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# **SYLLABUS**

- Some Basic Concepts of Chemistry: Matter and its nature, Dalton's atomic theory, Concept of atom, molecule, element and compound, Physical quantities and their measurement in chemistry, SI units, Precision and accuracy, Significant figures, Dimensional Analysis, Laws of chemical combinations, Isotopic (atomic) and molecular masses, Mole concept, Equivalent mass, Percentage composition, Empirical and molecular formulae, Stoichiometry and calculations
- Atomic Structure : Subatomic particles, Atomic models, Black body radiation, Photoelectric effect, Bohr's atomic model, Spectrum of hydrogen atom, Sommerfeld extension of Bohr's model, Dual character (de Broglie's relationship), Heisenberg's uncertainty principle, Shapes of orbitals, Quantum numbers, Concept of atomic orbitals as an electron wave function. • Rules for filling electrons in orbitals : Aufbau principle, Pauli's exclusion principle, Hund's rule, • Electronic configuration of elements, Extra stability of half-filled and completely filled orbitals.
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- Chemical Bonding and Molecular Structure : Chemical bond, Octet rule, Ionic Bond, Born Haber cycle, Lattice Energy, Covalent bond, Fajan's rule, Polar character in covalent bond, Bond parameters, Co-ordinate bond (Dative bond), Valance bond theory, Concept of hybridization, V.S.E.P.R theory, Diploe moment, Percentage of ionic character, Resonance, Molecular orbital theory, Hydrogen bond, Metallic bond, van der Waal's forces of attraction
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- Chemical Thermodynamics : Fundamentals of thermody -namics : System and surroundings, Extensive and intensive properties, and thermodynamic State functions. • First law of thermodynamics : Concept of work, heat, internal energy and enthalpy, Heat capacity, Molar heat capacity, Hess's law of constant heat summation, Enthalpies of bond dissociation, combustion, formation, neutralization, solution, hydration, fusion, vaporisation and sublimation, Carnot engine. • Second law of thermodynamics : Entropy, Spontaneity of processes, ΔS for universe, ΔG of the system as criteria for spontaneity, Gibb's Helmholtz equation, ΔG (standard Gibb's energy change) and equilibrium constant. • Third law of thermodynamics : Brief introduction.
- Equilibrium : Meaning of equilibrium, Types of equilibrium : Physical equilibrium and Chemical equilibrium • Law of mass action, Relation between Kc and Kp • Factors affecting equilibrium constant, Application of law of mass action, • Le Chatelier's principle and its applications • Ionic equilibrium : Electrolytes, Ostwald's dilution law, Acids and bases, Ionic product of water, The pH scale, Dissociation constant of weak acid and weak base, Buffer solution, Hydrolysis of Salts, Common ion effect, Solubility product, Acid base indicator, Theories of indicator.
- Redox Reactions : Redox reactions, Oxidation and reduction, Oxidising and Ruducing agents, Oxidation number, Rules for assigning oxidation number, Application of oxidation number, Types of Redox reactions, Balancing redox reactions.
- Hydrogen : Position of hydrogen in periodic table, Isotopes, Preparation, properties and uses of hydrogen, Hydrogen as a fuel, Classification of hydrides – ionic, covalent and interstitial, Physical and chemical properties of water, Soft water and Hard water, Heavy water, Hydrogen peroxide – preparation, properties and uses, Structure of hydrogen peroxide, Strength of hydrogen peroxide.

- s-Block Element (Alkali & 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 properties (such as ionization enthalpy, atomic and ionic radii) trends in chemical reactivity with oxygen, hydrogen and halogens, uses. • Preparation and properties of some important compounds: Sodium hydroxide, Sodium chloride, Sodium carbonate, Sodiumhydrogen carbonate, Sodium thiosulphate and Microcosmic salt. Biological importance of sodium and potassium. • Industrial use of Lime Slaked Lime, Limestone, Plaster of Paris and Cement. Biological importance of Mg and Ca.
- p-Block Elements (Boron Group Elements + Carbon Group Elements) : Group 13 elements : General introduction Occurrence Electronic configuration Variation of properties Oxidation states Trends in chemical reactivity Anomalous properties of first element of group –Boron Some important compounds, Boranes, Borax, Boric acids, Boron trifluoride, Aluminium chloride and Alums. Group 14 elements : General introduction Electronic configuration Occurrence Variation of properties Oxidation states Trends in chemical reactivity Anomalous properties of first elements of the group Carbon Allotropic forms Physical and chemical properties Uses of some important compounds, oxides Classification of carbides Silicon tetrachloride, Silicones and Silicates Zeolites.
- Organic Chemistry Some Basic principles and Techniques : • Tetravalency of carbon - Shapes of simple molecules, hybridisation (s and p-orbitals). Classification of organic compounds based on functional groups and those containing halogens, oxygen, nitrogen and sulplur, Nomenclature (trivial and IUPAC) of Organic compounds . Isomerism-Structural and stereoisomerism, 

  Mechanism of Organic reaction, Covalent bond fission - Homolytic and heterolytic fission, electrophiles and nucleophiles, Electronic displacement in covalent bond - Inductive effect, Resonance, resonance or mesomeric effect, electromeric effect and hyperconjugation. • Reaction intermediates : Carbocations, Free radicals and Carbanions. • Common types of organic reactions - Substitution, addition, elimination and rearrangement. • Purification and characterisation of organic compounds Qualitative analysis, Quantitative analysis Hydrocarbon: • Aliphatic Hydrocarbons, Alkanes: Nomenc-lature,
- Hydrocarbon: Aliphatic Hydrocarbons, Alkanes: Nomenc-lature, isomerism and conformations, Preparation, Physical properties, Chemical properties • Alkenes : Nomenclature and isomerism, Stability of Alkene, Preparation, Physical properties, Chemical properties • Alkynes : Nomenclature, Preparation, Physical properties, Chemical properties • Aromatic Hydrocarbons : Nomenclature of arenes, Synthesis of benzene, Structure of benzene, Properties of benzene :Mechanism of electrophilic substitution, Halogenation, Nitration, Sulphonation, Friedel Craft alkylation, acylation, Directive influence of functional group in mono substituted benzene. • Carcinogenicity and Toxicity
- Environmental Chemistry : Environmental pollution : Atmospheric, water and soil. • Atmospheric pollution :Tropospheric and stratospheric • Tropospheric pollutants : Gaseous pollutants : Oxides of carbon, nitrogen and sulphur, hydrocarbons, their sources, harmful effects and prevention, greenhouse effect and global warming, acid rain. • Particulate pollutants : Smoke, dust, smog, fumes, mist, their source, harmful effects and prevention. • Stratospheric pollution : Formation and breakdown of ozone, depletion of ozone layer, its mechanism and effects. • Water pollution : Major pollutants, such as phosphagens, organic wastes and chemical pollutants, their harmful effects and prevention. • Soil pollution : Major pollutants such as pesticides (insecticides,herbicides and fungicides), their harmful effects and prevention. • Strategies to control environmental pollution, Green Chemistry.
- Solid State : Crystalline solids, Amorphous solids, Classification of solids, Space lattice and Unit cell, Types of crystal lattices, Calculation of number of particles in a unit cell, Calculation of packing fraction, Close packing in the crystals, Calculation of Density of unit cell,

# .....Contd Syllabus

Radius ratio, Classification of ionic structures, Bragg's equation, Imperfections in a crystal, Electrical properties of solids, Band theory of metals, conductors, semiconductors, insulators, Magnetic properties of solids, Dielectric properties in solids.

- Solutions: Types of solutions. Different methods for expressing concentration of solution : Molarity, Molality, Normality, Mole fraction, Percentage (by volume and mass both) Solubility of gases in liquids, Henry's law and its applications. Vapour pressure of solutions and Raoult's law, Ideal solutions and non-ideal solutions,
   Atomic mixtures Colligative properties of dilute solutions : Relative lowering of vapour pressure, •Elevation of boiling point, Depression of freezing point, Osmotic pressure Determination of molecular mass using colligative properties, Abnormal value of molar mass, van't Hoff factor and its significance.
- Electrochemistry : Electrolytic and metallic conduction, Conductance in electrolytic solutions, Molar conductivities and their variation with concentration, Kohlrausch's law and its application, Farady's law of electrolysis. • Electrochemical cells, Electrolytic and Galvanic cells, Electrode potential, Different types of electrodes, Half cell and cell reactions, Electromotive force (e.m.f.) of galvanic cell Electrochemical series, Nernst equation and its applications, Relationship between cell potential and Gibb's energy change, Concentration cell, Dry cell and lead accumulator, fuel cell, corrosion
- Chemical Kinetics : Rate of a chemical reaction, Factors affecting the rate of reactions, (Concentration, temperature, pressure, catalyst),Rate law, Rate constant and its units; Order of reaction and Molecularity of reaction (Elementary and complex reactions), Differential and integrated forms of zero and first order reactions, their characteristics and half lives, Pseudo first order reaction, Determination of order of a reaction, Effect of temperature on rate of reaction. Activation energy and its calculation, Arrhenius equation
- Surface Chemistry : Adsorption : Physisorption and chemisorption and their characteristics, Factors affecting adsorption of gases on solids,Freundlich adsorption isotherms, Langmuir adsorption isotherm • Catalysis : Homogeneous and heterogeneous, Theories of catalyses, Enzyme catalysis and its mechanism • Colloidal state : Distinction among true solutions, colloid and suspensions, Classification of colloids -Lyophilic, lyophobic, multimolecular, macromolecular and associated (micelles), Preparation and properties of colloids, Tyndall effect, Brownian movement, Electrophoresis, Dialysis, Coagulation and Flocculation, Emulsions and their types, Applications of Colloids
- General Principles and Processes of Isolation of Metals : Modes of occurrence of elements in nature. • Minerals ores.• Steps involved in the extraction of metals : • Concentration • Reduction (chemical and electrolytic method)• Refining with special reference to the extraction of Al, Zn and Fe. • Thermodynamic and electrochemical principles involved in the extraction of metals. • Ellingham diagram
- The *p*-Block Elements : Group 15 elements : General introduction, electronic configuration, oxidation states, trends in physical and chemical properties, preparation, properties and uses of dinitrogen, ammonia, nitric acid, oxides of nitrogen (structure only), phosphorus, allotropic forms, compounds of phosphorus, phosphane, oxides of phosphorous, oxoacids (elementary idea only). • Group 16 elements : General introduction, electronic configuration, oxidation states, occurrence, physical and chemical properties and uses of dioxygen, ozone, sulphur oxide, oxoacids of sulphur (structures only) sulphuric acid and Hypo • Group 17 elements : General introduction, electronic configuration, oxidation states, occurrence, trends in physical and chemical properties, compounds of halogens, interhalogen compounds and oxoacids of halogens (structures only) preparation, and properties and uses of chlorine, hydrochloric acid, . • Group 18 elements : General introduction, electronic configuration, occurrence, trends in physical and chemical properties, Preparation, properties, structure and uses of compounds of Xenon.
- The *d*-Block and *f*-Block elements General introduction of *d*- and *f*-block elements, electronic configuration, occurrence and characteristics, General trends in properties of the first row transition elements. • Physical properties of the transition

elements : Atomic radii, Ionisation enthalpy, Oxidation states, Colour, Magnetic properties, Catalytic behaviour, Complex formation, Interstitial compounds, Alloy formation • Preparation, properties and uses of  $K_2Cr_2O_7$  and KMnO<sub>4</sub>.

- Coordination Compounds : Introduction to coordination compounds, Ligands, Coordination number, Classification of Ligands denticity, chelation, Spectrochemical series, Effective atomic number (EAN), Werner's theory, IUPAC nomenclature of mononuclear coordination compounds, Valence bond theory (VBT), Isomerism, Crystal field theory (CFT), Colour and magnetic properties, Organometallic compounds and their Application, Importance of coordination compounds (in qualitative analysis, extraction of metals and in biological systems)
- Haloalkanes and Haloarenes : 

   Haloalkanes : Nomenclature, General methods of preparation, Physical and chemical properties, Nature of C–X bond, Mechanism of substitution reaction, Elimination reaction •Uses : Environmental effects of dichloromethane, trichloromethane, iodoform and carbon tetrachloride • Haloarenes : General methods of preparation, Physical and chemical properties (Nature of C–X bond), Benzyne mechanism
- Organic Compounds Containing Oxygen, Alcohols, Phenols and Ethers : • Alcohols : Classification, Nomenclature and Structure, Methods of preparation, Physical and chemical properties, Mechanism of dehydration reaction, Identification of 1°, 2°, 3° alcohols, Uses with special reference to methanol and ethanol • Phenols : Classification and Nomenclature, Methods of preparation, Physical and chemical properties, Electrophilic substitution reactions, Uses of phenols • Ethers : Nomenclature, Structure, Methods of preparation, Physical and chemical properties, Uses of ethers.
- Aldehydes, Ketones and Carboxylic acids : Introduction, Nomenclature, Structure of carbonyl group, General methods of preparation , Relative reactivities of carbonyl compounds, Chemical properties of aldehydes and ketones (Nucleophilic addition, addition-elimination, oxidation and reduction), Name reaction and mechanism, Distinction between aldehydes and ketones, Carboxylic acids, Classification, Nomenclature, Structure of carboxylic acid, General methods of preparation, Properties, Derivatives of carboxylic acids.
- Amines : Nomenclature, classification, structure, Methods of preparation, Physical and chemical properties, Basicity of Amines, Uses, Identification of primary, secondary and tertiary amines, Cyanides and isocyanides, Diazonium salts : Preparation, chemical reactions and importance in synthetic organic chemistry
- Biomolecules : General introduction and importance of Biomolecules • Carbohydrates : Classification : Aldoses and Ketoses, Monosaccharides (glucose and fructose) and D.L. configuration, Oligosaccharides (Sucrose, lactose and maltose), Polysaccharides(starch, cellulose and glycogen) • Protein : Elementary idea of a – amino acids, peptide bond, polypeptides, primary, secondary, tertiary and quaternary structure, (qualitative idea only), denaturation of proteins, enzymes. • Vitamins : Classification and functions. • Hormones : Elementary idea • Nucleic acids : General constitution of DNA and RNA. Biological function of nucleic acids.
- Polymers : Polymers, monomers, degree of polymerisation, Classification of polymers, General methods of polymerisation (addition, condensation) copolymerisation, Some important polymers : Natural and synethetic like polyethene, nylon, polyesters, bakelite, rubber, biodegradable and nonbiodegradable polymers, Average molecular weight of polymers.
- Chemistry in Everyday Life: Chemistry of medicines : Drugs, Chemotherapy, Classification of drugs, Analgesics, tranquilizers, Narcotics, Aypnotics, Sedatives, antidepressants, antimicrobials, Sulpha drugs, antibiotics, antiseptics, disinfectants, antacids, antihistamines antifertility drugs - their meaning and common examples. • Chemicals in food : Preservatives, artificial sweetening agents - common examples and elementary idea of antioxidants. • Cleansing agents : Soaps and detergents, their cleansing action.

# **Chemistry Mnemonics**

### Some Basic Concepts in Chemistry

### 1. Metric System

The Great Morning King Henry Doesn't Usually Drink chocolate Milk Mixed with Natural Powder

The  $\rightarrow$  Tera (×10<sup>12</sup>) Great  $\rightarrow$  Giga (×10<sup>9</sup>) Morning  $\rightarrow$  Mega (×10<sup>6</sup>) King  $\rightarrow$  Kilo (×10<sup>3</sup>) Henry  $\rightarrow$  Hecto (×10<sup>2</sup>) Doesn't  $\rightarrow$  Deca (×10) Usually  $\rightarrow$  Unit (×1) Drink  $\rightarrow$  Deci (×10<sup>-1</sup>) Chocolate  $\rightarrow$  Centi (×10<sup>-2</sup>) Milk  $\rightarrow$  Milli (×10<sup>-3</sup>) Mixed with  $\rightarrow$  Micro (×10<sup>-6</sup>) Natural  $\rightarrow$  Nano (×10<sup>-9</sup>) Powder  $\rightarrow$  Pico (×10<sup>-12</sup>)

2. Pure Substance : Fixed Composition (Cu, Ag)

### Lets Pray Dear

Law of conservation of Mass-Lets(Lavoiser) Law of definite proportion-Pray(Proust) Law of Multiple proportions-Dear(Dalton)

### **States of Matter**

### 1. Gas Law's

### PTV

(letters that touches are directly proportional & letter don't are indirectly proportional)

 $[P \propto T], [V \propto T], \left[P \propto \frac{1}{V}\right]$ 

### 2. Constant terms in Gas Laws

```
Paid TV Can Be Good
```

Constant terms  $\rightarrow$  Pressure (P) Temp (T) Volume (V) Gas Law  $\rightarrow$  Boyle's (Gay-Lussac's)

3. Ideal Gas Behavior

```
PLIGHT
```

High temp & Low pressure to achieve ideal Gas behavior  $PL \rightarrow Pressure Low$ IG  $\rightarrow$  Inert Gas  $HT \rightarrow High Temp$  4. Kinetic Theory of Gas

### Mother SPEAKS

 $M \rightarrow$  Motion (Gas Particle are in Random Motion) S  $\rightarrow$  Size (negligible size of particle to total volume)

 $P \rightarrow Pressure (Pressure exerted due to Collision with walls of container) E \rightarrow Elastic Collision$ 

 $A \rightarrow Attractive forces are not present$ 

 $K \rightarrow K.E \propto Temp$ 

 $\mathsf{S} \to \mathsf{Speed}$  (Distribution of speed of particles remain const.)

### 5. Crystal System

**C**u **T**e **M**OTHe **R** 3224

```
.
```

```
Unit Cell - Cubic, Tetragonal, Monoclinic,
Orthorhombic, Triclinic, Hexagonal, Rhombohedral
Edge Length - a=b=c, a=b\neq c, a\neq b\neq c, a\neq b\neq c,
a\neq b\neq c, a=b=c
Axial Length - \alpha=\beta=\gamma, \alpha=\beta=\gamma, \alpha=\beta\neq\gamma, \infty=\beta=\gamma,
\alpha\neq\beta\neq\gamma, \alpha=\beta\neq\gamma, \alpha=\beta=\gamma
No. of Bravias Lattice - 3, 2, 2, 4, 1, 1, 1
```

### 6. Edge Length

**TOM H**andpicked **T**ag (**HT**) of **C**lass

Representative (CR)

Triclinic, Orthorhombic, Monoclinic  $(a \neq b \neq c)$ Hexagonal, Tetragonal  $(a=b\neq c)$ Cubic, Rhombohedral (a=b=c)

### 7. Axial Angles

TORC Has More (HM) Twists (T)

Tetragonal, Orthorhombic, Rhombohedral, Cubic  $(\alpha = \beta = \gamma)$ Hexagonal, Monoclinic  $(\infty = \beta \neq \gamma)$ Triclinic  $(\infty \neq \beta \neq \gamma)$ 

### **Atomic Structure**

### 1. Atomic No & Mass No

### APEMAN

Atomic No. = No. of Protons = No. of Electrons Mass No. = Atomic No. + No. of neutrons

### 2. Isotopes, Isobars & Isotones

```
Bring Top Talented MAN (BTT MAN)
```

Atoms having same Isobars  $\rightarrow$  Mass Number Isotopes  $\rightarrow$  Atomic Number Isotones  $\rightarrow$  Neutrons

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### MNEMONICS



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### MNEMONICS

13.	Lanthanides	2.	H–Bonding
	Ladies Can't Put Needles Properly in Slot- machines. Every Girl Tries Daily However,		is <b>FON</b> ( Fun)!
	Every Time You'd be lose		Fluorine, Oxygen, Nitrogen
	La, Ce, Pr, Nd, Pm, Sm, Eu Gd, Tb, Dy, Ho, Er, Tm,	3.	Hardness of Water
Ger	Yb, Lu neral Principles and Process of Isolation of		CM is temporarily hard with Head Clerks (HC) but permanently
	Metal		Temporary hardness due to $Mg(HCO_3)_2$ , $Ca(HCO_3)_2$
1.	Process of Metallurgy		Permanent hardness due to $MgCl_{2'}$ Ca $Cl_{2'}$ MgSO <sub>4</sub> , CaSO <sub>4</sub> Hard with civil servants (CS) Cl <sup>-</sup> , SO <sup>2-</sup>
I			hydrogen Carbonate (HCO <sub>3</sub> ⁻) s-block elements
	(a) Concentration of Ore (b) Isolation	4.	Group I Elements
	(c) Purification	<b>P</b>	Little Nasty Kids Ruts Cats Far
2.	Concentration of Ore Honest Man Feeling Low (HMFL)	-	Lithium (Li), Sodium (Na), Potassium (K), Rubidium (Rb), Caesium (Cs), Francium (Fr)
Ŧ		5.	Group II Elements
	(a) Hydraulic Washing (b) Magnetic Separation	e e e e e e e e e e e e e e e e e e e	Beer Mug Can Snape Bar's Reputation
	(c) Froth Floatation Method (d) Leaching	<b>T</b>	Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba), Badium (Ba)
3.	Conversion to Oxide	6	Costner Kellner Cell
Ħ	CRAP	0.	Compart Modified Soil (CMS) Oxidised
•	Calcination $\rightarrow$ Absence of O <sub>2</sub> Roasting $\rightarrow$ Presence of O <sub>2</sub>		Cement Houned Soli (CHS) Oxidised
4.	Ores		Cathode Mercury (Hg) on which
	MISH		Socium ion(Na+) is oxidised $ACC \rightarrow reduced \rightarrow Anode of carbon on which Cl- is$
Ű	Prime Minister Going China		reduced
	Iron ores $\rightarrow$ Magnetite, Iron pyrites, Siderite,	7.	Properties of Birch Reagent
	Copper ores $\rightarrow$ Copper pyrites, Malachite, Copper	<b>P</b>	Roman People Can Commute (RPCC)
	Glance, Cuprite	■	
	Hydrogen, s & p-Block Elements		(Na/Li + liq.NH <sub>3</sub> ) –
	Hydrogen		Coloured)
1.	Isotopes of Hydrogen		p-block elements
	Pro-Diabetic Treatment PDT)	8.	Group 13 Elements
•	(1)		BAGIT
	Protium $\begin{pmatrix} -H \\ 1 \end{pmatrix}$		Boron (B), Aluminium (Al), Gallium (Ga), Indium (In), Thallium (Tl)
	Deuterium $\begin{pmatrix} 2\\ 1 \end{pmatrix}$	9.	Group 14 Elements
	$\begin{pmatrix} 1 \\ H \end{pmatrix}$		Can Simple Germans Surprise Public
			Carbon (C), Silicon (Si), Germanium (Ge), Tin (Sn), Lead (Pb),

(13)

14 OSWAAL TOPPER'S HANDBOOK, CHEMISTRY 10. Borax bead Test Multiple Program Combined (MPC) for Your Growth (FYG). New Boys Get (NBG) 1. 3d-Series Common Boys Room (CBR) for Combining Desktop Drawing (CDD) Oxidising Flame Reducing Flame lon  $Mn^{+2}$ Pink Colour less MPC  $\rightarrow$ Fe<sup>+2</sup> /Fe<sup>+3</sup> Yellow Green  $\rightarrow$ FYG Ni<sup>+2</sup> Brown Grey NBG  $\rightarrow$ Cu+2 Blue Red CBR  $\rightarrow$ Co+2 Deep Blue Deep Blue  $\rightarrow$ CDD p-block elements 11. Group 15 Elements New Police Assigns Subordinate Bikram on duty Nitrogen (N) Phosphorus (P) Arsenic (As) Antimony (Sb) Bismuth (Bi) 12. Group 16 Elements Old Saharanpur Seems Terribly Polluted Yb, Lu Oxygen (O) Sulphur (S) Selenium (Se) Tellurium (Te) Polonium (Pu) 13. Group 17 Elements First Class Biryani In Australia Fluorine (F) Chlorine (Cl) |-| = |Bronine (Br) lodine (I) Astatine (At) 14. Group 18 Elements He Never Arrived; Karan exited with Rohan  $OH^- = Of$  $H_2O = He$ Helium (He) Neon (Ne) Argon (Ar) Krypton (Kr) Xenon (Xe) CO = Compounds

# d & f block elements and Coordination Compounds

Scary Tiny Vicious Creatures are Mean; Females Come to Night Club Zen Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn

### 2. 4d-Series

Yes S(z)ir, Nob Most Technicians Rub Rod's Pale Silver Cadillac

Y, Zr, Nb, Mo,Tc, Ru, Rh, Pd, Ag, Cd

### 3. 5d-Series

- Late Harry Took Walk, Reached Office In
- Pajamas After an Hour

La..... Hf, Ta, W, Re, Os, Ir, Pt, Au, Hg

### 4. Lanthanides

Ladies Can't Put Needles Properly is Slotmachines. Every Girl Tries Daily, However, Every Time You'd be Lose

La, Ce, Pr, Nd, Pm, Sm, Eu, Gd, Tb, Dy, Ho, Er, Tm,

### 5. Spectrochemical series

I Bought Some Copies to Study Fundamental of Chemistry He Natured Excellence in Necessary Coordination Compounds

 $||^{-} < Br^{-} < SCN^{-} < Cl^{-} < S^{2-} < OH^{-} < C_{2}O_{4}^{-2-} < H_{2}O < C_{2}O_{4}^{-2-}$  $NCS^{-} < EDTA^{4-} < NH_3 < CN^{-} < CO$ Br<sup>-</sup> = Brought SCN<sup>-</sup> = Some  $Cl^{-} = Copies to$  $S^{2-}$  = Study F = Fundamental  $C_2 O_4^{2-}$  = Chemistry NCS<sup>-</sup> = Natured EDTA<sup>4-</sup> = Excellence in  $NH_3 = Necessary$ CN<sup>-</sup> = Coordination

(14)



		EARS
		<ul><li>(a) Elimination</li><li>(b) Addition</li><li>(c) Rearrangement</li><li>(d) Substitution</li></ul>
1	6.	Structural Isomerism
	T	Poor Farmer Managing Crops (PFMC)
		<ul><li>(a) Position</li><li>(b) Functional Group</li><li>(c) Metamerism</li><li>(d) Chair</li></ul>
	7.	Optical Isomerism
		GO
		(a) Geometrical (b) Optical
ł		Hydrocarbons and Halogen Derivatives
	1.	m-directing Group
	Ţ	Queen Elizabeth Second's Navy Commands, Controls
		Quarternary amino (-NR <sub>3</sub> *) Ester (-COOR) Sulphonic acid (-SO <sub>3</sub> H) Nitro (-NO <sub>2</sub> ) Carbonyl (-CHO) Carboxyl (-COOH) Cyano (-CN)
	2.	o, p–directing
		АНА АНА Р
		Alkyl (–R) Halogen (–X) Alkoxyl (–OR) Amino (–NH <sub>2</sub> , NHR, NR <sub>2</sub> ) Hydroxyl (–OH) Amide (–CONH <sub>2</sub> ) Phenyl ( $C_6H_5$ )
1	3.	SN1 reaction
	Ĩ	CURT-I
		Carbocation Intermediate Unimolecular Reaction Racemic mixture is obtained Two step process Ist order kinetics
(1	5)	

5. Types of Organic Reaction

#### 16 OSWAAL TOPPER'S HANDBOOK, CHEMISTRY 4. Chirality ⊢он — -N=N-CANS (orange dye) $Chiral \rightarrow Non-Super imposable mirror Images$ Achiral $\rightarrow$ Super imposable Mirror Images (yellow dye) **Organic Compound Containing Oxygen Polymers and Biomolecules Detection test** 1. 1. Disaccharides TASte FAAR IMLy Non-reducing SGF **TASte** $\rightarrow$ Tollen's test, Aldehyde group, Silver Mirror Sucrose $\rightarrow$ Glucose + Fructose **FAAR** $\rightarrow$ Fehling's test, Aliphatic Aldehyde, Red-Brown ppt Non-Reducing Sugar **IMLY** $\rightarrow$ lodoform test, Methyl group linked to **Essential Amino Acids** 2. **PVT TIM HALL** –Ľ– ′Yellow ppt (Phenylalanine, Valine, Threonine, Tryptophan, 2. Common Names of Carboxylic Acid Isoleucine, Methionine, Histidine, Arginine, Frog Are Polite, Being Very Courteous Leucine, Lysine) 3. Fat soluble Vitamins $\rightarrow$ Vitamin K, E, D, A Formic, Acetic, Propionic, Butyric, Valeric, Caproic **KEDA** 3. Dicarboxylic Acid Rest all Vitamins are water Soluble Oh My, Such Good Apple Pie, Sweet As 4. DNA & RNA Sugar G3Cinema AT 2PM Oxalic, Malonic, Succinic, Glutaric, Adipic, Pimelie, Subric, Azelaic, Sebacic DNA A=T, G=C (2 H-bonds b/w Adenine & Thymine 4. Clemmenson and wolf Reaction 3 H-bonds b/w Guanine & Cytosine) G=C A=T Can Zebra Woo Nightingale Also, GCAT **Analytical Chemistry and Chemistry in** Reaction to convert –C– to alkane Everyday life (to remember regents of reaction) Clemmen $\rightarrow$ son $\rightarrow$ Zn–Hg/HCl 1. Artificial Sweetening Agents Wolf $\rightarrow$ Reaction $\rightarrow$ NH<sub>2</sub>-NH<sub>2</sub>/OH<sup>-</sup> ASSA **Organic Compounds Containing Nitrogen** 1. Carbylamine test Aspartame, Saccharin, Sucrolose, Alitame Also, Assac Sue Ali PAFSI (Say Pepsi) 2. Antiseptic & Disinfectants Primary amine gives Foul smell of Isocyanide with Bitter Chlor CHCl<sub>3</sub>+KOH Amine Smell $RNH_{2}+CHCl_{2}+KOH \rightarrow RNC+KCl+H_{2}O$ Bithionol, Terpineol, Chloroxylenol 2. Coupling Reaction 3. Antacids DSPO DAY (Say, DeSPO DAY) His Interaction Presented by lime Ran (Say Simran) Diazonium Salt + Phenol $\rightarrow$ Orange dye Interaction of Histamine prevented by limetidine, Diazonium Salt + Aniline $\rightarrow$ Yellow dye

NH.

Ranitidine



### CHAPTER

# SOME BASIC CONCEPTS OF CHEMISTRY

Homogeneous

Organic Compounds

Heterogeneous

# **Chapter Objectives**

Matter and its nature, Dalton's atomic theory, Concept of atom, molecule, element and compound, Physical quantities and their measurement in chemistry, SI units, Precision and accuracy, Significant figures, Dimensional Analysis, Laws of chemical combinations, Isotopic (atomic) and molecular masses, Mole concept, Equivalent mass, Percentage composition, Empirical and molecular formulae, Stoichiometry and calculations

### **STUDY MATERIAL**

### I. Concept Clarified :

➢ Matter and its Nature :



Metalloids

Inorganic Compounds

### Separation of Mixtures :

Metals

Methods for the separation of constituents from the mixtures can be described as following :

- (i) Filtration : When one constituent in the mixture is soluble in a particular solvent and the other is not.
- (ii) Distillation : When constituents in the mixture have different boiling points.

Non-metals

- (iii) Crystallisation : Solids having different solubilities in a particular solvent present in the mixture.
- (iv) Sublimation : Volatile solids are separated by heating from non-volatile solids when present in the mixture.
- (v) Chromatography : One or more constituents are separated from the mixture by utilising the property of difference of adsorption on a particular adsorbent.
- (vi) Gravitational separation: When constituents present in the mixture have different densities.
- (vii) Magnetic separation : Constituents having magnetic properties are separated from the mixture by using a magnet.
- (viii) Atomolysis : This is a method by which gaseous component is separated from gaseous mixture based upon different rates of diffusion of gases through a porous material.
- Dalton's Atomic Theory (1803) : Dalton proposed the first atomic theory and postulates of this theory are :
   (i) Atom is considered as hard, dense, smallest indivisible particle of matter which is indestructible, that is, it can neither be created nor destroyed.

- (ii) Atoms of the same element have identical properties and mass. Different elements have different For more details, properties and mass.
- (iii) Atom is the smallest portion of matter which takes part in chemical combination.
- (iv) Different atoms combine in simple ratio of whole numbers through chemical reaction to form compounds.

### Atoms and Molecules :

**Atom :** The smallest particle of an element which may or may not be capable of independent existence is known as atom.

Atoms of iron, copper, gold, platinum etc. can exist freely but hydrogen, oxygen, nitrogen etc. exist as H<sub>2</sub>, O<sub>2</sub>, N<sub>2</sub>.

**Molecule :** The smallest particle of an element or a compound which can exist freely in nature is called **molecule**. Molecules

Homoatomic or Homonuclear Examples :  $H_{2}$ ,  $O_{2}$ ,  $N_{2}$  etc.

Heteroatomic or Heteronuclear Examples :  $H_2O$ ,  $NH_3$ ,  $CO_2$  etc.

### Physical Quantities and their Measurement in Chemistry

**Physical Quantities :** The quantities which come across during scientific observations. A physical quantity can be expressed in terms of two parts :

(a) numerical value, *i.e.*, number and

(b) its unit.

:. Magnitude of physical quantity = numerical value × unit. For conversion,  $n_1u_1 = n_2u_2$ 

- Units are of two types :
- (a) Fundamental / Base units

(b) Derived units :

(a) Fundamental / Base units : The 11<sup>th</sup> General Conference of Weights and Measures (GCWM) in 1960 introduced the International System of Units (in French *Le System International d'Units* abbreviated as SI) The SI system has seven basic units of physical quantities as follows :

Physical Quantity	Abbreviation	Name of Unit	Symbol
Length	1	metre	m
Mass	т	kilogram	kg
Time	t	second	S
Electric current	Ι	ampere	А
Thermodynamic temperature	Т	kelvin	K
Amount of the substance	п	mole	mol
Luminous intensity	I <sub>v</sub>	candela	cd

(b) Derived units : The units which are obtained by combination of base units.

Some commonly used physical quantities and their derived units are given in the following table :

Physical Quantity	Definition	Unit	Symbol
Volume	Length cube	Cubic metre	m <sup>3</sup>
Density	Mass per unit volume	kilogram per cubic metre	kg m <sup>-3</sup>
Pressure	Force per Unit area	pascal (newton per sq. metre)	$Pa = Nm^{-2}$ $= kgm^{-1}s^{-2}$
Work, energy	Force × distance	joule	$J = Nm$ $= kgm^2s^{-2}$
Electric charge	Current×time	coulomb	C = As
Potential difference	Energy difference of 1 joule per coulomb charge	Volt	$V = kgm^{2}s^{-3}A^{-1} = JA^{-1}s^{-1} = JC^{-1}$
Electric resistance	Potential difference per current	Ohm	$\Omega = VA^{-1}$
Electric conductance	Reciprocal of resistance	Ohm <sup>-1</sup>	$\Omega^{-1} = AV^{-1}$





### Precision and Accuracy

(i) **Precision :** It refers to the closeness of various measurements for the same quantity and hence close to their average value.

Precision = Individual value – Average value

The precise value depends upon the (i) measuring device used and (ii) the skill of operation.

- (ii) Accuracy : It refers to the average value of different measurements as close to the true value, the measurement is said to be accurate.
- (iii) Significant Figures (S.E): These are the total number of meaningful digits in a number including the last digit whose value is not certain.

The last digit of a number is not certain due to error and error is observed for the measuring instrument and the human skill.

- (iv) Rules for determination of S.F.
- (i) All non zero digits are significant.

Example :	Figure	S.F.
	567.9	4
	7.9	2

(ii) The right zeroes from the decimal or after non zero digits are significant.

Example :	Figure	S.F.
	34.00	4
	345.00	5

(iii) Zero to the left of the first non zero digit is not significant.

Example :	Figure	S.F.
	0.04	1
	0.0078	2

(iv) Zeros present between two non zero digits are significant.

Example :	Figure	S.F.
	84.001	5

(v) When zeros are present at the end of a number and these zeros are not to right of a decimal point, these zeros may or may not be significant. In this type of measurement, scientific notation such as  $N \times 10^n$  is used.

Where N = a number with single non zero digit to the left of the decimal point.

n = an integer called exponent.

Example :	Figure	S.F.
	$6.023 \times 10^{23}$	4
	$6.07 \times 10^{5}$	3
	$1.603 \times 10^{-24}$	4

Addition or Subtraction : In case of addition or subtraction, the final answer should not have more digits in the right of the decimal point than either of the original number.

44.442 4.9 15.0

 $64.342 \Rightarrow 64.3$  (The result should be reported only upto one decimal point.)

**Multiplication and Division :** In case of multiplication and division, the final answer should be reported with no more significant figures than those in the measurement with the few significant figures. For example,

$$5.4 \times 5.32 = 28.728$$

Since 5.4 has two significant figures therefore, the result = 28.7

### Rounding off Numbers :

- (a) To remove right most digit which is more than 5, the preceding number is added by 1. For example, 4.487 can be written as 4.49.
- (b) If right most digit is less than 5, the preceding number remains unchanged. For example, 9.453 can be written as 9.45.

### Dimensional Analysis :

This method is used when during calculation, it is required to convert units from one system to another. Examples :  $1 L = 1000 mL = 1000 cm^3 = 10^{-3} m^3 = 1 dm^3$ 

1 Newton =  $10^5$  dynes per square metre 1 bar =  $10^5$  pascals per square metre 1 nm =  $10^{-9}$  m =  $10^{-7}$  cm 1 Å =  $10^{-10}$  =  $10^{-8}$  cm 1 amu = 931.5 MeV



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### Laws of Chemical Combinations :

The combination of elements to form compounds is governed by the following laws :

(i) Law of conversation of mass (Lavoisier – 1789) : The mass is neither created nor destroyed in a chemical reaction or during a physical change, the total mass of the reactants is equal to that of the products.



(ii) Law of definite proportions (Proust – 1799) : A chemical compound, irrespective of its source, always combines same elements in a fixed proportion by weight.

[For example :  $H_2O$  always contains H and O in the ratio by weight = 1 : 8]

(iii) Law of multiple proportions (Dalton – 1803): When two elements combine to form two or more compounds, the weight of one element that combines with a fixed weight of other element(s) bear a simple ratio of whole numbers to one another.

**Examples** :



Here weights of oxygen which combine with a fixed weight of carbon (12 parts) bear a simple ratio = 16:32 = 1:2

(iv) Law of reciprocal (or equivalent) proportions (Richter – 1792): The weights of two or more different elements which separately combine with a definite weight of a third element, are either the same as, or simple multiples or submultiples of the weights of these different elements when they combine to form a new compound.



- (v) Gay Lussac's Law of gaseous volume (Gay Lussac 1808): In a chemical reaction involving gases, the volume of the reactants bear a simple ratio with each other and also with the product's volume measured under the same condition of temperature and pressure.
- (vi) Avogadro's Law (Avogadro 1811) : At N.T.P., all gases contain equal number of molecules. The volume of one mole of gas at N.T.P is 22.4 L and it contains Avogadro's number of molecules ( $N_A = 6.023 \times 10^{23}$ ). Number of molecules in 1 cm<sup>3</sup> of an ideal gas at S.T.P. is Loschmidt  $N_0$  with value 2.688 × 10<sup>19</sup>.
- > Atomic and Molecular masses : The atomic mass of an element indicates the number of times an atom of that element is heavier than  $\frac{1}{12}$  the part of the weight of an atom of carbon-12 isotope.

Atomic mass = 
$$\frac{\text{Mass of an atom of the element}}{\text{Mass of an atom of C - 12 isotope}} \times 12$$

[Since the atomic mass of an element is a ratio, it has no unit so it is only a number.]

- (i) **Physical scale of atomic mass :** The scale of atomic mass with O<sup>16</sup> isotope, relative mass 16.0000 taken as standard of reference is known as physical atomic mass.
- (ii) Chemical scale of atomic mass : The scale of atomic mass with naturally occurring three isotopes of oxygen (with atomic mass 16, 17 and 18) having relative mass 16.00447 taken as standard of reference is known as chemical atomic mass.

$$\frac{\text{Physical atomic mass}}{\text{Chemical atomic mass}} = \frac{16.00447}{16.00000} = 1.000279$$

- (iii) Gram atomic mass (GAM) or gram atom : When the atomic mass of an element is expressed in grams, it is called gram atomic mass or gram atom. 1 g atom of an element also expresses the mass in g of  $6.023 \times 10^{23}$  atoms of the element (N<sub>A</sub>).
- (iv) Atomic mass unit or a.m.u : The unit of the actual mass of element is called atomic mass unit or a.m.u or u.



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 $1 \text{ a.m.u} = \frac{1}{12} \times \text{mass of } 1 \text{ carbon atom}$  $=\frac{1}{12} \times \frac{12}{6.023 \times 10^{23}}$ 

- (v) Molecular mass : Molecular mass is the number which indicates how many times the mass of one molecule substance is heavier than  $\frac{1}{12}^{\text{th}}$  the mass of an atom of carbon-12 isotope.
- (vi) Gram molecular mass (GMM) : When the molecular mass of a substance is expressed in grams, it is called gram molecular mass.
- (vii) Formula mass and Gram formula mass : Ionic compounds such as NaCl, CsCl, ZnS, Na<sub>2</sub>O, CaF<sub>2</sub> etc. do not contain discrete entities but exist as ions closely packed together in a three dimensional space. Each ion is surrounded by a number of oppositely charged ions. For example, in NaCl, each Na<sup>+</sup> ion is surrounded by six Cl<sup>-</sup> ions and vice versa and in ZnS, each Zn<sup>2+</sup> ion is surrounded by four S<sup>2-</sup> ions and vice versa.

So in such cases, instead of molecular mass, formula mass is used.

When formula mass of a substance is expressed in grams, it is called gram formula mass.

### Mole Concept:

The mole was introduced by Ostwald around 1900 and its meaning in Latin language is 'heap' or 'pile'. One mole is the amount of the substance which contains as many elementary entities or particles as there are atoms in exactly 0.012 kg of the C<sup>12</sup> isotope.

[In case of a mole elementary entity must be specified.]

Formulae of moles can be expressed as follow :

(i) 
$$n = \frac{V}{22.4}$$
 (ii)  $n = \frac{PV}{RT}$  (iii)  
(iv)  $n = \frac{Wt.}{M_N / At.Wt.}$  (v)  $n = \frac{N}{N_A}$ 

### Determination of Atomic Mass :

(1) Dulong and Petit's rule : Except few elements (C, B, Si, Be), the atomic heat of all solid elements at room temperature is same and it is equal to 6.4 (approximately). For more details, Atomic heat = Atomic mass  $\times$  Specific heat = 6.4

Atomic mass =  $\frac{6.4}{\text{Specific heat}}$ [approx.]

Correct atomic mass = Equivalent mass  $\times$  Valency

### (2) Specific heat method :

 $C_p$  = specific heat of a gas at constant pressure,  $C_v$  = specific heat of a gas at constant volume, С

$$\gamma = \frac{C_p}{C_v}$$
,  $\gamma$  is a constant,  $\gamma = 1.66$  for mono atomic gas,  $\gamma = 1.41$  for diatomic gas,  $\gamma = 1.33$  for triatemic gas.

triatomic gas

Atomic mass of gaseous element is calculated as :



> Equivalent Mass : The equivalent mass of an element is the number of parts by weight of the element which combines or replaces from a chemical combination 8 parts by mass of oxygen, 1.008 parts by mass of hydrogen or 35.5 parts by mass of chlorine.

Equivalent mass can be calculated as :

- (i) Equivalent mass of element  $= \frac{\text{Atomic mass}}{2}$ Valency
- (ii) Equivalent mass of acid  $= \frac{\text{Molecular mass of acid}}{\text{Basicity of acid}}$
- (iii) Equivalent mass of base =  $\frac{Molecular mass of base}{Molecular mass of base}$ Acidity of base



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(iv) Equivalent mass of salt =  $\frac{\text{Formula Mass}}{(\text{Valency of cation / anion}) \times (\text{No. of cations / anions})}$ 

Formula mass

(v) Equivalent mass of oxidizing agent  $=\frac{\text{FOITIGHA ITABS}}{\text{No. of electrons gained per molecule}}$ 

(vi) Equivalent mass for reducing agent  $= \frac{\text{Formula mass}}{\text{No. of electrons lost per molecule}}$ 

(vii) Equivalent mass of radical = Formula mass of radical No. of unit charge

### Percentage Composition :

The percentage of different elements present in a compound is known as percentage composition.

- (a) Mass % of an element =  $\frac{\text{Mass of the element in the compound}}{100} \times 100$ Molar mass of the compound
- (b) For solutions
  - (i) % by mass  $\left(\frac{w}{W}\right) = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$

[Since, x% by mass means 100 g solution contains x g solute,  $\therefore$  the solution contains (100 – x)g solvent.]

(ii) 
$$\% \left(\frac{w}{V}\right) = \frac{\text{mass of solute}}{\text{Volume of solution}} \times 100$$

(iii) 
$$\% \left(\frac{v}{V}\right) = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

(iv) Molarity (M) = 
$$\frac{\text{Number of moles of solute}}{\text{Volume of solution (in litre)}}$$

(v) Molality (m) = 
$$\frac{\text{Number of moles of solute}}{\text{Mass of solvent (in kg)}}$$

(vi) Parts per million (ppm) = 
$$\frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$$

(vii) Formality (F) = 
$$\frac{\text{No. of formula units}}{\text{Volume of solution (in litre)}}$$

(viii) Normality =  $\frac{\text{Gram equivalent of solute}}{\text{Volume of solution(in litre)}}$ 

### Empirical and Molecular Formulae :

### **Determination of Empirical formula**

(i) First of all % by wt. of each element present in the given molecule is calculated.

(ii) Determine the relative number of different kinds of atoms by the following relation :

No. of atoms = 
$$\frac{\% \text{ wt. of element}}{\text{Atomic wt.}}$$

- (iii) The number obtained in the step (ii) is divided by the smallest number among them to get simple atomic ratio.
- (iv) Multiply the figures, so obtained by a suitable integer in order to get a whole no. ratio (if necessary).
- (v) The simple atomic number ratio gives the empirical formula of a compound.

### Determination of molecular formula :

(i) By the above described method, empirical formula is first calculated.

(ii) 
$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$
 [Molecular formula = (Empirical formula)<sub>n</sub>, from this relation where = 1, 2, 3,...etc.] is determined.

(iii) Molecular formulae =  $n \times$  Empirical formula

### > Stoichiometry :

- It is a quantitative relationship between the reactants and products in a reaction.
- It determines that how many moles of each reactant combines to form how many moles of products.
- Stoichiometric coefficients are the numbers written before reactants and products in a balanced equation.
  - e.g.  $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$ 
    - 1 mol 1 mol 1 mol
- The volume of one mole of any gas occupies 22.4 litre volume
- Steps for Stoichiometric Calculation :
  - (i) Write down the balance equation for the reaction of given substances.
  - (ii) Predict the substances whose mass or volume is given or to be calculated.
  - (iii) Write down the atomic mass or molecular mass or moles or molar volume of each substances of reaction.
  - (iii) On the basis of above steps calculation is done.
- Limiting Reagent : The reactant that is completely consumed when a reaction goes to completion.

$$\therefore$$
 L.R. =  $\frac{\text{given}(\text{may moles, volume or molecules})}{(3 + 1)}$ 

stoichiometry co-efficient

Least value indicates the L.R.

### II. Important Formulae :

(1)	(n)	Number of atoms present in 12 g of C <sup>12</sup> isotope. Amount of any substance containing $6.023 \times 10^{23}$ particles/atom. This number is called as Avogadro's number ( $N_A$ ).
(2)	$\frac{1}{6.023 \times 10^{23}} = 1.66 \times 10^{-24} \mathrm{g}$	Mass of one atomic unit (u)
(3)	$M = \frac{W}{1.66 \times 10^{-24}}$	M = Atomic or molecular mass W = Weight of an atom or molecule in grams
(4)	$n = \frac{W}{M}$	n = number of moles W = Weight in grams M = Molecular or atomic mass
(5)	$n = \frac{V}{22.4}$	V = Volume of gas at STP in $dm^3$
	$n = \frac{\text{Number of atoms / molecules}}{N_A}$	$N_A = 6.023 \times 10^{23}$
	$n = \frac{\text{Number of particles}}{N_{\text{A}}}$	
(6)	$\begin{array}{l} \text{Atomic mass} = \frac{6.4}{\text{Specific heat}} \end{array}$	Dulong and Petit law
(7)	$M = 2 \times Vapour density$ $PV = nRT$	Molecular mass = M P = Pressure in atm, V = Volume in litre, $n$ = No. of moles R = 0.0821 L atm mol <sup>-1</sup> K <sup>-1</sup> , T = Temperature in Kelvin
(8)	$\frac{\text{Mass of element}}{\text{Molecular mass}} \times 100$	Percentage of element
(9)	$\frac{\text{GAM of an element}}{N_{A}}$	Mass of one atom of an element $N_A = 6.023 \times 10^{23}$
(10)	$\frac{\text{GMM of substance}}{6.023 \times 10^{23}}$	Mass of one molecule of a substance
(11)	$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$	Molecular formula = $n \times$ Empirical formula





# **Chapter Objectives**

- Subatomic particles, Atomic models, Black body radiation, Photoelectric effect, Bohr's atomic model, Spectrum of hydrogen atom, Sommerfeld extension of Bohr's model, Dual character (de Broglie's relationship), Heisenberg's uncertainty principle, Shapes of orbitals, Quantum numbers, Concept of atomic orbitals as an electron wave function.
- > Rules for filling electrons in orbitals : Aufbau principle, Pauli's exclusion principle, Hund's rule,
- > Electronic configuration of elements, Extra stability of half-filled and completely filled orbitals.

### STUDY MATERIAL

### I. Concept Clarified :



### Sub-atomic particles

### (a) Electron :

- 1. It was discovered by the study of cathode rays and was named by G. J. Stoney.
- **2.** In 1897, the value of  $\frac{e}{m}$  (*i.e.*, charge/ mass ratio) was determined by J. J. Thomson as
  - $1.758820 \times 10^8$  C/g for whatever be the material of electrodes, the value remaining same.
- 3. Electron is common universal constituent of all atoms.
- 4. The charge was determined by Millikan oil drop method as  $\sim 1.602 \times 10^{-19}$  coulombs or  $4.803 \times 10^{-10}$  e.s.u.

5. Mass of electron, 
$$m_e = \frac{e}{e / m_e}$$

$$= \frac{1.602 \times 10^{-19}}{1.758820 \times 10^{11}} \text{kg} = 9.1094 \times 10^{-31} \text{kg}$$

= 0.000548 a.m.u or 1/1837 times of proton.

- **6. Radius :** 10<sup>-15</sup> cm.
- **7. Density** :  $2.17 \times 10^{17}$  g/cc

8. Mass of electron in motion = 
$$\frac{\text{Ratio of mass of electron}}{\sqrt{\left[1 - \left(\frac{u}{c}\right)^2\right]}}$$

where, u = velocity of electron, c = velocity of light

- But when u = c, the mass of moving electrons  $= \infty$ . But no body can move faster than light.
- 9. 1 mole of electron have 1 Faraday (96500 coulomb) of charge.

### (b) Proton :

**1.** The nature and existence of proton was established by the discovery of canal rays or positive rays in 1886 by Goldstein.

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- **2.** Specific charge : It is  $9.58 \times 10^4$  coulomb/g, determined by Thomson and Rutherford, named as proton, *e/m* varies with nature of gas and it is maximum when H<sub>2</sub> is used.
- 3. Mass : It has mass  $1.6725 \times 10^{-27}$  kg or 1.00728 a.m.u. which is 1837 times heavier than an electron.
- **4.** Charge on 1 mole of proton is 96500 coulomb or 1 Faraday.
- 5. Volume : It is approximately  $1.5 \times 10^{-38}$  cm<sup>3</sup>.
- (c) Neutron :
  - 1. Neutron was discovered by James Chadwick in 1932.

₄Bε

2. Chadwick bombarded Be atoms sheets with high speed  $\alpha$ -particles as follows :

$$e^9 + {}_2He^4 \rightarrow {}_6C^{12} + {}_0n^1$$

- **3. Mass :** It is a massive particle having mass  $1.6752 \times 10^{-24}$  g or  $1.6752 \times 10^{-27}$  kg or 1.00866 a.m.u
- 4. **Density** :  $1.5 \times 10^{14}$  g/cc
- 5. Specific charge : e/m = 0
- 6. Disintegration : Isolated neutron is unstable and disintegrates into proton, electron and nutrino.

### Other Fundamental Particles :

Positron (Positive electron)	Mesons	Anti-proton	Neutrino and Antineutrino
<b>Symbol</b> : ${}^{0}_{+1}e$	$\pi^+$ , $\pi^-$ or $\pi^0$	$^{1}_{-1}p$	_
Discoverer : Anderson	Yukawa	Segre and Weigland	Fermi
<b>Charge :</b> 1.6 × 10 <sup>-19</sup> C	+, - or 0	$-1.6 \times 10^{-19} \mathrm{C}$	≈ 0
Mass : Same as electron	≈ 200 times that of electron	Same as proton	≈ 0

### Some Important Informations :

1. Atomic number : The number of proton present in an atom is known as atomic number. It is represented by *Z*. Moseley gave the relationship between frequencies of the characteristic X-ray of an element with atomic number as

$$\sqrt{\nu} = a \left( Z - b \right)$$

where, *a* and *b* both are constant, with *b* as a screening constant and *a* as a proportionality constant.

- 2. Mass number : It is sum of protons and neutrons present in nucleus (as nucleons) of an element and it is denoted by A. Mass number is always a whole number.
- **3. Isotopes** (Given by Soddy) : Atoms of the same element having same *Z* but different A due to different number of neutrons present in the nucleus. For example; C<sup>12</sup>, C<sup>13</sup> and C<sup>14</sup>.
- 4. Isobars : Atoms of different elements with different Z but same A. For example;  $_{19}K^{40}$  and  $_{20}Ca^{40}$ .
- **5. Isotones** : Atoms of different elements with different *Z* and A but same number of neutrons. For example;  ${}_{6}C^{14}$  and  ${}_{8}O^{16}$ .
- 6. Isodiaphers : Atoms of different elements having the same difference of the number of neutrons and protons.

Element	p	п	е	( <i>n</i> – <i>p</i> )
5 B <sup>11</sup>	5	6	5	1
<sub>6</sub> C <sup>13</sup>	6	7	6	1

7. Isosters : Different molecules which have same number of atoms and electrons.

	$CO_2$	N <sub>2</sub> O
Atoms :	1 + 2 = 3	2 + 1 = 3
Electrons :	$6 + 8 \times 2 = 22$	$7 \times 2 + 8 = 22$

8. Isoelectronic : Atoms/ molecules/ ions having the same number of electrons.

e.g. : (i)	<sub>10</sub> Ne		$_{11}$ Na <sup>+</sup>	
e	- 10		11 – 1 =	10
e.g. : (ii)	$\mathrm{NH}_4^+$	$CH_4$	$Mg^{2+}$	$H_2O$
e <sup>-</sup>	7 + 4 - 1	6 + 4	12 - 2	2 + 8
	= 10	= 10	= 10	= 10

### > Atomic Models :

- 1. Thomson's atomic model : Thomson proposed that atom is a uniform sphere of positive charge in which electrons are distributed randomly within the sphere.
  - This model is known as "Plum pudding model" or "Raisins pudding model" or "Watermelon model".

### Drawbacks:

- (i) It is a static model and does not consider the movement of electrons.
- (ii) Mass of atom is considered to be evenly spread within the atom.
- > Rutherford's atomic model : Rutherford suggested atomic model based on the ' $\alpha$ -rays scattering' For more details, experiment. scan the code



Doubly ionised He atom

### **Observations** :

- (i) Most of the  $\alpha$ -particles passed straight through the gold foil with little or no deflection.
- (ii) Small number of  $\alpha$ -particles are deflected backward through angles greater than 90°.
- (iii) Very less number of  $\alpha$ -particles (one out of 20,000) deflected in the opposite direction at an angle of 180°.

(iv) The scattering of 
$$\alpha$$
-particles  $\alpha = \frac{1}{\sin^4(\theta/2)}$ 

### **Conclusions :**

particles 
$$\alpha = \frac{1}{\sin^4(\theta/2)}$$

- (i) Most of the space in an atom is empty.
- (ii) As only few  $\alpha$ -particles retraced the path ( $\theta = 180^\circ$ ), the positive charge and the mass of an atom is concentrated at a small place called nucleus.
- (iii) Electrons move around the nucleus in a circular path called orbit.
- (iv) Electrons and nucleus are held together by electrostatic force of attraction.

### Drawbacks of Rutherford Model :

(1) According to classical theory of electromagnetism whenever a charged particle is subjected to acceleration around an opposite charge, it emits radiation continuously. Therefore, an electron in an orbit will emit radiations, the energy carried by radiation comes from the electronic motion and ultimately go into a spiral motion and fall into the nucleus within  $10^{-8}$  s. But this does not truly happen. Thus, this theory cannot explain the stability of an atom.



(2) According to him the spectra of an atom should be continuous but the actual observed spectrum consists of well defined lines of definite frequencies (discontinuous).

### Black body radiation :

An ideal emitter of radiation is a black body and has the following characteristics :

- (a) It emits the same type of radiation that it absorbs.
- (b) It can absorb radiations of all frequencies.
- (c) It emits radiations of all frequencies and at a given temperature emits more radiation than other non-ideal emitters.





At a given temperature, intensity of radiation emitted increases with decrease in wavelength, it reaches a maximum value at a given wavelength and then starts decreasing with further decrease in wavelength.

### > Photoelectric Effect :

When a clean metallic surface is irradiated by a monochromatic light of proper frequency, electrons are emitted from surface. The phenomenon of the ejection of electrons from a metal surface is known as photoelectric effect.

### Salient Features :

- (i) If the frequency of incident radiation is below a certain, minimum value (threshold frequency), no emission takes place.
- (ii) The kinetic energy of the electrons emitted is independent of the intensity of light. The kinetic energy of the electrons increases linearly with the frequency of incident light radiation.

For each metal, there is a characteristic minimum frequency which is threshold frequency  $(v_0)$  and below this frequency photoelectric effect is not observed.

When frequency of light is equal or above the threshold frequency, electrons are emitted. For a metal, threshold frequency is equal to its work function.

 $\therefore$  W<sub>0</sub> =  $hv_0$  and E = hv

$$\Rightarrow$$
 KE = E - W<sub>0</sub> =  $hv - hv_0$ 

 $\Rightarrow$  KE =  $h(v - v_0)$ 

where,  $W_0$  = Work function

E = Energy of incident light

 $v_0$  = Threshold frequency

 $h = \text{Planck's contant} (6.626 \times 10^{-34} \text{ Js})$ 

If m is the mass and v is the velocity of the ejected electron,

$$KE = \frac{1}{2}mv^2 = h(v - v_0)$$

> Bohr's Atomic Model :

Bohr's postulates : The main points of Bohr's atomic model are as follows :

(1) Electrons revolve around the nucleus in some selected but not just in any orbit, that is orbits of electrons are discrete rather than continuous. Thus, the certain selected orbits, are known as stationary orbits or permissible orbits. In a permissible orbit, the angular momentum of an electron is an integral multiple of  $h/2\pi$ .

Angular momentum =  $mvr = n\left(\frac{h}{2\pi}\right)$ 

where, *m* is the mass of the electron, v = velocity of the electron, r = radius of the orbit, h = Planck's constant and *n* is a positive integer.

- (2) Every selected orbit is associated with some definite amount of energy and so long as electron revolves around the nucleus it does not emit energy, that is, the energy remains constant. These selected orbits are also called closed orbits.
- (3) When an electron jumps from higher energy level to the lower energy level, it emits radiation in the form of photon or quanta and *vice versa*.









where 
$$E_2 - E_1 = hv$$
  
 $\Delta E = hv$   
 $E_1, E_2 = Energies of lower and higher orbits respectively.$   
 $h = Planck's constant$   
 $v = Frequency of radiation$ 

### Some Important Results from Bohr's model of atom :

The following equations were derived by Niels Bohr for an atom or ion containing one electron :

S.No.	Property	Formulae	Numerical value
1.	Radius	$r_n = \frac{n^2 h^2}{4\pi^2 m Z e^2}$	$r_n = 0.529 \left(\frac{n^2}{Z}\right) A^\circ \text{ or } r_n = \frac{n^2}{Z} a_0$ where, $a_0 = 0.529 A^\circ$
2.	Energy of orbits : (i) Kinetic energy (K.E.)	K.E. = $\frac{2\pi^2 m K^2 Z^2 e^4}{n^2 h^2}$	K.E. = $13.6 \left(\frac{Z^2}{n^2}\right) \text{ eVatom}^{-1}$ = $+21.8 \times 10^{-19} \left(\frac{Z^2}{n^2}\right) \text{ J atom}^{-1}$
	<b>(ii)</b> Potential energy (P.E.)	P.E. = $-\frac{4\pi^2 m K^2 Z^2 e^4}{n^2 h^2}$	$P.E. = -27.2 \left(\frac{Z^2}{n^2}\right) \text{ eVatom}^{-1}$
	<b>(iii)</b> Total energy E <sub>n</sub> =K.E. + P.E.	$\mathbf{E}_n = -\frac{2\pi^2 m Z^2 e^4}{n^2 h^2}$	$E_n = -13.6 \left(\frac{Z^2}{n^2}\right) \text{ eVatom}^{-1}$ $= -21.8 \times 10^{-19} \left(\frac{Z^2}{n^2}\right) \text{ Jatom}^{-1}$ $E_n = -\text{K.E.} = \frac{\text{P.E.}}{2}$
3.	Velocity of electron $(v_n)$	$v_n = \frac{2\pi Z e^2}{nh}$	$v_n = 2.188 \times 10^{-6} \left(\frac{Z}{n}\right) \text{ms}^{-1}$
4.	Time period of revolution of electron (T)	$T_n = \frac{n^3 h^3}{4\pi^2 m K^2 Z^2 e^4} \left(T_n = \frac{2\pi r_0}{v_r}\right)$	$T_n = 1.5 \times 10^{-16} \left(\frac{Z}{n}\right) s^{-1}$
5.	<b>Orbital frequency</b> $v = \frac{1}{T_n}$	$v = \frac{1}{T_n} = \frac{4\pi^2 m K^2 Z^2 e^4}{n^3 h^3}$	$v = \frac{1}{T_n} = 6.66 \times 10^{15} \left(\frac{Z^2}{n^3}\right) s^{-1}$

### Limitations of Bohr's Model :

- (i) This model fails to explain the spectrum of elements having only one electron.
- (ii) This model fails to explain the splitting of spectral lines under magnetic field (Zeeman effect) and electric field (Stark effect).
- (iii) This model fails to explain the three dimensional model of an atom.
- (iv) It also fails to explain quantization of angular momentum.
- (v) The de Broglie concept and Heisenberg's uncertainty principle can not be explained by this model.

### Hydrogen Spectrum :

When an electron jumps from a higher energy state to a lower energy state, a bluish light is emitted. When a ray of this light is passed through a prism, a spectrum of several isolated lines are obtained. The wavelength of these line spectra lies in the visible, UV and IR region, these lines are grouped into different series as follows :





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ų.

When an electron jumps from any higher state to the ground state (or lower state), the series of spectral lines are obtained according to following equations :

$$\bar{\mathbf{v}} = \mathbf{R}_{\mathrm{H}} \mathbf{Z}^{2} \left( \frac{1}{n_{1}^{2}} - \frac{1}{n_{2}^{2}} \right)$$

where,  $n_1 =$  lower energy state,  $n_2 =$  higher energy state,  $R_H =$  Rydberg's constant,  $\overline{v} =$  Wave number  $R_{\rm H}$ = 109677 cm<sup>-1</sup> (experimentally measured value),  $R_{\rm H}$ = 109737 cm<sup>-1</sup> (theoretical value). Total number of lines when an electron is in excited state,  $n = n_2$  to deexcited as  $n = n_1 \cdot n_2 > n_1$  is given by

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

For transition n = any to orbit n = 1

Total no. of lines =  $\frac{n(n-1)}{2}$ 

### Sommerfeld's Extension of Bohr's Model :

When lines spectra of hydrogen were passed through a high resolving power spectroscope, it was found that finer spectra of hydrogen over beam was obtained. To explain the finer spectra of hydrogen, Sommerfeld proposed that an electron always moves in an electrical orbit with different ellipticity in addition to the circular orbits. An electron has two degrees of freedom in an elliptical path.

- (a) Radial angle
- (b) Radial distance

The motion of an electron is controlled by two quantum numbers in an elliptical path.

- (a) Radial quantum number  $(n_r)$
- **(b)** Azimuthal quantum number (K) (Again,  $n = n_r + K$ )

 $\frac{n}{K} = \frac{\text{Length of the major axis}}{\text{Length of the minor axis}}$ 

When  $n > K \rightarrow$  Elliptical path But when  $n = K \rightarrow \text{Circular path}$ 



The azimuthal quantum number (K) has been replaced by l; l' is defined as l=K-1, it denotes the energy associated with angular momentum of the electron and it is quantized as follows :

$$mvr = \frac{h}{2\pi}\sqrt{l(l+1)}$$

Thus, to explain the finer spectra of hydrogen, Sommerfeld introduced the concept of subshells.

### Limitations of Sommerfeld's Theory :

- (1) It could not predict the correct number of observed fine structure of spectral lines.
- (2) It fails to explain the spectra of multi electron atoms.

### Dual Character :

- 1. In 1924, de-Broglie, a French physicist suggested that like light, matter also has dual character. This proposal gave birth to a new theory known as **wave mechanical theory of matter**. According to him, matter has wave as well as particle character.
- **2.** According to de-Broglie, wavelength ( $\lambda$ ) of an electron is inversely proportional to its momentum (*p*).

$$\lambda \propto \frac{1}{p} \Rightarrow \lambda \propto \frac{1}{mv}$$
,  $[p = \text{momentum} = \text{mass}(m) \times \text{velocity}(v)]$   
Again,  $\lambda = \frac{h}{p}$ ,  $[h = \text{Planck's constant}]$ 

3. Derivation of de-Broglie Relation :

According to Planck's equation,

$$\mathbf{E} = h\mathbf{v} = \frac{hc}{\lambda} \tag{1}$$

where, E = energy, h = Planck's constant, v = frequency of radiation, c = velocity of light,  $\lambda =$  wavelength According to Einstein's equation, (Mass energy relationship)

$$E = mc^2$$
 [m = mass]

Equating both (1) and (2), we get

$$\frac{hc}{\lambda} = mc^2 \implies \lambda = \frac{h}{mc} = \frac{h}{p}$$
$$\implies \lambda = \frac{h}{mv} = \frac{h}{p} \quad \text{[For smaller particle other than light, } c \text{ is replaced by } v\text{]}$$

It is evident from the above equation that  $\lambda$  decreases with increasing of mass of 'm'. The wavelength of a fast moving object like a tennis ball or an aeroplane is very low as their mass is heavier.

### Some important relations concerned with de-Broglie concept.

• de-Broglie wavelength in terms of kinetic energy (K.E.) :

K.E. = 
$$\frac{1}{2}mv^2$$
 or  $mv = \sqrt{2m \text{KE}}$ 

 $\lambda = \frac{h}{mv}$  or  $\lambda = \frac{h}{\sqrt{2m.\text{KE}}}$ 

Again

When a charged particle such as an electron is accelerated with a potential of V, then its K.E. may be as follows :

$$K.E. = \frac{1}{2}mv^{2} = eV$$
  

$$\Rightarrow m^{2}v^{2} = 2meV$$
  

$$\Rightarrow mv = \sqrt{2meV}$$
  

$$\Rightarrow \frac{h}{\lambda} = \sqrt{2meV} \quad \left[As \ \lambda = \frac{h}{mv}\right]$$
  

$$\Rightarrow \lambda = \frac{h}{\sqrt{2meV}}$$

• For a charged particle of charge 'q',  $\lambda = \frac{h}{\sqrt{2qmeV}}$ 

· de-Broglie wavelengths associated with charged particles :

(i) For electronic particle : 
$$\lambda = \frac{12.27}{\sqrt{V}} A^{\circ}$$

[V = accelerating potential of the particles]

...(2)

- (ii) For proton :  $\lambda = \frac{0.286}{\sqrt{N}} A^{\circ}$
- (iii) For  $\alpha$ -particle :  $\lambda = \frac{0.101}{\sqrt{V}} A^{\circ}$

· For uncharged particles : Neutrons

$$\lambda = \frac{h}{\sqrt{2Em}} = \frac{6.626 \times 10^{-34}}{\sqrt{2 \times 1.67 \times 10^{-27} \,\mathrm{E}}} = \frac{0.286}{\sqrt{\mathrm{E(eV)}}} \,\mathrm{A}^{\circ}$$

For gas molecules :  $\lambda = \frac{h}{\sqrt{3mkT}}$ , [k= Boltzmann constant]

### > Angular Momentum of Bohr Electron from de-Broglie Equation :

Tiny particles like revolving electron must have wave character. For an electron wave to be in phase, the circumference of the Bohr's orbit should be an integral multiple of the wavelength of an electron.

$$\therefore n\lambda = 2\pi r$$
$$\therefore \lambda = \frac{2\pi r}{n} \qquad \dots (1)$$

where, r = Radius of the orbit, n = an integer

Again from de-Broglie equation,

$$\lambda = \frac{h}{mv} \qquad \dots (2)$$

where, m = Mass of an electron, v = Velocity of electron From (1) & (2) we get

$$\frac{2\pi r}{n} = \frac{h}{mv}$$

$$mor(\text{Angular momentum}) = \frac{nh}{2\pi}$$

The above equation is Bohr's postulate of angular momentum.

### Heisenberg's Uncertainty Principle :

It is impossible to measure simultaneously, precisely both the position and the momentum |For more details, (or velocity) of a small (microscopic) moving particle with absolute accuracy or certainty. Mathematically it can be expressed as

$$\Delta x \cdot \Delta p \ge \frac{h}{4\pi}$$
$$\Rightarrow \Delta x \times \Delta v \ge \frac{h}{4\pi m}$$

scan the code

where,  $\Delta x$  = Uncertainty in position,  $\Delta p$  = Uncertainty in momentum,  $\Delta v$  = Uncertainty in velocity, *m* = Mass of microscopic particle.

### Significance of Heisenberg's Uncertainty Principle

It cancelled the existence of definite path or trajectories of electrons or other microscopic particles.

### Quantum Numbers :

1. Principle quantum number (*n*) :

(i) It was given by Bohr.

=

(ii) The values of *n* lies between 1 to  $\infty$ *i.e.*,  $n = 1, 2, 3, 4 - - - \infty$ ,

- (iii) It gives the size and energy of the orbit.
- (iv) The maximum number of For more details, electrons in an orbit =  $2n^2$ scan the code
- (v) Total number of orbitals =  $n^2$
- (vi) Angular momentum of a revolving electron is *mvr* =

### (iii) It gives the shape of orbitals. (iv) The value of *l* ranges from 0 to (*n*–1) for each

principal energy levels. The sub-levels are as follows : Value 0

(ii) It gives the angular momentum of an electron.

Value 
$$0$$
 1 2 3  
Subshell s p d f

(v) Angular momentum of an electron in an orbital h

$$mvr = \frac{n}{2\pi}\sqrt{l(l+1)}$$

2. Azimuthal quantum number (*l*) :

(i) It was given by Sommerfeld.

- (vi) Total number of electrons in a sub-orbit = 2(2l + 1).
- (vii) Total number of orbitals in a sub-orbit (2l + 1).
- 3. Magnetic quantum Number (*m*) :
  - (i) It was given by Lunde.
  - (ii) It determines orientation of orbitals, that is, the direction of electron density.
- (iii) Permissible values of m = all integral values from -l to + l including '0'.
- (iv) Maximum number of electrons in an orbital =2 (with opposite spin) For l = 1  $m = \pm 1$  0 1

For 
$$l = 1$$
,  $m = +1$ , 0, 1  
 $p_{x'} \quad p_{y'} \quad p_z$   
For  $l = 2$   $m = +2$   $+1$  0  $-1$ ,  $-2$   
 $d_{x^2-y^2} \quad d_{xz} \quad d_{z^2} \quad d_{yz} \quad d_{xy}$ 

- (v) Z = component of orbital angular momentum of electron, this is also quantized.
  - $\therefore \qquad L_z = m \left(\frac{h}{2\pi}\right) \quad [m = \text{Magnetic quantum} \\ \text{numbers}] \\ = \sqrt{l(l+1)}$

- 4. Spin quantum number :
  - (i) It was given by Goudsmit and Uhlenbeck
  - (ii) It gives the spin of electron.

(iii) Values of 
$$s = \pm \frac{1}{2}$$
  
 $\left( +\frac{1}{2} \text{ (clockwise) and } -\frac{1}{2} \text{ (anticlockwise)} \right)$ 

(iv) Total spin in an atom

 $=\frac{1}{2}$  × number of unpaired electrons.

(v) Spin angular momentum of electron

$$\mathcal{L}_s = \sqrt{s(s+1)} \times \frac{h}{2\pi}$$

(vi) It gives the spin only value of magnetic momentum  $\mu_s = \sqrt{n(n+2)}$ ,

### Concept of Atomic Orbitals as One Electron Wave Function :

In Schrödinger wave equation,  $\psi$  has no physical significance but  $\psi^2$  has physical significance.  $\psi^2$  represent the probability of finding electron around nucleus in space. When the probability of finding electron is maximum (90%), it is called as orbital. The variation of  $\psi$  and  $\psi^2$  at different distances (*r*) from the nucleus is shown in the following figures :



(a) Variation of (a) orbital wave function  $\Psi$ 

(b) Probability density  $\Psi^2$  with dense of the electron from the nucleus 1s and 2s orbitals.

A region or space where probability of finding an electron is zero is called a node and when finding an electron is maximum is called a peak.

Number of radial node / Nodal surface / spherical node = n - l - 1Number of angular node / Nodal plane = l total nodes = n - 1

### Rules for Filling Electrons in Orbitals :

• Aufbau principle : Electrons are filled up progressively to the various orbitals in their order of increasing energy in subshells.

The increasing order can be determined on the basis of (n + l) rule as follows (n = principle quantum number l = Azimuthal quantum number)

(a) Orbital energies increase, as (n + l) increases.

(b) If there are two orbitals with same value of (n + l), the one with the smaller *n* has the lower energy.

The approximate order is :  $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s \cong 4d < 5p < 6s \dots$ 

• **Pauli's Exclusion principle :** No two electrons in the same atom can have all the four quantum numbers equal. Based on the principle the following conclusions are drawn :

(i) Maximum number of electrons in an orbital is two.

- (ii) In any shell number of orbits is  $n^2$ . So, maximum number of electrons in any shell is  $2n^2$ .
- (iii) So, s, p, d, f subshells contain a maximum of 2, 6, 10, 14 electrons respectively.

### ➢ Hund's Rule :

The rule may be stated as

- (i) Ground state of an atom contains maximum number of unpaired electrons (within the same subshell).
- (ii) Electrons in different orbitals of the same energy will have parallel spins.

The ground state electronic configuration of  $2p^3$  is I not II or III.



### Electronic Configuration of Elements :

Electronic configuration of any orbital of an element can be represented by the notation  $nl^x$  where, n = Number of the principal shell, l = Symbol of subshell or orbital (*s*, *p*, *d*, *f*), x = Number of electrons present in an orbital.

Exceptional stability of partially and fully filled orbitals due to :

- (i) Symmetrical distribution
- (ii) Maximum exchange of energy

### II. Important Formulae :

Important formulation derived from Bohr's atomic model.

1. 
$$r_n = \frac{n^2 h^2}{4\pi^2 \text{Ke}^2 m^2}$$
$$\text{K} = \frac{1}{4\pi\epsilon_0} = 9 \times 10^{-7} \text{Nm}^2/\text{C}^2$$
$$r = \frac{n^2}{Z} \times 0.529 \text{ A}^\circ (r = \text{Bohr's orbit of electrons})$$

**2.** Energy (E) of electrons in  $n^{\text{th}}$  orbit

$$E = -\frac{Z^2}{n^2} \times \frac{2\pi^2 m K^2 e^4}{h^2}$$
$$= -\frac{Z^2}{n^2} \times 21.79 \times 10^{-19} \text{ J / atom}$$
$$= -\frac{Z^2}{n^2} \times 13.6 \text{ eV}$$
$$= -\frac{Z^2}{n^2} \times 1312 \text{ kJ / mol}$$
$$E_n = \frac{Z^2 R_E}{n^2}$$
$$R_E = -13.6 \text{ eV}(\text{ Rydberg energy})$$

3. 
$$E_n = \frac{E_1}{n^2}$$
  
 $E_n = E_1 \times \frac{Z^2}{n^2}$ , for hydrogen like species.

**4.** Velocity of electrons in a particular shell in orbit can be calculated as

$$v = \sqrt{\frac{\mathrm{K}e^2}{mr}}$$
  
where, K =  $\frac{1}{4\pi\varepsilon_0} = 9 \times 10^9 \,\mathrm{Nm^2/\,C^2}$ 

$$v = \frac{Z}{n} \times 2.188 \times 10^8 \,\mathrm{cm/s}$$

5. Potential energy of electrons in a particular shell

K.E. = 
$$\frac{1}{2} \frac{KZe^2}{r} = + \frac{13.6}{n^2} Z^2 eV$$
  
Total energy (TE) =  $-\frac{1}{2} \frac{KZe^2}{r}$   
TE =  $\frac{1}{2}$  PE  
TE =  $-KE$ 

6. Number of resolutions/s by an electron in a shell  $-\frac{\text{Velocity}}{\sqrt{2}} - \frac{\nu}{\sqrt{2}} - \frac{E_1}{\sqrt{2}} + \frac{2}{\sqrt{2}}$ 

7. Frequency of electrons in 
$$n^{\text{th}}$$
 orbit

$$=\frac{v}{2\pi r}=\frac{6.62\times10^{15}}{n^3}Z^2$$

**8.** Period of revolution of electrons in  $n^{\text{th}}$  orbit  $(T_n)$ 

$$T_n = \frac{2\pi r}{v_n} = \frac{1.5 \times 10^{-16}}{Z^2} n^3 s$$
$$T_n \propto \frac{n^3}{Z^2}$$

9. Ionization potential

$$= E_{\infty} - E_n$$
$$= 0 - \left( -\frac{Z^2}{n^2} \times 13.6 \text{ eV} \right)$$
$$= \frac{Z^2}{n^2} \times 13.6 \text{ eV}$$

$$=\frac{Z^2}{n^2} \times 21.79 \times 10^{-19}$$
 J/atom

10. Bohr's theory of atom is applicable to only species such as,  $Li^{2+}$ ,  $He^+$ .

**11.** 
$$\frac{1}{\lambda} = \overline{\upsilon} = R_{\rm H} \times z^2 \left[ \frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$$

where,  $R_H = Rydberg's$  constant

**12.** Maximum number of lines produced when an electron returns from  $n^{\text{th}}$  orbit to ground state

$$=\frac{n(n-1)}{2}$$

**13.** Total number of lines when an electron is in an excited state,  $n = n_2$ , to deexcited state to  $n = n_1$  is

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

**14.** According to de-Broglie, matter has particle as well as wave nature.

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

$$\lambda = \frac{h}{\sqrt{2\text{KE}.m}} \quad [\text{K.E.} = \text{Kinetic Energy}]$$

- **15.** According to Heisenberg's uncertainty principle  $\Delta x \cdot \Delta p \ge \frac{h}{4\pi}$
- **16.** Principal quantum number (n) = 1, 2, 3, ...
- **17.** Azimuthal quantum number l = 0, 1, 2, ..., (n-1)
- **18.** Magnetic quantum number (m) = -l to +l
- **19.** Spin quantum number  $(m_s) = +\frac{1}{2}$  or  $-\frac{1}{2}$
- 20. Orbital angular momentum

$$\sqrt{l(l+1)}\frac{h}{2\pi} = \sqrt{l(l+1)h}$$

- **21.** Number of subshells in a shell = n
- **22.** Number of maximum orbitals in a shell =  $n^2$
- **23.** Number of maximum orbitals in a shell = (2l + 1)
- **24.** Number of radial nodes = n l 1
- **25.** Number of angular nodes = l
- **26.** Plot of radial probability function  $(4\pi r^2 R^2)$







# **Chapter Objectives**

- Significance of classification, Brief history of the development of periodic table, Modern periodic law and the present form of periodic table, Classification of elements on the basis of their electronic configuration
- Periodic trends in properties of elements : Atomic radii, Ionization enthalpy, Electron gain enthalpy, Electronegativity, Valency, Diagonal relationship, Effective nuclear charge and shielding effect
- > Nomenclature of elements with atomic number greater than 100.

### **STUDY MATERIAL**

### I. Concept Clarified :

### > Significance of Classification of Elements :

At present 118 elements are known and recently discovered elements are not natural but are synthetic, *that is*, man made. It is difficult to remember all the properties of these elements and their compounds. For this reason scientists classify the elements in such a form where one can study them in an organized manner. The classification and arrangement of elements based on common properties to make a systematic study is called as Periodic Table.

### > Brief History of the Development of Periodic Table :

[a] Lavoisier Classification (1789) :

Lavoisier classified the elements into metals and non-metals. **Drawback** 

Lavoisier failed to place the metalloids in his classification.

[b] Proust's Hypothesis (1815) :

All elements are made up of H atom

Atomic weight of element =  $n \times$  (Atomic weight of one H atom)

[Atomic weight of H=1 and n = any whole number]

### Drawbacks

(i) Every element can not be formed by the combination of H-atom.

(ii) It fails to explain the fractional atomic weight.

### [c] Dobereiner's Triad Rule (1817) :

Atomic weights of second element is arithmetic mean of atomic weight of the first and third elements in the group. The three elements have similar chemical properties.





S.N.	Known as Triad	1 <sup>st</sup> element	2 <sup>nd</sup> element	3 <sup>rd</sup> element	Arithmetic mean
1.	Triad At. wt.	Li 7	Na 23	K 39	$\frac{7+39}{2} = 23$
2.	Triad At. wt.	Cl 35.5	Br 80	I 127	$\frac{35.5 + 127}{2} = 81.25$

The following triad also follows the above rule :

(Ca, Sr, Ba), (K, Rb, Cs), (P, As, Sb), (S, Se, Te), (H, F, Cl), (Sc, Y, La).

### Drawback

The rule is not applicable for d and f – block elements.

### [d] Newland's Octave Rule (1865) :

John Newland (English chemist) started his studies with hydrogen (lowest atomic weight) and ended at thorium which was 56<sup>th</sup> element.

The elements are arranged in order of their increasing atomic weights, the properties of every eighth element is the repetition of the first one like the eighth note of a musical scale.

### Newland's arrangement of elements into octaves

Notes of Music	sa(do)	re(re)	ga(mi)	ma(fa)	pa(so)	dha(la)	ni(ti)
	Н	Li	Be	В	С	Ν	0
Elements	F	Na	Ma	Al	Si	Р	S
	C1	Κ	Ca	Cr	Ti	Mn	Fe
	Co and Ni	Cu	Zn	Y	In	As	Sc
	Br	Rb	Sr	Ce & La	Zr	_	_

Drawbacks

- (i) This rule is applicable only up to Ca.
- (ii) Position of H along with F, Cl is not justified due to differences in chemical properties.
- (iii) The discovery of the inert gas gives this rule a death blow.

### [e] Lothar Meyer's Curve [1869] :

Lothar Meyer plotted a graph between atomic volume and atomic masses of the elements and observed that the elements having similar properties occupy similar positions on the curve.

The following observations can be drawn from the curve :

- (i) The most strongly electropositive, that is, the alkali metals (Li, Na, K, Rb and Cs.) having largest atomic volumes, occupy the peaks on the curve.
- (ii) The less strongly electropositive, i.e., the alkaline earth metals (Be, Mg, Ca, Sr and Ba) occupy the descending position on the curve.
- (iii) Metalloids (B, Si, As, Te, At etc.) and transition metals occupy minima, i.e., the bottom part of the curve.
- (iv) The most electronegative, i.e., the halogens (F, Cl, Br and I) occupy positions on the ascending position of the curve before inert gases.

Lothar Meyer suggested that the atomic weights of different elements are a periodic function which later on became the base of Mendeleev's periodic table. The practical utility of this graph is not in the satisfactory level because the various position of the curve to keep up in mind is not an easy task.



[f] Telluric Screw (or Helix) (1862) :

A three dimensional periodic table proposed by French geologist A.E.B. de Chancourtois.

### [g] Mendeleev's Periodic Table (1869) :

**Mendeleev's periodic law :** The physical and chemical properties of different elements are the periodic function of their atomic weight.

### Characteristics of Mendeleev's Periodic Table :

- (i) The periodicity is based on atomic weight.
- (ii) He placed 63 elements in his table. Inert gases were not discovered by then.
- (iii) In Mendeleev's periodic table, the horizontal rows are known as periods and the vertical columns are known as groups.
- (iv) In this periodic table, there are 7 periods and 8 groups.
- (v) Each group upto VII is divided into subgroup A and B respectively.
- (vi) In VIII group, there are nine elements in three rows.
- (vii) The elements possessing the same group have similar properties.

### Merits of Mendeleev's periodic table :

- (1) Systematic classification of elements : Mendeleev arranged the known elements in order of their atomic weights considering the fact that elements with similar properties should lie in the same vertical column, i.e., group.
- (2) Correction of doubtful atomic weights : Atomic weight = Valency × Equivalent weight, from this relation he corrected atomic weights of some elements.

Be has certain similarities with Al, Be must be trivalent and equivalent weight of Be was found to be 4.5. So atomic weight of Be would be  $4.5 \times 3 = 13.5$ , and there was no space in Mendeleev's table for this element. Hence, he asserted that Be must be bivalent and correct atomic weight as  $4.5 \times 2 = 9.0$  and for this there was a space between Li and B.

U, Be, In, Au, Pt are the elements in which corrections were done in atomic weight.

(3) **Prediction of new elements :** Mendeleev predicted the properties of those missing elements from the known properties of the other elements in the same group.

He predicted the properties of Gallium (Eka aluminium), Scandium (Eka Boron), Germanium (Eka Silicon), Technetium (Eka Manganese)

Mendeleev supplied valuable informations in the discovery of some trace elements like Pm (61), Re (75), Hf (72), Po (84), At (85), Fr (87). Therefore, it might be said that Mendeleev's periodic table is a renaissance in chemistry.

### Demerits of Mendeleev's Periodic Table :

- (a) Anomalous position of hydrogen : Hydrogen resembles both the group I A (alkali metals) and group VII B (halogens). So the position of hydrogen is anomalous or controversial.
- (b) Position of isotopes : Isotopes have different atomic weights but they were placed in the same group.
- (c) Anomalous pairs of elements : The principle of periodic law, that is, the increasing order of the atomic weights was not maintained in four pairs of elements which violated the periodic law. These are :

1	(	)	(		(	)	(	`
	Te	Ι	Ar	Κ	Co	Ni	ll Th	Pa
	127.61	126.92	39.94	39.1	58.94	58.69	232.12	231.0
1								

(d) Like elements were placed in different groups : There were some elements with almost identical properties but they were placed in different groups.

Cu Hg Pt Au IB IIB VIII IB

- (e) Unlike elements were placed in same group.
- (f) **Position of Lanthanides and Actinides :** 14-Lanthanide and 14-Actinide elements were not placed in proper positions in the main frame of the periodic table but they were placed in two separate rows at the bottom of the periodic table.
- (g) Different atomic weights but in the same position : The principle of the periodic table has been violated in group VIII where Fe, Co, Ni in 4<sup>th</sup> period, Ru, Rh, Pd in 5<sup>th</sup> period, Os, Fr, Pt in 6<sup>th</sup> period were placed together.
- (h) Cause of periodicity : Mendeleev could not explain the cause of periodicity among the elements.

### > Modern Periodic Law and the Present Form of Periodic Table :

**Moseley (1913) :** When a metal is bombarded with a strong beam of electrons, X-rays are produced. He observed that the square root of the frequency of X-ray  $(\sqrt{\upsilon})$  is directly proportional to the number of nuclear charge of metal.



$$\sqrt{\upsilon} = a(z-b)$$

a and b both are constants where b is a constant and it is known as screening constant, a is a proportionally constant, (Nuclear charge of a metal equals to the atomic number).

When we plotted a graph  $(\sqrt{\upsilon})$  versus atomic number (z), he got a straight line but the graph  $\sqrt{\upsilon}$  For more details, scan the code

versus atomic weight is not a straight line. So, Moseley related the properties of elements with their atomic number and stated the new periodic law.

**Moseley's Periodic Law or Modern Periodic Law :** 'The physical and chemical properties of different elements are periodic functions of their respective atomic number'.



### Bohr's Table :

(i) It follows Bohr's scheme based on electronic configuration and atomic number.

(ii) There are 7 periods (horizontal rows) and 18 groups (vertical columns).

(iii) 18 vertical columns are named according to IUPAC as 1<sup>st</sup> to 18<sup>th</sup> group.

(iv) The co-relation between the groups in modern and long form of periodic table is shown as below:

### PERIODIC TABLE

s. (Mai	-Block El n-Group	ements • Elements	(6	Ĥ	tend	ded o	or Lo:	ng Fc	rm c	of the	Peri	odic '	Iable						(18)
↓ (	(I)		1																Cero
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	1.007	(Z) 6												()   	[4) (J				0026
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	<u>.</u>	4												<u>n</u>	q	<u> </u>	x	ע	10
	2 Li	Be											·	В	υ	Z	0	щ	Ne
	6.94	1 9.0122											10.	81 12.(	011 14.0	07 15.9	999 18.	998/20	.179
	11	12					d-Blo	ck Elen	nents				1	3	14	15	16	17	18
	3 Na	Mg	į		į		lransit `_`	ion ele	ments	() (		1			Si		S	Ū	Ar
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	37	38	39	40	41	42	4	44	4	4	6 4	7 4	<b>x</b>	61	20	51	52	53	54
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			232.(	04 231	.04 23	8.03	237)	(244)	(243)	(247)	(247)	(251)	(252)	(257)	(258)	(259)	(262)		



### > Periods :

**I Period :** Shortest period. Contains only two elements  ${}_{1}H_{2}He$ .

II Period : Contains 8 elements.

<sub>3</sub>Li, <sub>4</sub>Be, <sub>5</sub>B, <sub>6</sub>C, <sub>7</sub>N, <sub>8</sub>O, <sub>9</sub>F, <sub>10</sub>Ne

III Period : Contains 8 elements.

<sub>11</sub>Na, <sub>12</sub>Mg, <sub>13</sub>Al, <sub>14</sub>Si, <sub>15</sub>P, <sub>16</sub>S, <sub>17</sub>Cl, <sub>18</sub>Ar

[N.B : II and III periods are known as short periods]

IV Period : Contains 18 elements.

<sub>19</sub>K, <sub>20</sub>Ca, <sub>21</sub>Sc, <sub>22</sub>Ti, <sub>23</sub>V, <sub>24</sub>Cr, <sub>25</sub>Mn, <sub>26</sub>Fe, <sub>27</sub>Co, <sub>28</sub>Ni, <sub>29</sub>Cu, <sub>30</sub>Zn.......... <sub>36</sub>Kr.

V Period : Contains 18 elements.

<sub>37</sub>Rb - - - - - - <sub>54</sub>Xe

IV and V periods are called as long periods.

**VI period :** Contains 32 elements starting with  ${}_{55}$ Cs and ending with  ${}_{86}$ Ra. This period contains representative, transition and inner transition elements. This period is known as the longest period.

**VII period :** It is incomplete period and contains 32 elements as in VI period. Starting from  ${}_{87}$ Fr to  ${}_{92}$ U, all the elements are naturally occurring but are radioactive with very short half-lives. The rest of the actinides are also radioactive but are synthetically prepared in laboratories.

### ➤ Groups:

Group 1 : Alkali metals	Group 15 : Pnictogens (Fertilizer producing elements)
General electronic configuration : <i>ns</i> <sup>1</sup>	General electronic configuration : $ns^2 np^3$
Group 2 : Alkaline earth metals	Group 16 : Chalcogens (Ore - forming elements)
General electronic configuration : <i>ns</i> <sup>2</sup>	General electronic configuration : $ns^2 np^4$
Group 13 : Boron family	Group 17 : Halogens (Salt forming elements)
General electronic configuration : $ns^2 np^1$	General electronic configuration : $ns^2 np^5$
Group 14 : Carbon family	Group 18 : Noble gases (Inert gases)
General electronic configuration : $ns^2 np^2$	General electronic configuration : <i>ns</i> <sup>2</sup> <i>np</i> <sup>6</sup>

### > Classification of Elements on the Basis of their Electronic Configuration :

On the basis of electronic configuration, the elements are divided into four blocks, i.e., *s*, *p*, *d*, *f*. This division is based upon the name of the orbital which receives the last electron.

**1.** *s* – **block elements** : The elements of groups 1 and 2 which have general outermost electronic configuration  $ns^{1-2}$  (n = outermost shell). [Where n = 2 - 7] are called *s*-block elements.

### General characteristics of s – block elements

- (i) They are soft metals with low melting and boiling points.
- (ii) They are highly electropositive and have low ionization enthalpies.
- (iii) They have valency + 1 (for alkali metals) and +2 (for alkaline earth metals) respectively.
- (iv) These are reactive elements, metallic as well as reactivity increases as we move down the group.
- (v) The compounds of *s*-block elements are mainly ionic except Be compounds.
- (vi) Most metals (except Be and Mg) in this block impart characteristic colour to the flame.

(vii) s-Block elements are strong reducing agents and are good conductor of heat and electricity.

**2.** *p***-Block elements :** The elements of groups 13, 14, 15, 16, 17 and 18 (excluding He) which have general electronic configuration  $ns^2 np^{1-6}$  (n = outer most shell) [where n = 2 - 7] are called *p*-block elements.

### General characteristics of *p*-block elements :

- (i) *p*-block elements contain both metals and non-metals. Metallic character decreases from left to right along the period and metallic character increases from top to bottom within a group.
- (ii) Their ionisation enthalpies are relatively higher comparing with s-block elements.
- (iii) They mostly form covalent compounds.
- (iv) Some elements show variable oxidation states in their compounds.
- (v) Their reducing character increases from top to bottom in a definite group and their oxidising character increases from left to right along a period.
- 3. *d* block elements : The elements which are present between *s* and *p* block elements having general outermost electronic configuration  $(n 1) d^{1-10} ns^{1-2} [n = \text{outermost shell}]$  are called *d* block elements. Characteristics of *d*–block elements :
  - (i) These are hard, malleable and ductile elements with high melting and boiling points.
  - (ii) They are good conductors of heat and electricity.



For more details,

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- (iii) Ionization enthalpies lie between *s* and *p* block elements.
- (iv) They form both ionic and covalent compounds which are generally coloured and paramagnetic in nature.
- (v) They show variable oxidation states.
- (vi) They form complexes and alloys.
- (vii) They are mostly used as catalyst.
- (viii) Zn, Cd, Hg do not show most of the properties of transition elements.
- **4.** *f***-block elements** : Lanthanides (rare earth elements) and actinides are both inner transition elements which have general electronic configuration as  $(n 2) f^{1-14} (n 1) s^2 p^6 d^{0-1} ns^2$  [n = outer most shell] are called *f*-Block elements.

### General characteristics of *f* - Block elements :

- (i) They are hard with high boiling and melting points.
- (ii) They show variable oxidation states and their compounds are generally coloured.
- (iii) These elements also form complex compounds.
- (iv) All the elements of actinide series are radioactive.

### > Periodic Trends in Properties of Elements :

### • Atomic Radii :

Atomic size is a very important property of atom. Many physical and chemical properties are explained by considering the atomic size, i.e., the atomic radii.

Atomic radius is defined as the distance from the centre of the nucleus to the outermost shell containing electrons.

### Difficulties in determining atomic radius :

- (i) Practically isolated single atom is quite impossible.
- (ii) There is no sharp boundary of an electron surrounding the atom, that is why probability of finding electrons even in large distances may never be zero.
- (iii) In different bonded state, the atomic radius can be different.

### Types of Atomic Radii :

(i) Covalent Radii : One half of the distance between the nuclei of two like atoms in homoatomic molecule bonded together by a single covalent bond.





(ii) Metallic Radii : One half of the distance between the nuclei of two adjacent atoms in the metallic crystal lattice which is closely packed.

Again according to electron sea model, the force of attraction between mobile electrons and the positive kernel is called metallic bond. As in a metallic lattice the valence electrons are mobile therefore they are weakly repelled by the metal ions or kernels. Therefore, the metallic radius is always greater than the covalent radius.

(iii) Van der Waal's Radii : It is defined as one half of the distance between the nuclei of two adjacent identical atoms belonging to two neighbouring molecules of an element in solid state.



Covalent radius = 99 pm

### Covalent and Van der Waal's radius of chlorine

[Van der Waal's radius > Metallic radius > Covalent radius.]

### Variation of Atomic Radii in Periodic Table :

- (a) Variation along period : Moving from left to right across period, atomic radii of atoms progressively decrease as the value of charge of nucleus is increased.
- (b) Variation along group : Atomic radii of atoms increase on moving from top to bottom within the group as a new shell is added to each succeeding element.

- (iv) Ionic Radii : Charged species of particles are called ions. Positively charged species of ions is called cation and negatively charged species of ions is called anion.
  - Positive ion is always smaller than that of corresponding atom.
  - Negative ion is always larger than that of the corresponding atom.
  - The cation with higher charge will be smaller than the less higher positive charge. For example, same atom with +2 charge will be smaller than atom with +1 charge.
  - The anion with higher charge will be larger than anion with smaller charge. For example, same atom with -2 charge will be bigger than atom with -1 charge.
  - For isoelectronic species, ionic radii  $\propto \frac{1}{7}$

### • Ionization Potential or Ionization Enthalpy :

The minimum amount of energy required to remove the outermost electron, i.e., the most loosely bound one, from its isolated gaseous atom in its lowest energy state, i.e., the ground state to form a gaseous cation is known as ionization energy / ionization potential. (As energy is expressed in electron volts, so it is also known as ionization potential.)

A(g) +E $\rightarrow$  $A^+(g)$  + $e^-(g)$ isolatedEnergy requiredGaseous cationRemoved electrongaseous atom(ionization potential) $A^+(g)$  $A^+(g)$  $A^+(g)$ 

The amount of energy required to remove 2nd, 3rd, 4th electron from isolated gaseous cation to form a corresponding gaseous greater charged cation is called as 2nd, 3rd, 4th ionization potential respectively.

A<sup>+</sup> (g) 
$$\frac{-e^{-}}{+E = IP_2}$$
 A<sup>2+</sup>(g)  $\frac{-e^{-}}{+E = IP_3}$  A<sup>3+</sup>(g).....

### **Factors affecting Ionization Potential :**

(i) Magnitude of the nuclear charge :

From Bohr's theory  $\Delta E \propto Z^2$ , i.e., greater is the nuclear charge, greater will be ionization potential.

(ii) Atomic radius : Ionization potential is inversely proportional to the atomic radius.

I.P. 
$$\propto \frac{1}{\text{Atomic radii}}$$

- (iii) **Penetration effect of subshell :** Ionisation potential increases with the increasing of penetration effect of electrons. An *s*-electron will move closer (due to its maximum ellipticity) than *p*, *d*, *f*. So within the same shell, the penetration effect decreases in the order s > p > d > f, therefore if the penetration effect is more, ionization potential will be high.
- (iv) Shielding or screening effect : For multi electron atom, inner shell electrons act as screen or shield between the nucleus and the electron in outermost shell. This is called a shielding or screening effect. More is the number of electrons in inner shells, greater is the shielding effect and smaller the force of attraction between nucleus and outermost shell's electron and lesser is the ionization potential values.
- (v) Fully and half filled electronic configuration : Fully or half filled orbital is energetically stable (due to maximum symmetry and exchange energy) and consequently ionisation potential will be higher.

Variation of I.P. in a period : On moving from left to right in a period , I.P. value of elements increases as effective nuclear charge increases.

**Variations of I.P. in a group :** On moving down from top to bottom in a definite group, the effective nuclear charge is almost constant but the number of shells increases, hence, as the size increases I.P. progressively decreases. **Applications of Ionization Potential :** 

- (1) Metallic character : Metallic character  $\propto \frac{1}{IP}$
- (2) **Relative reactivity :** Relative reactivity  $\propto \frac{1}{I.P.}$

For this reason inert gases are chemically inert due to high I.P.

(3) **Reducing power :** Reducing power  $\propto \frac{1}{I.P.}$ 

The order of reducing property for alkali metal increases down the group as I.P. decreases. Li> Na> K> Rb> Cs

### Electron Gain Enthalpy :

The amount of energy released when an electron is added to the outermost shell of an isolated gaseous atom in its lower energy state, i.e., the ground state to form a gaseous anion is called electron gain enthalpy.

[Electron affinity is measured at absolute zero. The relation between electron gain enthalpy ( $\Delta H_{eg}$ ) and electron affinity is as follows :

$$\Delta H_{eg} = -Ae - \frac{5}{2}RT$$

The electron gain enthalpy can be represented as follows :

$$X(g) + e^{-} \rightarrow X^{\ominus}(g)$$
  
anion  $\Delta H = -\Delta H_{eg}$ 

The  $1^{st}$  electron gain enthalpy is negative but  $2^{nd}$  electron gain enthalpy is positive because addition of second electron is opposed by electrostatic repulsion and hence energy is to be supplied for the addition of  $2^{nd}$  electron and the overall process becomes endothermic.

$$\begin{array}{ll} O + e^{-}(g) \to O^{-}(g) & \Delta H_{2eg} = -141 \text{ kJ} / \text{ mol} \\ O^{-} + e^{-} \to O^{2-}(g) & \Delta H_{2eg} = +780 \text{ kJ} / \text{ mol} \\ O + 2e^{-} \to O^{2-}(g) & \Delta H_{eg} = +639 \text{ kJ} / \text{ mol} \end{array}$$

Electron affinity always be less than the kinetic energy of the electron being added.

### Factors affecting Electron Gain Enthalpy (EGE) :

(i) Atomic size : If the atomic size increases the force of attraction between gaseous atom and incoming electron decreases and electron gain enthalpy becomes more positive.

$$\therefore$$
 E.G.E.  $\propto \frac{1}{\text{Atomic size}}$ 

- (ii) Effective nuclear charge : E.G.E.  $\propto$  Effective nuclear charge
- (iii) Screening or shielding effect ( $Z_{eff}$ ) : E.G.E.  $\propto \frac{1}{\text{Shielding effect}}$
- (iv) Half filled and completely filled orbitals : These orbitals are energetically stable. Hence, to add an electron, energy is to be supplied and the electron gain enthalpy decreases.

Variation of EGE in a period : On moving across a period, E.G.E. increases as effective nuclear charge increases. Variation of EGE in a group : On moving down a group, the atomic size increases , EGE decreases .

### Electronegativity :

It is defined as the relative power of an element to attract the shared electron pair in a covalent bond towards itself. **Factors affecting Electronegativity (E.N.) :** 

(i) Hybridization : In hybridised orbitals, greater the percentage of *s* - character, greater is the E.N.

Elements	$sp^3$	$sp^2$	sp
С	2.48	2.75	3.24
Ν	3.68	3.94	4.67

(ii) Atomic radius : E.N.  $\propto \frac{1}{\text{Atomic radius}}$ 

- (iii) Electron affinity : E.N.  $\propto$  E.A.
- (iv) Oxidation number : The electronegativity of an element increases with the increase of its positive oxidation state.

Periodic trends of electronegativity of elements :

In a period : On moving along a definite period from left to right, E.N. increases as atomic number increases.

**In a group :** On moving down in a definite group from top to bottom, atomic radius increases and E.N. decreases. **Application of Electronegativity :** 

(i) **Nature of bond :** Nature of bond formed depends on the differences of the electronegativity of the constituent atoms.

E.N. of elements		Interaction	Types of bond	
Α	В			
(i) Low	High	Transfer of e's	Ionic	
( <b>ii</b> ) High	High	Sharing of e's	Covalent	
<b>(iii)</b> Low	Low	Electron sea	Metallic	

(ii) Partial ionic character in a covalent bond : The ionic character in a covalent bond increases as electronegativity difference of a bond increases.

Example : H–F > H–Cl> H–Br> H–I

- (iii) Metallic and Non metallic character of elements : Non metal can easily accept an electron but metal loses electron readily. Therefore, higher is the E.N., higher will be the non metallic character and vice versa.
- (iv) Electronegativity and the nature of oxides of elements : The relative strength of acidic, basic and amphoteric oxides depends mainly upon E.N. of the elements.

Acidic oxide  $\rightarrow$  good acceptor of lone pair of electrons

Basic oxide  $\rightarrow$  good donor of lone pair of electrons

The oxides of 3rd period elements shows basic and acidic character as follows :

Na <sub>2</sub> O	MgO	$Al_2O_3$	$SiO_2$	$P_2O_5$	$SO_3$	Cl <sub>2</sub> O <sub>7</sub>
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Strongly basic Basic Amphoteric Weak acidic Acidic Strong acid Very strong acidic Basic nature decreases, acidic nature increases.

(v) Colour of salt : If the difference of EN becomes lesser, the colour of the compound get deepens.

Compounds		Colour
AgCl (30% ionic)	$\rightarrow$	White
AgBr (24% ionic)	$\rightarrow$	Light yellow
AgI (15% ionic)	$\rightarrow$	Yellow
Ag <sub>2</sub> S (4% ionic)	$\rightarrow$	Black

If the difference of EN is more, the compound is colourless.

- (vi) Bond strength : As the difference of EN between elements increases, bond distance decreases and bond strength increases.
- (vii) Hydrogen bond formation : Atoms with small size (F , N , O) and high electronegativity form hydrogen bonds easily.

### Measurement of EN :

(i) Pauling scale :

For AB molecule

$$X_{A} - X_{B} = 0.208 \left[ E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}} \right]^{1/2}$$

 $E_{AA}$  = Bond energy of AA molecules

- $E_{AB} = Bond energy of AB molecules$
- $E_{BB}$  = Bond energy of BB molecules
- (ii) Mulliken scale :

$$X_{A} = \frac{(IE)_{A} + (EA)_{A}}{2}$$
$$X_{A} = \frac{(IE)_{A} + (EA)_{A}}{2 \times 2.8}$$
 [When both have energy value in eV]

$$X_{A} = \frac{(IE)_{A} + (EA)_{A}}{2 \times 62.5}$$
 [When both have energy value in kcal unit]

(iii) Allred Rochow scale :

$$X_{\rm A} = \frac{0.359(Z_{\rm eff})_{\rm A}}{r_{\rm Cov}^2} + 0.744$$

 $Z_A$  = Actual nuclear charge of atom A (Atomic number)

$$Z_{\rm eff} = Z_{\rm A} - \sigma_{\rm A}^*$$

 $\sigma_{\Lambda}^{*}$  = Screening / Shielding constant of atom A

 $r_{A(Cov)}$  = Covalent radius of atom A



> Valency : Periodic classification of elements is actually according to their valency.

**In a period :** On moving from left to right in a period, the number of valence electrons increases from one to eight. From group-I to group-VIII, the valency of elements with respect to oxygen increases. The valency of elements belong to zero group is zero. Example : The valency of elements of period 3 with respect to oxygen increases as shown below.

Group	Ι	II	III	IV	V	VI	VII
Element	Na	Mg	Al	Si	Р	S	C1
Oxide	Na <sub>2</sub> O	MgO	Al <sub>2</sub> O <sub>3</sub>	SiO <sub>2</sub>	P <sub>2</sub> O <sub>5</sub>	SO <sub>3</sub>	Cl <sub>2</sub> O <sub>7</sub>
Valency	1	2	3	4	5	6	7

For transition elements, no general trend is observed in the valency of elements. These elements have variable valencies due to presence of vacant *d* - subshells. Similarly, no general trend is observed for inner transition elements due to same reasons.

In a group : The elements belonging to particular group have the same valency.

### Diagonal Relationship :

It exists between certain pairs of diagonally adjacent elements in the second and third periods of the periodic table.

In the first two short periods three pairs of elements which show close similarity in their chemistry.

$$\overset{\text{Li}}{\overset{\text{Na}}{\sim}}\overset{\text{Be}}{\overset{\text{Mg}}{\sim}}\overset{\text{Be}}{\overset{\text{Al}}{\sim}}\overset{\text{C}}{\overset{\text{C}}{\overset{\text{Si}}{\sim}}}$$

When we move from Li to Be, it raises the charge on the cation and further decreases the size of cations consequently Be attains a good degree of covalency. But when we move from Be to Mg in the same group, the size of the cation increases, but the charge remains the same. As a result ionic bond is favoured. The overall effect from Li to Na and then Mg possess similar deforming power, i.e., the diagonal relationship.

### > Effective Nuclear Charge and Shielding Effect :

why effect of one is ect from E the cancelled by other

(EN) Increases

Along diagonal EN

relationship, that is

is related by opposite



The actual charge felt by the valence shell electrons is called effective nuclear charge and repulsive force felt by the valence shell electrons from the electrons present in inner shells is known as screening or shielding effect. Effective nuclear charge can be find out by considering the following equation :  $Zeff = Z - \sigma$ 

Decreases

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Zeff – Effective nuclear charge $Z = Atomic number$	Electrons in orbitals	$\sigma$ Per electron of orbit,			
$\sigma =$ Slater's screening constant, and if values lies	$\downarrow$ (of shell) $\rightarrow$	п	( <i>n</i> –1)	( <i>n</i> –2)	( <i>n</i> – 3)
$0 < \alpha < Z$	s or p	0.35	0.85	1.00	
	d or f	0.35	1 00	1.00	

For 1s electron  $\sigma = 0.30$  (Z = 1e<sup>-</sup> system)

### IUPAC Nomenclature of elements with Atomic Number above 100

### CNIC (Commission on Nomenclature of Inorganic Chemistry) (Approved by IUPAC in 1994) :

This official name of elements is recommended for those elements which have atomic number greater than 100. This is based on the Latin words for atomic number of the elements.

According to the IUPAC system, the rules of naming the elements of atomic number more than 100 are as follows : (1) The root of the three digits in the atomic number of the elements are serially expressed with adding ium as follows :

i1		<b>h</b> ;	1	and	nont	how	cont	oct	
0	1	2	3	4	5	6	7	8	9

nil un bi tri quad pent hex sept oct enn (2) The symbol of the element is derived from the first letters from the root words.

Example :

Atomic Number	Numerical roots	IUPAC Name	Symbol
100	Un nil nil	Unninilium	Unn
101	Un nil un	Unnilunium	Unu
102	Un nil bi	Unnilbium	Unb
118	Un un oct	Ununoctium	Uuo
200	Bi nil nil	Binilnilium	Bnn

### **II. Important Points :**

- 1. Lowest melting point metal : Hg
- **2.** Highest melting point and boiling point metal : W (Tungsten)
- **3.** Lowest melting point and boiling point nonmetal : He
- 4. Notorious element : H
- 5. Lightest element : H
- 6. Least electropositive element : F
- 7. Highest electron affinity : Cl
- 8. Lowest electron affinity : Noble gases
- 9. Lowest ionization potential : Cs
- **10.** Highest ionization potential : He
- **11.** Highest electronegativity : F
- 12. Lowest electronegativity : Cs
- **13.** Most electropositive element : Cs
- **14.** Element with electronegativity next to fluorine : Oxygen
- **15.** Group containing maximum number of gaseous elements in periodic table : zero group (18)
- **16.** Total number of gaseous elements in periodic table : 11 (H, N, O, F, Cl, He, Ne, Ar, Kr, Xe, Rn)
- 17. Total number of liquid elements in periodic table : 6 (Ga, Br, Cs, Hg, Fr, Rb)
- **18.** Total number of radioactive elements in periodic table : 25
- 19. Liquid element of radioactive nature : Fr
- **20.** Volatile d block elements : Zn, Cd, Hg.
- **21.** Element containing no neutron : H
- 22. Most abundant element on earth : Oxygen
- **23.** Rarest element on earth : At
- 24. Most abundant metal on earth : Al

- 25. Element having maximum tendency for catenation : Carbon
- **26.** Non metal having highest melting point and boiling point : Carbon (Diamond)
- 27. Metal showing highest oxidation number : Os (+8)
- 28. Most electrovalent compound : CsF
- **29.** Most stable carbonate :  $Cs_2CO_3$
- 30. Strongest alkali (hydroxide) : CsOH
- 31. Smallest anion : H<sup>-</sup>
- **32.** Smallest atomic size : H
- **33.** Largest atomic size : Cs
- **34.** Largest anion :  $I^-$
- **35.** Smallest cation : H<sup>+</sup>
- 36. Heaviest naturally occurring element : U
- 37. Poorest conduction of electricity : Diamond
- 38. Hardest naturally occurring element : Diamond
- 39. Lightest solid metal : Li
- **40.** 90% of sun mass: H
- 41. Most abundant d block metal : Fe
- 42. Most abundant s block metal : Ca
- 43. Element having maximum isotopes : Sn(10)
- 44. Most stable element : Te
- **45.** Highest density (metals) : Os, Fr
- 46. Highest density (non metals) : Boron
- **47.** Strongest basic oxide :  $Cs_2O$
- 48. Best electricity conductor among metals : Ag
- **49.** Best electricity conduction among non metals : C (Graphite)
- 50. Most poisonous element : Pu (Plutonium)
- **51.** Liquid non metal : Br
- 52. Element kept in water : P

- 53. Element kept in kerosene : Na, Li, K
- **54.** Elements sublime on heating :  $I_2$
- 55. Noble metals : Au, Pt etc.
- 56. Amphoteric metals : Be, Zn, Al, Sn, Pb
- **57.** Amphoteric non metal : Si
- 58. Metalloid elements : B, Si, As, Te, Al, Ge, Sb
- **59.** Non metals having metallic lustre: Graphite, Iodine.
- **60.** First man made element : <sub>43</sub>Tc (Technetium)
- **61.** Smallest period : 1<sup>st</sup> (2 elements)
- **62.** Largest period in periodic table : 6<sup>th</sup> (32 elements)
- **63.** Largest group in periodic table : 3<sup>rd</sup> / III 13 (32 elements)
- **64.** First noble prize of chemistry was given to Jacobus Henricus Van't Hoff (1901) for the discovery of the laws of chemical dynamics and osmotic pressure in solutions.
- **65.** Promethium (61) was first synthetic radioactive element.
- **66.** Actinoids are radioactive elements, so present in very small quantity in earth. So, they are known as rare earth elements.
- **67.** Noble gases have largest atomic size due to large Van der Waal's radius.
- **68.** In alkali and alkaline carbonates, stability increases down the group.
  - $\begin{array}{l} \text{Li}_2\text{CO}_3 < \text{Na}_2\text{CO}_3 < \text{K}_2\text{CO}_3 < \text{Rb}_2\text{CO}_3 < \text{Cs}_2\text{ CO}_3.\\ \text{BeCO}_3 < \text{MgCO}_3 < \text{CaCO}_3 < \text{SrCO}_3 < \text{BaCO}_3. \end{array}$

**69.** In alkali carbonates solubility decreases down the group.

 $Li_2CO_3 < Na_2CO_3 < K_2CO_3 < Rb_2CO_3 < Cs_2CO_3$ [ $\Delta H_{hvd} > lattice energy$ ]

**70.** In alkaline carbonates, solubility decreases down the group.

 $BeCO_3 > MgCO_3 > CaCO_3 > SrCO_3 > BaCO_3$ [ $\Delta H_{hvd} > lattice energy$ ]

- Ra is ∞- emitter and produces two noble gases He and Rn.
- **72.** Bridge elements : Elements of 2<sup>nd</sup> period are also called as bridge elements. The properties of these elements are diagonally related with 3<sup>rd</sup> period of elements as follows.

2 <sup>nd</sup> period	Li 🗖	Be	B	C	N	0	F
3 <sup>rd</sup> period	Na	Mg	Al	<sup>∐</sup> Si	Ъ	<sup>⊾</sup> s	<sup>™</sup> Cl

- **73.** Typical elements : Properties of the elements of 3<sup>rd</sup> period, for example; Na, Mg, Al, Si, P, S, Cl belonging to structural group resemble the properties of the corresponding typical elements of the respective group.
- **74.** Highest density liquid : Hg (13.6 g/cc) and highest density solid : Os (22.6 g/cc)
- **75.** The elements after U (92) are known as transuranic on synthetic elements. Most of them were discovered by Enrico Fermi and Emilio Segrè.

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