more stable than any one of the two Kekule structure. Thus the resonance hybrid is .

90. Third law of thermodynamics : According to Nernst heat theorem ΔG and ΔH of a system becomes equal to each other at absolute zero temperature and approach each other asymptotically.

When $T = 0$, $\Delta G = \Delta H$

$$\text{i.e,}\ \underset{T\to 0}{\text{Lt}}\left[\frac{\partial\,\Delta H}{\partial t}\right]_p = 0,\ \underset{T\to 0}{\text{Lt}}\left[\frac{\partial\,(\Delta G)}{\partial T}\right]_p = 0 \qquad \text{i.e,}\ \underset{T\to 0}{\text{Lt}}\ \Delta S = 0 \qquad [\because\ dG = VdP - SdT]$$

$$\Rightarrow\ \underset{T\to 0}{\text{Lt}}\ \Delta C_p = 0$$

Statement of 3rd law : At absolute zero of temperature the entropy of every substance may become zero and it does become zero in case of perfectly crystalline solid.

91. (a) Enthalpy of formation, $\Delta_f H$ is the enthalpy change for formation of 1 mole of substance from its constituting elements. Let us write two thermochemical reactions,

Example - 1 $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O(\ell)\ \ \Delta H^0 = -285.8\,\text{kJ mole}$

$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O(\ell)\ \ \Delta H^0 = -571.6\,\text{kJ mole}$

The first case is heat of formation but second case is enthalpy of reaction.

Example - 2 $CaO(s) + CO_2(g) \longrightarrow CaCO_3(s),\ \Delta H = -178.3\,\text{kJ/mol}$

Here the heat released is called heat of reaction, ΔrH but not heat of formation, $\Delta_f H$ as $CaCO_3$(s) is not formed from the constituting elements.

In some cases heat of formation is same as heat of reaction, on formation of one mole of product.

e.g. $C_{(s)} + O_2(g) \longrightarrow CO_2(g)$

Here, $\Delta_r H = \Delta_f H$, as one mole of CO_2 (g) is formed from its constituting elements C and O_2; the heat content of these elements taken as zero.

$$\Delta_r H = \sum H_{\text{products}} - \sum H_{\text{reactants}}$$

(b) $CH_3COOH \longrightarrow CH_3COO^- + H^+ \qquad \Delta H = +1.9 \text{ kJ Mol}^{-1}$

$H^+ + OH^- \longrightarrow H_2O \qquad \Delta H = -57.1 \text{ kJ Mol}^{-1}$

$CH_3COOH + OH^- \longrightarrow CH_3CO\bar{O} + H_2O \qquad \Delta H = ?$

Heat of neutralisation = (1.9 - 57.1) kJ

= – 55.2 kJ / mol.

92. **Enthalpy of atomisation :**

It is defined as the enthalpy change for conversion of compounds and elements into free atoms in gaseous state. Here ΔH is positive i.e., endothermic.

e.g $H_{2(g)} \longrightarrow 2H_{(g)} \qquad \Delta H = +435 \text{ kJ / mol}$

It is the energy required in breaking 1 mole of bonds completely to get atoms in gaseous phase. Here (in case of H_2) the enthalpy atomisation is also same as bond dissociation enthalpy.

$CH_{4(g)} \longrightarrow C_{(g)} + 4H_{(g)} \qquad \Delta H = +1665 \text{ kJ / mol}$

93. ΔG and temperature. (Gibbs – Helmboltz equation)

$$G = H - TS$$

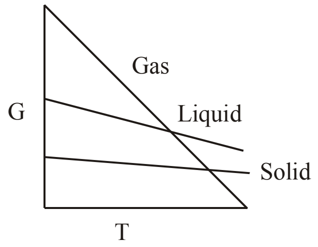
$$dG = dH - TdS - SdT$$

$$dG = d\cancel{U} + pd\cancel{V} + Vdp - T\cancel{d}s - SdT \qquad (\because H = U + pV)$$

$$\Rightarrow dG = Vdp - SdT \qquad TdS = dU + PdV$$

$$\left(\frac{\partial G}{\partial T}\right)_P = -S, \quad \left[\frac{\partial(\Delta G)}{\partial T}\right]_P = -(\Delta S)$$

i.e, S is positive then 'G' decreases when temperature increases at constant pressure.



$S_G > S_L > S_S$

More positive is the entropy more sharpen in the G.T. plot i.e., more -ve is the free energy.

Unit - VII

94. **Law of mass action (Norwegian chemists, Goldberg - Waage)**

Statement : Rate at which a substance reacts is proportional to its molar concentration (active mass) and the rate of chemical reaction is proportional to the product of the molar concentration of the reactants raised to the power equal to the stoichiometric coefficients.

$$A \longrightarrow \text{product}$$

rate of reaction $\propto [A]$

$$2A \longrightarrow \text{product}$$

$$A + A \longrightarrow \text{product}$$

rate of reaction $\propto [A]\,[A]$

$\propto [A]^2$ (Coefficient becomes the power of the concentration)

In ganeral :

$$aA + bB \longrightarrow \text{products}$$

According to the law of mass action,

rate of reaction $\propto [A]^a\,[B]^b$

$= k\,[A]^a[B]^b$

k is a proportionalilty constant called **rate constant.**

Application : Derivation of equilibrium constant, K

$$aA + bB \underset{k_2}{\overset{k_1}{\rightleftharpoons}} cC + dD$$

$k_1 =$ rate constant for forward

$k_2 =$ rate constant for backward reaction.

Now $r_f = k_1[A]^a[B]^b, \quad r_b = k_2[C]^c[D]^d$

At equilibrium $r_f = r_b$

Now $\dfrac{k_1}{k_2} = K_c = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$

* Law of chemical equilibrium

(Definition of K) : It is defined as, the ratio of mathematical product of molar concentration of the products to that of reactants with each concentration term is raised to the power by its coefficient in balanced chemical equation is a constant quantity at a given temperature.

→ If it's a gaseous homogeneous equilibrium the active mass can be expressed in partial pressure form.

So $K_p = \dfrac{p_c^c\ p_D^d}{p_A^a\ p_B^b}$

→ If the active mass expressed in mole fraction form.

$$K_x = \frac{X_C^c . X_D^d}{X_A^a . X_B^b}$$

Relationship between K_p & K_C **:** pV = nRT

$$p = \frac{n}{v}RT, \quad P = CRT$$

$$p_A = [A]RT, \quad p_A^a = [A]^a[RT]^a$$

Now $K_p = K_c.(RT)^{\Delta n}$

We know p = mole fraction × total pressure, p = partial pressure

$p_A = X_A.P, \; p_B = X.p.......$

$$K_P = \frac{X_C^c.p^c.X_D^d.p^d}{X^a.p^a.X_B^b.p^b} = K_X.\frac{p^{c+d}}{p^{a+b}} = K_X.p^{\Delta n} \quad \text{as} \quad \Delta n = (c+d)-(a+b)$$

95. $.K_p = K_C(RT)^{\Delta n} = K_x.p^{\Delta n}$

When $\Delta n = 0, \qquad K_P = K_C = K_X$

e.g. $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI$ (Homogeneous equilibrium)

$C_{(s)} + O_2(g) \rightleftharpoons CO_2(g)$ (Heterogeneous equilibrium)

96. Haber's synthesis of NH_3 :

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g), \; \Delta n = -2, \quad \Delta H = -92\,kJ$$

Factors :

- **(i)** Temperature : The reaction is exothermic but the reaction proceeds at 500^0 C i.e., an optimum temperature at which activation energy is surpassed favouring better yield of ammonia.
- **(ii)** **Pressure** : As per Le-chatelier's principle the reaction proceeds in a direction of decrease in volume or number of moles with rise of pressure. Here, the forward reaction proceeds with decrease in pressure i.e, very high pressure of 200-900 atm favours better yield of NH_3.
- **(iii)** **Concentration :** High concentration of pure and dry N_2(g) and H_2(g) favours forward reaction.

$$K_C = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Increase in concentration of N_2 and H_2 increases dr. value. As K_C is constant, the concentration of NH_3 must have to rise. Thus, the equilibrium shifts to right.

- (iv) **Catalyst :** Presence of catalyst does not alter 'K' value but it helps to attain the equilibrium quickly if it's a positive catalyst. Thus, finely divided iron as catalyst and Molybdenum as promoter lowers the activation energy favouring kinetics of the reaction.

97. Law of chemical equilibrium

For a general reaction, $aA + bB \rightleftharpoons cC + dD$,

$$K = \frac{[C]^c\,[d]^d}{[A]^a\,[B]^b}, \text{ the equilibrium constant.}$$

This mathematical equation is called law of chemical equilibrium :

The ratio of the mathematical product of the molar concentration of the products (resultants) to those of molar concentration of the reactants all raised to the power equal to their stoichimetric coefficients

$$NH_4HS_{(S)} \rightleftharpoons NH_{3(g)} + H_2S_{(g)}$$

We know $K_P = K_C\left(RT\right)^{\Delta n}$ $\Delta n = 2$

$= K_C (RT)^2$

$K_P = P_{NH_3} \cdot P_{H_2S}$, $K_C = \left[NH_3\right]\left[H_2S\right]$

98. Characteristics of equilibrium constant (K)

(i) It has a definite value for every reaction at a given temperature. However, it changes with temperature.

(ii) It does not depend on initial concentration / pressure / mole fraction but those at equilibrium state.

(iii) K value does not change with presence of catalyst, as catalyst affect the rate constants (k_1 and k_2) identically.

(iv) The equilibrium constant value is inversed if the equilibrium is reversed.

i.e., $H_2 + I_2 \rightleftharpoons 2HI$

$K_1 = [HI]^2 / [H_2][I_2]$

Equilibrium approached from backward reaction :

$2HI \rightleftharpoons H_2 + I_2$

$K_2 = \frac{[H_2][I_2]}{[HI]^2}$ Now $K_1 = \frac{1}{K_2}$

(v) The value of K predicts the extent of reaction. Higher is the 'K' value greater is the product i.e., reaction proceeds more in forward direction.

(vi) Unit of K. Unit varies from one to another.

e.g. $H_2 + I_2 = 2HI$ $\Delta n = 0$

K_c or K_p or K_x has no unit

$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$ $\Delta n = 1$

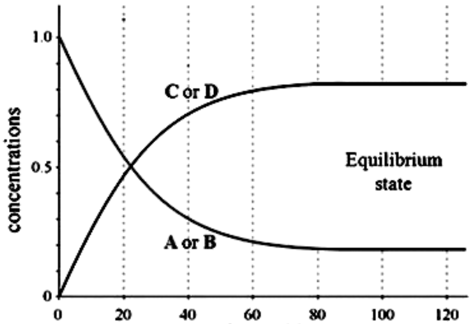
$K_c = mol/l$

i.e. in general unit $K_c = (mol / L)^{\Delta n}$

99. Characteristics / Features of equilibrium :

Chemical equilibrium and rate of reaction

$A + B \rightleftharpoons C + D$



Attainment of chemical equilibrium

Equilibrium (constancy of composition of A, B, C and D after some time)

Graph showing Attainment of equilibrium :

Characteristics of equilibrium : / Factors affecting equilibrium.

→ The state of equilibrium is possible only when the equilibrium is carried in closed vessel (for reactions involving gases, homogeneous or heterogeneous equilibrium)

→ At equilibrium the property of the system like concentration (intensity of colour), temperature and pressure remain unchanged.

→ Since equilibrium can shift towards product or towards reactant, it's dynamic in nature.

→ State of equilibrium can be approached from either side.

→ At equilibrium both opposing rates are identical. ($r_f = r_b$).

→ Catalyst if added never affect the state of equilibrium.

→ Addition of inert gas also affect the state of some equilibrium.

100. $2H_2O_{(g)} \rightleftharpoons 2H_{2(g)} + O_{2(g)} \ K_1$

$2CO_{2(g)} \rightleftharpoons 2CO_{(g)} + O_{2(g)} \ K_2$

$$K_1 = \frac{[H_2]^2 [O_2]}{[H_2O]^2} = 2.1\times10^{-3}$$

$$K_2 = \frac{[CO]^2 [O_2]}{[CO_2]^2} = 1.4\times10^{=132}$$

$CO_{2(g)} + H_{2(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)}$

$$K = \frac{[H_2O][CO]}{[CO_2][H_2]}. \qquad \frac{K_2}{K_1} = \frac{[CO]^2[O_2]}{[CO_2]^2} \times \frac{[H_2O]^2}{[H_2]^2 [O_2]}$$

$$K = \sqrt{\frac{K_2}{K_1}} = \frac{[H_2O][CO]}{[CO_2][H_2]} = \sqrt{\frac{1.4\times10^{-12}}{2.1\times10^{-3}}} = 2.58.$$

101.

$$\begin{array}{lccc} & PCl_5\ (g) \rightleftharpoons & PCl_{3(g)} & + Cl_{2(g)} \\ & 1\ \text{mol} & 0 & 0 \\ & 1-x & x & x \\ \text{Mole fraction} & \frac{1-x}{1+x} & \frac{x}{1+x} & \frac{x}{1+x} \quad \text{total number of moles } 1+x \\ \text{Partial pressure} & \frac{1-x}{1+x}P & \frac{x}{1+x}.P & \frac{x}{1+x}P \end{array}$$

$$K_P = \frac{P_{PCl_3} \cdot P_{Cl_2}}{P_{PCl_5}} = \frac{\frac{x^2\ P^2}{(1+x)^2}}{\frac{1-x}{1+x}P} = \frac{x^2P}{1-x^2}$$

As x is less $1-x^2$ is taken to be one.

Thus $x^2 p = \text{constant}$. $x \propto \frac{1}{\sqrt{P}}$

On increase of external pressure (P) the degree of de-composition decreases i.e., the equilibrium shift towards left. As per Lechatelier's principle on increase of pressure the equilibrium shifts in a direction of decrease of volume/no. of moles.

102. $NH_4HS_{(S)} \rightleftharpoons NH_{3(g)} + H_2S_{(g)}$

Initial : 1 0 0

At eqp : $1 - x$ $x/2x$ $x/2x$, total number of moles of gases = 2x

$$K_P = p_{NH_3} \cdot p_{H_2S} = \frac{x^2}{(2x)^2} = \frac{1}{4} \qquad \text{pressure} = 1\text{ atm.}$$

103. **Bronsted - Lowry acids and bases :** (Proton transfer theory)/ conjugate acid-base theory

Acid – proton donor

Base – proton acceptor

$$\underset{\text{acid}}{HCl} + \underset{\text{base}}{H_2O} \rightleftharpoons \underset{\text{conj acid}}{H_3O^+} + \underset{\text{conj base}}{Cl^-}$$

Acid – H⁺ = conj. base

Base + H⁺ = conj acid

→ **Features of this theory :**

(i) Acid base pairs that differ only by one proton are called conjugate acid base pairs.

Cl^- is the conjugate base HCl (acid)

(ii) Not only molecules but also ions can act as acid and base.

Acid
- molecule CH_3COOH, HCl, HNO_3 (proton donors)
- ions NH_4^+, H_3O^+, $NH_4^+ \longrightarrow NH_3 + H^+$
  $H_3O^+ \longrightarrow H_2O + H^+$

Base
- molecules $NaOH, NH_4OH$
- ions (all anions) $HCO_3^- \left(HCO_3 + H^+ \rightarrow H_2CO_3\right)$
  $CO_3^= + H^+ \longrightarrow HCO_3^-$

(iii) The conjugate base of a stronger acid is weaker & vice versa.

$$\underset{\text{(strong acid)}}{HClO_4} \longrightarrow \underset{\text{weak base (conj.)}}{ClO_4^- + H^+}$$

$$\underset{\text{(weak acid)}}{CH_3COOH} \longrightarrow \underset{\text{Strong base (conj.)}}{CH_3COO^- + H^+}$$

(iv) All Arrhenius acids are Bronsted acid but Bronsted acids are not Arrhenious acids e.g. NH_4^+ is not as Arrhenius acid. CO_3^{2-} is not a base as per Arrhenins, but it's is a Bronsted base.

$$CO_3^{=} + H^+ \longrightarrow HCO_3^-$$

(v) Neutralisation reaction takes place when acid base reaction takes place in equivalent amounts.

(vi) Acid only behaves acidic in a basic medium & vice versa.
e.g. HCl behaves acidic is water (base) but not in benzene (not a proton acceptor).

(vii) The equilibrium constant values (Ka or K_b) helps to compare the strength and predict weaker & strongeracids.

(viii) Amphoteric nature of H_2O can be explained. $H_2O(Base) + H^+ \longrightarrow H_3O^+$

$$H_2O(acid) \longrightarrow H^+ + OH^-$$

Limitations : Fails to explain the acidic character of CO_2, SO_2, AlF_3, BF_3... and basic character of ZnO, CaO etc.

Merits over Arrhenius theory :

i) H^+ ion of Arrhenius is aqueous solution present as $H_3\overset{+}{O}.(H_2O)_3$ i.e., $H_9O_4^+$

ii) NH_3 can act as base without having OH^- for $NH_3 + H^+ \rightarrow \overset{+}{N}H_4$

(Proton acceptor)

iii) In absence of water acid – base reaction can take place.

e.g. $NH_{3(g)} + HCl_{(g)} \rightarrow NH_4^+ + Cl^-, \left[NH_4Cl_{(S)}\right]$

(Proton acceptor) Proton donor

104. **Lewis theory :** Acid : electron acceptor, Base : electron donor.

→ **Lewis acid :**

(i) simple cations, Cu^{++}, Ag^+, Ca^{2+}

(ii) (electron deficient species). BF_3, $AlCl_3$, $ZnCl_2$, $FeCl_2$.........

(iii) Molecules with multiple bands CO_2, SO_2, SO_3 etc.
greater the s-character greater the acidic nature.

(iv) Molecules with vacant d-orbitals. $SiCl_4$, SiF_4

$$\underset{acid}{SiF_4} + \underset{Base}{2F^-} \longrightarrow [SiF_6]^{2-}$$

→ **Lewis bases :**

(i) All simple anions : CN^-, Cl^-, OH^-

(ii) Molecules with excess pair of e^-(s), $\ddot{O}H_2$, $\ddot{N}H_3$, $R\ddot{N}H_2$, $R\overset{..}{\underset{..}{O}}H$, $R-\overset{..}{\underset{..}{O}}-R$

$$\underset{\text{electron doner}}{R\ddot{N}H_2} + H^+ \longrightarrow R\overset{+}{N}H_3$$

→ **Acid-Base reaction (Lewis)**

$$F_3B + :NH_3 \longrightarrow F_3B \longleftarrow :NH_3$$

electron deficient

→ **Limitations :**

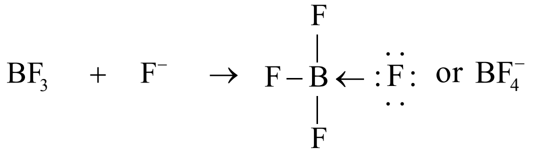
(i) Relative strength can't be predicted as it does not undergo ionisation.

(ii) Acid-base reactions are slow due to formation of dative bond.

(iii) Catalytic property of acids can't be explained by Lewis theory.

(iv) It can't explain the amphoteric nature of H_2O , HCO_3^- ions etc.

105. **Difference between Bronsted concept and Lewis concept of acid - base.**

Bronsted concept	**Lewis concept.**
* Acids are proton donors while bases are proton acceptors. e.g. Acid : $NH_4^+ \longrightarrow NH_3 + H^+$ Base : $Cl^- + H^- \rightarrow HCl$	* Acids are electron deficient i.e. tendency to accept electron, while bases have tendency to donate $\overline{e}$. Acid : $AlCl_3$ ($\overline{e}$ deficient) Base. :NH_3($\overline{e}$ rich)
* Acid-base reaction takes place by transfer of proton. $NH_3 + HCl \longrightarrow NH_4^+ + Cl^-$ Acid, HCl gives proton to NH_3 (Base) ionic bond formation takes place.	* Electrons transfer takes place forming co-ordinate- covalent bond.  $BF_3 + F^- \rightarrow F-\underset{F}{\overset{F}{B}} \leftarrow :\ddot{\underset{..}{F}}:$ or BF_4^- Lewis acid (Lewis base)
* Acid base reactions are fast	* Acid base reactions are slow.
* Such acids are used is different proton catalysed reaction. e.g. hydrolysis of ester.	* Lewis acids can't be used in proton catalysed reaction. However, these acids are used in forming electrophiles. e.g. $Cl-Cl + ACl_3 \rightarrow \overset{+}{Cl} + AlCl_4^-$
* Strength of acid and base can be calculated and compared with other acids comparing their K_a and K_b values respectively.	* Strength can't be found out.
* It can explain amphoteric nature of water (proton donor as well as proton acceptor).	* It fails to explain amphoteric nature of HCO_3^-, H_2O etc.

106. **Factors affecting acid strength.**

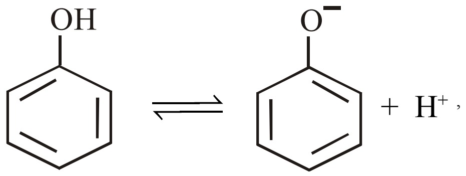
(i) Bond strength of H–A. Weaker is the bond strength or greater is the bond length, stronger is the acid.

e.g. HI > HBr > HCl > HF

(ii) Polarity of H–A bond. If the electronegativity of A increases, the polarity of HA bond increases increasing the acid strength.

$$\overset{2.5}{C}H_4 < \overset{3.0}{N}H_3 < H_2\overset{3.5}{O} < H\overset{4.0}{F}$$

(iii) Stability of A^- (conjugate base) Greater is the stability of the conjugate base greater is the ease of proton release & more is the acidic nature. e.g. acidic nature of phenol and alcohol.



$ROH \rightleftharpoons R\bar{O} + H^+$

less stable
due to + I effect

more stable due
to resonance

(iv) Increase of positive oxidation state :

More is the positive O.S. less is the size & greater is the $\bar{e}$ accepting tendency & greater is the acidic nature.

$$H\overset{+7}{Cl}O_4 > H\overset{+5}{Cl}O_3 > H\overset{+3}{Cl}O_2 > H\overset{+1}{Cl}O$$

In case of identical O.S. greater is the electrongativity greater is the acidic nature.

$$HClO_4 > HBrO_4 > HIO_4$$

Among acids $H\,Sb\,F_6$ is the strongest super acid, $SbF_5 + HF$.

(v) More s-character of multiple bond :

More is the s-character more is tendency to accept $\bar{e}$ & more is the acidic nature. e.g. molecules sp hybridised bond becomes more acidic.

$$\underset{(sp)}{HC \equiv CH} > \underset{(sp^2)}{CH_2 = CH_2} > \underset{sp^3}{CH_3 - CH_3}$$

s-character : 50 % 33.3% 25%

(vi) –I - effect : Presence of $\bar{e}$ withdrawing group causes $\bar{e}$ deficiency increasing tendency to attract $\bar{e}$ & more is the acidic nature.

$$ClCH_2COOH > CH_3COOH$$

(vii) Value of Ka : On theoretical consideration, higher is the Ka (Dissociation constant of acid) value greater is the acidic nature.

107. Buffer solution :

Statement : It is a solution, whose pH value does not change appreciably by addition of a small amount of either acid or alkali from outside i.e. buffer solution preserves its pH value.

→ **Types of buffer :** (a) simple buffer (b) Mixed buffer

Simple buffer : Salts of weak acid & weak base

e.g. NH_4CN, CH_3COONH_4

Mixed buffer : (i) acidic buffer - solution containing weak acid and its salt (conjugate base)

e.g. one mole CH_3COOH + one mole of CH_3COONa (CH_3COO^-) having pH < 7 (4.75).

(ii) Basic buffer - equimolar mixture of weak base & its salt (Conjugate acid)

One mole of NH_4OH + one mole of NH_4Cl (NH_4^+), pH>7 (9.25)

→ **Buffer action :** pH of the buffer solution almost remains unchanged on addition of acid or alkali.

(i) **Simple buffer** $CH_3C\bar{O}O\overset{+}{N}H_4$ ·

$$CH_3CO\bar{O} + \overset{+}{N}H_4 \begin{cases} \xrightarrow{H^+\text{added}} CH_3CO\bar{O}(aq) + H^+(aq) \rightarrow CH_3COOH(aq) \\ \xrightarrow{O\bar{H}\text{added}} NH_4^+ + O\bar{H} \rightarrow NH_4OH(\text{weak base}) \end{cases}$$

On adding few drops of acid (H^+) or alkali (OH^-) weak acids, CH_3COOH or weak base, NH_4OH are formed respectively.

The formation of these in the solution never changes the pH appreciably as these are weak with less K_a and K_b values.

(ii) **Acidic buffer :**

$$\underbrace{\begin{matrix} CH_3COOH + CH_3CO\bar{O} \\ HA + A^- \end{matrix}} \begin{cases} \text{Drops } H^+ \text{ (acid) added } CH_3CO\bar{O} + H^+ \longrightarrow \underset{\text{weak acid}}{CH_3COOH} \\ \text{drops of } OH^- \text{ (alkali added from outside)} \\ CH_3COOH + OH^- \longrightarrow \underbrace{CH_3COO^- + H_2O}_{\text{Never affect pH}} \end{cases}$$

(iii) Basic buffer, $\underbrace{NH_4OH + NH_4^+}_{\text{solution}}$

$$\longrightarrow \begin{cases} \text{Drops of } H^+ \text{ (acid) added } NH_4OH + H^+ \longrightarrow \underbrace{\underset{\text{weak acid}}{NH_4^+ + H_2O}}_{\text{Never affect pH}} \\ \text{drops of } OH^- \text{ (base / alkali) added} \\ NH_4^+ + OH^- \longrightarrow NH_4OH(\text{weak base}) \\ \text{Never increases pH value appreciably} \end{cases}$$

108. pH of Buffer solution : (Henderson - Hassel balch equation)
Calculation of pH of buffer solution

(i) Acidic buffer $\underset{(\text{Weak acid})}{HA} \rightleftharpoons H^+ + A^-$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$\Rightarrow [H^+] = K_a \times \frac{[HA]}{[A^-]} \quad \Rightarrow -\log[H^+] = -\log K_a - \log\frac{[HA]}{[A^-]}$$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

i..e, $\boxed{pH = pK_a + \log\frac{[salt]}{[acid]}}$

$[A^-]$ = mostly the concentration of salt

[HA] = conc^n of weak acid never dissociates much to increase the conc^n of $[A^-]$ or decrease the conc^n [HA]

This is called Henderson - Hasselbalch equation or simply Henderson equation.

For equimolar solution i.e. [salt] = [acid]

$pH = pK_a$

For alkaline buffer, Henderson's equation will be -

$$\text{Derive} \rightarrow \ \text{pOH} = \text{pK}_b + \log\frac{[\text{Salt}]}{[\text{Base}]}$$

$$\text{pH} = 14 - \text{pOH} = 14 - \text{pK}_b - \log\frac{[\text{Salt}]}{[\text{Base}]}$$

→ **Applications :**

* In qualitative and quantitative analysis.

* Industrial process such as manufacture of paper, dyes, inks ....

* In Agriculture, medicine etc.

109. Classification of salt :

A salt is formed by replacement of proton of the acid by metal ion or cation like NH_4^+ etc.

General classification -

(a) Simple salt –

(i) Normal salt, NaCl

(ii) Acid salt $NaHSO_4$ Partial neutralisation of poly basic acid.

$$\underset{}{NaHSO_4} + NaOH \longrightarrow \underset{\text{normal salt}}{Na_2SO_4} + H_2O$$

(iii) Basic salt (to react with acid forming normal salt)

$$\underset{\text{Basic salt}}{BiOCl} + 2HCl \rightarrow \underset{\text{Normal salt}}{BiCl_3} + H_2O$$

(b) Complex salt . $K_4[Fe(CN)_6]$

Fe^{2+} and CN^- lose their identity.

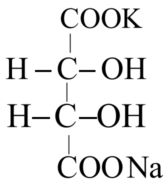
(c) Double salt. $K_2SO_4.\ Al_2(SO_4)_3.\ 24H_2O$ Potash alum

$(NH_4)_2\ SO_4.FeSO_4.6H_2O$ (Mohr. salt)

(d) Mixed salt $Na(NH_4)HPO_4$ – micro cosmic salt.

Replacement of H^+ by different cations of paoly basic acid.

Rochelle salt.



$$\begin{array}{c} COOK \\ | \\ H-C-OH \\ | \\ H-C-OH \\ | \\ COONa \end{array}$$

(Sodium potassium tartarate)

→ **Classification on the basic of hydrolysis :**

(i) Salts of strong acid strong base, K_2SO_4

(ii) Salts of strong acid and weak base, NH_4Cl

(iii) Salts of weak acid and strong base, CH_3COONa

(iv) Salts weak acid & weak base, CH_3COONH_4

110. NH_4Cl is a salt of strong acid (HCl) and weak base (NH_4OH).
So on hydrolysis it becomes acidic i.e., pH < 7.

$$NH_4^+ + Cl^- + H_2O \rightleftharpoons NH_4OH + H^+ + Cl^-$$

i.e. $NH_4^+ + H_2O \rightleftharpoons NH_4OH + H^+$, it shows the increase of H^+ ions on hydrolysis of NH_4^+.

On calculation, $\text{pH} = 7 - \frac{1}{2}(\text{pK}_b + \log C)$

K_b = dissociation constant of weak base.
C = concentration of the oslution.

As pH is less than 7, the solution is acidic. on the other hand, CH_3COONa is a salt of weak acid, CH_3COOH and strongbase, NaOH. So on hydrolysis the $[OH^-]$ increases rendering the solution alkaline.

$$CH_3COO^- + Na^+ + H_2O \rightleftharpoons CH_3COOH + Na^+ + OH^-$$

$$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$$

The increased concentration of OH^- increases the pH value rendering the solution alkaline.

i.e., $pH = 7 + \frac{1}{2}(pKa + \log C) \Rightarrow pH > 7$

111. Ammonium acetate is a salt of weak base, NH_4OH and weak acid, CH_3COOH. On hydrolysis.

$$CH_3COO^- + NH_4^+ + H_2O \rightleftharpoons CH_3COOH + NH_4OH$$

$$K_h = \frac{[CH_3COOH][NH_4OH]}{[CH_3COO^-][NH_4^+]}, \quad = \frac{K_w}{K_a \cdot K_b}$$

As $K_w = [H^+][OH^-]$, $Ka = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$, $K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]}$

$$\begin{array}{cccc} CH_3COO^- + & NH_4^- + H_2O & \rightleftharpoons CH_3COOH + & NH_4OH \\ C & C & O & O \\ C - Ch & C - Ch & Ch & Ch \end{array}$$

h = degree of hydrolysis (per mole)

$$K_h = \frac{\cancel{C}h \times \cancel{C}h}{\cancel{C^2}(1-h)^2} = \frac{h^2}{(1-h)^2}$$

Considering $1 - h \simeq 1$, $K_h \simeq h^2$

$$h = \sqrt{\frac{K_w}{K_a . K_b}}$$

We known $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

$$Ka = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}.$$

Substituting the respective concentration, $K_a = \frac{C(1-h).[H^+]}{C.h} = \frac{[H^+]}{h}$

$$[H^+] = Ka \times h \quad = Ka \times \sqrt{\frac{Kw}{Ka.Kb}}$$

$$[H^+] = \sqrt{\frac{Ka. Kw}{K_b}}$$

Taking log and multiplying -ve through out

$$-\log\left[H^+\right] = -\frac{1}{2}\log Kw - \frac{1}{2}\log K_a + \frac{1}{2}\log K_b$$

$$\boxed{pH = 7 + \frac{1}{2}\left(\log K_b - \log Ka\right)}$$

* **Observation :**

(i) $Ka > K_b$ $pH < 7$ (acidic).

i.e, $Kb > Ka$, $pH > 7$. (alkaline),

i.e, $K_a \simeq Kb$, $pH = 7$ i.e, neutral. If acid and base are of almost identical strength.

(ii) The pH of the solution as well as the degree of hydrolysis of the salt, 'h' is independent of concentration of the salt solution.

112. **Classification of cations into groups** (Qualitative analysis). The cations are classified into different groups which depends on the K_{sp} (solubulity product) values of the respective sparingly soluble salt. This is based on the fact that the ionic product value must have to exceed the K_{sp} value to get precipitate. To get the K_{sp} values appropriately comparable the significance of common ion effect is considered. The presence of common ion supresses the dissociation of the reagent / ion such that the ionic product is so maintained to exceed the K_{sp}, giving the precipitate.

Let us consider different groups :

Group - I radicals $\left(Pb^{2+}, Ag^+, Hg_2^+\right)$

Groupreagent : dil HCl

Precipitate formed : $PbCl_2$, AgCl, Hg_2Cl_2

On adding few drops of dil HCl to solution containing above cations (basic radicals) precipitation of respective chloride occurs as $\left[Ag^+\right]\left[Cl^-\right] > K_{sp}$ for AgCl.

The other cations except group - I never precipitate, as the ionic product never exceeds, K_{sp} values. K_{sp}, solubility products of other chlorides are much higher.

Group - II radicals $\left(Pb^{2+}, Cu^{2+}, H_g^{2+}, Cd^{2+}, As^{3+}, Sb^{3+}, Bi^{3+} \text{ etc.}\right)$

Group reagent	H_2S in presence of dil HCl
Group precipitate :	Sulphides e.g. PbS, CuS, HgS, CdS etc.

$$H_2S \rightleftharpoons 2H^+ + S^{2-}$$

$$HCl \longrightarrow H^+ + Cl^-$$

Here H^+ acts as a common ion which suppresses the disciation of H_2S reducing the concentration of $[S^{2-}]$. The ionic product of $\left[S^{2-}\right]$ and metal ion must have to exceed the lower K_{sp} of sulfides of group II cation.

e.g. $\left[Cu^{2+}\right]\left[S^{2-}\right] > K_{sp}$ for CuS.

Group - III$_A$ $\left[Fe^{3+}, Cr^{3+}, Al^{3+}\right]$

Group reagent. NH_4Cl (s), NH_4OH (aq)

Group precipitate - Hydroxides of Fe^{3+}, Cr^{3+} and Al^{3+}

$$NH_4OH \rightleftharpoons NH_4^+ + OH^-$$

$$NH_4Cl \longrightarrow NH_4^+ + Cl^-$$

Presence of common ion, NH_4^+ reduces the $[OH^-]$

$$\left[Fe^{3+}\right]\left[OH^{-}\right]^{3} > K_{sp} \text{ for } Fe(OH)_3$$

$\left[OH^{-}\right]$ is so managed that it will help to exceed lower K_{sp} values of $Fe(OH)_3$, $Cr(OH)_3$, $Al\ (OH)_3$ but not that higher K_{sp} of $Zn\ (OH)_2$.

Group - III$_B$ $\left(Zn^{2+}, Co^{2+}, Ni^{2+}, Mn^{2+}\right)$

Group reagent : $NH_4Cl_{(s)}$ till saturation
NH_4OH till ammoniacal / alkaline
H_2S (g) till precipitation.

Group precipitate : ZnS (white); CoS, NiS (Black); MnS (buff)
Here $H_2S_{(g)}$ in passed, is alkaline medium as a result $[S^{2-}]$ increases. It is sufficient enough the high or K_{sp} values of respective sulfides.

$$H_2S + 2OH^{-} \longrightarrow 2H_2O + S^{2-}$$

$$\left[Zn^{2+}\right]\left[S^{2-}\right] > K_{sp} \text{ for } ZnS$$

Group IV $\left(Ba^{2+}, Sr^{2+} \text{ and } Ca^{2+}\right)$

Group reagent - NH_4Cl (s), NH_4OH (aq), $(NH_4)_2\ CO_3$ (Aq)
Group - precipitate $BaCO_3$, Sr CO_3, $CaCO_3$

Due to presence of commonion, $N\overset{+}{H}_4$ ion, the ionisation of $(NH_4)_2\ CO_3$ is supressed i.e., $\left[CO_3^{2-}\right]$ is reduced. The lower concentration of CO_3^{2-} is sufficient enough to have the ionic product value which will exceed the lower K_{sp} of carbonates of Ba^{2+}, Sr^{2+} and Ca^{2+} as result respective precipitation occurs.

$$\left[Ba^{2+}\right]\left[CO_3^{2-}\right] > K_{SP} \text{ for } Ba\ CO_3.$$

Group - V $\left(NH_4^{+}, Mg^{2+}, Na^{+}, K^{+}\right)$ **has, no such common group reagent and individually tested.**

113. 200 ml of 0.1 M $HCl = \dfrac{200 \times .01}{1000} = 0.02$ moles HCl.

$$300 \text{ ml of } 0.1 \text{ M NaOH} = \frac{300 \times 0.1}{1000} = 0.03 \text{ mole NaOH}$$

0.02 mole HCl $\equiv$ 0.02 mole of NaOH
Residual NaOH = 0.01 mol

Total volume = 500 ml   Molarity of the solution $= 0.01 \times \dfrac{1000}{500}$

500 ml of 0.02 M $= 0.02$ M

Now the 500 ml of 0.02 M is made upto 1000 ml
So the molarity of dilute solution $500 \times 0.02 = 1000 \times \text{Molarity}$

$$\Rightarrow \quad \text{Molarity} = 0.01 \text{ M}$$

or. 0.01 mole NaOH present in 1L i.e., molarity = 0.01 M

$$\left[OH^{-}\right] = 0.01 \text{ M}$$
$$= 10^{-2} \text{ M}$$

pOH = 2, pH = 12

114. **Reaction quotient and equilibrium constant :**

Definition : For the reaction , $aA + bB \rightleftharpoons cC + dD$

At any stage of reaction (not at equilibrium), the ratio, $\frac{[C]^c[D]^d}{[A]^a[B]^b}$ is termed as reaction quotient, Q_C

similarly of gaseas equilibrium $Q_p = \frac{p_C^c \cdot p_D^d}{p_A^a \cdot p_B^b}$

(i) If Q = K, the reaction is at **equilibrium.**

(ii) If Q > K, Q will tend to decrease to become equal to K i.e., the reaction will proceed in **backward direction.**

(iii) If Q < K, Q will tend to increase. As result the reaction will proceed in **forward** direction..

115. **Equilibrium constant and Temperature :**

With rise of temperature the rate of forward reaction and rate of back ward reaction change i.e., k_f and k_b (rate constants of two opposing reactions) value change differently with rise or fall in temperature. Since K=kf/k_b. changes with temperature. For endothermic **increases** ($k_f > k_b$) and K value for exothermic reaction decreases ($k_b > k_f$) with **rise** in temperature. Temperature does not affect a reaction for which $\Delta_r H = 0$.

From van't equation (integral form), we know that

$$\log\frac{K_2}{K_1} = \frac{\Delta_r H^0}{2.303R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$$

K_1 equilibrium constant at temperature T_1 .
K_2 equilibrium constant at temperature T_2
$\Delta_r H^0$ = standard enthalpy change of reaction
R = Gas constant, 8.314 J/mol/K

116. **Common ion effect :** Let us study the dissociation of NH_4OH (weak base) in presence of NH_4Cl (NH_4^+).

$$\overleftarrow{\qquad\qquad}$$

$$NH_4OH \rightleftharpoons NH_4^+ + OH^-$$

$$NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$$

Here N^+H_4 is common to both & the dissociation is supressed.

$K_b = \frac{\uparrow[NH_4^+][OH^-]}{\uparrow[NH_4OH]}$, Increase Conc^n of N^+H_4 increases the Nr-value as K_b is a constant, conc of $[NH_4OH]$ must have to increase proportionately i.e, the equilibrium shifts towards left. pH decreases.

Thus, the presence of common ion supresses the ionisation of a weak electrolyte.

Application :

(i) Solubility of spaingly soluble salt decreases in presence common ion.

(ii) Salt can be separated from saline water on passing dry $HCl_{(g)}$ through it where Cl^- acts as a common ion. Increasing concentration of Cl^- helps to exceed its K_{SP} value and separates out.

$$\left[Na^+\right]\left[Cl^-\right] > K_{S_p}\,(NaCl)$$

$HCl_{(g)}$ is passed as its highly soluble and [Cl^-] increases as volume of the solution never changes on passing $HCl_{(g)}$.

(iii) Salting out effect : Soap can be settled out on adding saturated solution of NaCl.

$$\longleftarrow\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-$$

$$\underset{\text{Soap}}{C_{17}H_{35}COO^-N_a^+} \rightleftharpoons \underset{\text{(Stearate)}}{C_{17}H_{35}COO^-} + Na^+$$

$$NaCl(aq) \longrightarrow Na^+ + Cl^-,$$ Here Na^+ acts as common ion.

Common ion effect suppresses the dissociation.

(iv) Qualitative analysis :

Example :

(i) Precipitation group II radicals (Pb^{2+}, Hg^{2+}, Cu^{++}, Cd^{++}.........), on passage of $H_2S_{(g)}$ in presence of dil. HCl. In acidic medium (H^+ common ion) the [S^{2-}] is reduced.

$$H_2S \rightleftharpoons 2H^+ + S^{2-}$$

$$HCl_{(aq)} \longrightarrow H^+ + Cl^-$$

$$\left[Cu^{++}\right]\left[S^{2-}\right] > K_{SP}\left(Cus\right)$$

Measurable concentration $S^=$ helps to precipitate sulfids of Group II but not sulfides of group III, IV etc; which have higher K_{SP} values.

(ii) Precipitation of Group III_A, Fe^{3+}, Al^{3+} and Cr^{3+} radicals as hydroxides. Reagent : NH_4Cl and NH_4OH. Due to common ion $\left(NH_4^+\right)$ concentration of OH^- is reduced and helps to exceed the lower K_{SP} values of hydroxides of Fe^{3+}, Al^{3+} and Cr^{3+}.

Group III_B $\left[Zn^{2+}, Co^{2+}, Ni^{2+}, Mn^{2+}\right]$ as sulfides and Group IV $\left[Ba^{2+}, Sr^{2+}, Ca^{2+}\right]$ radical, as carbonates also obtained. It is explained in the light of common ion effect and K_{SP} values.

117. K_{SP} for CuS is 6.3×10^{-36}, K_{SP} for ZnS = 1.6×10^{-24}

i.e., K_{SP} for ZnS >> CuS. This indicates to precipitate group II less [S^{2-}] is required which can be achieved on passage of H_2S is presence of HCl where H^+ acts as a common ion. The dissociation of H_2S is supressed.

$$H_2S \rightleftharpoons 2H^+ + S^2$$

$$\longleftarrow\!\!-\!\!-\!\!-$$

$$HCl \longrightarrow H^+ + Cl^-$$

On the other hand K_{SP} of Group III_b sulfides (eg. Zn S) is much higher than CuS.

This indicates the requirement of high [S^{2-}]. A given concentration of Zn^{2+} requires high concentration S^{2-} as a result the ionic product, $\left[Zn^{2+}\right]\left[S^{2-}\right]$ exceeds, K_{SP} for ZnS.

The high concentration of S^{2-} can be achieved if H_2S is passed in alkaline medium.

$$H_2S \rightleftharpoons 2H^+ + S^{2-}$$

$$NH_4OH \rightleftharpoons N\overset{+}{H}_4 + OH^-$$

$$H_2S + 2OH^- \rightleftharpoons 2H_2O + S^{2-}$$

$$-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!-\!\!\longrightarrow$$

Equilibrium shifts towards right providing higher concentration of S^{2-} i..e, [S^{2-}] >>.

118.

	Valency.		**Oxidation number (O.N)**
(i)	It is the number of $\overline{e}$ that an atom can transfer or share mutually to form bonds.	(i)	It is the residual charge left on an atom when present in a combined state.
(ii)	Valency is a number not having any +/– sign. e.g. H_2O. Oxygen combines with 2 H-atoms, hence valency of oxygen is two.	(ii)	As it refers to itonge, O.N. can either be +ve or – ve. e.g. O.N. of 'O' is $H_2O = (-2)$
(iii)	Valency can't be fraction	(iii)	O.N. can be zero and even fraction. $\overset{0}{C}\overset{+2}{H_2}\overset{-2}{Cl_2}$, $\overset{+}{K}\overset{-1}{O_2}$ O.N. of C = 0, O.N. of $O = -\frac{1}{2}$
(iv)	In some cases valency of an element is fixed. e.g. C is tetravalent in most of its compounds.	(iv)	O.N. of an element may be different in different compounds. eg. ${}^{-4}\underline{C}H_4$, $\underline{\overset{-6}{C_2}}\overset{+6}{H_6}$, $\overset{+2}{\underline{C}}\overset{+1}{H}\overset{-3}{Cl_3}$, $\overset{+4}{C}\overset{-4}{O_2}$
(v)	Valency of an element is though variable but limited. e.g. valency of N is 3 or 5	(v)	Same element has wide variation of oxidation number. e.g. $\underset{(+2)}{\overset{+2}{N}\overset{-2}{O}}$, $\underset{(+1)}{\overset{+2}{N_2}\overset{-2}{O}}$, $\underset{(+4)}{\overset{+4}{N}\overset{-4}{O_2}}$, $\underset{(+3)}{\overset{+6}{N_2}\overset{-6}{O_3}}$, $\underset{+5}{\overset{+10}{N_2}\overset{-10}{O_5}}$

119. (i) Oxidation Number

$$-8\times 1$$

$$\overset{0}{M}g + H^+ + \overset{+10}{2N}O_3^{-} \longrightarrow Mg^{2+} + \overset{+2}{N_2}\overset{-2}{O} + H_2O$$

$$+2\times 4$$

$$4Mg + 10H^+ + 2NO_3^- \longrightarrow 4\ Mg^{2+} + N_2O + 5H_2O$$

Ion-electron

$$\left[Mg \longrightarrow Mg^{2+} + 2e^-\right]4$$

$$2NO_3^- + 10H^+ + 8e^- \longrightarrow N_2O + 5H_2O$$

$$4Mg + 2NO_3^- + 10H^+ \longrightarrow 4Mg^{2+} + N_2O + 5H_2O$$

(ii) Oxidation Number

$$-6$$

$$\overset{12}{Cr_2}O_7^{2-} + 14H^+ + 6Fe^{2+} \longrightarrow \overset{+6}{2Cr^{3+}} + 6Fe^{3+} + 7H_2O$$

$$+1\times 6$$

Ion-electron

$$\begin{array}{r} Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O \\ 6Fe^{2+} \longrightarrow 6Fe^{3+} + 6e^- \\ \hline Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \longrightarrow 2Cr^{3+} + 6Fe^{3+} + H_2O \end{array}$$

(iii) Oxidation Number

$$\overbrace{2\overset{+7}{Mn}O_4^- + Br^- + H_2O \longrightarrow 2\overset{+4}{Mn}O_2}^{-3\times 2} + \underbrace{\overset{(+5)}{Br}O_3^-}_{+6} + 2OH^-$$

Ion-electron

$$\begin{array}{r} [MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2 + 4OH^-]\times 2 \\ Br^- + 6OH^- \longrightarrow BrO_3^- + 3H_2O + 6e^- \\ \hline 2MnO_4^- + H_2O + Br^- \longrightarrow 2MnO_2 + 2OH^- + BrO_3^- \end{array}$$

120. Cell notation : $Zn\,|\,\overset{2+}{Zn}_{(Aq)}\,||\,A\overset{+}{g}(Aq)\,|\,Ag$

(i) –ve electrode, Zn | Zn^{2+}, Current flows for RHE(+) to LHE (–)

(ii) electron / ion is the carrier.

Electron flows from left hand electrode (LHE), Anode to RHE, cathode (outer cell), in innercell ions flow.

(iii) LHE : $Zn \longrightarrow Zn^{2+} + 2\overline{e}$ (oxidation)

RHE : $Ag^+ + e^- \longrightarrow Ag$ (reduction)

(iv) Zn is oxidised, hence a reducing agent $AgNO_3$ is the oxidised agent as Ag^+ reduces to Ag, hence an oxidising agent.

121. (i) $AgNO_3 \longrightarrow Ag^+ + NO_3^-$

At anode (+), $Ag + NO_3^- \longrightarrow AgNO_3 + e^-$

(Silver electrode dissolves)

At cathode (–), $Ag^+ + e^- \longrightarrow Ag$

(silver deposists or the electrode gains weight)

On passage of 1F of electricity

1 gram eqvt of silver deposits at cathode, while 1 gram eqvt. of silver dissolves froming silver nitrate.

(ii) With Pt-electrodes:

$$AgNO_3 \longrightarrow Ag^+ + NO_3^-$$

$$H_2O \longrightarrow H^+ + OH^-$$

At anode : $4OH^- \longrightarrow O_2\uparrow + 2H_2O + 4e^-$ / (or) $2H_2O \longrightarrow O_2 + 4\overset{+}{H} + 4e^-$

At cathode : $Ag^+ + e^- \longrightarrow Ag$ as $E^0_{Ag^+|Ag} >> E^0_{H^+|\frac{1}{2}H_2}$

$O_2(g)$ liberates at anode and Ag deposits / separates at cathode.

(iii) $H_2O \rightleftharpoons H^+ + OH^-$

At cathode : $2H^+ + 2e^- \longrightarrow H_2$

At anode : $2OH^- \longrightarrow \frac{1}{2}O_2\uparrow + H_2O + 2e^-$[or $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

$$H_2O \xrightarrow{\text{electrolysis}} H_2\uparrow + \frac{1}{2}O_2\uparrow$$

One mole of water on electrolysis liberates 1 mole of $H_2(g)$ and $\frac{1}{2}$ mole of oxygen i.e., in the volume ratio $H_2 : O_2 = 1 : \frac{1}{2} \Rightarrow 2:1$

(iv) $CuCl_2 \longrightarrow Cu^{2+} + 2Cl^-$

At anode : $2Cl^- \longrightarrow Cl_2 + 2e^-$

At anode : $Cu^{2+} + 2e^- \longrightarrow Cu$ $\quad E^0_{Cu^{2+}\left|\substack{Cu\\(+0.34)}\right.} > E^0_{H^+\left|\substack{\frac{1}{2}H_2\\(0.00)}\right.} > E^0_{H_2O\left|\substack{OH^-\\(-0.83V)}\right.}$

(v) At Cu anode : $Cu + 2Cl^- \longrightarrow CuCl_2 + 2e^-$ Copper electrode dissolves.

At cathode : $Cu^{2+} + 2e^- \longrightarrow Cu$ Copper electrode gains weight.

122. (i) SO_2 a reducing agent. It reduces excess Cl_2 to Cl^-.

$$2H_2O + SO_2 \longrightarrow SO_4^{2-} + 4H^+ + 2e^-$$
$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$
$$SO_{2(g)} + 2H_2O + Cl_{2(g)} \longrightarrow SO^{2-}_{4\,(aq)} + 4H^+_{(aq)} + 2Cl^-_{(aq)}$$

(ii) Chlorine, sulfur, phosphorus undergoing disproportion in alkaline medium :

$$\overset{0}{Cl_2} + 2NaOH \longrightarrow Na\overset{-1}{Cl} + Na\overset{+1}{Cl}O + H_2O$$

$$\overset{(0)}{S_8}(s) + 12OH^- \longrightarrow 4\overset{(2-)}{S} + 2\overset{(+2)}{S_2}O_3^{2-} + 6H_2O$$

$$P_4(s) + 3OH^{(-)} + 3H_2O(\ell) \longrightarrow \overset{(-3)}{P}H_{3(g)} + 3H_2\overset{(+1)-}{PO_2}$$

(iii) $2NH_3 + \frac{5}{2}O_2 \longrightarrow 2NO + 3H_2O$

34 g   80 g   60 g

10 g $\frac{80}{34}\times 10 = 23.52$ (oxygen)

Given : Oxygen 20 gram (Less)   80 gram oxygen reacts with 34 gram NH_3

20 gram oxygen reacts with $\frac{34}{4} = 8.5$ gm NH_3

Limiting reagent = Oxygen (20 gram)

80 gram oxygen given 60 g NO

Given, 20 gram oxygen will give $60 \div 4 = 15$ gram (NO).

(iv) Metal showing variable oxidation state can show disproportionation e.g. Mn, Cu and Ga.

$$\overset{+1}{Hg_2}\,Cl_2 \xrightarrow{\Delta} \overset{0}{Hg} + \overset{+2}{Hg}Cl_2$$

$${}^{+3}Mn_2O_3 \longrightarrow \overset{+2}{Mn}O + \overset{+4}{Mn}O_2$$

$$2Cu^{+1}(aq) \longrightarrow Cu^{(2+)}(Aq) + Cu^0$$

$$3Ga^+(aq) \longrightarrow Ga^{3+}(Aq) + 2G\overset{0}{a}$$

(v) $E^0_{\underset{(0.80V)}{Ag^+|Ag}} > E^0_{\underset{(0.34V)}{Cu^{2+}|Cu}}$

As silver has higher reduction potential it is a better oxidant can oxidise both benzaldehyde and acetaldehyde. But Cu^{2+} cann't oxidise benzaldehyde to benzoic acid. It is a less powerful oxidising agent due to its low reduction potential. That is the reason why

$C_6H_5CHO + 2Cu^{2+} + 5(OH)^-$ does not react.

However C_6H_5CHO can be oxidised by $[Ag(NH_3)_2]^+$ due to higher $E^0_{Ag^+|Ag}$ value. Greater is the reduction potential better is the oxidising power.

$$C_6H_5CHO + 2[Ag(NH_3)_2]^+ + 3OH^- \rightarrow 2Ag + C_6H_5COO^- + 4NH_3 + 2H_2O$$

(vi) Reasons :

(a) $KMnO_4$ in neutral medium OH^- ions are produced. No need of adding alkali / acid from out side i..e, $MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2 + 4(OH^-)$

(or) $2KMnO_4 + H_2O \longrightarrow 2\ MnO_2 + 2\ KOH + 3(O)$

(b) $KMnO_4$ with alcohol and toluene forms a homogeneous solution in which reaction proceeds faster.

123. $KMnO_4$ acts as an oxidant in three media.

Acidic : $MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$ Eq. mass $= \frac{1}{5}$ Mol. mass.

Alkaline : $MnO_4^- + e^- \longrightarrow MnO_4^{2-}$,   Eq. mass = Mol. mass

Neutral : $MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2 + 4OH^-$ Eq. mass $= \frac{\text{Mol. mass}}{3}$

Thus, less consumption $\left(\frac{1}{5}\text{mol. mass}\right)$ in preparation normal solution in acidic medium for estimating iron, copper etc.

124. **Factors goverining electrode potential.**

The tendency of an electrode to lose or gain electrons is called electrode potential. Since each electrode represents a half cell, therefore, electrode potential is also called potential for half cell. The electrode potential and hence cell potential depends on –

(i) Conc^n :

$M^{n+} + ne^- \longrightarrow M_{(s)}$

Nernst equation $E_{M^{n+}|M} = E^0_{M^{n+}|M} - \frac{2.303}{nF}\log\frac{1}{\left[M^{n+}\right]}$

$$= E^0_{M^{n+}|M} + \frac{2.303}{nF}\log\left[M^{n+}\right]$$

It shows if $\left[M^{n+}\right]$ increases $E_{M^{n+}|M}$, the reduction potential of the electrode increases. $E^0_{M^{n+}|M}$ is constant at a given temperature (298 K), $\left[M^{n+}\right] = 1M$.

(ii) The nature of the metal and its ions.

The electrode potential (oxidation / reduction) varies from element to element,

e.g. $E^0_{Cu^{2+}|Cu} = +0.34V$ $E^0_{Zn^{2+}|Zn} = -0.76V$

It dependds on nature of the element in respect of

(a) Heat of atomisation / Sublimation

(b) Ionisation enthalpy

(c) Hydration enthalpy

(iii) **Temperature** Temperature affects the electrode potential / cell potential depending on energetics of the reaction (electrode / cell). If the equilibrium reaction is endothermic the potential increases with temperature. However, standard potential, E^0_{Cell} is found out at 298 K.

(iv) **Pressure** In case of gaseous electrode (eg $H^+ | \frac{1}{2} H_2$) the electrode potential can be changed with increasing the pressure of the gas. However, the standard electrode potential is measured on passing the gas at 1 atm and 298 K into solution containing ions (1 M)

125. **Resemblance with alkalimetals:**

(a) Identical valence shell configuration.

(b) electropositive character

$$HCl \longrightarrow H^+ + Cl^-$$

$$NaCl \longrightarrow Na^+ + Cl^-$$

(c) Oxidation state (I) $\overset{+}{H}Cl^-, Na^+Cl^-, Cs^+Cl^-$

(d) Combination with nonmetals -

Oxide	H_2O,	N_2O,	K_2O	....
Halide	HCl,	NaCl,	KCl	....
Sulfide	H_2S,	N_2S,	K_2S	....

(e) Discharged at cathode

$$H^+ + e^- \longrightarrow \frac{1}{2}H_2$$

$$Na^+ + e \longrightarrow Na \text{ (metal) from fused NaCl}$$

(f) Reducing character-

$$Fe_3O_4 + 4H_2 \xrightarrow[\Delta]{} 3Fe + 4H_2O$$

$$B_2O_3 + 6K \longrightarrow 2B + 3K_2O$$

126. **Resemblance with halogen-**

(a) Electronic configuration - one electron short of duplet.octet.
(stable configuration of noble gas)

(b) Electronegative character - $H + e^- = H^-$ (hydride)
$X + e^- \rightarrow X^-$ (halide)

(c) ΔiH, comparable with halogens but higher than alkalimetals.

	Li	H	F
ΔiH (kJ/mol)	520	1312.0	1681

(d) Liberation at anode on electrolysicphysics (e.g. electrolysis NaH(*l*).

(e) Atomicity (diatomic) H_2, X_2

(f) O.N., –1 O.N. (H^-, X^-)

(g) Combination with metal- $Na + \frac{1}{2}H_2 \rightarrow NaH$, $Na + \frac{1}{2}Cl_2 \rightarrow NaCl$

(h) Nonmetalic nature H(nonmetal) H (nonemetal)
Li (metal) F (non-metal)

(i) Formation of covalent bond CH_4 CCl_4

(j) Replecement of Hydrogen by halogen CH_4 CH_3Cl

(k) Physical state : H_2(g), F_2(g), Cl_2(g)

127. **Difference from alkali metal & halogen-**

(a) Hydrogen can both act as oxidant and reductant $(H - e \rightarrow H^+)$ but alkalimetal reducing agent, halogens oxidant.

(b) Hydrogen has one electron (no neutron) but alkalimetals & halogens are polyelectronic and also contain more number of neutrons.

(c) In the molecular state hydrogen has two (nuclear) isomers o–H_2 and p–H_2.

(d) Nature of oxids -

Li/Na/K	$\boxed{H}$	Cl/Br/I
Basic (Li_2O)	Neutral (H_2O)	Acidic (Cl_2O_7)

(e) Absence of lone pair of electron $H:H$, $:\ddot{\underset{\cdot\cdot}{Cl}}-\ddot{\underset{\cdot\cdot}{Cl}}:$

(f) Nature of Compounds

(Ionic Compound) LiF , HF (covalent compound)

128. Similarty of 'H' with carbon :

'H' resembles carbon in the following respects :

(a) Half outer shell

H- $1s^1$ $\frac{1}{2}\times 2 = 1(K-\text{shell})$

C - $1s^2 2s^2\, 2p^2$ $\frac{1}{2}\times 8 = 4(L-\text{shell})$

(b) Electronegativity Both have almost similar electronegativity. H = 2.1, C = 2.5

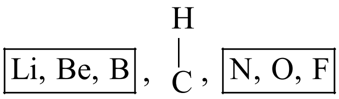
(c) Formation of covalent compounds

e.g. HCl, CCl_4

C, H covalent compounds are huge in number

e.g. Hydrocarbons (saturated, unsaturated, alicyclic, aromatic etc.)

(d) Thomson's placement of 'H' in periodic table.



$\boxed{\text{Li, Be, B}},\ \overset{\text{H}}{\underset{\text{C}}{|}},\ \boxed{\text{N, O, F}}$

129. **Isotopic effect**

Isotopes of an element have same atomic number and have identical electronic configuration. Thus exhibit same chemical properties but difference masses account for variation in rate of reaction, equilibrium constant (for reversible reaction). They also differ in their physical properties such as m.p., b.p, latent heat of fusion, evaporation, sublimation etc. The difference in properties on account of the mass difference in isotopes of an element is known as isotopic effect.

On changing the isotopes is chemical reaction the reaction mechanism and kinetics can be studied.

130. Nascent 'H' It is hydrogen i.e., formed in situ (reaction system) also called newly born. When it comes out of the system, it appears as molecular, H_2.

Preparation

(a) alcoholic sodium : $C_2H_5OH + Na \longrightarrow C_2H_5O\,Na + H$

(b) Sodium - amalgam with water.

$$Na-Hg+H_2O \longrightarrow NaOH+H+Hg$$

Reducing properties It is a powerful reducing agent much better than H_2 (dihydrogen).

(i) when hydrogen gas (H_2) is passed through acidified $KMnO_4$ solution, its pink colour is not discharged.To the same solution if some pieces of zinc is added pink colour gets discharged after some lime. This is because of nascent hydrogen which just generated by the action of Zn and dilute sulfuric acid.

$$KMnO_4 + H_2SO_4 + H_2\left(Zn + H_2SO_4\right) \longrightarrow \text{No change of colou}$$

$$2KMnO_4 + 3H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5[O]$$

$$\left[Zn + H_2SO_4 \longrightarrow ZnSO_4 + 2(H)\right] \times 5$$

$$5[O] + 10[H] \longrightarrow 5H_2O$$

$$\overline{\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad}$$

$$2KMnO_4 + 8H_2SO_4 + 5Zn \longrightarrow K_2SO_4 + 2MnSO_4 + -8H_2O + 5ZnSO_4$$

(ii) $FeCl_3$ (yellow) with nascent hydrogen (Zn+HCl) form Ferrous chloride (Light green).

$$Zn + 2HCl \longrightarrow ZnCl_2 + 2(H)$$

$$\left[FeCl_3 + H \longrightarrow FeCl_2 + HCl\right] \times 2$$

$$\overline{\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad}$$

$$Zn + 2FeCl_3 \longrightarrow ZnCl_2 + 2\ FeCl_2$$

131. **Atomic hydrogen (H)** It is best prepared by passing molecular hydrogen through tungsten electric arc (2000 – 3000°C) at low pressure. The dissociation is endo thermic.

$$H_2 + 433\,kJ = H + H\ \left(\Delta H = +433\,kJ/mol\right)$$

This form of hydrogen is more reactive than ordinary and nascent hydrogen.

Stability. The life period of 'H' is only 0.03 second, i.e., it is extremely unstable and reunites readily to form molecular hydrogen (H_2) with liberation of large amount of heat raising temperature to 4000 to 5000°C. This is how atomic hydrogen torch works.

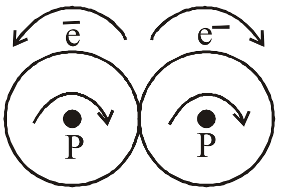
$$H + H \rightarrow H_2 + 433\ kJ$$

It reduced CO to formaldehyde

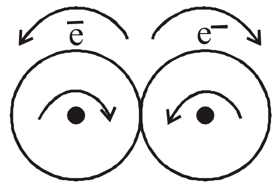
$$\overset{+2}{C}\overset{-2}{O} + 2H^0 \longrightarrow \overset{+1}{H}\overset{0}{C}\overset{+1}{H}\overset{-2}{O}\ \text{(Formaldehyde).}$$

132. **Isomers of hydrogen** : $o–H_2$ and $p.H_2$

In $o–H_2$ two nuclei (protons) spins in the same direction while $p\text{-}H_2$ the protons in opposite direction, however electronic spin will be in opposite direction in both the forms.



Orthohydrogen (parellal nuclearspin)

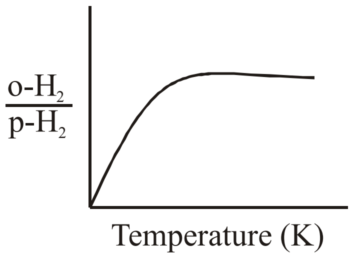


Para hydrogen (antiparallel nuclear spin)

$p\text{-}H_2$ (pure) can be prepared by adsorbing ordinary hydrogen in activated charcoal in quartz vessel at 20 K for 3 to 4 hours.

$o–H_2$ (pure) cann't be prepared.

$p–H_2$ converts to $o–H_2$ with rise of temperature. At normal temperature ordinary hydrogen is a mixture of 75% $o–H_2$ and 25 % $p–H_2$. At liquid air temperature $\dfrac{o - H_2}{p - H_2} = 1$.



Characteristics –

(i) o–H_2 has higher intenal energy than p–H_2

(ii) o–H_2 is more stable than p–H_2 & hence p–H_2 has the tendency to change to o–H_2.

(iii) The two forms differ in physical properties like m.p, b.p, thermal conductivity, specific heat etc.

(iv) These have identical chemical properties.

133. **Hydrides –**

Almost all elements except groups 18 (noble gases), 7 (Mn group), 8 (Fe group), 9 (Co-group) react with hydrogen under suitable condition form hydrides with general formula, MHx

Classification :

(a) **Ionic hydrides :** H_2 reacts with more electro +ve group 1 and 2 producing stoichiometric hydride e.g. NaH, CaH_2 etc.

Properties :

(i) These are crystalline, non volatile and does not conduct electricity is solid form.

(ii) Molten hydride liberated hydrogen at anode.

$$2H^- \longrightarrow H_2 + 2e^-$$

(iii) On action with water $H_2(g)$ liberates.

$$CaH_2(s) + 2H_2O(\ell) \longrightarrow Ca(OH)_2(aq) + 2H_2(g)$$

(iv) These hybrides are better reducing agents.

(b) **Covalent molecular hydrides :** p-block elements mostly with dihydrogen forms covalent hydride e.g. CH_4, NH_3, H_2O, HF on Lewis concept these are of

(i) electron - deficient hydride e.g. borane

(ii) electron precise hydride e.g. CH_4

(iii) electron-rich hydride e.g. : NH_3, OH_2, HF

(c) **Interstitial hydrides** / Non-stoichiometric

d and f block elements except metals of group 7, 8, 9 usually form metal hydride. These are non-stoichiometric being deficient in hydrogen. e.g. La $H_{2.87}$. Hydride lattice differs from the metal itself.

Properties : -

(i) These are hard having metallic lusture, conducting electricity with magnetic property.

(ii) Density of hydride is lower than the metals as crystal lattice expands due to inclusion of hydrogen.

134. Heavy water (D_2O) and ordinary water (H_2O)

Heavy water (D_2O)	**Ordinary water (H_2O)**
(i) It is an oxide of Deuterium (heavy hydrogen, ^2_1H)	(i) Oxide of ordinary hydrogen $\left(^1_1H\right)$
(ii) Density at 293 K 1.017 gm/cm^3	(ii) 0.998 gm / Cm^3
(iii) M.P. 276.8 K	(iii) 273 K
(iv) B.P 374.4 K	(iv) 373 K
(v) Solubility of NaCl in gram/100 gm water at 298 K is 30.5	(v) 35.9
With metal $2D_2O + Na \longrightarrow 2NaOD + D_2$	$2H_2O + Na \longrightarrow 2NaOH + H_2$
Kinetics slow rate of reaction Exchange reaction. $HCl + D_2O \longrightarrow DCl + HOD$	Faster reaction rate – Both D_2O and H_2O almost give similar type of reaction only replacing 'D' for 'H'.

135. **Concentration of Hydrogen peroxide**

(i) The hydrogen peroxide contain water as impurity which can be separated by distillation under reduced pressure (10-15 mm). Water distils at 303–313 K leaving H_2O_2 (99%).

(ii) The last traces of water can be removed by cooling the solution in freezing mixture consisting dryice. Crystals of H_2O_2 separates out, dried and fused to get pure H_2O_2.

Preservation : $2H_2O_2 \xrightarrow[\text{light/etc}]{\text{metalion/dust/OH}^-} 2H_2O + O_2$

To prevent decomposition –

(i) It is to be stored in coloured paraffin wax coated plastic or teflon bottle. Glass bottle is not used as it can catalyse the decomposition

(ii) Glycerine, acetalinide, phosphoric acid etc. (negative catalyst) must be used to check further decomposition of Hydrogen peroxide.

136. The volume of hydrogen peroxide is the volume of O_2 (g) liberated at NTP per litre of it.

e.g. ' x vol H_2O_2' = 1 litre of H_2O_2 gives x litre of O_2 at NTP.

Normality of '10 vol. H_2O_2'.

1 litre of it gives 10 L of O_2 at NTP

$2\,H_2O_2 \longrightarrow 2H_2O + O_2$

68 gram — 22.4 L at NTP

$\frac{68}{22.4} \times 10$ of H_2O_2 — 10 L at NTP

Thus $\frac{68\times10}{22.4}$ gram of H_2O_2 present in 1 L of '10 vol. H_2O_2' equivalent mass of H_2O_2 is 17 gram / eqvt.

1 L contains $\frac{68\times10}{22.4}$ gram $= \frac{68\times10}{22.4\times17}$ gram eqvt.

Normality = 1.78 N

Thus Normality of 'x volume H_2O_2' = $\frac{68\ x}{22.4 \times 17}$ N

$$\text{Molarity} = \frac{68\ x}{22.4 \times 34}\,\text{M}$$

$$\text{Strength or gram/L} = \frac{68\,x}{22.4}$$

Here x L of O_2 liberates at NTP per litre of 'x vol. H_2O_2'

137. . $2NaHCO_3 + H_2O_2 \longrightarrow Na_2O_2 + 2\,CO_2 + 2\,H_2O$

*
$$\begin{array}{r}
H_2O_2 + 2e^- \longrightarrow 2OH^- \\
Mn^{2+} + 4OH^- \longrightarrow MnO_2 + 2H_2O + 2e^- \\
\hline
H_2O_2 + Mn^{2+} + 2OH^- \longrightarrow MnO_2 + 2H_2O
\end{array}$$

*

$$H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$$

$$\underline{2\left[Fe(CN)_6\right]^{4-} \longrightarrow 2\left[Fe(CN)_6\right]^{3-} + 2e^-}$$

$$H_2O_2 + 2H^+ + 2\left[Fe(CN)_6\right]^{4-} \longrightarrow 2\left[Fe(CN)_6\right]^{3-} + 2H_2O$$

138. **Hydrogen economy :** One proposed way to meet the need for new energy sources is to burn hydrogen as fuel in industry and power plants and possibly also in homes and motor. This proposal is referred to hydrogen economy.

Advantages :

(i) Dihydrogen releases large quantity of heat on combustion. On mass to mass basis, dihydrogen can release more energy than gasoline.

(ii) On burning water is formed, no pollution takes place. So it's an ecofriendly fuel.

In India, since october 2005, CNG is mixed with H_2 (5%) for use as four-wheeler fuel.

139. H_2O_2 as oxidant : $H_2\ O_2 \longrightarrow H_2O + [O]$

(i)

$$PbS + 4(O) \longrightarrow PbSO_4$$

$$\underline{\text{Black} \qquad\qquad \text{White}}$$

$$PbS + 4H_2O_2 \longrightarrow PbSO_4 + 4H_2O$$

(ii) It oxidises ferrous sulfate solution (green) to ferric sulfate (yellow)

$$H_2\ O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$$

$$\underline{Fe^{2+} \longrightarrow Fe^{3+} + e] \times 2}$$

$$H_2O_2 + 2H^+ + 2Fe^{2+} \longrightarrow 2Fe^{3+} + 2H_2O$$

H_2O_2 as reductant :

– In neutral medium it reduces $Cl_2(g)$ to Cl^- ion.

$$H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$$

$$\underline{Cl_2 + 2H^+ + \xrightarrow{2e^-} 2HCl}$$

$$H_2O_2 + Cl_2 \longrightarrow 2HCl + O_2$$

– acidic medium dichromate (orange colour) is reduced to $Cr^{3+}(Cr_2\ (SO_4)_3$–green).

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2C_r^{3+} + 7H_2O$$

$$\underline{3 \times \left[H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-\right]}$$

$$Cr_2O_7^{2-} + 3H_2O_2 + 8H^+ \longrightarrow 2Cr^{3+} + 3O_2 + 7H_2O$$

– In alkaline medium ferricyamide is reduced to ferrocyamide.

$$H_2O_2 + 2OH^- \longrightarrow O_2 + 2H_2O + 2e^-$$

$$\underline{\left[Fe(CN)_6\right]^{3-} + e^- \longrightarrow \left[Fe(CN)_6\right]^{4-}}$$

$$H_2O_2 + 2OH^- + 2\left[Fe(CN)_6\right]^{3-} \longrightarrow 2\left[Fe(CN)_6\right]^{4-} + O_2 + 2H_2O$$

140. (i) When an etherical solution of H_2O_2 solution shaken with acidified solution of $K_2Cr_2O_7$, blue colour appears in the ether-layer due to formation of chromium pentoxide (CrO_5)

$$H_2O_2 \longrightarrow H_2O + O\big] \times 4$$

$$\underline{K_2Cr_2O_7 + H_2SO_4 + 4(O) \longrightarrow K_2SO_4 + 2CrO_5 + H_2O}$$

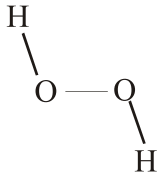
$$K_2Cr_2O_7 + H_2SO_4 + 4H_2O_2 \longrightarrow K_2SO_4 + 2CrO_5 + 5H_2O$$

(ii) $_{15}P = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Although 'P' exhibits +3 and +5 O.S. it fails to form PH_5. This is due to high $\Delta_a H$ value of hydrogen and $\Delta_{eg} H$ value of Hydrogen. These two values do not favour to exhibit the highest oxidation of P and consequently the formation of PH_5. However it can form PH_3.

(iii) H_2O_2 is not linear.

The two O–H bond are in two different planes with dihedral (inter planar) angle 111.5^0 (in gas phase), in solid phase at 110 K is 90.2^0.



(iv) Water as a universal solvent. Water has ability to dissolve most of the (ionic) compounds due to

(a) it has high dielectric constant (78.39) which weakens the inter ionic force of attraction.

(b) it's a polar molecule due to high electronegativity of oxygen and bent molecular shape. These polar molecules surrounds the ions by stronger ion-dipole force as a result the ions get hydration energy. The lattice breaks the ionic compounds become soluble.

(c) The covalent molecules dissolve in water due to its tendency to form intermolecular H-bond (dipole - dipole force) e.g. NH_3, amines, alcohols, sugar etc.

(v) **(Water break down) Hydrolysis (solvolysis)**

Hydrolysis	**Hydration (Solvation)**
— Double decomposition reaction with water as one of the reactant $CH_3COONa + HOH \rightleftharpoons CH_3COOH + NaOH$	— Ions surrounded by water molecules by ion - dipole force. e.g. Na^+(Aq). Cl^- (Aq) It is not a chemical process.
— Occurs with bond cleavage. $NH_4^+ + H_2O \longrightarrow NH_4OH + H^+$ $CN^- + H_2O \longrightarrow HCN + OH^-$	— No new compound due to non cleavage of H_2O molecule.
— New compounds are formed No hydrates and dehydrates is formed.	— Here hydrates are formed which may form dehydrates on heating. $CuSO_4 . 5H_2O$ (blue) $\longrightarrow$ $CuSO_4$ (White) (Anhydrous) The hydration depends on ionic size, charge, presence of vacant orbitals, where water behaves as ligand.

Unit - X

141. (a) Anomalous behaviour of first element of group 1 and 2. Reasons for the difference :

(i) Small atomic or ionic radii

(ii) high electrogenative and ΔiH value.

(iii) high polarising power of cation i.e. charge / unit area = $\dfrac{\text{ionic charge}}{\frac{4}{3}\pi\,(\text{ionic radius})^2}$ thus polarising power $\propto \dfrac{\text{ionic change}}{(\text{its radius})^2}$

(iv) absence of d-electrons in its valence shell

e.g. Li & Be differs from other elements of group 1 and group 2 respectively.

(v) Acting as powerful reducing agent even if high ΔiH value. This is due to greater $\Delta_{hyd}H$ value as ions can easily hydrated out of ion-dipole $\left(\because\!O\begin{smallmatrix}\diagup H\\ \diagdown H\end{smallmatrix}\right)$ attraction.

$E^0_{Li^+_{(aq)}|Li}$ and $E^0_{Be^{2+}_{(aq)}|Be}$ values are highly negative.

(b) Anomalous behaviour of Lithium :

(i) Lithium forms monoxide, Li_2O while other forms peroxide and superoxide.

(ii) Li (OH) a weak alkali decomposes forming Li_2O.

(iii) $Li_2\,CO_3$ is thermally unstable. $Li_2\,CO_3 \rightarrow Li_2O + CO_2$

(iv) $2\,Li\,NO_3 \rightarrow Li_2O + 2NO_2 + \frac{1}{2}O_2$

$$NaNO_3 \longrightarrow NaNO_2 + \frac{1}{2}O_2$$

(v) Lithium forms nitride.

$6\,Li + N_2 \longrightarrow 2\,Li_3N$ other alkali metals don't.

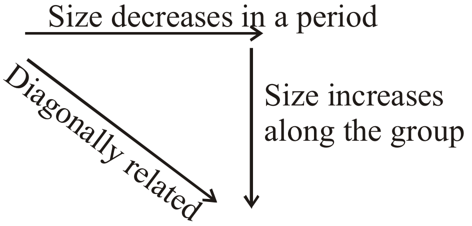
142. **Diagonal relationship :**

$$\begin{tikzcd} Li \arrow[dr] & Be \arrow[dr] & B \\ Na & Mg & Al \end{tikzcd}$$

This similarities are weaker than the similarities found within the group.

Reasons : (i) Similarity in electropositive character (electronegativity values)

(ii) Similarity in atomic atomic/ions radii



(iii) Comparable polarising power.

Example : (i) Li and Mg form nitrides

(ii) Li_2O and MgO never form superoxide.

(iii) $Li_2 CO_3$ and $MgCO_3$ decompose librating CO_2.

(iv) Li & Mg do not form solid bicarbonate

(v) Li Cl and $MgCl_2$ soluble in alcohol due to covalent nature.

(vi) Li Cl amd $MgCl_2$ are deliquescent. The hydrates are $LiCl.2H_2O$ and $MgCl_2.8H_2O$

143. Due to higher hydration enthalpy $\text{Li} \longrightarrow \text{Li}^+(\text{Aq}) + \bar{e}$ is easier providing much lower reduction potential, $E^0_{\text{Li}^+(\text{Aq})|\text{Li}} = -3.05\text{V}$ and hence a powerful reducing agent.

144. Lithium has higher $\Delta_i H$ than Na and K. It is the lightest metal, floates over kerosine, but dips inside wax in order to avoid its reaction with air it is kept inside wax. Wax is denser than kerosine.

145. Due to bigger size, these metals have lower $\Delta_i H$ which gradually decreases on moving down the group. The energy from the Bunsen flame is sufficient to excite the $(e^- s)$ of these metals to higher energy level. The excited state is quite unstable and therefore when these excited $\bar{e}$s come back to original energy level, they emit energy which falls in visible region of electromagnetic spectrum & they appear coloured.

e.g. Na imparts golden yellow white K imparts pale violet. Due to lower $\Delta_i H$ value of K, the excitation absorbs lower energy of the visible spectrum emitting ligher energy (complementary) i.e., in the violet zone.

146. When radiation strikes the surface of the alkali metals electrons get emitted. The emission becomes easier because of their low $\Delta_i H$ value. The flow of these photo electrons exhibit photo electric effect. Cs has lowest $\Delta_i H$ value hence shows photoelectric effect to the maximum extent.

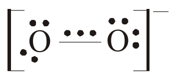
$$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

'$h\nu_0$' is the minimum energy necessary to expel the $\bar{e}$ from the metal surface.

147. (i) Because of the small size, Li^+ has a strong positive field around it which attract negative charge so strongly that it does not permit the O^{2-} (oxide ion) to combine with another oxygen atom to form peroxide O_2^{2-}. On the other hand small cation stabilities small anion. As $[-o-o-]^{2-}$ is a bigger anion can form sodium peroxide because Na^+ is bigger, Due to bigger size it has weaker positive field than Li^+. Such weaker +ve field around Na^+ cann't prevent O^{2-} to combine with another oxygen atom to form peroxide (O_2^{2-}). K^+, Rb^+, Cs^+ ions have still weaker +ve field due to bigger size and can stabilise bigger superoxide, $\bar{O}_2$ anion and form superoxide.

$$\underset{(\text{Oxide})}{\overset{2-}{\text{O}}} \xrightarrow{\frac{1}{2}O_2} \underset{\text{Peroxide}}{O_2^{2-}} \xrightarrow{O_2} \underset{\text{Superoxide}}{2O_2^-}$$

Weaker positive field on $K^+/Rb^+/Cs^+$ cannot prevent even peroxide to combine with O_2 to form superoxide.

Superoxides are coloured due to presence of one unpair $\bar{e}$. , three $\bar{e}$s two centre bond.

(ii) M—OH bond in hydroxides is very weak and can easily ionises to $M^+ + OH^-$ ions. The ionic nature of metal hydroxides increases on moving down the group with decrease of $\Delta_i H$ values. More is the ionisation more is the basic nature of the hydroxides.

Thus, LiOH < NaOH < KOH < RbOH < CsOH

(iii) (a) LiF is less ionic due to small cationic size of Lithium. Smaller the cation more is the polarising power (Fajan's rule) The covalent character : LiF < LiCl < LiBr < LiI

(b) Smaller cation and anion influence higher lattice energy.

(iv) $Na + (x+y)NH_3 \longrightarrow Na^+(NH_3)_x + \overline{e}(NH_3)_y$

This ionisation makes the solution conducting and the ammoniated $\overline{e}$ transits in visible spectral zone emitting colour in blue zone.

(v) Conducting power of the ions depends on its mobility on application of electric field. In aqueous medium the polar water molecules hydrate the alkalimetal cations making it bulkier and hence the mobility decreases. Smaller the cation greater is the hydration due to ion-dipole attraction, more volumenous the ion and less is the conductance.

Conductance : $Li^+(Aq) < Na^+(Aq) < K^+(Aq) < Rb^+(Aq) < Cs^+(Aq)$

(vi) $Be(2s^2)$, $Mg(3s^2)$ due to smaller size experience greater nuclear influence on the outer electrons as a result electronic transition does not occur in visible spectral zone. Further the high $\Delta_i H$ value needs higher energy for excition which does not occur in presence of Bunsen flame.

148. **Sodium hydroxide** (Caustic soda)

Principle - Electrolysis of sodium chloride solution (brine) with mercury cathods and carbon anode. Thus in Castner - Kellner Cell at cathode :

$$Na^+ + e^- \xrightarrow{Hg} Na - \text{amalgam}$$

At anode $Cl^- \longrightarrow \frac{1}{2}Cl_2 + e^-$

$$Na-Hg + 2H_2O \longrightarrow 2NaOH + Hg + H_2$$

It is deliquescent and on exposure converts to Na_2CO_3.

$$2NaO + CO_2 \rightarrow Na_2CO_3$$

Uses - (i) Manufacture of different salts.

(ii) Lab chemicals.

(iii) In the purification of bauxite.

149. **Baking soda** It is prepared by passing CO_2 through saturated solution of sodium carbonate.

$$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

Compsosition of baking powder : Sodium bicarbonate and tartaric acid.

150. **Biological function of Na and K**

(a) Maintaining osmotic pressure : Na and K function to maintain the normal osmotic pressure of the different body fluids & thus protect the body against excessive loss of fluids

(b) Maintaining proper viscosity of blood : Chlorides of -Na and K present in blood plasma help in keeping the globulin in physical solution and regulating the degree of hydration of plasma proteins, which is very vital in maintaining proper viscosity of blood.

(c) Secretion of digestive fluids :Gastric HCl is derived from NaCl present in blood, while the base in pancreatic juice and bile is derived form blood Na^+ and K^+.

Na^+, K^+ differ quantitatively in their ability to activate enzymes in their transport mechanism & in penetratig cell membrane.

151. $2Na + 2NH_3 \longrightarrow NaNH_2 + H_2$

$$2K + 2CH_3OH \longrightarrow 2CH_3OK + H_2$$

$$2Na\,HCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2$$

$$\begin{array}{c} \left[Cl_2 + 2NaOH \longrightarrow NaCl + NaOCl + H_2O\right] \times 3 \\ 3NaOCl \xrightarrow{\Delta} 2\,NaCl + NaClO_3 \\ \hline 3Cl_2 + 6\,NaOH \longrightarrow 5\,NaCl + NaClO_3 + 3\,H_2O \end{array}$$

152. Na (other alkalimetals also) on exposure to atmosphere (air and moisture) get converted into oxides, hydroxides (hygroscopic) finally to carbonate. By such sodium (silvery white metal) gets tarnished when exposed to air and moisture.

$$\begin{array}{l} 4Na + O_2 \longrightarrow 2\,\cancel{Na_2O} \\ 2\,\cancel{Na_2O} + 2\cancel{H_2O} \longrightarrow 4\,\cancel{NaOH} \\ 4\,\cancel{NaOH} + 2\,CO_2 \longrightarrow 2Na_2\,CO_3 + 2\cancel{H_2O} \\ \hline 4Na + O_2 + 2CO_2 \longrightarrow 2Na_2\,CO_3\,(\text{white powder}) \end{array}$$

153. Preparation of $Na_2CO_3.10H_2O$ (washing soda) Ammonia soda process: (Solvay process)

$$NH_{3(aq)} \xrightarrow{CO_2} (NH_4)_2\,CO_{3(aq)} \xrightarrow{CO_2\,(\text{excess})} NH_4HCO_3\,. \quad \xrightarrow{NaCl} NaHCO_3 \xrightarrow[\Delta]{} Na_2CO_3$$

K_2CO_3 is not prepared by solvay process as it's more soluble in water and does not precipitate when CO_2 is passed. $NH_4HCO_3 \xrightarrow{KCl} KHCO_3$ (highly soluble difficult to precipitate.)

154. The existence of divalent ion in aqueous solution is due to greater enthalpy of hydration of the divalent ion which counter balance the $\Delta_i(H)_2$ i.e., $M^+_{(g)} \longrightarrow M^{2+}_{(g)} + e^-$

Due to smaller size and high change (+2), these cations have greater tendency to be hydrated & greater $\Delta_{hyd}H$.

e.g. $\Delta_{hyd}\,Li^+(\text{size } 76 \text{ pm}) = -506\,kJ/mol$

$$\Delta_{hyd}\,Mg^{2+}(\text{size } 72 \text{ pm}) = -1921\,kJ/mol$$

$$\Delta_{hyd}\,Mg^{+}(\text{size } 72 \text{ pm}) = -365\,kJ/mol$$

On thermodynamic calculation

$$\Delta H_{MgCl} = -74\,kJ/mole, \quad \Delta H_{MgCl_2} = -774\,kJ/mole$$

i.e., $MgCl_2\,(Aq)$ is much more stable than MgCl(Aq)

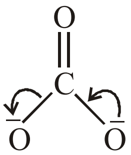
This is reason why group-2 cations are dipositive.

155. Nature of carbonates and sufates of group-2 :

Solubility of sulfates / carbonates of : Be > Mg > Ca > Sr > $BaCO_3$ and $BaSO_4$ are insoluble.

The solubility of ionic compounds depends on lattice energy and hydration energy. In these cases the lattice energy remains almost constant because of bigger anions. The small increase in catonic size down the group (Be ro Ba) does not make any difference. However, the hydration energy form Be^{2+} to Ba^{2+} decreases appreciably as the size increases down the group. Hence the solubilities of carbonates / sulfates decrease from Be to Ba. Stability of carbonate / sulfate increases from Be to Ba.

$M_2CO_3 \xrightarrow{\Delta} MO + CO_2$ M = Be, Mg, Ca, Sr or Ba



$\longrightarrow \quad O=C=O+O^{2-}$

Greater is the polarisation of CO_3^{2-}, the easier is to weaken a C – O bond in CO_3^{2-} and form CO_2 and oxide on heating. The polarising power gradually decreases from Be^{2+} to Ba^{2+} on increase of cationic size.

Thus in Be CO_3, CO_3^{2-} gets more polarised by Be^{2+} liberating CO_2 easily.

$Be\,CO_3 \xrightarrow[<373K]{\Delta} BeO + CO_2\uparrow$, while

$BaCO_3 \xrightarrow{1633K} BaO + CO_2$, as CO_3^{2-} gets less polarised by Ba^{2+} (bigger cationic size)

The stability of metal carbonates follows in the order.

$BeCO_3 < MgCO_3 < CaCO_3 < SrCO_3 < BaCO_3$

$MSO_4 \xrightarrow{\Delta} MO + SO_3\uparrow$. $\xrightarrow[\text{Stability increaces}]{BeSO_4 < MgSO_4 < CaSO_4 < Sr_2SO_4 < BaSO_4}$

156. **Anomalous behaviour of Be.**

(i) Small size, high electro negativity, high ΔiH value high polarising power hence mostly form covalent compounds.

(ii) Be has maximum covalency state four but other can have six due to presence of d-orbitals.

Be (4) $1s^2\,2s^1\,2p_x^1\,2p_y^0\,2p_z^0$ (Four valence state)

e.g. $BeF_2 + 2F^- \longrightarrow [BeF_4]^{2-}$ tetrafluorido beryllate (II)

(iii) BeO and $Be(OH)_2$ are **amphoteric** while others are increasingly basic.

(iv) Be_2C on action with water gives methane while MgC_2 give C_2H_2 (acetylene)

(v) Be due to its low reactivity fails to react with water and acid while other form $H_2(g)$.

157. (i) $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$

milkiness

(ii) $CaCO_3 + H_2O + CO_2 \longrightarrow Ca(HCO_3)_2$

Soluble / milkiness diappears.

(iii) $Ca(HCO_3)_2 \xrightarrow{\Delta} CaCO_3 + CO_2\uparrow + H_2O$

turbidity

(iv) $CaSO_4 + 2H_2O \longrightarrow CaSO_4.2H_2O$

(v)
$$\left.\begin{array}{l} 3Mg + N_2 \xrightarrow{\Delta} Mg_3N_2 \\ 2Mg + O_2 \longrightarrow 2MgO \end{array}\right\} \text{white ash}$$

(vi) $Mg_3N_2 + 6H_2O \longrightarrow 3\,Mg(OH)_2 + 2NH_3 \uparrow$ (pungent odour)

$MgO + H_2O \longrightarrow Mg(OH)_2$

158. **Diagonal relationship with Al :**

Some of the similarities:

(i) $Be(OH)_2$ dissolves in excess of alklai to give beryllate ion, $[Be(OH)_4]^{2-}$ just as Al gives $\left[Al(OH_4)\right]^-$, aluminate ion.

(ii) Be and Al have Cl^- bridged structure in vapour phase. Both the chlorides are soluble in organic solvents and are strong Lewis acids. They are used as Friedel Crafts catalyst.

(iii) Be and Al ions can fom complexes $[BeF_4]^{2-}$, $[AlF_6]^{3-}$ because of identical polarising power.

(iv) BeO and Al_2O_3 are amphoteric in nature. These are high melting insoluble solids.

(v) Be_2C and Al_4C_3 on hydrolysis form methane (g).

159. **Biological importance of Mg and Ca :**

An adult body contains about 25 g of Mg and 1200 g of Ca compounds.

Functions :

(i) Chlorophylll, the green pigment of plant, contains magnesium, which makes photosynthesis possible.

(ii) All enzyme that utilise ATP in phosphate transfer requisite Mg as Co-factor.

(iii) About 99% of body calcium present in bones and teeth.

(iv) Calcium plays an important role in clotting of blood and stabilisation of protein structure.

(v) Calcium finds its use in the functing of neuro - muscular and interneuronal systems.

Unit - XI

160.	**Property**	**Boron**		**Aluminium**
(i)	Nature	Non metal	-	metal
(ii)	Nature compounds	Forms covalent compound only	-	Forms both covalent and ionic compounds.
(iii)	Action with steam.	Deoes not decompose	-	It decomposes liberating H_2. $Al_{(s)} + 3H_2O(g) \rightarrow Al(OH)_3 + 3H_2$
(iv)	Action with Mg	It forms compound as boron has higher electronegativity. $3Mg + 2B \rightarrow Mg_3B_2$	-	As 'Al' is a metal with $Mg^{(2\%)}$ it form an alloy, Magnallium.
(v)	Valence state	$B \rightarrow 1s^2\,2s^2\,2p^1$ can have maximum covalency state four. $[BF_4]^-$	-	$Al_{13} = 1s^2\,2s^2\,2p^6\,3s^2\,3p^1\,3d^0$ Due to presence of d-orbitals it can have covalency state six. $[AlF_6]^{3-}$
(vi)	Nature of oxide	B_2O_3 is acidic	-	Al_2O_3 is amphoteric

161. Diagonal relationship

$$\begin{tikzcd} B \arrow[r, "\text{Size decreases}"] \arrow[dr] & C \arrow[d, "\text{Size increases}"] \\ Al & Si \end{tikzcd}$$

(i) Atomic radius (along the period) decreases from B to C while increases C to Si (in the group). By such B and Si have comparable atomic size.

(ii) Electronegativity value : B(2.0) and Si (1.8) have almost identical electronegativity value.

(iii) Ionic potential (charge / radius) B (0.073) and Si (0.074) have identical ionic potential value and having same polarising power and tendency to form covalent compounds.

Similarity in properties :

(a) Both are non-metals hence bad conductors of heat & electricity.

(b) With hydrogen both forms gaseous hydrides, B_2H_6 and SiH_4.

(c) Both form stable acidic oxides.

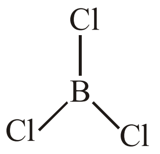
$$B_2O_3 + 3H_2O \longrightarrow 3H_3BO_3, SiO_2 + H_2O \longrightarrow H_2SiO_3 \quad \text{(metasilicic acid)}$$

(d) Both the oxides can be reduced by Mg to give the element.

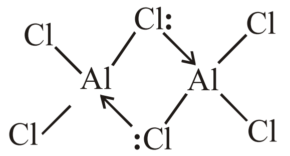
$$B_2O_3 + 3Mg \longrightarrow 2B + 3MgO,\ SiO_2 + 2Mg \rightarrow Si + 2MgO$$

162. (i) 'B' due to its smaller size cannot co-ordinate with four large halide ions whereas Al due to its bigger size is capable of co-ordinating with four halide.

The BCl_3 is a single unit cann't dimerise while $AlCl_3$ exists as Al_2Cl_6 (dimer) in the vapour state.

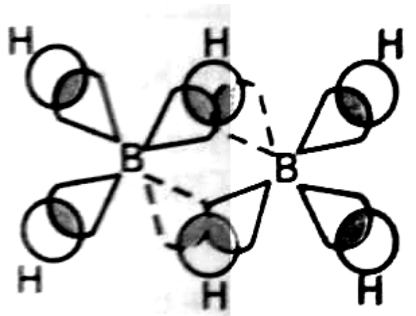


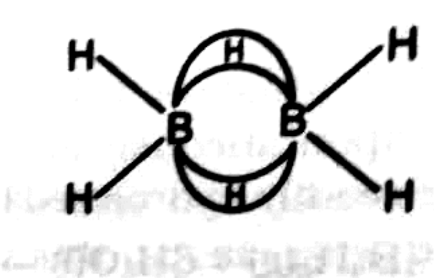
sp^2 - hybridised (Three bonds)



sp^3 hybridised (Four bonds)

(ii) Diborane (B_2H_6)





- * There are four terminal $2\bar{e}$-2 centre bond B–H bond.
- * There present two non-traditional $2\bar{e}$ three centre bonds i.e. BHB bridge bond / banana bond
- * Each 'B' atom uses its sp^3 hybrids. The terminal B–H bonds are normal and formed on overlapping of sp^3 of B and s-orbital of H (σ-bond).

  The $2\bar{e}$ three centre bridge bonds (2) are formed by 2 sp^3 hybrid orbital of two B with one $\bar{e}$ and one 1s orbital of H with one $\bar{e}$.
- * Thus, diborane is an electron deficient molecule $(12\,\bar{e}s)$ and acts as a Lewis acid.

  All bonds would have been normal if there present 14e(s) like $H_3C–CH_3$ (C_2H_6)

(iii) $_{81}T\ell \Rightarrow [Xe]4f^{14}\,4d^{10}\,6s^2\,6p^1$

Due to inert pair effect $T\ell$ has it compounds in +1 and +3 O.S. It has stable (+1) O.S. due to inert pair effect. The s^2 $\overline{e}$ (s) are reluctant to take part in bond formation. Thus $T\ell Cl$ is more stable than $T\ell Cl_3$. Because

(a) The (ns^2) penetrate to (n-1)d electrons and thus closer to nucleus & more effectively pulled towards nucleus & less available for bonding.

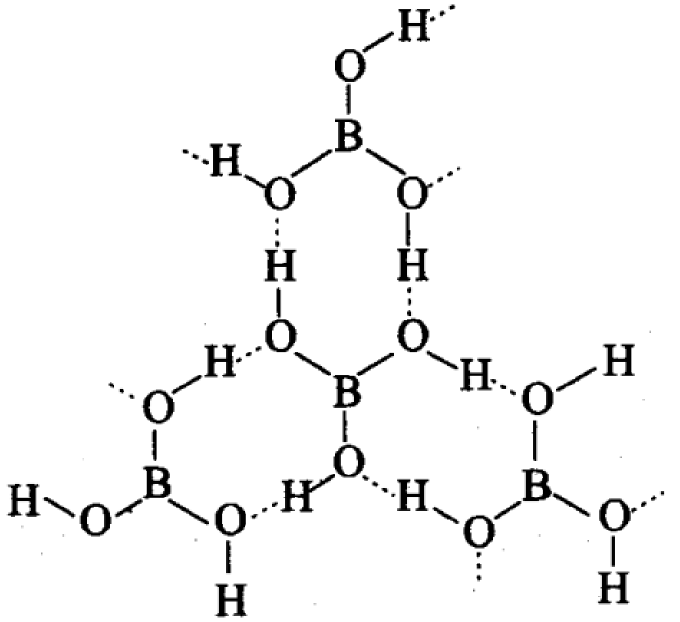
(b) Poor screening / ineffective shielding of d and 'f' electrons also accounts for this effect.

(iv) Borax on a platinum loop when heated forms a colourless glassy bead. When coloured salt of cobalt is heated on this bead dark blue colour is obtained due to formation cobalt meta borate.

$$Na_2B_4O_7.\ 10\,H_2O \xrightarrow{\Delta} Na_2\,B_4\,O_7 \xrightarrow{\text{red heat}} 2\,Na\,BO_2 + \underset{\text{(glassy bead)}}{B_2O_3}$$

e.g. $$\underset{\text{red}}{CoSO_4} + B_2O_3 \longrightarrow SO_3 + \underset{\text{dark blue}}{Co(BO_2)_2}$$

(v) Ortho boric acid, H_3BO_3 or $B(OH)_3$ where BO_3 units are joined by H–bonds as follows



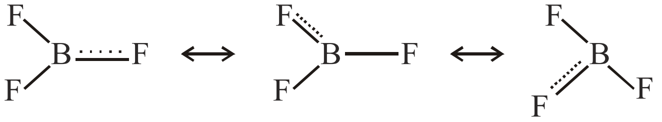
Structure of boric acid. the dotted lines represent hydrogen bonds.

It is a very weak acid and ionises as monobasic acid. Although it's not a proton donor, it behaves as a Lewis acid by accepting an $\overline{e}$ pair. $B(OH)_3$ is electron deficient i.e. in sextet state.

$$B(OH)_3 + H_2O \rightleftharpoons [B(OH)_4]^- + H_3O^+$$

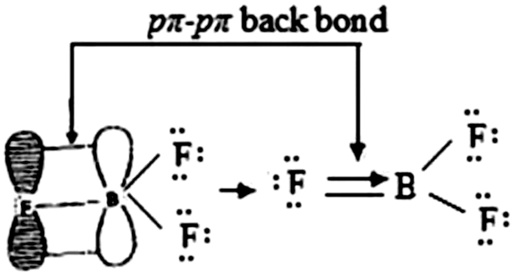
Here H_3BO_3 accepts an $\overline{e}$-pair from OH^- (i.e., from H_2O)

(vi) BF_3 is the weakest Lewis acid inspite of high electronegativity F (–I effect). This is due to the tendency of halogen (F) atom to back donate its lone pair of $p-\overline{e}$ (s) to empty p-orbital. (unhybridised) of boron atom through $p_\pi - p_\pi$ bonding. By such back donation and $\pi-$bond formate BF_3 can be represented as a resonance hybrid of three structures :



With increase of size of halogen atom the side wise overlapping and back donation tendency decreases as a result increases the electron deficiency & acidic nature. Thus

$BF_3 < BCl_3 < B\ Br_3 < BI_3$

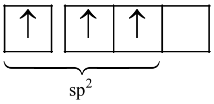


The formation of $p\pi = p\pi$ back bonding prominent in BF_3

* $p_\pi - p_\pi$ back bonding decreasing the electron deficiency at B .

* sp^2 hybridisation with a pure vacant p-orbital $\perp^r$ to it.

* ${}_9F\ 1s^2\ 2s^2\ 2p^5$, possesses filled p-orbital to form π -bond with vacant p-orbital of boron.



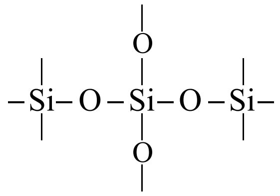
(vii) HNO_3 is an oxyacid which oxidises Al to aluminium oxide forming a protective layer on the metal (Al) surface when such Al dips inside any reagent it fails to react i.e., makes it passive. Thus, surface oxidation prevents corrosion also.

$$2HNO_3 \longrightarrow H_2O + 2NO_2 + [O]$$

$$2Al + 3[O] \longrightarrow Al_2O_3$$

(viii) Al form tripositive cation, Al^{3+}. Due to smaller size and higher charge, it has higher polarising power i.e., Cl^- gets deformed making $AlCl_3$ polar. Thus, $AlCl_3$ is covalent as per Fajan's rule.
Covalent character : $AlF_3 < AlCl_3 < AlBr_3 < AlI_3$

(ix) $O = C = O$



(a)	CO_2 is a discrete molecule.	(a)	It is polymeric having a gaint structure.
(b)	Due to smaller size of C. it can form C–O π-bond	(b)	No π bond tendency as Si atom is bigger. Si–O σ bonds are formed three - dimensionally.
(c)	CO_2 exists as gas because of its smaller size and held by weaker vanderWalls force	(c)	It exists as solid for there exists stronger Si–O covalent force.

(x) Carbon compounds are large in number. The chemistry of 'C' constitute a branch of chemistry, organic chemistry,

Reason :

(i) Carbon is having the highest catenating / self linking property due to greater C–C bond energy (355 kJ/mol)

(ii) Due to smaller size there exists $p_\pi - p_\pi$/ multiple bond. e.g. C=C, $C \equiv C$, $C = O$, $C = N$, $C \equiv N$

163. In diamond each C- is tetrahedrally bonds to four other C-atom through hybrid - hybrid $\sigma -$ bond where C – C bonds are three dimensionally arranged by stronger C–C bond. Hence it's very hard and the hardest of all substance. Due to its hardness it's used for cutting and polishing.

On the other in graphite another crystalline allotrope of 'C' where it is sp^2 hybridised and each 'C' is joined to three other C-atoms in an array of hexagons having layer, one sheet above another. Each layer of hexagons are weakly bonded (vander Waals force) and soft. Due to soft ness these are used as lubricants in heavy machinaries.

164. **Anomalous behaviour of carbon.**

(i) Smaller size and high electro-vity, high ΔiH value.

(ii) It has maximum O.N. +4 and minimum –4.

Other elements can have maximum covalency-6. e.g. $\left[SiF_6\right]^{2-}$

(iii) It fails to form complex due to absence of d-orbitals. ($C=1s^2\,2s^2\,2p^2$)

(iv) Due to smaller size it can form multiple bond i.e. $p_\pi p_\pi$ bond.

$C = C$, $C \equiv C$, $C = O$, $C = S$, $C \equiv N$.

(v) Catenation. (Self linking property) This is due to greater strength of C – C bonds. The catenation decreases down the group. C>>Si > Ge = Sn. Pb does not show catenation.

Due to property of catenation and $p_\pi p_\pi$ bond formation, carbon is able to show allotropic form and a large number of compounds known.

165.

		Diamond	**Graphite**
*	Colour :	Colourless	Dark grey
*	Appearance :	Transparent Crystal	Lustrous
*	Sp. gravity	3.5	2.5
*	Purity	100%	95 - 99 %
*	Hardness	Hardest of all with very high thermal conductivity	Soft and mark paper
*	Electrical conductivity	Bad conductor	Good conductor due to presence delocalised unhybridised p-electron.
*	C–C–bond length	Equal (1.54 $\overset{0}{A}$)	In the hexagonal layer all C-atoms are equally spaced (1.42 $\overset{0}{A}$) but there is difference of distance between two successive layers (3.4 $\overset{0}{A}$). This accounts for its softness.

*	Nature of hybridisation	sp^3, all are σ-bonds	sp^2 each C- having 3-σ bonds and one π-bond.

166. **Fullerenes**

This is the third crystalline allotrope of carbon (other two being diamond and graphite) with molecular formula C_{2n} $(n \geq 30)$. Most common is C_{60} which is also called bucky ball i.e., spheroidal shape.

Preparation : graphite $\xrightarrow[\text{He or Ar}]{\text{Electric are}}$ vapourised carbon $\xrightarrow{\text{Condensed}}$ Fullerene soot

$$\xrightarrow[\text{Over } Al_2O_3]{\text{Chromato graphy}} C_{60} + C_{70}$$

Unlike other two crystalline forms fullerene, the only pure form of carbon (because of absence of dangling edges or surface bonds) that dissolves in organic solvents. C_{60} being dissolved in toluene looks purple where C_{70} is orange red.

Structure The most stable fullerene, C_{60} looks like a soccer ball (bucky ball) containing 20 six membered rings and 12- five membered rings. 6-membered rings are fused with other six and five membered rings, but the five membered rings are connected to only six membered rings.

C_{70} fullerene resembles a rugby ball with 12 five membered rings and 25 six membered rings. Here also no- five membered rings are adjacent to each other.

Properties :

(i) Being covalent soluble is organic solvents.

(ii) With group 1 alkali metal it forms solids such as K_3C_{60} (Potassium fulleride) behaving as super conductor below 18 K.

(iii) It also forms platinum complex.

Uses : (a) Radio isotopes encapsulated C_{60} can be used in cancer therapy and AID therapy.

(b) During petro-chemical refining in catalytic amount.

167.

(a)	Dry ice	(vi)	Solid carbondioxide
(b)	Cryolite	(iv)	$3\ NaCl.\ AlCl_3$
(c)	Boron nitride	(vi)	isostructural with graphite
(d)	Carborundum	(ii)	silicon carbide
(e)	Borax	(iii)	Tincal
(f)	Corundum	(ii)	Aluminium oxide

168. $A \Rightarrow$ Carbon $\quad 3C + CaO \longrightarrow CaC_2 + CO$

$B \Rightarrow CaC_2(s)$

$C \Rightarrow CH \equiv CH \quad CaC_2 + 2H_2O \rightarrow Ca(OH)_2 + C_2H_2 \uparrow$

$D \Rightarrow Ag_2C_2 \quad C_2H_2 + 2AgNO_3 \xrightarrow{NH_3(\ell)} Ag_2C_2 + 2HNO_3$

169. Silicones : Silicones are long chain organo silicon polymers . Si along with its bond with ‘O’ linked to alkyl and phenyl satisfying its tetravalency.

Preparation :

(i) Preparation of alkyl chlorosilane

$$Si + 2CH_3Cl \xrightarrow[\text{Catalyst}]{\text{Cu}} (CH_3)_2SiCl_2$$

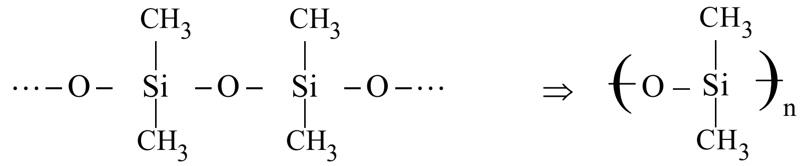
(ii) Hydrolysis of alkyl chlorosilane

$$(CH_3)_2\,SiCl_2 + 2H_2O \longrightarrow (CH_3)_2\,Si(OH)_2 + 2HCl$$

dimethyl dihydroxy silane

(iii) Condensation of alkyl hydroxy silane eliminating water molecules leads prepare a silicone, polymer with $-O-\overset{|}{\underset{|}{Si}}-O-$ linkage. ,

e.g.



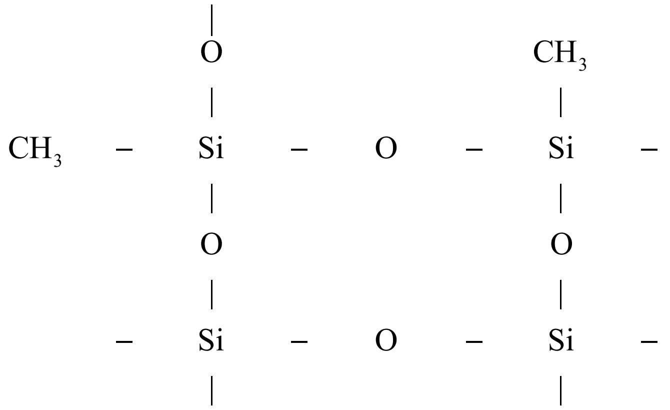
$$\cdots-O-\overset{CH_3}{\underset{CH_3}{Si}}-O-\overset{CH_3}{\underset{CH_3}{Si}}-O-\cdots \Rightarrow \left(O-\overset{CH_3}{\underset{CH_3}{Si}}\right)_n$$

Formation of cross linked silicone polymer –

$$SiCl_4 + CH_3MgCl \longrightarrow (CH_3)SiCl_3 + MgCl_2$$

Hydrolysis of trialkyl trichloro silane and subsequent condensation give rise to cross linked polymer

$$CH_3\,SiCl_3 + 3H_2O \longrightarrow CH_3\,Si(OH)_3 + 3HCl$$



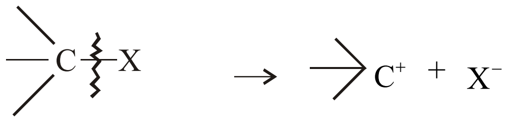
Uses : (i) Chain polymers are used as silicone oil & rubbers. The oils are used as greases, varnishes and resins.

(ii) Water repellants and good insulators.

Unit - XII

170. (i) Carbocation : When a covalent bond linked to C with more electronegative atom / group, breaks up by heterolysis the more electronegative atom takes using the $\overline{e}$ pair while C losing its $\overline{e}$ acquires positive charge.

Such organic ions carrying a positive charge are called carbonium ions or carbocations.



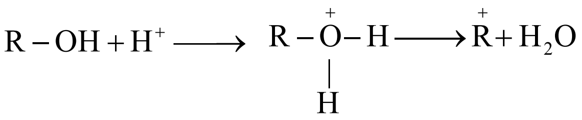
X = more electro –ve atom.

Formation :

(i) By heterolysis – $(CH_3)_3\,C\overset{\frown}{-}Cl \longrightarrow (CH_3)_3\,\overset{+}{C} + Cl^-$

(ii) Protonation of alkene / alcohol.

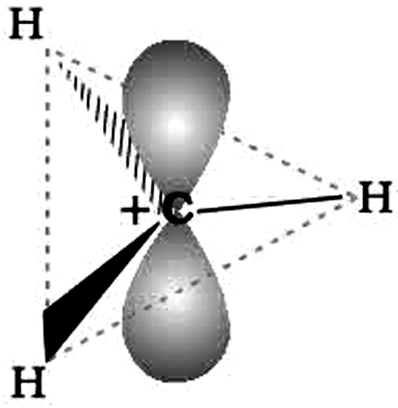
$$CH_2 = CH_2 + H^+ \longrightarrow \overset{+}{C}H_2 - CH_3$$



$$R - OH + H^+ \longrightarrow R - \overset{+}{\underset{\underset{H}{|}}{O}} - H \longrightarrow \overset{+}{R} + H_2O$$

Structure : C contains three $\bar{e}(s)$ as C^+ undergoing sp^2 hybridisation with a pure empty 'p' orbital.

$$C^+ \quad \underbrace{\boxed{\uparrow}\ \boxed{\uparrow}\boxed{\uparrow}}_{sp^2}\boxed{} - \text{empty}$$



Shape of methyl radical

Stability:

(a) $R_3C^+(3^0) > R_2\overset{+}{C}H(2^0) > R - \overset{\oplus}{C}H_2 > \overset{\oplus}{C}H_3$

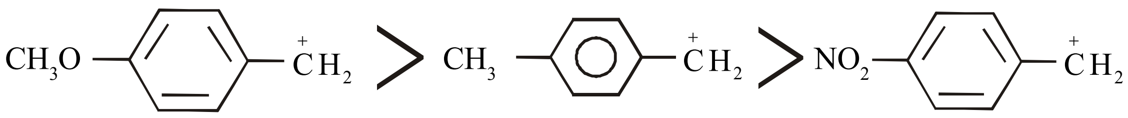
Increased +I effect increases stability (Less reactivity). This can also be explained by hyper-conjugation.

(b) Resonance stabilises greater than inductive effect (+I)

e.g. Tri phenyl methyl > benzyl > allyl > 3^0 cations

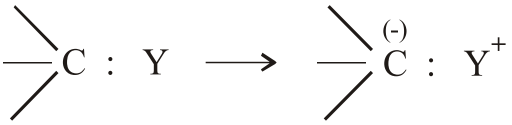
$$CH_2 = CH - \overset{\oplus}{C}H_2 \longleftrightarrow \overset{\oplus}{C}H_2 - CH = CH_2$$

Presence of electron releasing group in the phenyl ring makes the carbocation more stable while presence of EWG decreases stability.



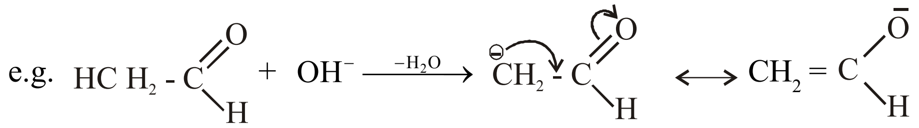
p-methoxybenzyl. p-methyltolyl – p-nitrotoly –

(ii) **Carbanion.** When C bonded to less electronegative atom (< 2.4) undergoes heterolysis the shared pair shifts to more electronegative 'C' rendering a –ve charge.



Formation : Taking away $\alpha-H$ by a base

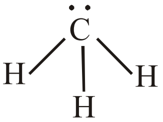
e.g.



Structure :

Here C-has five $\bar{e}$(s) in its valence shell, isoeelctronic with 'N' and undergoes sp^3- hybridisation forming 3-σ bonds leaving a lone pair in sp^3 hybrid orbital.

$\bar{C}\quad \underbrace{\boxed{\uparrow\downarrow}\ \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}$



(Pyramidal)

Stability :

(i) – I effect improves stability of carbanion

$$Cl-\bar{C}H_2 > \bar{C}H_3 > R\,\bar{C}H_2\left(1^0\right) > R_2\bar{C}H\left(2^0\right) > R_3\bar{C}\left(3^0\right)$$

The electron with drawing groups like $-NO_2$, $-CN$, $-COOC_2H_5$, $>C=O$ increases stability.

(ii) Resonance stabilises carbanion.

Thus, $\left(C_6H_5\right)_3\bar{C} > \left(C_6H_5\right)_2\bar{C}H > C_6H_5\bar{C}H_2 > CH_2=CH-\bar{C}H_2 > \bar{C}H_3$

(iii) Increased s-character of 'C' stabilises carbanion.

$$R-C\equiv\bar{C} \quad > \quad R_2CH=\bar{C}H \quad > \quad R_3C-\bar{C}H_2$$

s-character → 50% 33.3% 25 %

(iii) Free radicals : The neutral, unstable, reactive unpaired electron formed by homolytic cleavage of covalent bond are called free radicals.

Formation : (i) $Cl\bullet\wr\bullet Cl \xrightarrow{h\nu} Cl\bullet + \bullet Cl$

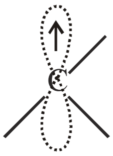
(ii) Free radical with a neutral molecule :

$$Cl\bullet + H\bullet\!-\!\bullet CH_3 \longrightarrow HCl + \bullet CH_3$$

Structure of alkyl free radical

Here 'C' with 3-σ bonds and one unpair $\bar{e}$ is undergoing sp^2 hybridisation with an $\bar{e}$ in pure p-orbital perpendicular to the trigon.

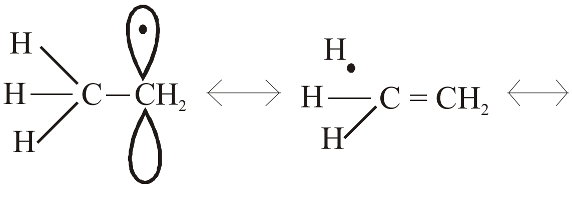
$C^{\bullet}\quad \boxed{\uparrow}\ \underbrace{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}_{sp^2}$



Stability :

(i) $R_3C^0\left(3^0\right) > R_2\dot{C}H\left(2^0\right) > R\dot{C}H_2\left(1^0\right) > \bullet CH_3$ (super primary)

The stability can be explained by hyper conjugation. More is the hyper conjugative structure more is the stability.



.... Three hyperconjugative structures.

(ii) Resonance stabilitises Benzyl, allyl free radicals.

e.g. $(C_6H_5)_3\dot{C} > (C_6H_5)_2\dot{C}H > C_6H_5-\dot{C}H_2 > CH_2=CH\dot{C}H_2 > (CH_3)_3\dot{C}$..........

$$CH_2=CH-\dot{C}H_2 \longleftrightarrow \dot{C}H_2-CH=CH_2$$

171. **Inductive effect** (Transmission effect) the permanent $\overline{e}$ displacement along a carbon chain due to presence of an atom / group with difference of $\overline{e}$ affecting tandency (releasing or with drawing) in place of 'H' atom of C–H bond. This effect is represented by an arro-whead in the middle of the covalent bond pointing in the direction of electron displacement ($\rightarrow$).

Types : – I effect : (Due to presence of $\overline{e}$ attracting group)

e.g. $C - \underset{\delta\delta\delta^+}{C} \rightarrow \underset{\delta\delta^+}{C} \rightarrow \underset{\delta^+}{\overset{2.5}{C}} \rightarrow \underset{\delta^-}{\overset{3.0}{Cl}}$

Cl (due to its $\overline{e}$ withdrawing tendency) attracts $\overline{e}$ from C– making δ^+ which influences shifting of $\overline{e}$ from the adjacent 'C' making it $\delta\delta^+$ (still smaller +ve influence). This influence of σ-electron shift is not seen after 3rd C-atom in the chain.

Groups influencing – I effect.

$$-NO_2 > -CN > -COOH > -F > -Cl > -Br > -I$$

$$> -O-R > -OH > -C_6H_5 > -CH=CH_2 > -H.$$

+I effect : This occurs when the chain is ended with electron releasing (pushing) group (Y).

$$\underset{4}{C}- \underset{3}{\overset{\delta\delta\delta-}{C}} - \underset{2}{\overset{\delta\delta-}{C}} - \underset{1}{\overset{\delta-}{C}} \leftarrow \overset{\delta+}{Y}$$

Groups influencing +I effect :

$$(CH_3)C_3- > (CH_3)_2CH > CH_3-CH_2- > CH_3- > D- > -H$$

Characteristics :

(a) It is a permanent effect.

(b) It is operative only in sigma bonds.

(c) Only $\sigma - e(s)$ are involved.

(d) Only $\overline{e}(s)$ are partially displaced rendering δ^+ and δ^- charges.

(e) It is transmitted along the –C–C– chain.

(f) As the length of the chain increases the inductive effect decreases.

Relative strength of acidic and basic nature :

As we know acids are electron deficient species more is the –I effect more is the acidic strength, on the other hand baseses have tendency to donate $\overline{e}(s)$ i.e. $\overline{e}$ rich species. Thus +I effect increases basic strength.

Example : Acids $F \leftarrow CH_2COOH > Cl \leftarrow CH_2COOH > Br \leftarrow CH_2COOH > I \leftarrow CH_2COOH$

$$\xleftarrow{-\text{I effect increases}}$$

+I effect decreases acidic strength

e.g. $HCOOH > CH_3COOH > C_2H_5COOH$

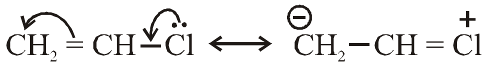
Bases : $(C_2H_5)_2\ddot{N}H > CH_3CH_2\ddot{N}H_2 > CH_3\ddot{N}H_2 > \ddot{N}H_3 > Cl \leftarrow \ddot{N}H_2$

$+$ I effect $\qquad\qquad\qquad\qquad\qquad\qquad$ $-$I effect

172. +R effect (Positive resonance / mesomeric effect)

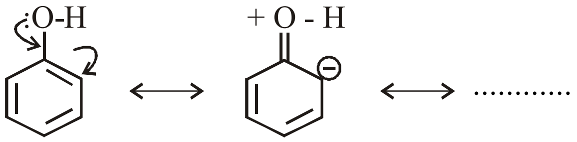
Groups ($\overline{e}$ releasing group) which donate $\overline{e}$ to conjugated system are said to have +R effect.

$-\ddot{O}R, -\ddot{O}H, -\ddot{N}H_2, -\ddot{N}HR, -\ddot{N}R_2, -\ddot{C}l, -\ddot{B}r$ etc.



$$CH_2 = CH - \ddot{C}l \longleftrightarrow \overset{\ominus}{C}H_2 - CH = \overset{+}{C}l$$

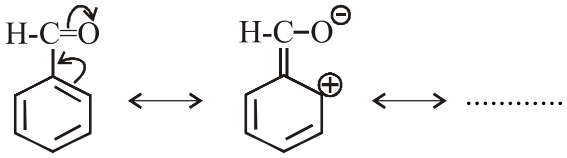
This is reason why vinylchloride is less reactive than CH_3-CH_2-Cl in respect of S_N reaction.



This makes phenol weakly acidic.

$-$R effect :

Groups which withdraws $\overline{e}$ (s) from the conjugated system towards themselves are said to exhibit $-$R effect.

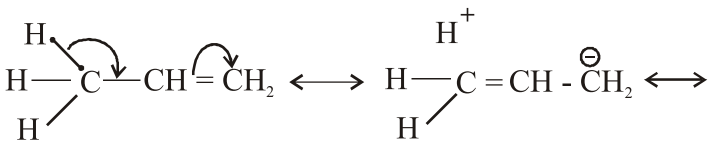


These electron with drawing groups are ————

— NO_2, — CN, ——COOH, ——CHO, ——COCl, ——$CONH_2$, ——COOR etc.

173. Hyper conjugation / No-bond resonance / Baker - Nathan effect.

The alkyl group wih atleast one H–atom on the α – C-atom, attached to unsaturated carbon atom, $\rangle C = C \langle$, carbocation or carbanion or free radical the σ - electron of C—H bond involved in delocalisation with π-electron.

e.g.  similar for other two $\alpha-H$ atoms.

Propene has three hyper conjugative structures.

In the resonating structure there is no definite bond between the C-atom and one of the α -H atoms, hence it's also known as no-bond resonance.

Stability of alkene :

More substituted alkenes are more stable i.e.

$$(CH_3)_2C=C(CH_3)_2 > (CH_3)_2C=CH-CH_3 > CH_3-CH=CH-CH_3 >$$

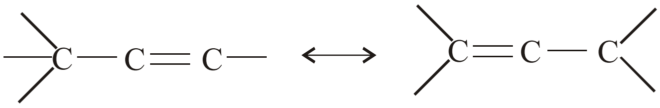
No. of α -H atoms : 12 $\qquad$ 9 $\qquad$ 6

$$CH_3-CH=CH_2 > CH_2 = CH_2$$

3 $\qquad$ 0

More is the $\alpha - H$ - atoms more is the hyper conjugative structures and more is the stability.

The bond lengths are also affected with conjugation and stability e.g. in propene the C-C bond length differs from C – C (1.54 $\overset{0}{A}$) and C = C 1.334, ^{0}A .



1.46 $\overset{0}{A}$ 1.353 $\overset{0}{A}$

Due to partial duble bond haracter C—C disance is 1.46 ^{0}A instead of 1.54 ^{0}A similarly due to single bond character the bond length increases from 1.334 to 1.353 ^{0}A.

174. **Detection of N is organic compounds (Lassaigne test)**

The organic compound is fused with shining sodium metal and fusion extract is boiled with ferrous sulfate and then acidified with conc. sulfuric acid. The formation of prusian blue colour confirms the pressence of 'N', the proceedure of CHN system is followed. If 'N' is absent procedure of CHO system is to be followed.

$$Na + C + N \xrightarrow[\text{fused}]{\Delta} NaCN \text{ (extracted with water)}$$

$$CN^{-} + Fe^{2+} \longrightarrow \left[Fe\,(CN)_6\right]^{4-}$$

$$Fe^{2+} \xrightarrow{\text{Conc.}H_2SO_4} Fe^{3+} + e^{-}$$

$$3\left[Fe(CN)_6\right]^{4-} + 4Fe^{3+} \xrightarrow{x\,H_2O} Fe_4\left[Fe\,(CN)_6\right]_3 \bullet x\,H_2O$$

Ferric ferrocyanide. (Prusian blue)

175. Types of organic reactions

There four main types of organic reactions :

(i) Substitution reaction

(ii) Elimination reaction

(iii) Addition reaction

(iv) Rearrangement reactions

(i) **Substitution reaction** : In these reactions atom / groups of atoms bonded to 'C' is replaced (substituted) by another atom / group without change in the C-skeleton.

These are : (a) Free radical substitution reaction.

e.g. Chlorination methane (alkane)

$$Cl_2 \xrightarrow{h\nu} 2Cl.\ \text{(chain initial)}$$

$$CH_4 + \bullet\, Cl \longrightarrow \bullet CH_3 + HCl\ (\text{chainpropagation})$$

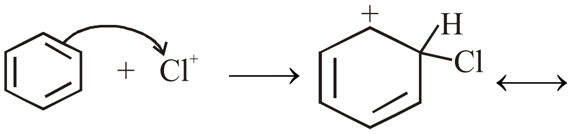
$$Cl..Cl + .CH_3 \longrightarrow CH_3Cl + .Cl$$

(b) Electrophilic substitution. e.g. chlorination of benzene.

As benzene is rich of $\pi\ \overline{e}(s)$ susceptable to be attacked by electrophile (e.g. Cl^{+}, $\overset{+}{N}O_2$, SO_3)

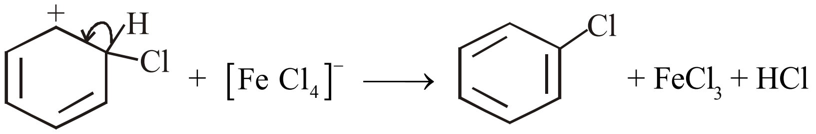
Formation of electrophile $Cl - Cl + FeCl_3 \longrightarrow [Fe\ Cl_4]^{-} + Cl^{+}$

Attack of electrophile



.... Stabilised by resonance

Deprotonation



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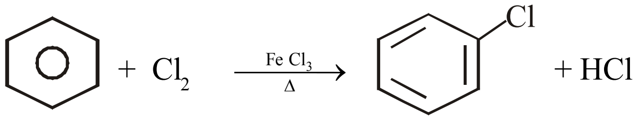


(c) Nucleophilic substitution reaction. (S_N)

$$CH_3-CH_2Br+KOH \longrightarrow CH_3-CH_2OH + KBr,\ OH^-$$ is the nucleophile

$$CH_3OH + HBr \longrightarrow CH_3Br + H_2O$$

$$R\,CH_2OH + H^+ \longrightarrow RCH_2\overset{+}{O}H_2 \longrightarrow R\overset{+}{C}H_2 + H_2O$$

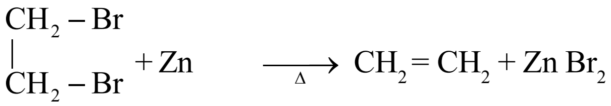
$$R\overset{+}{C}H_2 + Br^- \longrightarrow RCH_2Br$$

(ii) **Elimination reaction :** Saturated molecule (sp^3) (dehydration, dehydrohalogenation or dehalogenation) on elimination produces unsaturated molecule (sp^2/sp hybridised C)

Dehydration : $CH_3-CH_2-OH \xrightarrow[170^0\,C]{Conc.H_2SO_4} CH_2=CH_2+H_2O$

Dehydrohalogention $CH_3–CH_2Br \xrightarrow[KOH]{alcoholic} CH_2=CH_2+HBr$

Dehalogenation :



$$\begin{array}{l}CH_2-Br\\ |\\ CH_2-Br\end{array} + Zn \xrightarrow[\Delta]{} CH_2=CH_2+Zn\,Br_2$$

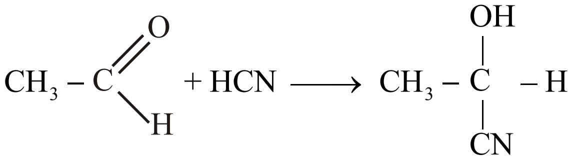
(iii) **Addition reaction :** The reagent adds to unsaturated double or triple bonded substrate forming saturated molecule. This is of three types :

(a) Free radical addition reaction

$$CH_3-CH=CH_2+HBr \xrightarrow[Peroxide\,effect]{R_2O_2} CH_3-CH_2-CH_2Br$$

1 – Bromopropane

(b) Nucleophilic addition reaction (Ad_N)



$$CH_3-C(=O)H + HCN \longrightarrow CH_3-\underset{CN}{\overset{OH}{C}}-H$$

acetaldehyde cynohydrin

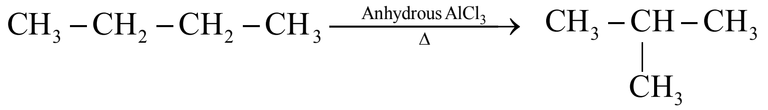
(c) Electrophilic addition reaction (Ad_E)

$$CH_2=CH_2+HBr \longrightarrow CH_3-CH_2Br$$

(Here H^+ is the electrophile)

(iv) Rearrangement reactions :

The reactions which involve the migration of an aton or group from one site to another site within the molecule resulting a new molecular structure are known as re-arrangement reactions. The new compound is just a structural isomer.



$$CH_3-CH_2-CH_2-CH_3 \xrightarrow[\Delta]{\text{Anhydrous AlCl}_3} CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_3$$

n-butane isobutane

$$NH_4CNO \xrightarrow[\Delta]{} H_2N-CO-NH_2 \text{ (urea)}$$

Ammonium cyanate

176. **Types of reagents :**

In organic reaction : substrate (substance taken) + reagent (added to it) → Product (s)

The attacking reagents are of two types :

(a) Electrophilics : (electro loving species) These are electrondeficient species attacking electron rich site.

Example : (i) Positively charged species. e.g. H_3O^+, Cl^+, Br^+, I^+, $\overset{+}{N}O_2$, $\overset{+}{N}O$ etc.

(ii) Neutral species with electron deficiency.

(Lewis acids) BF_3, $AlCl_3$, $FeCl_3$, $ZnCl_2$, $:CCl_2$ (carbene), SO_3 etc.

(b) Nucleophiles : (Nucleus loving) These electron donating species (Lewis bases) i.e., rich in $\bar{e}(s)$ attack at low electron density site of the substrate.

Example : (i) Negatively charged species : X^-, OH^-, $N\bar{H}_2$, $C\bar{N}$ etc.

(ii) Neutral molecule with unshared pair(s) of $\bar{e}$.

$$.\ddot{O}H_2,\ :NH_3,\ R-\ddot{\underset{..}{O}}-H,\ R-\ddot{\underset{..}{O}}-R,\ R\ \ddot{\underset{..}{O}}\ H, R\text{–}S\text{–}R, RNH_2, R_2NH, RMgX \text{ etc}$$

177. Acid taken 50 ml of 0.5 M H_2SO_4 (=1.N)= $\frac{50\times1}{1000}=0.05$ gram eqvt.

Residual acid = 60 ml of 0.5 = $\frac{60\times0.5}{1000}=0.03$ mol = 0.03 gram eqvt.

0.05 – 0.03 = 0.02 gram eqvt of acid reacts with 0.02 gram eqvt NH_3 formed = 0.02 gram eqvt NH_3

$$[\ 2NH_3+H_2SO_4 \longrightarrow (NH_4)_2SO_4$$

$$\frac{2}{2}\text{ mole of }NH_3 \equiv \frac{1\text{ mole}}{2}\ H_2SO_4 \text{ i.e, 1 gram eqvt. of } NH_3 \equiv \text{1 gram eqvt. of } H_2SO_4.]$$

1 gram eqvt. of NH_3 = 1 mole of NH_3 ≡ 14 gram (N)

0.02 gram eqvt of NH_3 contains 0.28 gram N

mass of the sample = 0.50 gram.

Thus, 0.5 gram sample contains 0.28 gram of N

100 gram sample contains $\frac{0.28}{0.50}\times100=56$

∴ % Nitrogen in the sample = 56.

178. It is a case of free radical substitution reaction hence does not occur in dark.

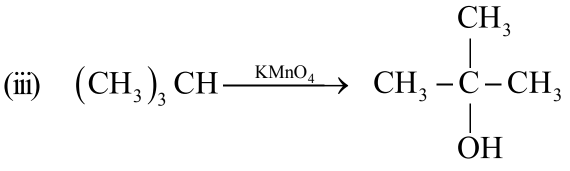
Reactivity of alkane $3^0 > 2^0 > 1^0 > . CH_3$ (See type of organic reactions)

Reactivity of halogen F > Cl > Br > I iodination is slow and reversible. Hence carried out in presence of oxidant like HNO_3 / HIO_3.

179. Controlled oxidation :

(i) $2CH_4 + O_2 \xrightarrow{Cu/523K/100\ atm} 2\ CH_3OH$

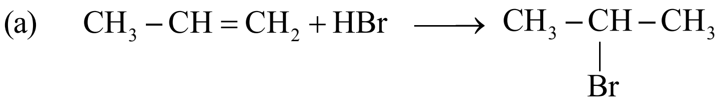
(ii) $CH_4 + O_2 \xrightarrow{Mo_2O_3} HCHO + H_2O$



(iii) $(CH_3)_3CH \xrightarrow{KMnO_4} CH_3-\overset{\displaystyle CH_3}{\underset{\displaystyle OH}{C}}-CH_3$

2-methyl propane t-butyl alcohol

180. Hydrobromination of propene.



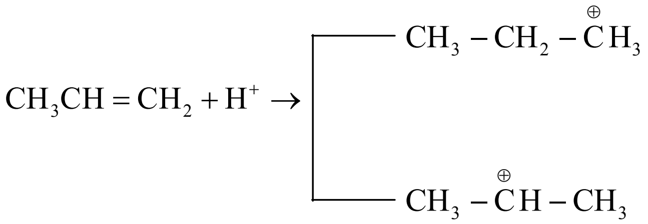
(a) $CH_3-CH=CH_2+HBr \longrightarrow CH_3-\underset{\displaystyle Br}{CH}-CH_3$

2-Bromo propane

Markownikoff's rule (Addition reaction)

When a polar molecules add to unsymmetrical alkene, the negative part of the addendum goes to the C-atom of the C=C, that contains less number of 'H' atom. Here, Br is the negative part and prefers to bond at second C-atom of the propene.

Mechanism : (AdE)



$CH_3CH=CH_2+H^+ \rightarrow CH_3-CH_2-\overset{\oplus}{C}H_3$ / $CH_3-\overset{\oplus}{C}H-CH_3$

2^0 – Carbonium being more stable yields more

$CH_3-\overset{+}{C}H-CH_2+Br^- \longrightarrow CH_3-CH(Br)-CH_3$ (2- Bromopropane)

(b) **Peroxide effect** (Kharasch effect)

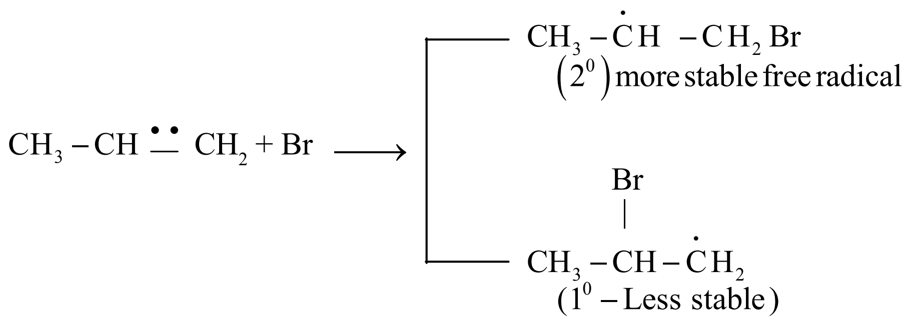
$CH_3-CH=CH_2+HBr \xrightarrow[\text{(Organic peroxide)}]{R_2O_2} CH_3-CH_2-CH_2Br$

1 – Bromopropane as major product

Mechanism : This is a case of free radical addition reaction.

$R-O \wr O-R \xrightarrow{homolysis} RO.$

$RO. + HBr \longrightarrow ROH + Br.$



$$CH_3-CH \overset{\bullet\bullet}{-} CH_2+Br \longrightarrow \begin{cases} CH_3-\dot{C}H-CH_2Br \quad \left(2^0\right)\text{more stable free radical} \\ CH_3-\underset{}{\overset{Br}{\overset{|}{C}H}}-\dot{C}H_2 \quad (1^0-\text{Less stable}) \end{cases}$$

$$CH_3-\dot{C}H-CH_2Br+HBr \longrightarrow \underset{\text{major product}}{CH_3-CH_2-CH_2Br}+\dot{B}r$$

181. **Acidic nature of alkyne :-**

Alkyne with terminal triple bond, $R-C\equiv CH$ behaves acidic and reacts with sodium liberating hydrogen and forming sodium salt.

Reason. In alkyne $(-C\equiv CH)$ carbon forms triple bond and undergoes sp- hybidisation where s-character is 50%. s-$\overline{e}(s)$ are closer to nucleus & hence increases the $\overline{e}$- accepting tendency (electronegativity).

Acidic nature : $HC\equiv CH > CH_3-C\equiv CH >> CH_3-C\equiv C-CH_3$

$$HC\equiv CH > CH_2=CH_2 > CH_3-CH_3$$

	$HC\equiv CH$	$CH_2=CH_2$	CH_3-CH_3
	sp	sp^2	sp^3
s-character :	50%	33.3%	25%

Salt fomation with strong conjugate base like NH_2^- or Na in $NH_3(\ell)$.

(i) $CH\equiv CH+Na \xrightarrow[-\frac{1}{2}H_2]{NH_3(\ell)} NaC\equiv CH \xrightarrow{Na} NaC\equiv CNa+\frac{1}{2}H_2$

Colourless

(ii) $HC\equiv CH+2AgNO_3 \xrightarrow{NH_3(\ell)} \underset{\text{white}}{AgC\equiv CAg\downarrow} + 2HNO_3$

$$NH_3+HNO_3 \rightarrow NH_4NO_3$$

(iii) $H-C\equiv CH+Cu_2Cl_2 \xrightarrow{NH_3(\ell)} \underset{\text{red}}{Cu\,C\equiv CCu\downarrow} + 2HCl$

$$NH_3+HCl \longrightarrow NH_4Cl$$

Applications :

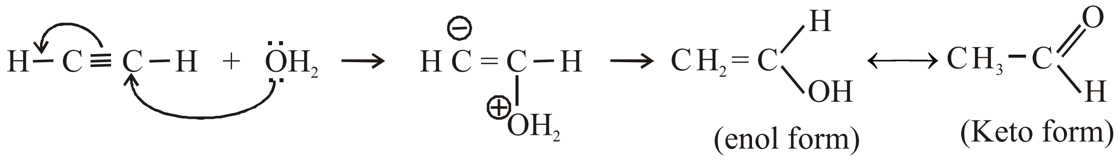
(a) But-1-yne and But - 2- yne can be distinguished from each other by the nature of salt formation.

(b) Ascending C-no. in the chain.

e.g $CH\equiv CNa+RBr \longrightarrow HC\equiv C-R+NaBr$ on such S_N reaction C-atom can be increased in the chain forming symmetrical $R-C\equiv CR$ and unsymmetrical alkyne $R-C\equiv C-R^{/}$ or $H-C\equiv C-R$ can be formed.

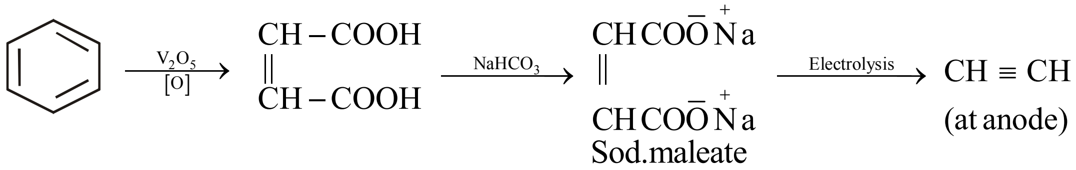
182. (i) When acetaldehyde is warmed with dilute sulfuric acid in presence of mercuric sulfate acetaldehyde is formed. This is a case nucleophilic addition reaction (Ad_N) due to sp. hybridised C, that makes C to accept electron from the nucleophile.

$$H-C\equiv CH+H_2O\xrightarrow[HgSO_4(60^0C)]{H_2SO_4}CH_3-CHO$$



(ii) $CH_2=CH_2\xrightarrow[(CCl_4)]{Br_2}BrCH_2-CH_2Br\xrightarrow{alc.KOH}H-C\equiv CH$

(iii)



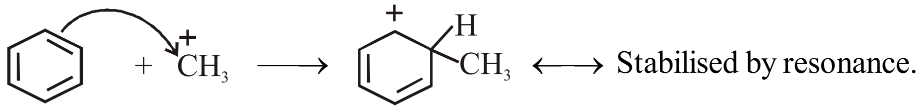
183. Fridel Craft's reaction : (S_E)

(i) Alkylation When benzene is treated with alkyl halide R–X (e.g. CH_3Cl) in presence of anhydrous $AlCl_3$ Alkyl benzene (Toluene) is formed.

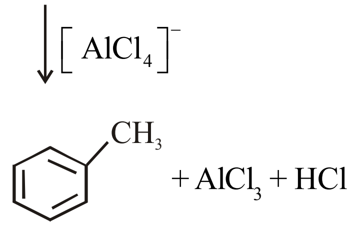
Formation of E^+: $CH_3Cl+AlCl_3\text{ (Anhy.)}\longrightarrow\overset{+}{C}H_3+Al\overline{Cl_4}$

(electrophite)

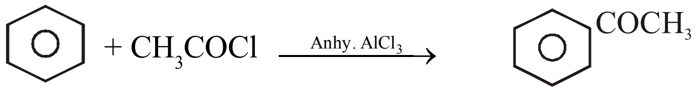
Attack of E^+ :



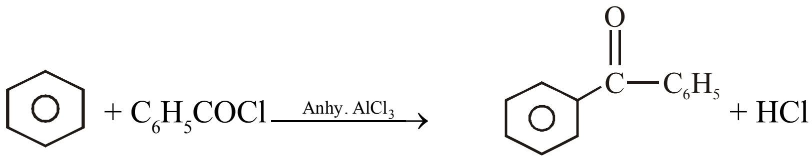
Deprotonation in presence of $[AlCl_4]^-$:



(ii) Acylation / benzoylation.



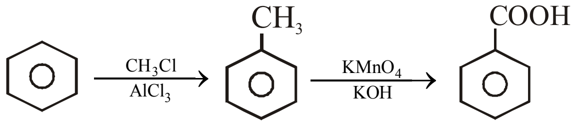
Acctophenone

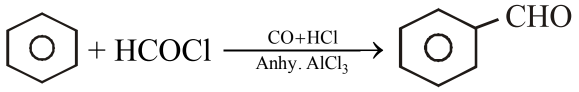


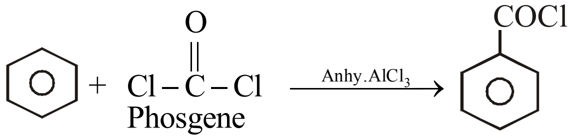
Benzophenone

Application : Ring substitution leads to form many synthetic products.

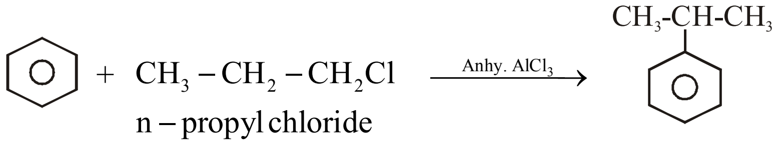
e.g.







Limitations i)



This process fails to give n-propyl benzene

$$CH_3-CH_2-CH_2Cl+AlCl_3 \longrightarrow CH_3-CH_2-\overset{+}{C}H_2+AlCl_4$$

1^0 (Less stable)

$$\Big\downarrow \text{rearranges}$$

$$CH_3-\overset{\oplus}{C}H-CH_3$$

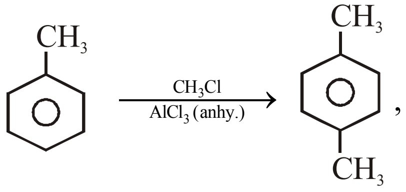
stable attacking electrophile (2^0)

ii) Presence of electron withdrawing group which deactivates the substitution (attack of electrophilic)

e.g. Nitro-benzene fails to undergo F–C reaction.

iii) Aryl substitution is not possible. Hence halide taken must be alkyl halide. C–Cl bond has π-character is aryl chloride.

iv) Polyalkylation may take place as mono-alkylation activates the ring.

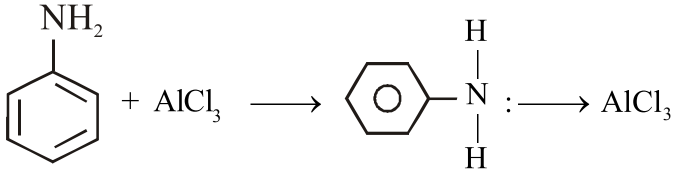


$-CH_3$ (activating group)

v) Ring having groups ($-NH_2$, $-NHR$, $-NR_2$) fails to undergo F.C. reaction.

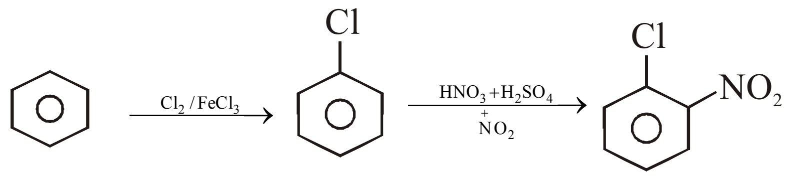
The $\bar{e}$ - present on 'N' (Lewis base) forms dative bond with anhy. $AlCl_3$ (Lewis acid) which fails to form electrophile.

e.g.

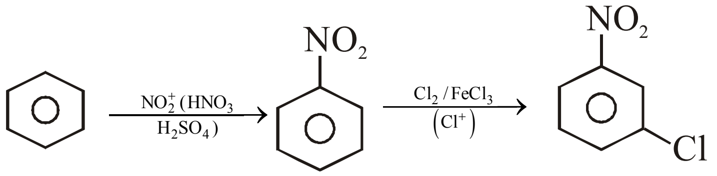


Adduct. (Donor-acceptor bond)

184.



Activating group Cl– is o/p - directing.



185. * Activating groups (Electron releasing group)

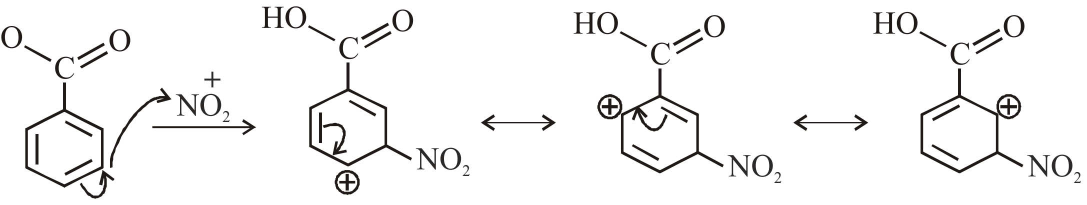
The groups that makes the benzene ring more $\overline{e}$ rich as a result easily attacked by electron deficient electrophile (E^+) are – OCH_3, – CH_3, –OH, –NH_2, –NHR, –NR_2, –X etc. and these are ortho, para directing .

* Deactivating groups (Electron withdrawing groups). The electron density on ring decreases disfavouring electrophilic attack. Here the incoming group (E^+) gets substituted at meta position.

e.g. $-CHO, -COR, CN, -NO_2, -CF_3, -\overset{+}{N}R_3$

186. In Benzoic acid, –COOH is a deactivating group and meta directing. So on nitration (Conc. HNO_3 + conc. H_2SO_4) forms m-nitrobenzene acid.

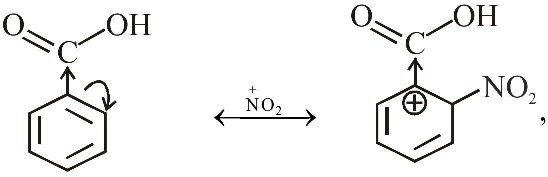
Mechamism:



All these canonical structures are stable on substitution of–NO_2 at meta possition as (+) charge is not shown at C. atom of the ring attached to withdrawing / deactivating group (–COOH)

On the contrary substitution at o–/p- position makes the canonical structures unstable.

e.g.



, +ve charge on the C- of the ring bonded to electron withdrawing – COOH group. Thus, substitution at o– position makes the canonical structure unstable. Thus formation of o-nitrobenzoic acid is negligible..

Unit - XIV

187. CO is a strong ligand. It bonds to haemoglobin of blood to form carboxyhaemoglobin which is about 300 times more stable than oxygen- haemoglobin complex. In blood, when concentration of carboxyhaemoglobin reaches about 3-4%, oxygen carrying capacity of blood is greatly reduced. This oxygen deficiency (hypoxia) results into headache, weak eyesight, nervousness and cardio-vascular disorder. This is the reason why CO is more dangerous, while CO_2 is required by green plants during photosynthesis liberating oxygen to the atmosphere.

188. Different sources of air pollution :

(i) Industrial chimney wastes (CO_2, CO, SO_2, N_2O_x, H_2S etc.)

(ii) Automobile exhaust (CO 77%, oxides of N 8%, unburnt hydrocarbon 15% and Noise pollution.

(iii) Burning of fossil fuels. (CO, CO_2, SO_2, N_2O_x, CH_4 etc.

(iv) Thermal power plants e.g. NTPC (SO_2, N_2O_x, flyash and hydrocarbon)

189. Chief sources of water pollution :

(a) Sewage and other waste

(b) Industrial waste (Waste from breweries, tanneriss, paper, textile) sugar, steel, paints etc.) These contain in organic substances like chlorides, carbonates, organic substances like phenol, alcohol etc., heavy metals like Lead, Hg, As and poisonous substances like cyanides, acetylene etc., colour producing dyes, acid and alkalis (rapidly changing pH of water affecting aquatic life.)

(c) Agricultural discharge, (excess pestcides, fertilisers)

(d) Radio-active wastes due to nuclear arm tests in ocean.
(e.g. Sr–90, Cs–137, C–14, H-3)

(e) shipping water pollution or oil pollution (Spellage of oil, affoshore oil drilling)

190. Detrimental effects of soil pollution :

(i) Industrial wastes containing toxic metals like Hg, Pb, Cd, As, Cu etc kill bacteria & baneficial microorganism in the soil.

(ii) Pesticides, being non-biodegradible seep into ground water through soil and contaminate public water supply system.

(iii) Dumped urban waste spreads several chronic diseases.

(iv) Excess fertiliser in the soil, inhibit synthesis of vitamin-C and cartene in vegetables and fruits.

(v) The radioactive wastes into the food chain and causes the disruption of physiological process and metabolic change.

(vi) Pathogenic bacteria in soil act as carrier of a number of chronic disease like cholera, typhoid, dysentry etc. from soil to man & vice versa.

191. (i) Depletion of ozone layer due to $Cl^{\bullet}$ from CFC is commonly known as ozone hole. A unique set of atmospheric conditions is responsible for ozone hole over south pole (Antarctica) in summer season, NO_2 and CH_4 react with chlorine monoxide and chlorine atom forming chlorine (Cl_2), prevents ozone depletion.

$$ClO^{\bullet} + NO_2(g) \longrightarrow ClONO_2(g) \qquad (1)$$

$$Cl^{\bullet}(g) + CH_4(g) \longrightarrow {}^{\bullet}CH_3 + HCl(g) \qquad (2)$$

$$ClONO_2(g) + H_2O(g) \longrightarrow HOCl(g) + HNO_3(g) \qquad (3)$$

$$ClONO_2(g) + HCl(g) \longrightarrow Cl_2(g) + HNO_3(g) \qquad (4)$$

In winter, special type of clouds called polar stratopheric clouds (PSC) formed over Antarctica which provide surface on which chlorine nitrate ($ClONO_2$) formed gets hydrolysed (3) and subsequently form $Cl_2(g)$ (eqn. 4)

When sunlight returns to Antarctica in the spring, the sun's wamnth breaks up clouds and HOCl and Cl_2 are photolysed to give (Cl.), responsible for ozone deplection.

$$HOCl(g) \xrightarrow{hv} .OH(g) + .Cl(g)$$

$$Cl_2(g) \xrightarrow{hv} .Cl(g) + .Cl(g)$$

$$Cl^{\bullet}_{(g)} + O_{3(g)} \xrightarrow{hv} ClO. + O_2(g)$$

(ii) Taj Mohal is a monument of marble stone which reacts with acid rain losing its lustre. Variety of human activities (burning of fossil fuels, metallurgical operation, burning of coal and oil in power plant, vehicular pollution etc.) emits the oxides of sulfur and nitrogen in the atmosphere. SO_2 and NO_2 after oxidation and reaction with water, are major contribution to acid rain.

$$2SO_{2(g)} + O_{2(g)} + 2H_2O(\ell) \longrightarrow 2H_2SO_4(Aq.)$$

$$4\,NO_{2(g)} + O_{2(g)} + 2H_2O_{(g)} \longrightarrow 4\ HNO_3(Aq)$$

$$CaCO_{3(s)} + H_2SO_4(Aq) \longrightarrow CaSO_4(s) + CO_{2(g)} + H_2O$$

The acid rain reacts with marble ($CaCO_3$) causing damage to this wonderful monument.

192. Chloro fluorocarbon (CFCs compounds or freons are non-reacive, non-flamable, non-toxic organic molecules used in refrigeation, production of plastic foam, cleaning computer parts etc. Once CFCs are released in the atmosphere, they mix with normal atmospheric gases and eventually reach the stratosphere, where they get broken down by powerful U.V. radiation, releasing chlorine free radical (Cl.)

$$CF_2\ Cl_2(g) \xrightarrow{U.V.} Cl.(g) + .CF_2Cl(g)$$

Cl. reacts with ozone (stratosphere), thus destroying ozone layer, which prevents U.V. rays reaching the earth surface.

$$.Cl(g) + O_3(g) \longrightarrow ClO. + O_2(g)$$

$$ClO.(g) + O(\text{gaseous atomic oxygen}) \longrightarrow Cl.(g) + O_2(g)$$

Thus, Cl. radicals are continuously regenerated & breaking ozone.

193. Missing of green house gases :

Green house gases trap the UV rays coming from Sun and heat up the earth i.e., responsible for global warming. In the absence of green house gases, the earth would not be able to retain any heat which is essential for survival of most of the species on earth i.e, without green house gases the global average temperature will decline making the present habitation questionable.

194. Green chemistry is defined as the sustainable, safe and non-polluting chemical science which enable chemists to manufacture products with minimum consumption of materials and energy producing minimum waste i.e., it's an environment friendly process.

Example : Elimination / replacement of hazardous chemicals :

H_2O_2 is used for the purpose of bleaching clothes in the process of laundry in place of hazardous, carcinogenic tetrachloride.

195. Incineration involves the burning of combustible wastes including house hold waste, chemical waste and biological waste at high temperature.

Importance :

(i) It is a part of waste management.

(ii) This type of waste treatment often used in the production of electricity. It reduces the mass of waste to 76%.

(iii) The produced ash on burning can be used as land-filling material.

Demerits : The gases produced cause environmental pollution. The ash produced might contain toxic heavy metals causing serious health hazards. Mercury, dioxins, lead and other pollutants come from burning waste. Dioxin produced in the incinerators (plant) is a cancer forming chemical.

◆◆◆

Group - C

(Long Questions)

1. State and explain law of equivalent proportion . [4

Water contains 11.11% of hydrogen, ammonia contains 82.35% nitrogen and nitrogen trioxide (N_2O_3) contains 63.15% oxygen. Show that result explains above law. [3

2. Discuss Bohr's atomic model. What are its merits and limitations ? [3 + 2 + 2

3. Explain different quantum numbers.

Show that no two $\bar{e}(s)$ of an atom can have four quantum number values identical. [5+2

4. Explain (i) $\bar{e}$ does not exist in the nucleus.

(ii) 2nd and 3rd shell contain 8 and 18 electrons respectively.

(iii) It is possible to predict the probability of locating an electron in a given regions.

[3 + 2 + 2]

5. Long form periodic table is completely based on electronic configuration. Explain. How does $\Delta_i H$ value change in the periodic table. [5 + 2

6. Write notes on

(i) Electron gain enthalpy (ii) Electronegativity [4 + 3

7. (a) Discuss Born-Haber cycle to calculate the lattice energy of NaCl.

(b) Show that

(i) Between two atoms there present only one σ – bond.

(ii) SO_4^{2-} and PO_4^{3-} are of identical shope. [4+3]

8. Write notes [4 + 3]

(a) Hybridisation.

(b) Hydrogen bonding

9. State and explain VSEPR theory. How does it predict the geometry of large number of molecules ?

[4 + 3]

10. Discuss VBT. How does it help to explain the formation of dihydrogen. [4 + 3]

11. Why does real gas deviate from ideal behaviour. Derive vander Waal's equation of state. [2 + 5]

12. State and explain Hess's law of constant summation of heat. Calculate the heat of formation of methane. Given : Heat od combustion of C, hydrogen and CH_4 are –393.5, –511.8 and –890.3 kJ respectively.

[2 + 2 + 3]

13. What is bond energy ? How is it calculated ? Mention the factors affecting bond energy. [2+2+ 3]

14. Write notes on (i) Law of mass reaction.

(ii) Arrhenius theory of acid an bases. [4 + 3]

15. State and explain Lechatelier's principle with reference to Haber's synthesis of ammonia. [1 + 6]

16. Write notes on (i) Buffer solutions

(ii) Common ion effect [4 + 3

17. What are redox reactions ? Discuss its applications. 1.575 g of $(COOH)_2.x\,H_2O$ crystals are dissolved in water and the volume made upto 250 ml. 8.34 ml of this solution is required for complete neutralisation of 10 ml of N/12 KOH. Calculate the water of crystallisation. [2 + 2 + 3

18. How is pure hydrogen peroxide prepared from (i) Sulfuric acid (ii) by Merck's process

Discuss its action with (a) aq. KI (b) Ferric sulfate is alkaline medium. [4 + 3]

19. Discuss the trends in physical properties of Group 1.

Write down their chemical reactivity. [5 + 2]

20. Group -2

21. Group -13

22. Group -14

23. Write notes on

(i) Isomerisation of organic compounds.

(ii) Wurtz reaction. [4 + 3]

24. Discuss the electrolytic methods of preparation of ethane, ethene and ethyne.

How can you get ethene from ethanol and vice versa ? [4.5 + 2.5]

25. Write notes on :

(i) Ozonolysis (ii) Polymerisation. [4 + 3

26. (a) What are the possible reactions in stratosphere ? [4

(b) The environment is getting warmer day by day - explain. [3

◆◆◆

ANSWERS

Group - C (Long Questions)

1.

Statement of the law	(2)
Explanation	(2)
Problem	(3)

In H_2O	11.11% H	0-88.89 %	H : O = 1 : 8
In NH_3	82.35% N	H-17.65%	
In N_2O_3	63.15% O	N 36.85%	

H : 0 = 1 : 8 , H : N = 1:4.7 , N : O = 1:1.7

In H_2O	1 part of H	8 parts of 'O'	
In NH_3	1-part of H	$\frac{82.35}{17.65}$ parts N = 4.7	(Part of H is kept fixed)

In N_2O_3 N : O = 36.85 : 63.15 = 1:1.7

Expected ratio of N:O = 4.7 : 8 = 1 : 1.7

Observed ratio of N : O (in N_2O_3) = 1 : 1.7

Hence, the law of reciproportion is verified.

2. **Bohr's atomic model :** Postulates [3], Merits any two [2], Draw backs (any two) [2]

Introduction : In order to overcome the draw backs of Rutherford's atomic model and to explain the origin of the atomic spectra Neil Bohr in 1913 gave the following postulates for an atom.

(i) The electrons are moving around the nucleus in circular paths having definite radii.

(ii) The electrons move around the nucleus in stationary orbits without losing-energy because energy in fraction of a quantum can not be lost (according to Plank's quantum theory).

(iii) Electrons can occupy any one such paths but it will occupy such a path for which its angular momentum is an integral multiple of $h / 2\pi$; i..e, Angular momentum, $m\,v\,r = n\frac{h}{2\pi}$

Where n = 1, 2, 3 .... etc.

(iv) Radius the path for 'H' like atom is, $r = \frac{n^2h^2}{4\pi^2me^2}$

(v) Energy of the path for 'H' like atom is $E = \frac{-2\pi^2\,me^4}{n^2\,h^2}$

(vi) Electrons can jump from one energy level to another which accounts for the origin of the spectra.

$$\Delta E = E_2 - E_1$$

(vii) Each orbit possess certain capacity to accomodate electrons. This is given by $2n^2$ rule, where 'n' is the number of path.

Value of n	Name of the path/ shell/orbit	Capacity of the shell

1	K - Shell	$2\times1^2=2$
2	L - shell	$2\times2^2=8$ and so an

(viii) The outer most shell (valence shell) will have maximum capacity of ‘8’ electrons.

(ix) The penultimate shell (n-1) will contain maximum 18 electrons.

Merits : (i) Explaining spectral lines : Electronic transitions from one orbit to another correspond to atomic spectra, the characteristic of an element.

(ii) Quantisation of electronic energy is key concept of Bohr’s theory principal quantum number (n) is introduced.

(iii) It explains the limitation of Rutherford model.

Draw backs :

(i) Bohr’s model is only applicable to hydrogen like atom i..e, it could explain the spectral behaviour of only one electron system like H, He^+, Li^{2+} etc.

(ii) The quantisation of angular momentum is purely arbitary one.

(iii) It could not explain “Zeeman’s effect”, “The splitting of a spectral line in to closely spaced cluster of spectral lines under the influence of an external magnetic field is known as Zeeman’s effect”.

(iv) It could not explain the “Stark effect”.

“ The splitting of a single spectral line into closely spaced cluster of spectral lines under the influence of an external electric field is called Stark effect”.

3. **Quantum Numbers :**

Electrons are identical with respect to their charge, mass and other properties. However, the energy content of the electrons differ from each other according to their existance in different shells. Since energy in quantised , we use a set of mathematical numbers to locate an electron and to distinguish between different electrons is known as “Quantum numbers”. They are of four types :

1. Principal Quantum number
2. Azimuthal or angular Quantum number
3. Magnetic Quantum number (Details- TB Page-133)
4. Spin Quantum number

Pauli’s exclusion principle.

According to Pauli “No two electrons of an atom will have four identical sets of quantum numbers”.

Explanation. **When we write $1s^2$,**

1st $\overline{e}$:	2nd $\overline{e}$:
$n=1$	$n=1$
$\ell=0$	$\ell=0$
$m=0$	$m=0$
$s=+\frac{1}{2}$	$s=-\frac{1}{2}$

When we represent in box form,

$\boxed{\downarrow\uparrow}$ $\boxed{\uparrow\uparrow}$ or $\boxed{\downarrow\downarrow}$

Correct incorrect

4. **(i) Absence of $\overline{e}$ in the nucleus**

(a) Rutherford scattering experiment : The large deflection of positively charged α-ray show negatively charged $\overline{e}$ is absent in the positively charged nucleus.

(b) Heisenberg's uncertainty principle :

Let us take $\overline{e}$ to be present in side nucleus of size (diameter) in the order 10^{-14}m (Δx), mass of $\overline{e} = 9.1\times10^{-31}$ kg, the minimum uncertainly in velocity (ΔV) will be ———

$$\Delta V = \frac{h}{4\pi}\times\frac{1}{\Delta x.m} \text{ (Heisenberg's uncertaintly principle } \Delta x.\Delta p = \frac{h}{4\pi}\text{ , } \quad \Delta p = m.\Delta V \text{)}$$

$$= \frac{6.6\times10^{-34}}{4\times\frac{22}{7}}\times\frac{1}{10^{-14}\times9.1\times10^{-31}}$$

$$= 6.\times10^{9} \text{ m/s}$$

This value is much higher than velocity of light $\left(3\times10^{8} \text{ m/s}\right)$ and hence not pssible to conceive the presence of $\overline{e}$ in the nucleus.

(ii) The number orbitals in a shell = n^2 Each orbital can have two $\overline{e}(s)$

So total number of $\overline{e}s.$ in a shell = $2n^2$

e.g. $n = 2$ no. of orbitals = one - 's'-orbital + three 'p' orbitals $= 4$ and no. of $\overline{e}$s = $2\times4 = 8$

n = 3 ☐ ☐☐☐ ☐☐☐☐☐ total 9 - orbitals

s(1) P(3) d(5)

and maximum no. of $\overline{e}$s = $2\times9 = 18$

(iii) Position of $\overline{e}$ is not certain like Bohr's model with definite radius in circular orbits around nucleus. Heisenberg contradicts such statement and expressed for minute particle like $\overline{e}$, the position and momentum cannot be determined simultaneously with accuracy such concept of uncertainty leads to define orbital i.e., the maximum probability of finding $\overline{e}$ around the nucleus in the three dimensional space e.g. s(spherical), p-(dumb-bell-shaped) d (double double dumb-bell shaped) etc. This is also supported by the wave mechanical concept (Schrodinger's wave equation) giving rise to radial probability function and angular probability density.

5. Long form periodic table basing on electronic configuration :

First period (2-elements)

As per Aufbau principle $\overline{e}$ enters first into 1s which can accomodate 2 $\overline{e}$ s, thus first period contain 2-elements H($1s^1$) and He($1s^2$)

2nd Period : (8 elements)

Here electron enters to 2s and 2p orbitals 2 elements in s-block and 6 elements in p-block. Thus, this is the second short period.

$$\left.\begin{matrix} {}_{3}\text{Li } 1s^2\ 2s^1 \\ {}_{4}\text{Be } 1s^2\ 2s^2 \end{matrix}\right\} 2-\text{s block} \qquad \left.\begin{matrix} {}_{5}\text{B } 1s^2\ 2s^2\ 2p^1 \\ \vdots \\ {}_{10}\text{Ne } 1s^2\ 2s^2\ 2p^6 \end{matrix}\right\}(6)\,\text{p}-\text{block}$$

3rd period (8 elements) having $\overline{e}$ (3 s) and (3p)

$$\left.\begin{matrix} {}_{11}\text{Na } 2,\ 8,\ 3s1 \\ {}_{12}\text{Mg } 2,\ 8,\ 3s^2 \end{matrix}\right\} \text{s}-\text{block (2)}$$

$$\left.\begin{array}{l} {}_{13}\text{Al}\ \ 2,\ 8,\ 3s^2 3p^1 \\ \quad | \\ {}_{18}\text{Ar}\ \ 2,\ 8,\ 3s^2\ 3p^6 \end{array}\right\} p-\text{block (6)}$$

4th period (18 elements) This is the first long period where $\overline{e}$(s) successively enter to 4s, 3d and 4p subshells.

$$\left.\begin{array}{l} {}_{19}\text{K}\ \ 2,\ 8,\ 8\ 4s^1 \\ {}_{20}\text{Ca}\ \ 2,\ 8,\ 8\ 4s^2 \end{array}\right\} s-\text{block (2)} \qquad \left.\begin{array}{l} {}_{21}\text{Sc}\ \ 2,\ 8,\ 3s^2\ 3p^6\ \ 3d^1\ 4s^2 \\ \ \vdots \\ {}_{30}\text{Zn}\ \ 2,\ 8,\ 3s^2\ 3p^6\ \ 3d^{10}\ 4s^2 \end{array}\right\} \begin{array}{c} d-\text{block} \\ 3d-\text{series (10)} \end{array}$$

$$\left.\begin{array}{l} {}_{31}\text{Ga}\ \ 2,\ 8,\ 18\ \ 4s^2\ 4p^1 \\ \ \vdots \\ {}_{36}\text{Kr}\ \ 2,\ 8,\ 18\ \ 4s^2\ 4p^6 \end{array}\right\} \begin{array}{c} p-\text{block} \\ (6-\text{elements}) \end{array}$$

5th period (18 elements) It is the second long period with 18 elements succesively entering into 5s, 4d and 5p subshells.

$$\left.\begin{array}{l} {}_{37}\text{Rb}\ \ 2,\ 8,\ 18\ \ 8\ \ 5s^1 \\ {}_{38}\text{Sr}\ \ 2,\ 8,\ 18\ \ 8\ \ 5s^2 \end{array}\right\} \begin{array}{c} s-\text{block} \\ (2-\text{elements}) \end{array} \qquad \left.\begin{array}{l} {}_{39}\text{Y}\ \ 2,\ 8,\ 18\ \ 4s^2\ \ 4p^6\ \ 4d^1\ 5s^2 \\ \ \vdots \\ {}_{48}\text{Cd}\ \ 2,\ 8,\ 18\ \ 4s^2\ \ 4p^6\ \ 4d^{10}\ 5s^2 \end{array}\right\} \begin{array}{c} d-\text{block} \\ (4d-\text{series}) \\ 10\ \text{elements} \end{array}$$

$$\left.\begin{array}{l} {}_{49}\text{In}\ \ 2,\ 8,\ 18\ \ 18\ \ 5s^2\ \ 5p^1 \\ \ \vdots \\ {}_{54}\text{Xe}\ \ 2,\ 8,\ 18\ \ 18\ \ 5s^2\ \ 5p^6 \end{array}\right\} \begin{array}{c} p-\text{block} \\ (6-\text{elements}) \end{array}$$

Sixth period (32 elements) This is the first longest period with 32 elements (6s block-2+4f bock 14 + 5d-block 10+5p-6) $_{55}$Cs 2, 8, 18, 18, 8, 6s^1

$_{56}$Ba 2, 8, 18, 18, 8, 6s^2.

$_{57}$La (exception) 2, 8, 18, 18 5s^2 5p^6 5d^1 6s^2. Here 57th $\overline{e}$ enters into 5d instead of 4f.

$$\left.\begin{array}{l} {}_{58}\text{C}_e\ \ 2,\ 8,\ 18,\ 4s^2\ 4p^6\ d^{10}\ 4f^2\ 5s^2\ 5p^6\ 5d^0\ 6s^2 \\ \vdots \\ {}_{71}\text{Lu}\ \ 2,8,18,\ 4s^2\ 4p^6\ 4d^{10}\ 4f^{14}\ 5s^2\ 5p^6\ 5d^1\ 6s^2 \end{array}\right\} \begin{array}{c} \text{4f series} \\ \text{(Lanthanides)} \end{array}$$

$$\left.\begin{array}{l} {}_{72}\text{Hf}\ \ 2,\ 8,\ 18,\ 32\ 5s^2\ 5p^6\ 5d^2\ 6s^2 \\ \ \vdots \\ {}_{80}\text{Hg}\ \ 2,\ 8,\ 18,\ 32\ 5s^2\ 5p^6\ 5d^{10}\ 6s^2 \end{array}\right\} \begin{array}{l} \quad d-\text{block} \\ \text{along with La there are} \\ 10\ \text{elements} \end{array}$$

$$\left.\begin{array}{l} {}_{81}\text{T}\ell\ \ 2,\ 8,\ 18,\ 32\ 5s^2\ 5p^6\ 5d^{10}\ 6s^2\ 6p^1 \\ \ \vdots \\ {}_{86}\text{Rn}\ \ 2,\ 8,\ 18,\ 32\ 5s^2\ 5p^6\ 5d^{10}\ 6s^2\ 6p^6 \end{array}\right\} \begin{array}{c} p-\text{block} \\ (6\ \text{elements}) \end{array}$$

Seventh Period . This is the second longest period (last period) in the periodic table containing 32 elements similar to 6th period, containing $\overline{e}$ in 7s, 5f, 6d and 7 p subshells.

$$\left.\begin{array}{l} {}_{87}\text{Fr } 2, 8, 18, 32\ 5s^2\ 5p^6\ 5d^{10}\ 6s^2\ 6p^6\ 7s^1 \\ {}_{88}\text{Ra } 2, 8, 18, 32\ 5s^2\ 5p^6\ 5d^{10}\ 6s^2\ 6p^6\ 7s^2 \end{array}\right\} \begin{array}{l} s-\text{block} \\ (2-\text{elements}) \end{array}$$

Here also ${}_{89}$Ac has exceptional configuration where 89th $\overline{e}$ enters to 6d instead of 5f.

${}_{89}\text{Ac } 2,\ 8,\ 18,\ 32\ \ 5s^2\ 5p^6\ 5d^{10}\ 6s^2\ 6p^6\ 6d^1\ 7s^2$

$$\left.\begin{array}{l} {}_{90}\text{Th } 2, 8, 18, 32\ 5s^2\ 5p^6\ 5d^{10}\ 5f^0\ 6s^2\ 6p^6\ 6d^2\ 7s^2 \\ {}_{91}\text{Pa } 2, 8, 18, 32\ 5s^2\ 5p^6\ 5d^{10}\ 5f^2\ 6s^2\ 6p^6\ 6d^1\ 7s^2 \\ \vdots \\ {}_{103}\text{Lr } 2, 8, 18, 32\ 5s^2\ 5p^6\ 5d^{10}\ 5f^{14}\ 6s^2\ 6p^6\ 6d^1\ 7s^2 \end{array}\right\} \begin{array}{l} 5f\ \text{series} \\ =14\,\text{elements} \end{array}$$

$$\left.\begin{array}{l} {}_{104}\text{Rf } 2, 8, 18, 32\ 5s^2\ 5p^6\ 5d^{10}\ 5f^{14}\ 6s^2\ 6p^6\ 6d^2\ 7s^2 \\ \vdots \\ {}_{112}\text{Cn } 2, 8, 18, 32\ 5s^2\ 5p^6\ 5d^{10}\ 5f^{14}\ 6s^2\ 6p^6\ 6d^{10}\ 7s^2 \end{array}\right\} \begin{array}{l} 6d\ \text{series} \\ 10\ \text{elements} \\ \text{includingAc.} \end{array}$$

$$\left.\begin{array}{l} {}_{113}\text{Nh (Nihonium) } 2, 8, 18, 32, 32, 18, 7s^2, 7p^1 \\ \vdots \\ {}_{118}\text{Og (Organesson) } 2, 8, 18, 32, 32, 18, 7s^2, 7p^6 \end{array}\right\} \begin{array}{l} p-\text{block} \\ =6\text{elements} \end{array}$$

* **Down the groups :**

The $\Delta_i H$ value generally decreases in moving down the group with increase of shell number, however exceptions are there due to effective nuclear charge $\left(\text{ENC} = Z^*\right)$ which increases due to less shielding of d & f electrons. This is the reason why $\Delta_i H\ \left({}_{82}\text{Pb}\right) > \Delta_i H\ {}_{50}\left(\text{Sn}\right)$ is group - 14

* Along the period the $\Delta_i H$ value generally increases from left to right with increase of nuclear charge and no change in shell number. However exceptions are due to stable electronic configuration. More stable (half-filled / full filled) is the configuration higher is the $\Delta_i H$ value.

e.g. In 2nd period $\Delta_i H_{(\text{Be})} > \Delta_i H\left({}_5\text{B}\right)$, $\Delta_i H_{({}_7\text{N})} > \Delta_i H_{({}_8\text{O})}$

(Full filled) half-filled

6. (i) Electron gain enthalpy : TB page - 166

Statement - [1]

Factors affecting - [2]

Significance [1]

(ii) Electronegativity . TB page - 167

Statement [1]

Variation in periodic table [1]

Significance. [1]

(i) Polarity of the bond

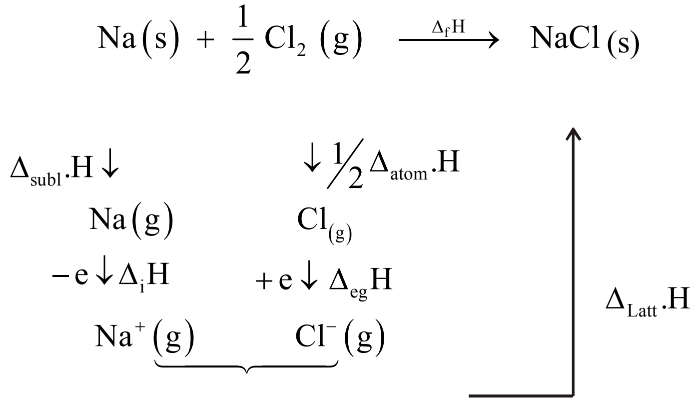
(ii) Nature of the bond e.g. $\Delta x > 1.7$ more ionic, Δx = difference of electronegativity.

(iii) % of ionic character

(iv) Distinction between metal and non-metal.

7. (a) Born-Haber cycle (Calculation of Lattice energy)

Based n first law of thermodynamics i.e., conservation of energy.



According to first law of thermodynamics / Hess's law

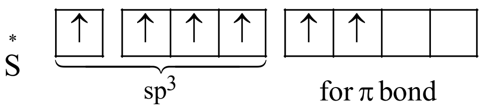
$$\Delta_f H = \Delta_{sub} H + \Delta_i H + \frac{1}{2}(\Delta_{atom} Cl_2) + \Delta_{eg} H + \Delta_{latt.} H$$

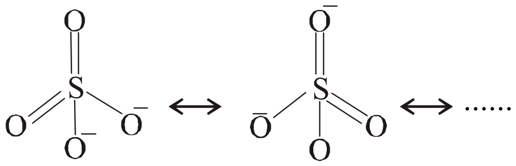
Substituting heat of formation (–ve), heat of sublimation of Na (+ve), ionisation of enthalpy of Na(g) (+ve), Bond dissociation of / mole of Cl (+ve), electron gain enthalpy Cl(g) (–ve), the lattice energy of $Na^+ Cl^-$ can be calculated.

$$Na^+_{(g)} + Cl^-_{(g)} \longrightarrow NaCl(s) + \text{energy(L.E)}$$

(b) (i) Between two nuclei there present only one internuclear axis and $\sigma-$ bond is formed by axial overlapping (s-s, p-p, s-p etc). Hence there is one $\sigma-$ bond between two atoms e.g. H–Cl.

(ii) * $\overset{+6}{S}O_4^{2-}$ (6 unpair electrons to form 4σ bonds and 2π bonds as o^-, o^- form σ bond and O, O form two double bonds $(2\sigma$ and $2\pi)$

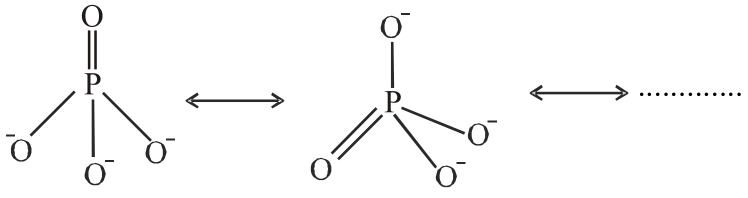




* PO_4^{3-} Total number of valence $\bar{e}s = 5 + (4 \times 6) + 3 = 32$

$32 \div 8 = 4$ quotient and '0' reminder i.e., it does not have lone pair.

Total number of stereo specific pairs = 4+0 = 4, i.e., it (P) under goes sp^3 hybridisation, three sp^3 hybrid orbitals single bonded to $\bar{O}$ and other 'O' forms double bond with P.



Thus, both SO_4^{2-} and PO_4^{3-} are tetrahedral in shape.

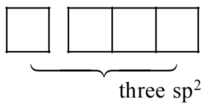
8. **(a)** **Hybridisation**

- * Definition.
- * Steps involving hybridisation.

(i) Excitation to have required number of unpair $\overline{e}$ s e.g. $\overset{+4}{S}\overset{-2}{O}_2$ where 'S' to be excited for availability of four unpair $\overline{e}(s)$

(ii) Mixing and re-orientation.

The pure orbitals mix together to have hybrid orbitals with identical shape and energy but reoriented in space to minimise repulsion among themselvs.

e.g. 

s-orbital mix with two 'P' orbitals to have threee sp^2 hybrid orbitals with s, p character in the ratio 1:2. These three sp^2 hybrid orbitals reorient at an angle of 120^0 trigonally.

- * Conditions of hybridisation –
- * Characteristics–

Example : sp^3 (CH_4).

Refer Text book Page - 207

(b) Hydrogen bond. Statement –
Condition for formation –
Types :
Consequence :

Refer text book Page 222

9. **VSEPR- Theory**

In a covalent molecule / ion there present two types of $\overline{e}$ pairs. One that is taken part in bonding (bond pairs), other i.e., unshared, not taking part in bonding. The $\overline{e}$ pairs surrounding the central atom repel one another and move apart from one another that there are less repulsion between them. As a result, the molecule has minimum energy and maximum stability.

Basic assumptions of valence shell electron pair repulsion theory :

(i) The shape of molecule containing only two atoms is always linear.

(ii) If the central atom is linked to similar atoms and surrounded by only bond pairs (bp) of $\overline{e}$ s the repulsions among them are similar. The shape of the species (molecule / ion) is symmetrical and having regular geometry. e.g. BF_3 (sp^2. trigonal) CH_4 (sp^3 tetrahedral)

(iii) If the central atom is linked to different atoms or surrounded by bond pairs as well as lone pairs (lp), the repulsion between them is different as a result the molecule has an irregular or distorted geometry. The order of repulsion : lp – lp > lp-bp > bp - bp

(iv) The $\overline{e}$ pairs (bp) determines the shape of the molecule. e.g. NH_3 though sp^3 hybridised but pyramidal. Presence of lone pair distort the shape, bond angle.

(v) The double and triple bond are treated as a single bonding group but affect the bond angle when present adjacent to single bond e.g. $POCl_3$ 111^0 in place 120^0.

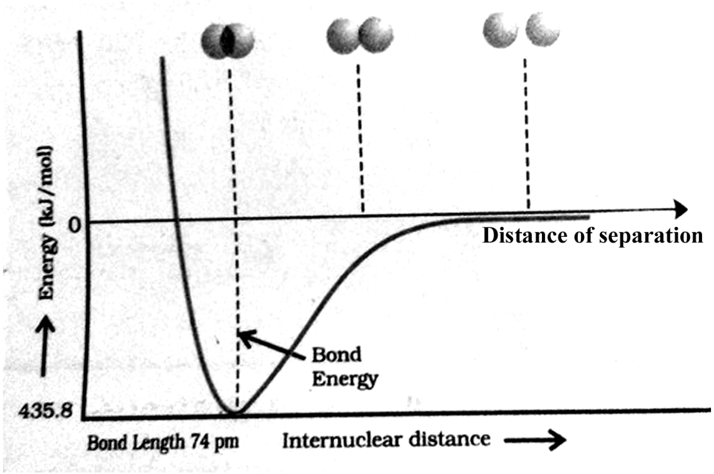
Conclusion :

VSEP number	**Geometry of the $\overline{e}$ pairs**	**Type of Molecule** **B=bond pair** **L=lone pair**	**Shape**	**Example**
2	Linear	AB_2	Linear	$BeCl_2$, CO_2
3	Trigonal	AB_3	Trigonal	BF_3
		AB_2L	Bent	$SnCl_2$
4	Tetrahedral	AB_4	Tetrahedral	CH_4
		AB_3L	Pyramidal	NH_3
		AB_2L_2	Bent	H_2O
5	Trigonal bipyramidal	AB_5	Trogonal bipyramidal	PCl_5
		AB_4L	See-saw	SF_4
		AB_3L_2	T-shaped	ClF_3
		AB_2L_3	Linear	XeF_2
6	Octahedral	AB_6	Octahedral	SF_6
		AB_5-L	Square pyramidal	IF_5
		AB_4L_2	Square planar	XeF_4
7	Pentagonal bipyramidal	AB_7	Pentagonal bipyramidal	IF_7
		AB_6L	Distorted octahedral	XeF_6

10. Valence bond theory (VBT)

Postulates - Refer Text book Page 198

Explanation of formation of H_2:



The potential energy curve for the formation of H_2 molecule as a function of internuclear distance of the H atoms. The minimum in the curve corresponds to the most stable of H_2

$$H(g)+H(g)\longrightarrow H_2(g)+435.8\,\text{kJ mol}^{-1}$$

The decrease in energy as shown in the graph.

11. Causes of deviation. [2]

Derivation of $\left(P+\frac{an^2}{v^2}\right)(V-nb)=nRT\ [5]$

Refer Text book page - 275 - 279

12. **Hess's law of constant summation of heat**

Statement - [2] Explanation [2] problem - [3]

(i) $C_{(s)}+O_2(g) \longrightarrow CO_2(g)$ $\Delta H=-393.5\ \text{kJ}$

(ii) $2H_{2(g)}+2\times\frac{1}{2}O_2(g)\longrightarrow 2H_2O(\ell)$ $\Delta H=-571.8\times 2\,\text{kJ}=-1143.6$

(iii) $CH_4(g)+2O_2(g)\longrightarrow CO_2(g)+2H_2O(\ell)$ $\Delta H=-890.3\ \text{kJ}$

[(i) + 2 [ii] - (iii)] $\Rightarrow$ $C_{(s)}+2H_2(g)=+CH_4(g)$

$= -(393.5+1143.6)+890.3$

$= -646.8\,\text{kJ/mol}$

13. **Bond energy**

Definition – [2]

Calculations [2]

Factor governing bond energy [3]

(3 - factors)

Refer Text book page - 333

14. **Law of mass action :**

(i) Statement and Application [2+2]

Refer Text book Page - 362

(ii) Arrhenius theory of acid and base -

Features :

Refer Text book Page -382

15. **Le-chatelier's principle**

Statement : If a system in equilibrium is subjected to a change in concentration, pressure or temperature then the equilibrium shifts in that direction, which tends to undo or reduce the effect of change (stress).

Explanation : The above stresses i.e. change in concentration, pressure or temperature are explained on the following equilibrium.

Habers process : $N_2(g)+3H_2(g) \rightleftharpoons 2NH_3(g)\ \Delta H=-92\,\text{kJ}$

(i) Effect of change of concentration :

$$K_C=\frac{[NH_3]^2}{[N_2][H_2]^3}$$

Increasing the concentration of dry and pure N_2 or H_2 or both the (dr) increases as K_C is constant at the studied temperature the (nr) must have to increase i.e., yield NH_3 increases and the equilibrium shifts towards right in order to realease (undo) the effect of concentration change.

(ii) Effect of change in pressure :

In this equilibrium, $\Delta n = -ve$ i.e., 2–(1+3) = –2 i.e., the reaction proceed in forward direction with decrease in volume (no. of moles)....

Hence increase of pressure $\left(p \propto \frac{1}{v}\right)$ shifts the equilibrium in the direction of decrease of volume.

i.e., the equilibrium shifts towards right yielding more of NH_3. Pressure applied 200–900 atom in Haber's process.

(iii) Effect of change of temperature :

Formation of NH_3 is exothermic and the reverse i..e, decomposition of NH_3 is endothermic. As per the Lechatelier's principle if such a system in equilibrium is heated externally the reaction proceeds in endothermic way i..e, the reaction proceeds in backward direction inoder to undo the effect of external heat. However, in Haber's process the reaction system initially heated upto an optimum temperature of 500^0C just to start the reaction (i.e. to surpass the potential barrier / activation energy.)

In addition to changes in concentration, pressure and temperature, the process is followed with catalyst (finely divided Fe in presence of Mo as promotor) to reduce the activation energy and favouring quick action. Catalyst does not affect the point of equilibrium, (K value).

16. **(i) Buffer solution :**

Statem ent : [2]
Types of buffer :
Example (Buffer action)] [2]

Refer Text Book page 398.

(ii) Common ion effect :

Statement. [1]
Applications (any two) [2]

Refer Text book Page - 406

17. Redox reaction : The reactions where oxidation and reduction take place simultaneously i.e. the reaction between oxidising agent and reducing agent.

Types of redox reaction :

a. Synthesis : $\overset{0}{C} + \overset{0}{O}_2 \longrightarrow \overset{+4}{C}O^{-4/2}$

Here C is oxidised from 0 to +4 while oxygen is rediced from 0 to –2. i.e. the reaction where element, change their oxidation number is a redox reaction.

b. Decomposition reaction :

$$\overset{+1}{K}\overset{+5}{Cl}O_3^{-6} \longrightarrow \overset{+}{K}\overset{-}{Cl} + 3/2\, O_2$$

In this reaction Cl is reduced from +5 to –1 oxygen is increased from –2 to 0.

c. Displacement reaction :

$$X + YZ \longrightarrow XZ + Y$$

(i) $$Zn^0 + \overset{+2}{Cu}SO_4 \longrightarrow \overset{+2}{Zn}SO_4 + \overset{0}{Cu}$$

Cu^{2+} is reduced while Zn is oxidised to $\overset{2+}{Zn}$.

(ii) $$\overset{0}{Ca} + 2\overset{0}{H_2}O \longrightarrow \overset{+2}{Ca}(OH)_2 + \overset{0}{H_2}$$

Ca is oxidised while H^+ is reduced.

d. Disproportionation reaction : Here same element undegoes oxidation and reduction simultaneously

$$\overset{+1}{H_2}\overset{-1}{O_2} \longrightarrow H_2\overset{-2}{O} + \frac{1}{2}\overset{0}{O_2}$$

Oxygen is recuced as well as oxidised.

$$Cl_2(g) + 2OH^-(Aq) \longrightarrow \overset{+}{Cl}O^-_{(aq)} + Cl^-(aq) + H_2O(\ell)$$

Here Cl is oxidised to hypochlorite & reduced to chloride ion.

Application :

(a) Volumetric analysis e.g. permanganometry, dichrometry, iodometry (where liberated iodine is titrated against sodium thiosulfate)

$$2Cu^{2+}(aq) + 4I^-(aq) \longrightarrow Cu_2I_2(s) + I_2(aq)$$

$$I_2(aq) + 2S_2O_3^{2-}(aq) \longrightarrow 2I^-(aq) + S_4O_6^{2-}(aq.)$$

(b) Electro-chemical cell

$$\underset{\text{Anode}}{Zn \overset{-}{|} Zn^{2+}(Aq)} \quad || \quad \underset{\text{Cathode}}{Cu^{2+}(Aq) \overset{\oplus}{|} Cu}$$

$$Zn + Cu^{++} \longrightarrow Zn^{2+} + Cu \quad \text{(Daniell cell)}$$

Electron flows from anode (oxidation) to cathode (reduction), causing flow of electric current from cathode (+) to anode (–).

(c) Extraction of metal :

$$Fe_2O_{3(s)} + 3\,C_{(s)} \longrightarrow 2Fe_{(s)} + 3CO(g)$$

(d) Redox reaction and supply of energy.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell) + \text{energy}$$

hydrocarbon

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6\,CO_{2(g)} + 6H_2O(\ell) + \text{energy}$$

carbohydrate

Problem :

gm/L of oxalic acid. xH_2O $= 1.575\times4$

$= 6.300$

Normality of this solution $= \frac{10\times N/12}{8.34} = 0.1\ (\text{appx})$

Eq. mass $= \frac{\text{gm}/\text{L}}{\text{normality}} = \frac{6.3}{0.1} = 63$

Oxalic acid is a dibasic acid . So its molecular mass = 126.

Now $(COOH)_2\ x\ H_2O = 90 + 18\ x = 126$

$x = 2$

∴ Water of crystallisation is 2.

18. Preparation of H_2O_2

(i) By electrolysis of 50% H_2SO_4 taking Pt-anode and graphite cathode, peroxy disulfuric acid is formed at anode which on treating with water H_2O_2 is separated on distillation at reduced pressure. H_2O_2, due to its low b.p. separates out leaving H_2SO_4 which can be recycled.

$$2H_2SO_4 \longrightarrow 2H^+ + 2HSO_4^-$$

At cathode. $2H^+ + 2\bar{e} \longrightarrow H_2\uparrow$

At anode (Pt) $2HSO_4^- \longrightarrow H_2S_2O_8 + 2\bar{e}$

$$H_2S_2O_8 + 2H_2O \longrightarrow H_2O_2 + 2H_2SO_4$$

(ii) Merck's process : Carbon dioxide is passed through a suspension of barium peroxide in icecold water and fittered to get H_2O_2 as filtrate.

$$Ba\ O_2 + H_2O + CO_2 \longrightarrow BaCO_3\downarrow + H_2O_2$$

Purification :

(i) Water is distilled out at 303 – 313 K under reduced pressure of 10–15 mm of Hg.

(ii) Last traces of water is removed on treating with freezing mixture containing dry ice and ether to give crystals of hydrogen peroxide.

Reactions : (a) H_2O_2 liberates violet vapours of iodine when treated with KI (aq).

$$\begin{array}{r}H_2O_2 + 2\bar{e} \longrightarrow 2OH^-\\ 2I^- \longrightarrow I_2 + 2e^-\\ \hline H_2O_2 + 2I^- \longrightarrow 2OH^- + I_2\end{array}$$

(b) H_2O_2 reduces ferric sulfate (yellow) to ferrous sulfate (green) in alkaline medium.

$$\begin{array}{r}H_2O_2 + 2\left(OH^-\right) \longrightarrow O_2 + 2H_2O + 2e^-\\ 2Fe^{3+} + 2e^- \longrightarrow 2Fe^{2+}\\ \hline 2Fe^{3+} + H_2O_2 + 2OH^- \longrightarrow O_2 + 2Fe^{2+} + 2H_2O\end{array}$$

19. **Trends in physical properties of group 1** and write down their chemical reactivity.

Alkali metals : These are having general valence shell configuration ns^1 (n = 2 to 7) with noble gas kernel (part of atom without valence $\bar{e}$). In aqueous medium these metals form strong alkali M (OH) M = Li, Na, K, Rb, Cs, *Fr.

* **Atomic and ionic radii :** In a period the the first element is an alkali metal having higher atomic radii but in a group the radii gradually increases on moving down the group with incease of shell number.

Ionisation enthalpy $(\Delta_i H)$. $\Delta_i H$ is positive i.e. endoergic. The value decreases on moving down the group with incease of atomic size. $\Delta_i H_2$ is very high as the single charged cation (M^+) has noble gas configuration and proton - electron ratio is high.

* **Melting and boiling point :** These metals are soft, loosely packed due to bigger size and having low m.p. All these metals have a single $\bar{e}$ in the valence shell hence binding in the crystallattice is low. The m.p. and b.p. decrease on moving down with increase of atomic size and loose packing in the lattice.

* **Characteristic flame colouration.**

All alkali metals / their salt impart different colour to the flame due to the low $\Delta_i H$ values. On heating the excitation takes place in the visible spectra & the complementary colour of the energy absorbed for $\bar{e}$ - transition is given by the metal salt.

e.g. for sodium due to its smaller size high frequency of the visible spectrum (violet) is absorbed for $\bar{e}$ excitation, and the complementary emitting colour falls in the yellow zone of the visible spectrum, hence looks yellow.

Density : Alkali metals have lower density, even lighter than water. However the densities increases on moving down the group (Li to Cs) as the mass increases more in comparison to the increase of volume. $\left(d = {}^{m}\!/_{v}\right)$

Chemical reactivity All these metals are highly reactive due to their smaller $\Delta_i H$ value. These are highly electropositive and more reactive and their compounds are mostly electrovalent.

Lithium mostly form covalent compounds due to smaller cation $\left(\overset{+}{Li}\right)$ with higher polarising power.

These elements are better reducing agents due to their lower $E^0_{M^+(Aq)|M}$ i.e reduction potential.

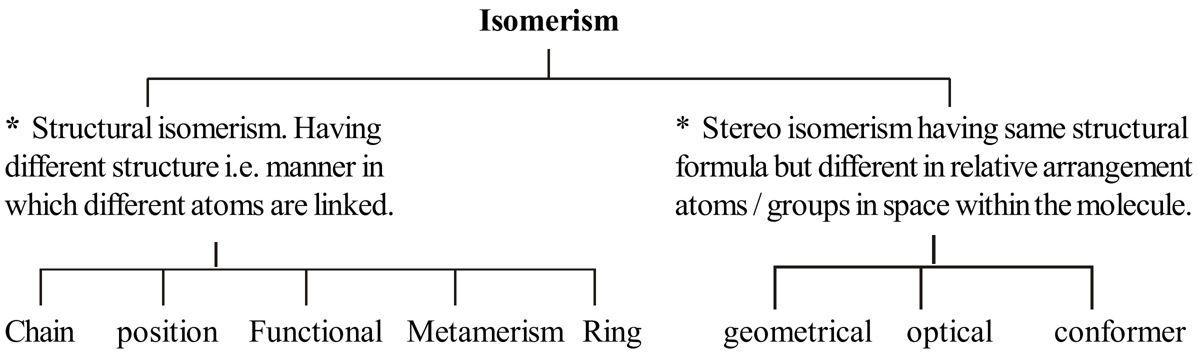
Among these $E^0_{Li^+(Aq)|Li^=}$ -3.05 V hence lithium acts as a powerful reducing agent.

20,21,22. Trends in physical properties of

(group-2) Refer Text book page. 513
(group-13) Refer Text book page. 529
(group - 14) Refer Text book page. 557

23. **Isomerism :** The phenomenon of existence of two or more compounds possessing the same molecular formula but different properties is known as isomerism. Such compounds are called isomers.

Flow chart :



Isomerism

* Structural isomerism. Having different structure i.e. manner in which different atoms are linked.

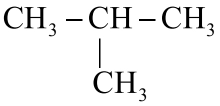
Chain position Functional Metamerism Ring

* Stereo isomerism having same structural formula but different in relative arrangement atoms / groups in space within the molecule.

geometrical optical conformer

* Chain isomerism Nucleus / chain changes.

$CH_3-CH_2-CH_2-CH_3$,
butane



$CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_3$

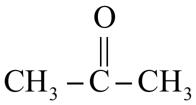
methyl propane (isobutane)

* Position isomerism. Position of double bond / triple bond / functional group changes without change of nucleus.

$CH_2=CH-CH_2-CH_3$
But – 1 – ene

$CH_3-CH=CH-CH_3$
But – 2 – ene

* Functional isomerism : Same molecular formula with different functional group.



$CH_3-\overset{\overset{\displaystyle O}{||}}{C}-CH_3$

Propanone

CH_3-CH_2-CHO

Propanal

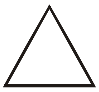
* Metamerism. Distribution of C-atoms about the functional group changes.

$CH_3-CH_2-O-CH_2-CH_3$
Diethyleether
(Ethoxy ethane)

$CH_3-O-CH_2-CH_2-CH_3$
1 – Methoxy propane
Methyl propylether

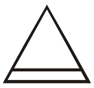
* An open chain compound with $C_n\ H_{2n}$ or $Cn\ H_{2n-2}$ formula can have ring isomers with $n \geq 3$.

C_3H_6 : $CH_3-CH=CH_2$,
Propene



cyclopropane

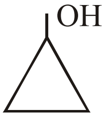
C_3H_4 : $HC\equiv C-CH_3$
Propyne



Cyclopropene

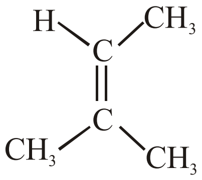
$Cn\ H_{2n}O$: $CH_2=CH-CH_2-OH$ (Prop-2-en-1-ol)

$C_3\ H_6O$

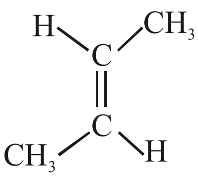


Cyclopropanol

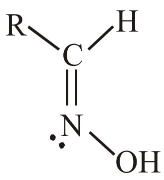
* Geometrical isomerism : Such isomers are obtained due to restriction in rotation of $\rangle C=C\langle$, $\rangle C=N-$.



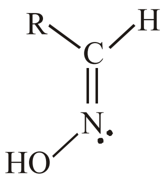
Cis. but-2 ene



trans-but-2-ene



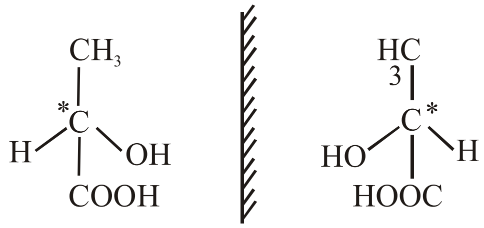
Syn-oxime



Antioxime

* Optical Isomerism . When four different atoms/groups are attached to a carbon tetrahedrally, the C- is called a chiral / asymmetric carbon. Such compounds rotate the plane polarised light differently i.e., they have different optical activity. If rotates to right it's called d-isomer while to the left it's called ℓ-isomers. Compounds having the same molecular formula, but different spatial arrangement of atoms or groups within the molecule and possessing the property of rotating the plane polarised light are known as optical isomers or enantiomers.

e.g.



d - lactic acid ℓ-lactic acid

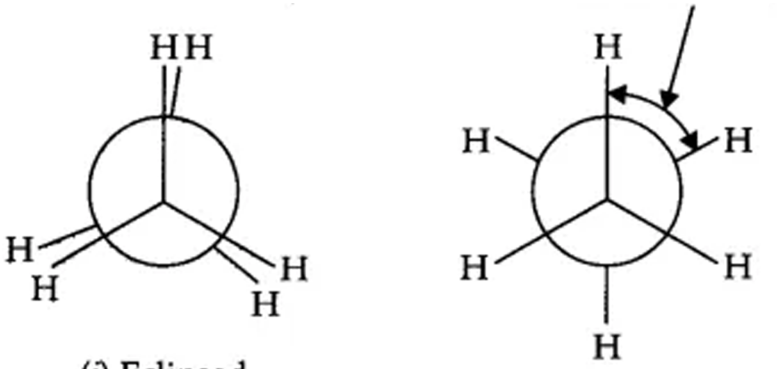
i..e, the compounds must have a chiral centre.

* The mirror image should not be super imposable.

* **Conformers / rotamers .** These are the isomers only differ in their energy occuring due to free rotation of C–C σ-bond (sp^3 hybridisation). Out of many conformations two extreme cases are choosen i.e., eclipsed form (of higher energy due to electronic repulsion) and staggered form (where atoms / groups are away from each other causing less repulsion and more stable).

Newman's representation of

(a) Eclipsed ethane (b) Staggered ethane



(*i*) Eclipsed (*ii*) Staggered

Newman's projections of ethane

(ii) Wurtz reaction :

Statement : When alkyl halide is refluxed in presence of sodium is dry ether medium higher alkane is formed.

Mechanism (a) $CH_3I + 2Na \longrightarrow \bar{C}H_3\overset{+}{N}a + NaI$

$$CH_3Na + CH_3I \longrightarrow CH_3 - CH_3 + NaI$$

(b) Free radical mechanism.

$$CH_3I + Na \longrightarrow \bullet CH_3 + NaI$$

$$\bullet CH_3 + \bullet CH_3 \longrightarrow CH_3 - CH_3$$

Merits : (i) Preparations of higher alkane

(ii) To ascend C-atom in chain.

(iii) Wurtz-Fittig reaction.

$$C_6H_5\,Cl + CH_3Cl \xrightarrow[\text{ether}]{2Na} \underset{\text{Toluene}}{C_6\,H_5\,CH_3} + 2NaCl$$

(iv) Fittig reaction :

$$2C_6H_5Cl \xrightarrow{Na/\text{dry ether}} C_6H_5 - C_6H_5 \text{ (biphenyl)}$$

Limitations : (i) Preparation of odd numbered alkane.

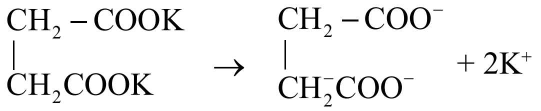
$$CH_3Br + C_2H_5Br \xrightarrow[\text{reflux}]{Na-\text{dry ether}} \begin{array}{l} CH_3 - CH_3 \text{ ethane} \\ CH_3 - C_2H_5 - \text{propane} \\ C_2H_5 - C_2H_5 - \text{butane} \end{array}$$

A mixture of alkanes are obtained.

24. Kolbe's reaction : (Electrolytic method of preparation of hydrocarbon :)

(i) Preparation of ethane – Electrolysis of CH_3COOK, at anode ethane is obtained.

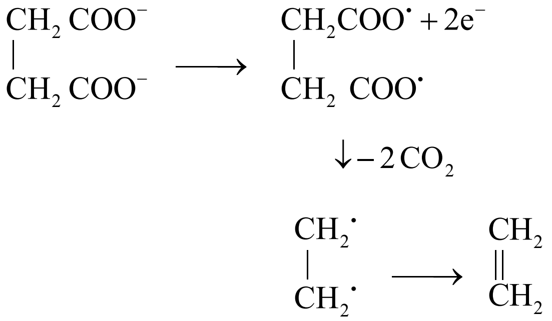
(ii) Preparation ethene : Electrolysis of saturated solution of sodium or potassium succinate gives ethene at anode.



$$\begin{array}{l} CH_2 - COOK \\ | \\ CH_2COOK \end{array} \rightarrow \begin{array}{l} CH_2 - COO^- \\ | \\ CH_2COO^- \end{array} + 2K^+$$

At cathode : $H^+ + e^- \longrightarrow \frac{1}{2}H_2$

At anode

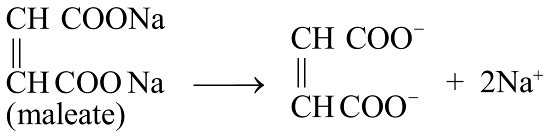


$$\begin{array}{l} CH_2\,COO^- \\ | \\ CH_2\,COO^- \end{array} \longrightarrow \begin{array}{l} CH_2COO^{\bullet} + 2e^- \\ | \\ CH_2\,COO^{\bullet} \end{array}$$

$$\downarrow -2\,CO_2$$

$$\begin{array}{l} CH_2^{\bullet} \\ | \\ CH_2^{\bullet} \end{array} \longrightarrow \begin{array}{l} CH_2 \\ \| \\ CH_2 \end{array}$$

(iii) Preparation of ethyne. Electrolysis of saturated solution of sodium maleate or sodium fumerate gives ethyne and CO_2 at anode and $H_2(g)$ at cathode.



$$\underset{\text{(maleate)}}{\begin{array}{l} CH\;COONa \\ \| \\ CH\,COO\,Na \end{array}} \longrightarrow \begin{array}{l} CH\;COO^- \\ \| \\ CH\,COO^- \end{array} + 2Na^+$$

Due to less reduction potential of Na^+, it does not get discharged at cathode.

$$E_{Na^+(Aq)|Na} = -2.71\,V$$

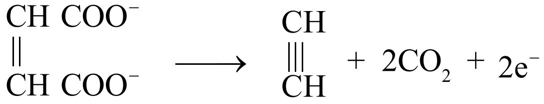
At cathode -

$$H_2O \longrightarrow H^+ + OH^-$$

$$H^+ + e \longrightarrow \frac{1}{2}H_2$$

$$H_2O + e \longrightarrow \frac{1}{2}H_2 + OH^-$$

At anode :

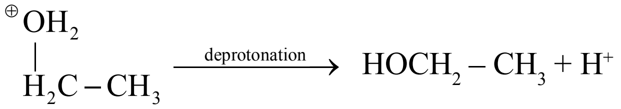


$$\begin{array}{l} CH\ COO^- \\ \| \\ CH\ COO^- \end{array} \longrightarrow \begin{array}{c} CH \\ ||| \\ CH \end{array} + 2CO_2 + 2e^-$$

(b) $CH_3-CH_2-OH \xrightarrow[170^0]{Conc.H_2SO_4} CH_2=CH_2+H_2O$

$$CH_2=CH_2+H_2O \xrightarrow[Ad_E]{Conc.\ H_2SO_4\left(H^+\right)} CH_3-CH_2OH$$

$$CH_2=CH_2+H^+ \longrightarrow \overset{+}{C}H_2-CH_3 \xrightarrow{OH_2}$$



$$\begin{array}{l} {}^{\oplus}OH_2 \\ | \\ H_2C-CH_3 \end{array} \xrightarrow{deprotonation} HOCH_2-CH_3+H^+$$

25. (i) Ozonolysis Refer Text Book Page - 766

(ii) Polymerisation-Refer Text Book Page - 767

26. (a) The possible reactions in the stratosphere.

Refer Text Book Page - 863

(b) Global warming.

Refer Text Book Page - 872

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MODEL SET

Answer all the bit questions of a particular question serially at one place to ensure effective valuation.

Group - A

1. Choose and write the correct answer from the choices given : [1 x 7]

(i) Which of the following is the molecular formula of hydrolith ?

(a) NaH (b) $LiAlH_4$ (c) CH_4 (d) CaH_2

(ii) Which of the following fail to exhibit structural isomerism ?

(a) C_2H_6O (b) C_2H_4O (c) C_3H_6 (d) C_2H_6

(iii) The solution with higher pH is __________.

(a) NaCl (Aq) (b) CH_3COONH_4 (aq) (c) $CuSO_4$ (Aq) (d) CH_3COONa (Aq)

(iv) Which of the following process is not exothermic ?

(a) Combustion of methane (b) Evaporation aclohol

(c) Neutralisation reaction (d) Synthesis of ammonia.

(v) For the reaction $C_{(s)} + O_2(g) \rightleftharpoons CO_2(g)$

(a) $K_p = K_c(RT)^2$ (b) $K_p = K_c(RT)^0$

(c) $K_p = K_c(RT)^{-1}$ (d) $K_p = K_c(RT)^{-2}$

(vi) Which of the following salt gives hydrogen gas at anode ?

(a) NaCl (aq) (b) KCl (melt) (c) NaH (melt) (d) CH_3COONa (satd.)

(vii) In which of the following pair of elements Mandeleef periodic classification of elements is violated?

(a) Te, I (b) K, Ca (c) Ni, Cu (d) Cl, Ar

2. Answer the following : [1 x 7]

(a) The highest magic number value in periodicity is ___________.

(b) The basicity of H_3BO_3 is __________.

(c) Lithium is diagonally related to ________.

(d) The IUPAC name of $CH_2 = CH - CO - C(CH_3)_3$ is ___________.

(e) The shape of ClF_3 is ____________ .

(f) The first transition element, Sc present in ______ period.

(g) The outer shell configuration of p-block element is ________.

(h) Arrange the elements of carbon family in order of decreasing $\Delta_i H$ value.

(i) Write the seven elements present in group - 18.

Group - B

(Answer any seven of the following) [2 x 7]

3. (a) 0.01 mole of Ca CO_3(s) is decomposed. Calculate the volume of CO_2(g) formed at 298 K and 1-atm. pressure.

(b) State law of multiple proportion.

(c) Equivalent mass of metal, M is ‘E’. Calculate the molecular mass of its compound MCl_x.

(d) Potassium permanganate acts as an oxidant in acidic, alkaline and neutral media. Show that its equivalent mass in alkaline medium is maximum.

(e) The equivalent mass of a metal is 32.2. Its specific heat is 0.095. Determine its atomic mass.

(f) Calculate the mass of the following.
   (i) 0.05 mole of sodium sulfate
   (ii) 280 ml of nitric oxide (g) at NTP.

(g) Write down the defect of Rutherford model of atom.

(h) Explain the visible atomic spectralines by Bohr's atomic model.

(i) Mention the different quantum numbers of the valence electrons of nitrogen atom.

(j) Draw the dot structure of ethene and ethyne.

4. Answer any seven of the following : [3 x 7]

(a) What is electron gain enthalpy ? Show that Chlorine has its higher value than fluorine.

(b) Discuss the exceptional electronic configuration of copper.

(c) How can you calculate the lattice energy of $MgCl_2$ using Born-Haber cycle.

(d) What is dipolemoment ? Can you calculate the percentage of ionic character of covalent bond ?

(e) Explain the formation of hydroen molecule in the light of valence bond theory.

(f) Distinguish between bonding and antibonding molecular orbitals. Out line π^* orbital.

(g) Show, BCl_3 and PCl_3 are not of identical shape.

(h) How does real gas deviate from ideal gas ?

(i) Define bond energy : Explain the factors affecting it.

(j) If the density of a gas at 27^0C and 1 atm. pressure is 1.5 gm/L, find its density at 127^0C and 4 atm. pressure

Group - C

(Answer the following) [7 x 3]

5. State and explain Hess's law.

Calculate the enthalpy of hydrogeneration of ethene.

Given that C=C, C–C, H–H and C–H bond dissociation energy 615, 347.3, 435.1 and 416.2 kJ respectively. [1 + 3 + 3]

6. State law of mass action. Deduce law of chemical equilibrium. Write down the significance of reaction quotient. [2 + 3 + 2]

7. Write notes on -
   (i) Buffer solution
   (ii) Hyper conjugation.

8. (a) Discuss the trends is the physical properties of Group-13 elements.
   (b) What are allotropes of carbon ? [3 + 4]

9. (a) Prepare :
   (i) acetaldehyde from acetylene.
   (ii) ethene from ethanol
   (iii) ethane by Kolbe's reaction [3]
   (b) Give reasons :
   (i) Propene with HBr in presence of organic peroxide gives 1- bromopropane.
   (ii) Acetylene is acidic in nature. [4]

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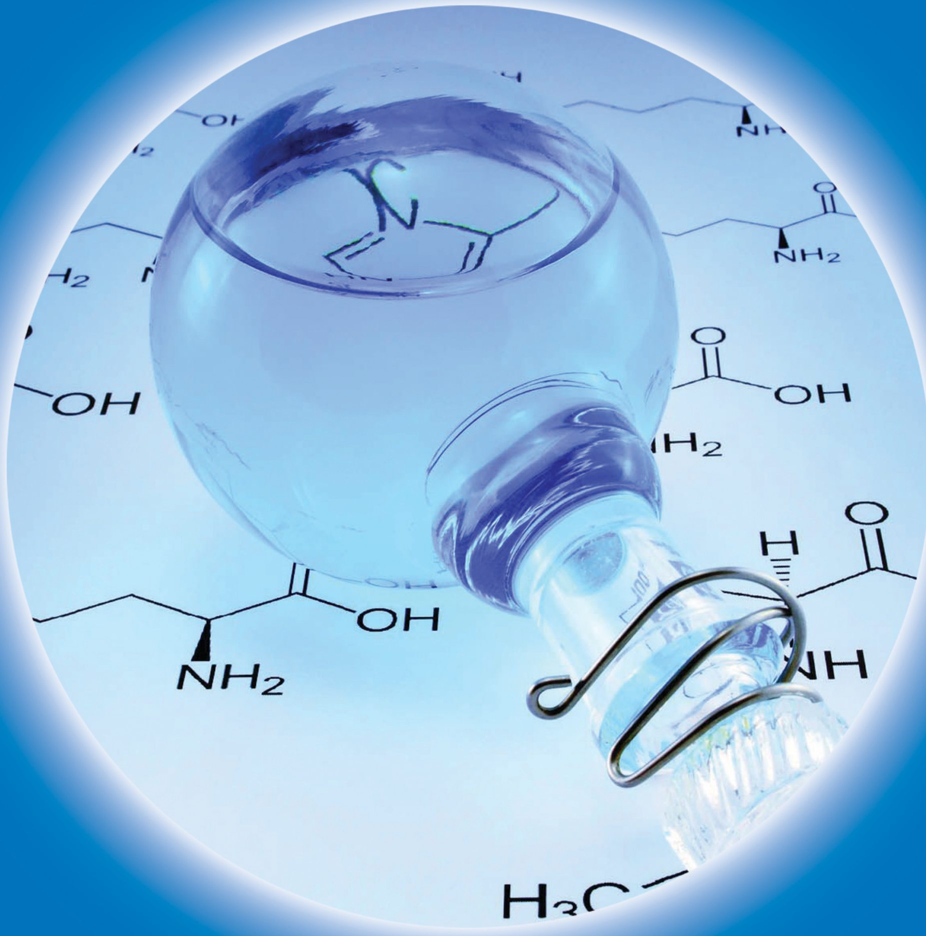
OH
NH2
OH
NH2
H
O
OH
NH2
NH
H3C