MIND MAP : LEARNING MADE SIMPLE CHAPTER - 1

Some Basic Concepts of Chemistry

Stoichiometry

Deals with the calculation of masses of the products and reactants involved in a reaction.

How To Balance A Chemical Equation:
Step 1: Write correct formulas of reactants and products.
Step 2: Balance number of C atoms.
Step 3: Balance number of H atoms.
Step 4: Balance number of O atoms.
Step 5: Verify the number of atoms of each element.

Atomic Mass Unit (amul): A mass exactly equal to one-twelfth the mass of one carbon-12 atom.
Molecular Mass: Sum of atomic masses of the elements present in a molecule.
Formula Mass: Sum of the atomic weights of atoms in a molecule. One Mole is the amount of a substance that contains as many particles/entities as there are atoms in exactly 12g (or 0.012 kg) of the ^{12}C isotope.
Molar Mass: Mass of one mole of a substance in grams.

Mass Percent $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$

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

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

Empirical Formula: Simplest number ratio of various atoms present in a compound

Molecular Formula: Shows exact number of different types of atoms present in a molecule of a compound.

How To Determine Empirical and Molecular Formula:
Step 1: Conversion of mass percent to grams.
Step 2: Convert into number moles of each element.
Step 3: Divide the mole value obtained above by the smallest number.
Step 4: Write empirical formula by mentioning the number after writing the symbols of respective elements.
Step 5: Writing molecular formula
(a) Determine empirical formula mass. Add the atomic masses of various atoms present in the empirical formula.
(b) Divide molar mass by empirical formula mass.
(c) Multiply empirical formula by n obtained above.

Limiting Reagent: Reactant which gets consumed and limits the amount of product formed.

Mole Fraction

Mole Fraction of A $= \frac{n_a}{n_a + n_b}$

Mole Fraction of B $= \frac{n_b}{n_a + n_b}$

Classification

Mixtures: Two or more substances present in any ratio. (air,tea)

- **Homogenous** : Uniform composition (Sugar solution, air)
- **Hetrogenous:** Non-uniform composition (Mixtures of salt and sugar.)

Pure Substance: Fixed composition (Cu, Ag)

- Elements: Contains one type of particles i.e., atoms, molecules (Na,Cu)
- Compounds: Two or more atoms of different elements (H_2O, NH_3)

Law of Conservation of Mass: Matter can neither be created nor destroyed. It was given by Antoine Lavoisier.
Law of Definite Propportions: A given compound always contains exactly the same proportion of elements. It was given by Joseph Proust.
Law of Multiple Porportions: If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratio of small whole numbers. It was given by Dalton.

(i) Gay Lussac's Law of Gaseous Volume: When gases combine or all produced in a chemical reaction they do so in simple ratio by volume provided all gases are at same temperature and pressure.
(ii)Avogadro Law: Equal volumes of gases at the same temperature and pressure should contain equal number of molecules.
(iii)Dalton's Atomic Theory:
- Matter consists of indivisible atoms.
- All the atoms of a given element have identical properties including identical mass.
- Compounds are formal when atoms of different elements combine in a fixed rate.
- Atoms are neither created nor destroyed in chemical reaction.

Importance of Chemistry

- In meeting human needs for food, health care products and other products required for improving quality of life.
- In diverse areas such as weather patterns, functioning of brain and operation of a computer.
- In chemical industries.

Matter: Anything which occupies space

(i) Solid: Particles are held very close to each other in an orderly fashion with no freedom of movement. — Have definite volume and shape.

(ii) Liquids: Particles are close and can move around. — Have definite volume but no definite shape.

(iii)Gases: Particles are far apart and their movement is easy and fast. — Neither have definite volume nor definite shape.

Properties of Matter:
(i) Physical: Properties measured/observed without changing the identity or composition of substance. (Colour, odour)
(ii) Chemical: Properties measured/observed when a chemical reaction occurs. (Acidity or Basicity, combustibility)

Measurement of Matter:
-International System of Units (SI): (Length-m)
-Mass is the amount of matter present in a substance.
-Weight is the force exerted by gravity on an object.
-Uncertainty of measurement: Range of possible values within which the true, value of the measurement lies.
Significant figures: Meaningful digits known with certainty.
-Precision: Closeness of various measurements for same quantity.
-Accuracy: Agreement of a particular value to true value of result.

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 2

Structure of Atom

Atomic Spectra

Types:

- **Emission Spectra:** Spectrum of radiation emitted by a substance that has absorbed energy.
- **Absorption Spectra:** It is like photographic negative of an emission spectra.
- **Line / Atomic Spectra:** Emission Spectra which do not show a continuous spread of wavelength from red to violet, rather they emit light only at specific wavelength with dark space between them.

$\bar{\nu} = 109677\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right) cm^{-1}$ where

$n_1=1,2$............$n_2=n_1+1, n_1+2$.........

Series	n_1	n_2	Spectral Region
Lyman	1	2,3.........	Ultraviolet
Balmer	2	3,4.........	Visible
Paschen	3	4,5.........	Infrared
Brackett	4	5,6.........	Infrared
Pfund	5	6,7.........	Infrared

Atomic Number (Z)

Number of protons in nucleus of an atom or Number of electrons in a neutral atom

Mass Number (A)

Number of protons (Z) + number of neutrons (*n*)

Isotopes

Atoms of same element having same atomic number but different mass number. (Isotopes of hydrogen) Protium $^1H_1$, Deuterium $^2D_1$ and Tritium $^3T_1$,

Isobars:

Atoms of different elements with different atomic number but same mass number. (Ca and Ar)

Discovery of Electron:

By J.J.Thomson Charge to mass ratio of electron $= 1.758820 \times 10^{11}$ C kg^{-1}
Charge an electron $= 1.6022 \times 10^{-19}$C
Mass of electron $= 9.1094 \times 10^{-31}$Kg

Discovery of Proton

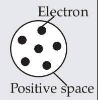
By Ernest Rutherford
Charge on proton $= +1.6022 \times 10^{-19}$C
Mass of Proton= 1.672×10^{-27} Kg

Discovery of Neutron

By James Rutherford
Charge on neutron = 0
Mass of neutron= 1.675×10^{-27} Kg

Thomson Model of Atom: (Watermelon Model)

Atom possesses a spherical shape in which the positive charge is uniformly distributed.



Rutherford's Nuclear Model of Atom

Postulates:
- Positive charge and most of the mass of atoms was densely concentrated in extremely small region (ie) nucleus.
- Nucleus is surrounded by electrons that move around the nucleus with high speed in circular path called orbits.
- Electrons and nucleus are held together by electrostatic forces of attraction.

Drawbacks:
- It cannot explain the stability of an atom.
- It does not say anything about the electronic structure of atoms.

Developments Leading To Bohr's Model of Atom:

Dual nature ie wave like and particle like of the electromagnetic radiation.
- Experimental results regarding atomic spectra.
- Wave nature of electromagnetic radiation:

It was given by James Maxwell.
Frequency (ν): Number of waves that pass a given point in one second. Unit – Hertz (Hz), Velocity of light = Frequency × Wave length
Wave number ($\bar{\nu}$): Number of wavelengths per unit length. Unit – m^{-1}
- **Particle nature of electromagnetic radiation:**
  Planck's quantum theory: E= hν
  Planck's constant (h) $= 6.626 \times 10^{-34}$ Js

Photoelectric Effect: It Was found By H.Hertz

It is the phenomenon of ejection of electrons from the surface of a metal when light of suitable frequency strikes on it.

Energy of orbitals:

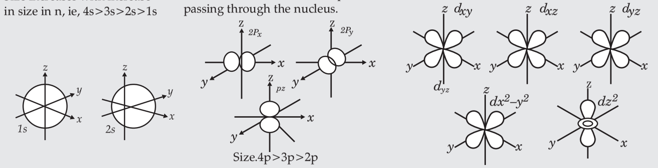
Lower the value of (*n*+l) for an orbital, lower is its energy.

Shapes of orbitals:

s Orbitals: Spherically symmetric. Size increases with increase in size in n, ie, 4s>3s>2s>1s

p Orbitals: Each p Orbitals consists of two sections called lobes on either side of plane passing through the nucleus.

d Orbitals – "Clover leaf" distribution. – Two angular nodes.



KE of ejected electron $h\nu=h\nu_0 + 1/2\, m_e v^2$

Filling of orbitals in atom

- Aufbau Principle: In the ground state of atoms, the orbital's are filled according to increasing energies.
- P auli Exclusion Principle: No2 electrons in an atom can have same set of four quantum numbers.
- Hund's rule: Pairing of electrons in the orbitals belonging to same subshell does not take place until each electron belonging to the subshell is singly occupied.

Electronic Configuration of atoms

(i) $s^a p^b d^c$......... notation
(ii) Orbital diagram

s p d

Quantum Numbers

(i) Principal quantum number (n) :n = 1,2,3,4.............. Shell = K, L, M, N.......
(ii) Azimuthal Quantum number : For given value of n ,l = 0 to n-1
(iii) Magnetic Quantum number (m) : for subshell with 'l' value m_l=2l+1
(iv) Spin Quantum Number (m_s): $+1/2\,(\uparrow), -1/2\,(\downarrow)$

Schrodinger Equation

Fundamental Equation was developed by Schrodinger as $\hat{H}\psi = E\psi$ = where $\hat{H}$= Hamiltonian

Quantum Mechanical Model of Atom:

- Dual behaviour of atom i.e., particle and wavelike. De Broglie equation: $\lambda = \frac{h}{mv} = \frac{h}{p}$
- Heisenberg's Uncertainty Principle: It is impossible to determine simultaneously, the exact position and momentum of an electron

$x.\; P_x \geq = \frac{h}{4\pi}$

$\Rightarrow$ v. $x = \frac{h}{4\pi}$

Bohr's Model for Hydrogen Atom

Postulates:
- Electron in H atom can move around the nucleus in a circular path of fixed radius and energy called as orbits. These orbits are arranged concentrically around the nucleus.
- Each of these orbits has a definite energy known as energy levels or stationary states.
- When an electron jumps from a lower energy level to higher one, some energy is absorbed.
- $v = \frac{E}{n} = \frac{E_2 - E_1}{n}$ Bohr's frequency rule.

Angular momentum of electron: $m_e vr = n\frac{h}{2\pi}$ n=1,2,3.....

Limitations:
- Unable to account for finer details of H atom. Spectrum observed by sophisticated spectroscopic techniques.
- Could not explain the ability of atoms to form molecules by chemical bonds.

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 3

Classification of Elements and Periodicity in Properties

Purpose

To ease out difficulty in studying individually the chemistry of all the elements and their compounds.

Genesis of Periodic Classification

- Law of Triads: Johann Dobereiner (1829)
- Law of Octaves: John Alexander Newlands (1865)
- Periodic Law: Dimitri Mendeleev and Lothar Meyer.
  It states that the properties of the elements are periodic function of their atomic weights.
- Modern Periodic Law: Henry Moseley (1913)
  It states that the physical and chemical properties of the elements are periodic functions of their atomic numbers. ; Horizontal rows – Periods, Vertical columns – Groups – are numbered from 1 to 18.

IUPAC Nomenclature of Elements with Atomic numbers > 100:

Derived from the atomic number of element using numerical roots for 0 and numbers 1-9 and "ium" is added at the end.

Properties of Elements based on Electronic Configuration

- Electronic Configuration is the distribution of electrons into orbitals of an atom.
- In periods: Number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.
- Group wise: Elements in same group have similar valence shell electronic configurations. Same number of electrons in outer orbitals and similar properties. These are classified into four blocks i.e., s-block, p-block, d-block and f-block.

Periodic Table Classification based on Types of Elements

s-Block Elements

- Group 1 (alkali metals) and Group 2 (alkaline earth metals)
- Outermost configuration is ns^1 or ns^2
- Reactive with low IE.
- Metallic character and reactivity increases down the group.

P- Block element

- Group 13 to 18.
- Also called as representatives or main group elements
- Outermost configuration varies from $ns^2 np^1$ to $ns^2 np^6$
- At the end of period are low reactive noble gases.
- Halogens and Chalcogens have high negative electron gain enthalpies.
- Metallic character increases down the group.

d- Block: Elements

- Group 3-12
- Outer configuration is $(n-1) d^{1-10} ns^2 np^6$
- Forms coloured ions.
- Exhibit variable valence, paramagnetism.
- Also called as Transition elements.
- Some are used as catalysts.

f -Block: Elements

- Also called as Inner Transition Elements.
- Contains Lanthanoids and Actinoids.
- Outer configuration ie $(n-2) f^{1-14} (n-1) d^{1-10} ns^{0-2}$
- All are metals.
- Actinoids are radioactive.

Periodic Trends in Properties of Elements

Properties	Group	Period
(a) Atomic Radius : Distance from the centre of the nucleus to the outermost shell containing electrons.	Increases	Decreases
(b) Electron Gain Enthalpy : Energy released when a neutral isolated gabelous atom accepts an electron from anion.	Becomes less negative	Becomes more negative
(c) Ionization Energy: The minimum amount of energy required to remove the electron from the outmost orbit of an isolated atom in gaseous state.	Decreases	Increases
(d) Electronegativity : Tendency of an atom to attract the shared pair of electrons towards it "Group"self.	Decreases	Increases
(e) Electron Affinity	Decreases	Increases
(f) Valency: Number of univalent atoms which combine with an atom of given element.	No Change	Increase from 1 to 4 and then decrease from 4 to 0.
(g) Metallic Character:	Increases	Decreases
(h) Non–Metallic Character:	Decreases	Increases

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 4

Chemical Bonding and Molecular Structure

Chemical Bond: Attractive forces which hold the various chemical constituents together in different chemical species.

Types:

(i) Covalent Bond: A chemical bond formed between two atoms by mutual sharing of electrons between them to complete their octet.

(ii) Ionic Bond: A chemical bond formed by complete transference of electrons from one atom to another acquire the stable nearest noble gas configuration.

Molecular orbital theory

Postulates:
- Electrons in a molecule are present in various molecular orbitals as electrons are present in atomic orbitals.
- Atomic orbitals of comparable energies and proper symmetry combine.
- Atomic orbitals is monocentric while a molecular orbital is polycentric.
- Number of molecular orbital formed is equal to number of combining molecular orbitals.
- Bonding molecular orbitals has low energy and high stability Types of MO: σ(Sigma), π (Pi), δ (Delta)

Chemical Bonding

Kossel Lewis approach to chemical bonding:
* Lewis pictured the atom as a positively charged 'kernel' and the outer shell accommodates a maximum of eight electrons.
- Lewis postulated that atoms achieve the stable octet when linked by chemical bonds.
- Kossel gave following facts:
* In the periodic table, highly electronegative halogens and highly electropositive alkali separated by noble gases.
* Formation of a negative ion from a halogen atom and a positive ion from an alkali metal atom is associated with gain and loss of electron by respective atoms.
* Negative and positive ions formed attain noble gas electronic configurations.
- Negative and positive ions are stabilized by electrostatic attraction.

Octet Rule: Atoms can combine either by transfer of valence electrons from one atom to another or by sharing of valence electrons to complete octet in their valence shells.

Lewis Dot Structure provides a picture of bonding in molecules and ions in terms of the shared pairs of electrons and the octet rule.

How To Write A Lewis Dot Structure:

Step 1: Add the valence electrons of the combining atoms to obtain total number of electrons.

Step 2: For anions, each negative charge means addition of one electron. For cations, each positive charge means subtraction of one electron from total number of valence electrons.

Step 3: Write chemical symbols of combining atoms.

Step 4: Least electronegative atom occupies central position.

Step 5: After accounting for shared pairs of electrons remaining are either utilized for multiple bonding or remain as lone pairs.

Formal Charge= (Total number of valence electrons in free atom) – (Total number of non-bonding electrons) – 1/2(Total number of bonding electrons)

Limitations Of Octet Rule:
- Shows three types of exceptions (ie) incomplete octet of central atom, odd-electron molecules and expanded octet.
- Does not account for the shape of molecules.
- Fails to explain stability of molecules.

Hydrogen Bond: Formed when the negative end of one molecule attracts the positive end of other.

Types:

(i) Intermolecular: Between two different molecules of same or different compounds.

(ii) Intramolecular : H atom is between two highly electronegative atoms.

Lattice Enthalpy

Energy required to completely separate one mole of a solid ionic compound into gaseous constituent ions.

Valence Shell Electron Pair Repulsion (VSEPR)

Postulates:
- Shape of molecule depends upon the number of valence shell electron pairs around central atom.
- Pairs of electrons in the valence shell repel one another.
- These pairs of electrons tend to occupy such positions in space that minimize repulsion.
- The valence shell is taken as a sphere with electron pairs localising on spherical surface at maximum distance from one another.
- A multiple bond is treated as if it is a single electron pair and the two or three electron pairs of a multiple bond are treated as a single super pair.
- When one or more resonance structures can represent a molecule, VSEPR model is applicable:
- **Decreasing order of repulsive interaction:**
  lp – lp > lp – bp > bp – bp

Valence Bond Theory: Given by L Pauling. It explains that a covalent bond is formed between two atoms by overlap of their half-filled valance orbitals, each of which contains one unpaired electron.

Orbital Overlap Concept: Formation of a covalent bond results by pairing of electrons in valence shell with opposite spins.

Types of Overlapping: (i) Sigma (σ) bond – end to end.
(ii) Pi (TT) bond – axis remain parallel to each other.

Hybridisation: Process of intermixing of orbitals of different energies resulting in formation of new set of orbitals of equivalent energies and shape.

Types of Hybridisation –(i) sp (ii) sp^2 (iii) sp^3

Bonding Molecular Orbitals: Addition of atomic orbitals.

Antibonding Molecular Orbitals: Substraction of atomic orbitals.

Bond Parameters

(I) Bond Length: Equilibrium distance between the nuclei of two bonded atoms in molecule.

(ii) Bond Angle: Angle between the orbitals containing bonding electron pairs around central atom in a molecule complex ion.

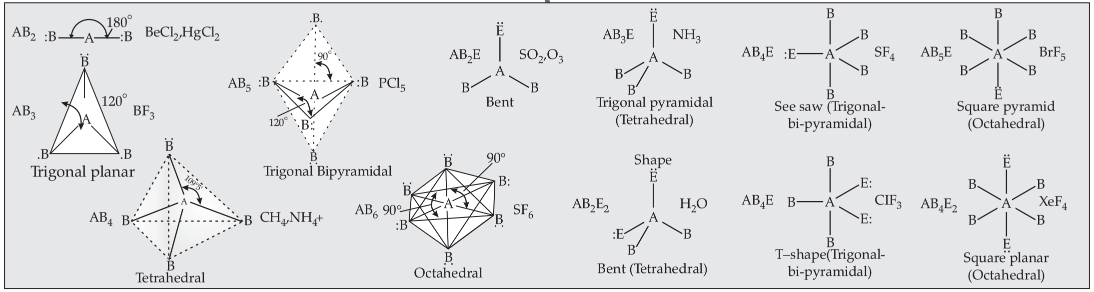
(iii) Bond Enthalpy: Amount of energy required to break one mole of bonds of particular type between 2 atoms.

(iv) Bond Order: Number of bonds between the two atoms of a molecule.

(v) Resonance Structures: Are a set of two or more Lewis structures that collectively describe the electronic bonding a single polyatomic species.

(vi) Dipole Moment: Product of the magnitude of the charge and distance between centres of positive and negative charge. u= Q x r

Hybridization



MIND MAP : LEARNING MADE SIMPLE — CHAPTER - 5

States Of Matter: Gases, Liquids & Solids

- **Intermolecular Forces are the forces of attraction and repulsion between interacting particles**
  - London dispersion forces are the weakest intermolecular forces present among non-polar atoms and molecules.
  - Dipole – dipole forces act between the molecules possessing permanent dipole.
  - Dipole – induced dipole forces are the attractive forces operate between polar molecules having permanent dipole.
  - Hydrogen bond is found in the molecules in which highly polar N–H, H–O and H–F bonds are present.
- **Thermal Energy is the energy of body arising from motion of its atoms or molecules**
- **Intermolecular forces vs thermal interaction**
  - Gas → Liquid → Solid: Predominance of intermolecular interactions
  - Gas ← Liquid ← Solid: Predominance of thermal energy
- **Gaseous State**
  - Physical Properties:
    - Gases are highly compressible.
    - Gases exert pressure equally in all directions.
    - Gases have much lower density than solids and liquids.
    - Volume and shape of gases are not fixed.
    - They mix evenly and completely in all proportions.
- **Gas Laws**
  - Boyle's Law (Pressure – Volume relationship): At constant temperature, the pressure of a fixed amount of gas varies inversely with its volume. $P_1V_1 = P_2V_2 = \text{Constant} \Rightarrow \frac{P_1}{P_2} = \frac{V_1}{V_2}$
  - Charle's Law (Temperature – Volume relationship) : At constant pressure, the volume of a fixed mass of a gas is directly proportional to the absolute temperature $\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{V}{T} = \text{Constant}$
  - Gay Lussac's Law (Pressure – Temperature relationship): At constant volume, pressure of a fixed amount of a gas varies directly with temperature. $\frac{P}{T} = \text{constant}$
  - Avogadro Law :(Volume – Amount relationship): Equal volume of all gases under the same conditions of temperature and pressure contain equal number of molecules. $V = kn;\ n = \frac{\text{Mass of gas}}{\text{Molar Mass}}$
- **Ideal Gas Equation**
  - $R = \frac{pV}{nT}$ R is gas constant / universal gas constant
  - $R = 8.314\ JK^{-1}\ mol^{-1}$
- **Viscosity**
  - $F = \eta A \frac{dx}{dz}$
  - Viscosity coefficient (η) is the force when velocity gradient is unity and the area of contact is unit area.
  - SI unit of viscosity coefficient is Newton second per square metre ($N\ s\ m^{-2}$) = Pa s (Pascal second)
- **Liquid State**
  - Intermolecular forces are stronger than gaseous state.
  - Boiling is the condition of free vapourisation throughout the liquid.
  - Normal b.p. is boiling point at 1 atm.
  - Standard b.p.is boiling point at 1 bar.
- **Liquifaction of Gases**
  - Volume of one mole of gas at critical temperature is called critical volume (V_C) and pressure at this temperature is called critical pressure (P_C)
- **Deviation from ideal gas behaviour**
  - Real gases show deviations from ideal gas law because molecules interact with each other.
  - Pressure exerted by the gas is lower than the pressure exerted by the ideal gas
  - $P_{ideal} = \underbrace{P_{real}}_{\text{Observed pressure}} + \underbrace{\frac{an^2}{V^2}}_{\text{Correction term}}$
  - Compressibility factor $Z = \frac{PV}{nRT}$
- **Postulates of Kinetic –Molecular Theory of gases**
  - A gas contains a large number of small particles called molecules. Size and mass of all molecules of each gas are identical.
  - There is no force of attraction between the particles of a gas at ordinary temperature and pressure.
  - Particles of a gas are always in constant and random motion.
  - Particles of a gas move in all possible directions in straight lines.
  - Collisions of gas molecules are perfectly elastic.
  - At any particular time, different particles in the gas have different speeds and hence different kinetic energies.
  - If a molecule has variable speed, then it must have a variable kinetic energy.
- **Dalton's Law of Partial Pressure**
  - Total pressure exerted by the mixture of non-reactive gases is equal to sum of the partial pressures of individual gases. $P_{total} = P_1 + P_2 + P_3$ .... (At constant T,V)
- **Density and molar mass at a gaseous substance**
  - $M = \frac{dRT}{\rho}$
- **Surface Tension**
  - is the force acting per unit length perpendicular to the line drawn on the surface of liquid.
  - Denoted by Gamma (γ) Unit: Nm^{-1}

MIND MAP : LEARNING MADE SIMPLE

CHAPTER - 6

Chemical Thermodynamics

Terms:

- Systems and surroundings: A system refers to that part of universe in which observations are made and remaining universe constitutes the surroundings.
- **Types of System:**

1. Open System: There is exchange of energy and matter between system and surroundings.
2. Closed System: There is no exchange of matter but exchange of energy is possible.
3. Isolated System: There is no exchange of energy or matter between the system and surroundings.
4. State of system : is described by its measurable or macroscopic properties.
5. Internal Energy : as a state function: It is the sum of chemical, electrical, mechanical or any other type of energy.
6. (a) Work:
Adiabatic process is in which there is no transfer of heat between system and surroundings.
$\Delta U = U_2 - U_1 = W_{ad}$
(b) Heat is change in internal energy of a system by transfer of heat from the surroundings to the system or vice-versa without expenditure of work. q is positive: Heat is transferred from surroundings to system.
q is negative: Heat is transferred from system to surroundings.
(c) General Case
First law of Thermodynamics: $\Delta U = q + w$ the energy of an isolated system is constant.

Applications

$\text{Work} = P_{ex}\Delta l = -P_{ex}(V_f - V_i) = \int_{V_i}^{V_f} P_{ex}\, dV$

Reversible Process: The process which can be reversed at any moment by an infinitesimal change.
Irreversible Process: Processes other than reversible process.
At constant temperature $W_{rev} = 2.303\, nRT \log \frac{V_f}{V_i}$
For adiabatic change, $q = 0$, $\Delta U = W_{ad}$

Enthalpy, H

ΔH is negative: Exothermic reactions.
ΔH is positive: Endothermic reactions.

Heat Capacity

$q = C \times m \times \Delta T = C\Delta T$

Entropy

$\Delta S_{Total} = \Delta S_{systems} + \Delta S_{surroundings}$
$\Delta = \frac{q_{rev}}{T}$

A spontaneous process is an irreversible process which may be reversed by some external agency

Lattice Enthalpy

is the enthalpy change when one mole of an ionic compound dissociates into its ions in gaseous state.

$\Delta_r G\theta = -RT \ln k$

Gibb's Energy

* $\Delta G = \Delta H - T\Delta S$
ΔG is <0, process is spontaneous.
* ΔG is >0, process is non-spontaneous

Enthalpy of Solutions

${}_{sol}H^{\theta}$ is the enthalpy change when one mole of a substance dissolves in a specified amount of solvent.

Hess's Law of Constant Heat Summation

If a reaction takes place in several steps then its standard reaction enthalpy is the sum of standard enthalpies of international reactions into which the overall reactions may be divided at the same temperature.

$\Delta_r H = \Delta_r H_1 + \Delta_r H_2 + \Delta_r H_3 \ldots\ldots\ldots$

$$\begin{array}{ccc} A & \xrightarrow{\Delta_r H_1} & C \\ {\scriptstyle \Delta_r H_1}\downarrow & & \uparrow{\scriptstyle \Delta_r H_3} \\ B & \xrightarrow{\Delta_r H_2} & D \end{array}$$

Enthalpy Change, $\Delta_r H$ – Reaction Enthalpy

$\Delta_r H = \sum a_i H_{products} - \sum_i b_i H_{reactants}$
(a) Standard Enthalpy of reactions is the enthalpy change for a reaction when all the participating substances are in their standard states.
(b) Enthalpy changes during phase transformations: Standard enthalpy of fusion / molar enthalpy of fusion, $\Delta_{fus}H^0$ is the enthalpy change that accompanies melting of one mole of a solid substance in standard state. Standard enthalpy of vaporization or molar enthalpy of vaporization. $\Delta_{vap}H^0$ is the amount of heat required to vaporize one mole of a liquid at constant temperature and under standard pressure.
(c) Standard molar enthalpy of formation $\Delta_f H^0$ is the standard enthalpy change for the formation of one mole of a compound from its elements in their most stable state of aggregation.

Measurement of ΔU and ΔH

In calorimetry, the process is carried out in a vessel called calorimeter, immersed in a known volume of a liquid.
(a) U measurements: The energy changes are measured at constant volume. No work is done.
(b) H measurements: In exothermic reaction, heat is evolved, so q_p and rH will be negative. In endothermic reaction, heat is absorbed, so q_p and rH will be positive.

Relationship Between C_p And C_v For An Ideal Gas

$C_p - C_v = R$

Intensive property

property whose value depends in the quantity or size of matter present. (Temperature density: Pressure).

Extensive Property

property which does not depend on the quantity or size of matter present in the system. Temperature density, pressure.

Standard Enthalpy of combustion

$\Delta_c H^{\theta}$ is the enthalpy change per mole of a substance when it undergoes combustion and all reactants and products are in standard state.

Enthalpy of atomization

$\Delta_a H^{\theta}$ is the enthalpy change on breaking one mole of bonds completely to obtain atoms in gas phase.

ΔrH^0

$\sum$bond enthalpies reactants – $\sum$bond enthalpies products

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 7

Equilibrium

Predicting the Extent of a Reaction
- $K_c < 10^{-3}$ reaction proceeds rarely.
- $K_c > 10^{3}$ reaction proceeds nearly to completion
- $K_c \rightarrow 10^{-3}$ to 10^{3}, reaction is at equilibrium

Acids and Bases
Acids turn blue litmus into red and liberate dihydrogen on reacting with some metals.
Bases turn red litmus into blue, taste bitter and feel soapy. Salts are formed when acids and bases are mixed in right proportion. Lewis Acid accepts an electron pair. Lewis Base donates an electron pair.
Acidic: $[H_3O^+] > [OH^-]$
Neutral $[H_3O^+] = [OH^-]$
Basic $[H_3O^+] < [OH^-]$

Factors Affecting Equilibra
- **Le Chatelier's principle:** A change in any of the factors that determine the equilibrium conditions of a system will cause the system to change in such a manner so as to counteract the effect of the change.
- **Effect of Concentration Change:** If concentration of reactants is increased, equilibrium will shift in forward direction and vice-versa.
- **Effect of Pressure Change:** No effect if moles of reactants and products are same. If there is change in number of moles , equilibrium will shift in the direction having smaller number of moles when the pressure is increased and vice-versa.
- **Effect of Inert Gas Addition:** If volume is constant, ie addition does not change partial pressure or molar concentrations, equilibrium is not affected.
- **Effect of Temperature change:** For exothermic, low temperature favours formation of reactants; for endothermic, high temperature favours formation of products.

Predicting the Direction of a Reaction
- If $Q_c < K_c$, net reaction goes from left to right.
- If $Q_c > K_c$, net reaction goes from right to left.
- If $Q_c = K_c$, no net reaction occurs

Calculating Equilibrium Concentrations
Step 1: Write the balanced equation for the reaction.
Step 2: Make a table that lists for each substance involved in reaction; Initial concentration, change in concentration, equilibrium concentration.
Step 3: Substitute the equilibrium concentrations into the equilibrium equation for the reaction and solve.
Step 4: Calculate equilibrium concentrations from the calculated value of concentration of substances.
Step 5: Check results by substituting them into the equilibrium equation.
Buffer Solutions are the solutions which reside change in pH on dilution or addition of acid or alkali.
Solubility Product (k_{sp}):
For a sparingly soluble salt, it is the product of molar concentration of ions raised to power equal to numbers of times each ion.

Under given set of conditions
if a chemical reaction is carried out in a closed container, a constancy in some observable properties like colour, pressure is observed. Such a state is referred as equilibrium state.

Equilibrium In Chemical Processes
Dynamic Equilibrium: Chemical reactions reach a state of dynamic equilibrium in which the rate of forward and reverse reactions are equal and there is no net change in composition.

Equilibrium Law or Law of Chemical Equilibrium
At a given temperature, the product of concentrations of the reaction products raised to the respective stoichiometric coefficient in the balanced chemical equation divided by the product of concentration of the reactions raised to their individual stoichiometric coefficients has a constant value: $aA + bB \rightleftharpoons cC + dD$

$$K_c = \frac{[C]^c + [D]^d}{[A]^a + [B]^b}$$

Relationship Between Equilibrium Constant K, Reaction Quotient Q And Gibbs Energy G
G is negative, reaction is spontaneous and proceeds in forward reaction.
G is positive, reaction is non-spontaneous and proceeds in reserve reaction.
G is 0, reaction has achieved equilibrium.
$K = e^{-\Delta G^\circ/RT}$
Arrhenius Concept of Acids and Bases: Arrhenius theory states that acids are substances that dissociates in water give hydrogen ions and bases are substances that produce hydroxyl ions.
Bronsted Lowry Theory states that acid is a substance that is capable of donating hydrogen ion and bases are substances capable of accepting a hydrogen ion.
•Weak acids have very strong conjugate bases.

Equilibrium Constant
for the reverse reaction is the inverse of the equilibrium constant for reaction in the forward direction.
Applications of Equilibrium Constants:
- Applicable only when concentrations of the reactants and products have attained constant value at equilibrium state.
- Value of K_c is independent of initial concentrations of the reactants and products.
- Equilibrium constant is temperature dependent.
  K_c reverse $= \frac{1}{K_c}$
- Equilibrium constant for a reacting is related to equilibrium constant of corresponding reaction.

Types

Homogeneous Equilibria
Reactants and products are in same phase.

Hetrogeneous Equilibria
Reactants and products are present in two or more phases.

Physical
Equilibrium set up in physical processes like evaporation of water, etc.
- Solid $\rightleftharpoons$ Liquid
- Liquid $\rightleftharpoons$ Gas
- Solid $\rightleftharpoons$ Gas

Characteristics:
- Possible only in a closed system at a given temperature.
- Both the reactions occur at same rate.
- All measurable property remains constant.
- When equilibrium is attained one of its parameter remains constant at given temperature.
- Magnitude of such quantities at any stage indicates the extent to which process has proceeded before reaching equilibrium.

Chemical
Speed of forward reaction equals that of backward reaction.
Characteristics:
- Can be achieved from either side.
- Dynamic in nature.
- There is no change in concentration of various species.
- Observable physical properties become constant.
- Equilibrium is attained rapidly by the presence of catalyst.

pH of a solution is negative logarithm to base 10 of the activity of hydrogen ion

Size increases
$HR << HCl << HB << HI \longrightarrow$
Size increases $\longrightarrow$
Acidic strength increases $\longrightarrow$

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 8

Redox Rections

- **Oxidation Number**
  - Oxidation state of an element in a compound ascertained according to a set of rules formulated on the basis that pair in a covalent bond belongs to electronegative element.
  - **Rules for O.N.**
    1. In elements, each atom bears an oxidation number of zero.
    2. For ions with only one atom, O.N. is equal to the charge on ion.
    3. O.N. of oxygen in most compounds is -2.
    4. O.N. of hydrogen is +1, except when it is bonded to metals in binary compound.
    5. Halogens have on O.N. of -1, when they occur as halide ions in their compound.
    6. Algebraic sum of O.N. of all the atoms in a compound must be zero.
- **Competitive Electron Transfer Reactions**

  $$\mathrm{Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)}$$

  Release of $2e^-$ (Zn → Zn^{2+}); Gain of $2e^-$ (Cu^{2+} → Cu)
- **Oxidation:** Addition of oxygen to an element/compound or removal of hydrogen from a substance
  - Loss of electrons
    - Reducing agent: Donor of electron(s) → Lowers O.N.
- **Reduction:** Removal of oxygen from a substance or addition of hydrogen to a substance
  - Gain of electrons
    - Oxidising agent: Acceptor of electron(s) → Increases O.N.
- **Types of Redox Reactions**
  1. Combination reactions: A+ B → C; Either A and B or both A and B must be in elemental form.
  2. Decomposition reactions: Leads to the breakdown of a compound into two or more components at least one of which must be in the elemental state $CaCO_3(s) \xrightarrow{\Delta} CaO(s)+CO_2(g)$
  3. Displacement reactions: An ion/ atom in a compound, is replaced by an ion/atom of another element, X + YZ → XZ +Y
     (a) Metal displacement: $CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$
     (b) Non-Metal displacement: Includes hydrogen displacement.
  4. Disproportionation reaction: An element in one oxidation state is simultaneously oxidised and reduced.
- **Application**
  - Study of electrode processes and cells.
- **Redox Reactions and Electrode Processes**
  - **Redox Couple:** Having together oxidised and reduced forms of a substance involved in oxidation/reduction half reaction.
  - **Electrode Potential:** Potential associated with each electrode.
  - **Standard Electrode Potential:** Electrode potential of concentration of each species is unity and carried out at 298K.
  - **Negative E°:** Stronger reducing agent than H^+/H_2 couple.
  - **Positive E°:** Weaker reducing agent than H^+/H_2 couple.
- **Balancing of Redox Reactions**
  - Two Methods
    - **O.N. Method**
      - **Step 1:** Write the correct formula for each reactant and product.
      - **Step 2:** Identify atoms undergoing change in O.N.
      - **Step 3:** Calculate increase / decrease in O.N per atom and for entire ion or molecule. If unequal, multiply by suitable number to make equal.
      - **Step 4:** Add H^+/OH^- ion to make total ionic charges of reactants and product equal.
      - **Step 5:** Equalize H^+ on two sides by adding water.
    - **Half-Reaction Method**
      - **Step 1:** Produce, unbalanced equation for the reaction in ionic form.
      - **Step 2:** Separate equation into two half-reactions.
      - **Step 3:** Individually balance atoms other than O and H.
      - **Step 4:** For reactions occurring in acidic medium, add H_2O to balance O atoms and H^+ to balance H atoms.
      - **Step 5:** Balance charges by adding electrons to one side of the half reaction.
      - **Step 6:** Add two half reactions and cancel the electrons on each side.
      - **Step 7:** Verify if equation has same type and number of atoms and same charges on both sides of the equation.

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 9

Hydrogen

Dihydrogen (H_2)

Occurrence: Most abundant in universe.

Preparation:

– Laboratory Method:

- $Zn+2H^+ \rightarrow Zn^{2+}+H_2$
- $Zn+2NaOH \rightarrow Na_2ZnO_2+H_2$

–**Commercial Method:**

- $2H_2O(l) \xrightarrow[\text{Traces of acid/base}]{\text{Electrolysis}} 2H_2(g)+O_2(g)$
- By electrolysing warm aqueous barium hydroxide solution between Nl electrodes.
- $C_nH_{2n+2}+nH_2O \xrightarrow[\text{Ni}]{1270\text{ K}} nCO+(2n+1)H_2$

–$CO+H_2$ is called water gas

–Cool Gasification : $C_{(s)}+H_2O_{(g)} \xrightarrow{1270\text{ K}} CO_{(g)}+H_{2(g)}$

–Water-gas shift reaction : $CO_{(g)}+H_2O_{(g)} \xrightarrow[\text{Catalyst}]{673\text{ K}} CO_{2(g)}+H_{2(g)}$

Physical Properties:

- Colourlers, odourless, tasteless and combustible.
- Lighter than air, insoluble in water.

Chemical Properties:

$H_{2(g)}+X_{2(g)} \longrightarrow 2HX_{(g)}$ (X=F, Cl,Br,I)

$2H_{2(g)}+O_{2(g)} \xrightarrow[\text{Heating}]{\text{Catalyst or}} 2H_2O(l)$ $\Delta H^\circ = -285.9$ KJmol^{-1}

$3H_{2(g)}+N_{2(g)} \xrightarrow[\text{Fe}]{673\text{ K, }200\text{ atm}} 2NH_{3(g)}$ $\Delta H^\circ = -92.6$ KJmol^{-1}

$H_{2(g)}+2M_{(g)} \longrightarrow 2MH_{3(s)}$ Where M= alkali metal

$H_{2(g)}+Pd^{2+}_{(aq)} \longrightarrow Pd_{(s)}+2H^+_{(aq)}$

$H_2+CO+RCH=CH_2 \longrightarrow RCH_2CH_2CHO$

$H_2+RCH_2CH_2CHO \longrightarrow RCH_2CH_2CH_2OH$

Uses:

- Synthesis of ammonia.
- In manufacture of vanaspati fat.
- In manufacture of bulk organic chemicals.
- For manufacture of metal hydrides
- Preparation of HCl.
- In metallurgical processes.
- As a rocket fuel
- In fuel cells

Heavy Water (D2O)

Preparation: By exhaustive electrolysis of water-

Uses:

- As a moderator in nuclear reactors.
- In exchange reactions for the study of reaction mechanisms.

Isotopes

Protium: Predominant form.($^1_1$H)

Deuterium: ($^2_1$H)

Tritium : Radioactive ($^3_1$H)

Properties

- Resembles alkali metals (lose one e⁻ to form unipositive ions).
- Resembles halogens (gain one e⁻ to form uninegative ion).
- Forms oxides, halides and sulphides.
- Very high ionization enthalpy.
- Does not possess metallic characters under normal conditions.
- Forms diatomic molecules.

Hydrides

Types:

(i) **Ionic/saline/saltlike:** Stoichiometric compounds of dihydrogen formed with s-block elements.

(ii) **Covalent/molecular:** Formation of molecular compounds from dihydrogen and p-block elements.

(iii) **Metallic/non-stoichiometric/interstitial:** Formed by d-block and f-block elements.

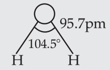
Water

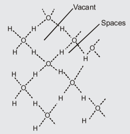
(Human body: 65%, Plants: 95%)

Physical Properties:

- Colourless and tasteless
- Presence of extensive H bonding.
- High freezing point, high B.P., high heat of vaporization, high heat of fusion

Structure:





Chemical Properties:

$H_2O_{(l)}+NH_{3(aq)} \rightleftharpoons OH^-_{(aq)}+NH_4^+{}_{(aq)}$

$2H_2O_{(l)}+2Na_{(s)} \longrightarrow 2NaOH_{(aq)}+H_{2(g)}$

$6CO_{2(g)}+12H_2O_{(l)} \longrightarrow C_6H_{12}O_{6(aq)}+6H_2O_{(l)}+6O_{2(g)}$

$P_4O_{10(s)}+6H_2O_{(l)} \longrightarrow 4H_3PO_{4(aq)}$

Difference: Hard water does not give lather with soap whereas soft water gives lather easily. Temporary hardness: Removed by boiling, Clark's method. (Due to presence of magnesium and calcium hydrogen carbonates)

Permanent hardness: Removed by treatment with washing soda, Calgon's method, Ion exchange method, synthetic resins methods. (Due to presence of soluble salts of Mg and Ca in the form of chlorides and sulphate)

Electronic Configuration

$1S^1$

Hydrogen Peroxide (H_2O_2)

$1S^1$

Preparation

- $BaO_2.8H_2O_{(s)}+H_2SO_{4(aq)} \rightarrow BaSO_{4(s)}+H_2O_{2(aq)}+8H_2O_{(l)}$
- $2HSO^-_{4(aq)} \xrightarrow{\text{Electrolysis}} HO_2SOOSO_3H_{(aq)} \longrightarrow 2HSO^-_{4(aq)}+2H^+_{(aq)}+H_2O_{2(aq)}$
- 2– ethylanthraquinol $\underset{H_2/Pd}{\overset{O_2(\text{air})}{\rightleftarrows}}$ H_2O_2+ Oxidised Product

Physical Properties

- Colourless, miscible with water.

Structure: Non-Planar

Chemical Properties:

$PbSO_{4(s)}+4H_2O_{2(aq)} \rightarrow PbSO_{4(s)}+4H_2O_{(l)}$

$HOCl+H_2O_2 \rightarrow H_3O^++Cl+O_2$

$Mn^{2+}+H_2O_2 \rightarrow Mn^{4+}+2OH^-$

$I_2+H_2O_2+2OH^- \rightarrow 2I^-+2H_2O+O_2$

Decomposes on exposure

Uses

- As hair bleach, disinfectant.
- Manufacture chemicals used in detergents.
- In industries as bleaching agent.
- In environmental chemistry.

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 10

S- Block Elements

Group 2 Elements (Alkaline Earth Metals)

- Atomic and Ionic Radii: Smaller than corresponding alkali in group,increases with increase in atomic number.
- I.E:IE_1 higher than corresponding Group 1 metals.
  IE_2 smaller than corresponding alkali metals.
- Hydration Enthalpies: Decreases with increase in ionic size down the group.

Physical properties:
- Slivery white, lustrous and relatively soft but harder than alkali metals.
- M.p. and b.p. higher than corresponding alkali metals.
- Electropositive character increases down the group.

Chemical properties:
- Be and Mg are kinetically inert to O and H_2O
- Mg is more electropositive and burns in air.
- Ca, Sr and Ba with air forms oxide and nitride.

$M + X_2 \longrightarrow MX_2 (X = F, Cl, Br, I)$

$2BeCl_2 + LiAlH_4 \longrightarrow 2BeH_2 + LiCl + AlCl_3$

$M + 2HCl \longrightarrow MCl_2 + H_2$

$M + (x+y)NH_3 \longrightarrow [M(NH_3)_x]^{2+} + 2[e(NH_3)_y]^-$

Uses:
- Be is used in the manufacture of alloys.
- Metallic Be is used for making windows of X–rays tubes.
- Mg-Al alloys are used in air craft construction.
- Ca in extraction of metals.
- Ra is used in radiotherapy.

Characteristics of Compounds of Alkaline Earth Metals:
- **Oxides and Hydroxides**
- Alkaline earth metals burn oxygen to form MO.
- All oxides except BeO are basic in nature

$MO + H_2O \longrightarrow M(OH)_2$

$Be(OH)_2$ is amphoteric in nature

Halides:
- Except for Be halides, all other halides are ionic.
- Tendency to form halide hydrates decreases gradually.
- Salts of oxoacids: Forms carbonates, sulphates and nitrates.
- Anomalous behavior of Be: Small atomic and ionic sizes, does not exhibit C.N. more than four, its oxide and hydroxide are amphoteric.
- Be shows diagonal relationship with Al.

Biological Importance Of Mg And Ca:
- All enzymes that utilise ATP in PO_4 transfer requires Mg as cofactor. Chlorophyll contains Mg. Ca is present in bones and teeth. Important in neuromuscular function, intraneuronal transmission and blood coagulation.

Important Compounds of Calcium

(i) CaO, Quick Lime

Preparations: $CaCO_3 \overset{heat}{\rightleftharpoons} CaO + CO_2$

Properties : White amorphous solid with m.p. 2870 K

$CaO + H_2O \longrightarrow Ca(OH)_2$, $CaO + CO_2 \longrightarrow CaCO_3$

(ii)$Ca(OH)_2$ Calcium hydroxide : Preparations: Additional of water to CaO.

Properties: white amorphous powder .

$Ca(OH_2) + CO_2 \longrightarrow CaCO_3 + H_2O$

(iii)$CaSO_4$. ½H_2O(Plaster of Paris)

$2(CaSO_4.2H_2O) \longrightarrow 2(CaSO_4).H_2O + 3H_2O$

Group 1 Elements (Alkali Metals)

- **Atomic and Ionic Radii:** Increases with increase in atomic number.
- **I.E.:** Decreases down the group.
- **Hydration Enthalpy:** Decreases with increase in ionic sizes.
- **Physical properties**:
- Silvery white, soft and light metals.
- Low m.p. and b.p.
- Alkali metals and their salts impart colour to an oxidizing flame.

Chemical Properties:

$4Li + O_2 \rightarrow 2Li_2O$; $2Na + O_2 \rightarrow Na_2O_2$; $M + O_2 \rightarrow MO_2$ (M=K, Rb, Cs)

$2M + 2H_2O \rightarrow 2M^+ + 2OH^- + H_2$

$2M + H_2 \rightarrow 2M^+H^-$

React vigorously with halogens to form ionic halides

$M + (x + y) NH_3 \rightarrow [M(NH_3)_x]^+ + [e(NH_3)_y]^-$

Uses:
- Li is used to make useful alloys.
- Li is used in thermonuclear reactions and making electrochemical cells.
- Na is used to make Na/Pb alloy.
- Liquid Na metal is used as coolant in nuclear reactors.
- KCl is used as fertilizer.
- Cs is used in devising photoelectric cells.

Characteristics Of Compounds Of Alkali Metals:

– Oxide and Hydroxides

Combustion in excess of air $\xrightarrow{+Li} Li_2O$; $\xrightarrow{+Na} Na_2O_2$; $\xrightarrow{+K/Rb/Cs} MO_2$

- Halides
- Alkali metal halides (MX) have high melting, colourless crystalline solids.
- Preparation: Reaction of oxide, hydroxide or carbonate with aq HX.
- High negative enthalpies of formation.
- Melting and boiling points: F >Cl> Br > I
- Soluble in water.

Salts of Oxo-Acids:
- Alkali metals form salts with all oxo-acids.
- Soluble in water and thermally stable.
- Stability of carbonates and hydrogencarbonates increases

Anomalies properties of Li: Due to

(i) exceptionally small size of its atom and ion.

(ii) High polarising power.

Biological Importance of Na and K:

Na ions participate in nuclear signals transmission, regulator of flow of water across cell membranes. K ions activate many enzymes and oxidation of glucose to produce ATP.

Electronic Configuration

ns^1 : Alkaline metals; ns^2 : Alkaline earth metals

Important Compounds of Sodium:

(i)Sodium Carbonate (Washing Soda): Preparation: By Solvay process $2NH_3 + H_2O + CO_2 \rightarrow (NH_4)_2CO_3$

$(NH_4)_2CO_3 + H_2O + CO_2 \rightarrow 2NH_4HCO_3$

$NH_4HCO_3 + NaCl \rightarrow NH_4Cl + NaHCO_3$

$2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

$2NH_4Cl + Ca(OH)_2 \rightarrow 2NH_3 + CaCl_2 + H_2O$

Properties: (a)White, crystalline solid. (b)Readily soluble in water

$Na_2CO_3 \cdot 10H_2O \xrightarrow{375K} Na_2CO_3 \times H_2O + 9H_2O$

$Na_2CO_3 \cdot H_2O \xrightarrow{>373K} Na_2CO_3 + H_2O$

Uses: Water softening, laundering, cleaning, manufacture, as laboratory reagent.

(i)Sodium Chloride *(Nacl)*

Preparation: Crude NaCl by crystallization of brine solution.

-Pure NaCl is obtained by dissolving crude salt in minimum water and filtered to remove insoluble impurities. Solution is saturated with HCl gas. Uses: As common salt.

(ii)Sodium Hydroxide (NaOH)

Preparation: By electrolysis of NaCl in Castner-Kellner cel.

Uses: In manufacture of soap, paper, petroleum refining.

(iii)Sodium Hydrogencarbonate ($NaHCO_3$)

Preparation: $Na_2CO_3 + H_2O + CO_2 \rightarrow 2NaHCO_3$

MIND MAP : LEARNING MADE SIMPLE

CHAPTER - 11

p - Block Elements

Group 14 Elements

[C, Si, Ge, Sn, Pb]

Electronic configuration: $ns^2 np^2$

Covalent Radius: Increase from C to Si and from Si to Pb a small increase is observed.

I.E.: Decreases down the group.

Electro negativity: More than group 13 elements. From Si to Pb same.

Physical Properties:

- Solid.
- C and Si are non- metals, Ge metalloid, Sn and Pb are soft metals.
- M.p. and b.p. higher than group 13 elements.

Chemical Properties:

- Oxidation State: +4, +2
- Forms oxides on heating in oxygen.
- C, Si, Ge, and Pb are not affected by water.

$Sn + 2H_2O \xrightarrow{\Delta} SnO_2 + 2H_2$

- Form halides of formula MX_2 and MX_4

Important Trends and Anomalous Behavior of Carbon:

- C differs from rest due to smaller size, higher electronegativity, high ionisation enthalpy and unavailability of d orbitals.
- C forms $p\pi$–$p\pi$ multiple bonds
- Shows Catenation

Allotropes of Carbon:

- Diamond
- Carbon
- Fullerenes

Uses of Carbon:

- Graphite composites are used in tennis rackets, fishing rods, aircrafts and canoes.
- Graphite is used for electrodes in batteries and industrial electrolysis.
- C black is used in black ink and as filter in automobile tyres.
- Diamond is used in jewellry.

Important Compounds of C and Si

- Carbon monoxide:
  $2C + O_2 \xrightarrow[\text{Limited air}]{\Delta} 2CO$
  $HCOOH \xrightarrow[\text{Conc.} H_2SO_4]{373K} H_2O + CO$
- Silicones: Group of organosilicon polymers with (R_2SiO) replanting unit.
- Carbon Dioxide:
  $C + O_2 \xrightarrow{\Delta} CO_2$
  $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$
- Silicates Structural 4- unit is SiO_4^{-4}
- Silicon Dioxide: Si is covalently bonded to 4 oxygen atoms.
- Zeolites: used as catalyst in petrochemical industries.

Water Gas

Or Synthesis Gas: Mixture of $CO+H_2$

Electronic configuration

ns^2np^{1-6} (except He)

Producer Gas

Mixture of $CO+N_2$

Group 13 Elements

[B, Al, Ga, In, Tl]

Electronic configuration: ns^2np^1

Atomic Radii: Atomic radius of Ga is less than Al.

I.E.: Decrease from B to Al is associated with increase in size. $\Delta_iH_1 < \Delta_iH_2 < \Delta_iH_3$

Electronegativity: Down the group, decreases from B to Al and then increases marginally.

Physical Properties:

- Non-metallic, hard and black coloured solid.
- Exists in many allotropic forms.
- B has high m.p. while others have low m.p.
- Density increases down the group.

Chemical Properties:

·Oxidation State: +1, +3

$2E_{(s)} + 3O_{2(g)} \xrightarrow{\Delta} 2E_2O_{3(s)}$

$2E_{(s)} + N_{2(g)} \xrightarrow{\Delta} 2EN_{(s)}$

$2Al_{(s)} + 6HCl_{(aq)} \longrightarrow 2Al^{3+}{}_{(aq)} + 6Cl^-_{(aq)} + 3H_{2(g)}$

$2E_{(s)} + 3X_{2(g)} \longrightarrow 2EX_{3(s)}$ $(X = F, Cl, Br, I)$

Trends and Anomalous Properties of Boron:

- Except in boron, $[M(OH)_4]^-$ and $[M(H_2O)_6]^{3+}$ exist in aqueous medium.
- Boron trifluoride reacts with Lewis bases.
- Maximum covalence of B is 4.

Important Compounds of Boron:

(i) Borax$(Na_2B_4O_7 . 10H_2O)$White crystalline solid

$Na_2B_4O_7 + 7H_2O \longrightarrow 2NaOH + 4H_3BO_3$

$Na_2B_4O_7.10H_2O \xrightarrow{\Delta} Na_2B_4O_7 \xrightarrow{\Delta} 2NaBO_2 + B_2O_3$

Borax bead test is used for identification.

(ii) Orthoboric acid (H_3BO_3): White crystalline solid, soapy touch

Preparation Properties:$Na_2B_4O_7 + 2HCl + 5H_2O \longrightarrow 2NaCl + 4B(OH)_3$

$B(OH)_3 + 2HOH \longrightarrow [B(OH)_4]^- + H_3O^+$

$H_3BO_3 \xrightarrow{\Delta} HBO_2 \xrightarrow{\Delta} B_2O_3$

(iii)Diborane (B_2H_6)

Preparation: $4BF_3 + 3LiAlH_4 \longrightarrow 2B_2H_6 + 3LiF + 3AlF_3$

$2NaBH_4 + I_2 \longrightarrow B_2H_6 + 2NaI + H_2$

$2BF_3 + 6NaH \xrightarrow{450K} B_2H_6 + 6NaF$

Properties: Colourless, toxic, b.p.: 180K, catches fire

$B_2H_6 + 3O_2 \longrightarrow B_2O_3 + 3H_2O \quad \Delta_CH^\theta = -1976 \text{ KJ mol}^{-1}$

$B_2H_6(g) + 6H_2O(l) \longrightarrow 2B(OH)_{3(aq)} + 6H_{2(g)}$

$B_2H_6 + 2NMe_3 \longrightarrow 2BH_3.NMe_3$

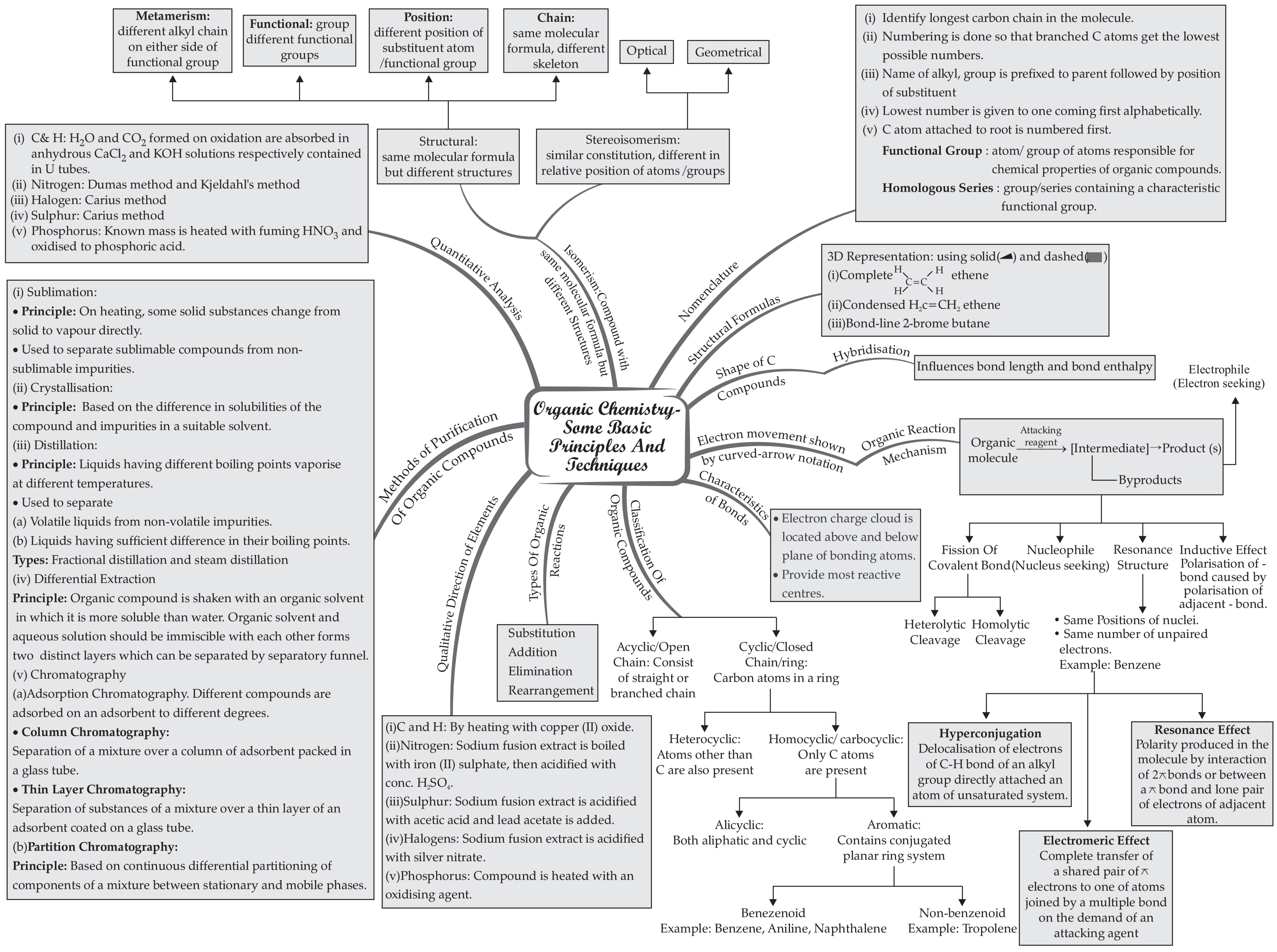
$B_2H_6 + 2CO \longrightarrow 2BH_3CO$

$3B_2H_6 + 6NH_3 \longrightarrow 3[BH_2(NH_3)_2]^+[BH_4]^- \xrightarrow{\Delta} 2B_3N_3H_6 + 12H_2$

Uses of Boron and Aluminium and their Compounds:

- Boron fibers are used in making bullet-proof vest and light material for aircraft.
- Borax is used in manufacturing heat resistant glasses.
- Al is used in packing, utensil making, construction and transportation industry.

MIND MAP : LEARNING MADE SIMPLE CHAPTER - 12



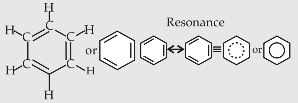
MIND MAP : LEARNING MADE SIMPLE CHAPTER - 13

Hydrocarbons

Compounds of Carbon and Hydrogen

Aromatic Hydrocarbon

- Types: Benzenoids – contain benzene ring, Non-benzenoids – does not contain benzene ring.
- Isomerism: Ortho (o-), Meta (m-), Para (p-) Resonance and
- Structure:



- Aromaticity: Planarity, complete delocalisation of the TT–electros in the ring, presence of (4n + 2) π electrons in the ring where n is an integer (n = 0,1,2,......) (Huckel rule)

Preparation:

- Cyclic polymerisation of ethyne

$C_6H_5COONa + NaOH \xrightarrow[\Delta]{CaO} C_6H_6 + Na_2CO_3$

$C_6H_5OH + Zn \xrightarrow{\Delta} C_6H_6 + ZnO$

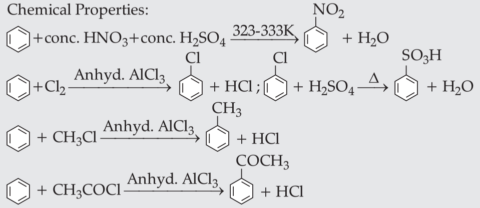
Mechanism of Electrophilic substitution reactions

(i) Generation of Electrophile
(ii) Formation of carbocation intermediate
(iii) Removal of proton

Physical Properties:

- Non-polar, usually colourless liquids or solids with characteristic aroma.
- Immiscible with water but miscible with organic solvents.
- Burns with sooty flame.

Chemical Properties:

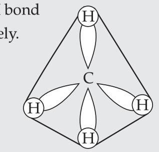


Classification

- **Saturated:** Contain C-C and C-H single bonds. (alkanes)
- **Unsaturated:** Contain C-C multiple bonds.
- **Aromatic**: Contain cyclic compounds

Alkene (C_nH_{2n+2})

- H-C-H bond angles – 190.5, C-C and C-H bond lengths are 154 pm and 112 pm respectively.
- Shows structural and chain isomerism.



Preparation :

$CH_2 = CH_2 + H_2 \xrightarrow{pt/pd/Ni} CH_3 - CH_3$

$CH_3Cl + H_2 \xrightarrow{zn,H^+} CH_4 + HCl$

Wurtz reaction :

$CH_3Br + 2Na + BrCH_3 \xrightarrow{Dryether} CH_3 - CH_3 + 2NaBr$

$CH_3COO^-Na^+ + NaOH \xrightarrow[\Delta]{CaO} CH_4 + Na_2CO_3$

$2CH_3COONa + 2H_2O \longrightarrow C_2H_6 + 2CO_2 + H_2 + 2NaOH$

- Physical Properties:

Non-polar, weak van der Waals forces, colourless, odourless.
B.P. increases with increases in molecular size.

Chemical Properties:

$CH_4 \xrightarrow[hv/-HCl]{+Cl_2} CH_3Cl + \xrightarrow[hv/-HCl]{+Cl_2} CH_2Cl_2 \xrightarrow[hv/-HCl]{+Cl_2} CHCl_3 \xrightarrow[hv/-HCl]{+Cl_2}$

$C_nH_{2n+2} + \left(\frac{3n+1}{2}\right)O_2 \longrightarrow nCO_2 + (n+1)H_2O$

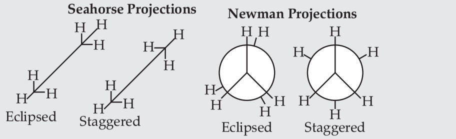
$2CH_4 + O_2 \xrightarrow{Cu/523K/100atm} 2CH_3OH$

$CH_4 + O_2 \xrightarrow[\Delta]{Mn_2O_3} HCHO + H_2O$

$CH_4(CH_2)CH_3 \xrightarrow{Anhyd.AlCl_3/HCl} CH_3-\underset{CH_3}{CH}-(CH_2)_2-CH_3 + CH_3-CH_2-\underset{CH_3}{CH}-CH_2-CH_3$

$CH_4 + H_2O \xrightarrow[\Delta]{Ni} CO + 3H_2$

$C_6H_{14} \xrightarrow{733\,K} C_6H_{12} + H_2$; $C_uH_8 + C_2H_6$; $C_3H_6 + C_2H_4 + CH_4$



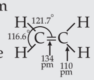
Alkenes (C_2H_{2n})

- Shows structural and geometrical isomerism

Preparation:

$RC \equiv CR' + H_2 \xrightarrow{Pd/C}$ R(H)C=C(H)R' cis-Alkene

$RC \equiv CR' + H_2 \xrightarrow{Na/Liquid\ NH_3}$ R(H)C=C(H)R' trans-Alkene



$H_3C - CH_2X \xrightarrow[\Delta]{Alc.KOH} H_2C = CH_2$

$CH_2Br - CH_2Br + Zn \longrightarrow CH_2 = CH_2 + ZnBr_2$

$H_3CCH_2OH \xrightarrow[\Delta]{conc.H_2SO_4} CH_2 = CH_2 + H_2O$

- Physical Properties :
  - Ethene is a colourless gas with faint sweet smell.
  - All other are colourless and odourless, insoluble in water but fairly soluble in non-polar solvents.
  - Increase in b.p. with increased molecular size.
- Chemical Properties:

$CH_2 = CH_2 + Br_2 \xrightarrow{CCl_4} BrCH_2 - CH_2Br$

$CH_2 = CH_2 + HBr \longrightarrow CH_3CH_2Br$

$CH_3\text{-}CH{=}CH_2 + HBr \xrightarrow{Markovnikov\ rule}$ $CH_3\text{-}\underset{Br}{CH}\text{-}CH_3$; $CH_3\text{-}CH_2\text{-}CH_2\text{-}Br$

$H_3C\text{-}\underset{CH_3}{C}{=}CH_2 + H_2O \xrightarrow{H^+}$ $(CH_3)(CH_3)C(CH_3)(OH)$

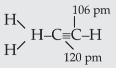
$CH_3 - CH = CH - CH_3 \xrightarrow{KMnO4/H^+} 2CH_3COOH$

$n(CH_2 = CH_2) \xrightarrow[Catalyst]{HighTemp/Pressure} \text{-}(CH_2 - CH_2)\text{-}_n$

$n(CH_3 - CH = CH_2) \xrightarrow[Catalyst]{HighTemp/Pressure} \text{-}(CH - CH_2)\text{-}_n$

Alkenes (C_nH_{2n-2})

- IUPAC name: replacing 'ane' by the suffix 'yne'.
- Shows position and chain isomerism



Preparation:

$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$

$CaO + 3C \longrightarrow CaC_2 + CO$

$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H_2$

$CH_2Br - CH_2Br + KOH \xrightarrow[-KBr,\,-H_2O]{alcohol} H_2C = CHBr \xrightarrow[-NaBr,\,-NH_3]{NaNH_2} CH \equiv CH$

Physical Properties:

- First three members are gases, next eight are liquids and higher ones are solids.
- Colourless, ethyne has characteristic odour and other are odourless.
- Lighter than water, immiscible with water but soluble in organic solvents.
- m.p, b.p. and density increase with increase in molar mass.

Chemical Properties:

$HC \equiv CH + Na \longrightarrow HC \equiv CNa + 1/2\,H_2$

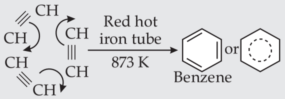
$HC \equiv CH + H_2 \xrightarrow{Pt/Pd/Ni} [H_2C = CH_2] \xrightarrow{H_2} CH_3 - CH_3$

$CH_3 - C \equiv CH + Br_2 \longrightarrow [CH_3CBr = CHBr] \xrightarrow{Br_2} CH_3\text{-}CBr_2\text{-}CHBr_2$

$HC \equiv C - H + HBr \longrightarrow [CH_2C = CH - Br] \longrightarrow \underset{CH_2}{CHBr_2}$

$HC \equiv CH + H_2O \xrightarrow[333\,K]{Hg^{2+}/H^+} [CH_2 = \underset{OH}{C} - H] \xrightarrow{Isomersation} CH_3 - \underset{O}{\overset{\|}{C}} - H$

Polymerisation



MIND MAP : LEARNING MADE SIMPLE CHAPTER - 14

Environmental Chemistry

Deals with the study of origin, transport, reactions, effects and fates of chemical species in the environment

- Troposphere: ~10 Km from sea level
- Stratosphere: Between 10 and 50 km above sea level

Soil Pollution

- •Pesticides (DDT)
- •Herbicidies
- •Industrial Waste:

(i)Biodegradable wastes are generated by cotton mills, food processing units, paper mills and textile factories.
(ii)Non-biodegradable wastes are generated by thermal power plants, integrated iron and steel plants, fertilizer industries, etc.
(iii)Control:
(i)Use of manures
(ii)Use of bio–fertilizers
(iii)Proper sewage system
(iv) Salvage and recycling

Green Chemistry

It is a way of thinking and utilizing the existing knowledge and principles of chemistry and other sciences to reduces the adverse impact in environment.

•In Daily life:
(i)Dry cleaning of clothes.
(ii)Bleaching of paper
(iii) Synthesis of chemicals

Strategies to Control

(i) Waste Management
(ii) Collection and Disposal

Biochemical Oxygen Demand (BOD)

Amount of oxygen required by bacteria to break down the organic matter present in a certain volume of a sample of water.

Greenhouse Effect

It is a naturally occurring phenomenon responsible for heating of earth's surface and atmosphere due to presence of certain gases in the atmosphere.

Global warming is the gradual and continuos increase in average temperature of surface of the earth due to increase in concentrations of greenhouse gases (CO_2, CFCs, CH_4)

Environmental Pollution

Effect of undesirable changes in our surroundings that have harmful effects on plants, animals and human beings)

Acid Rain (when pH of the rain water drops below 5.6.)

Water Pollution

Causes of water pollution:
(i)Pathogens which cause diseases.
(Escherichia coli, Streptococcus faecalis)
(ii)Organic Wastes: (Leaves, grass, trash)
(iii)Chemical Pollutants: Heavy metals, acids from mine drainage, salts from different sources, petroleum products, fertilizers and industrial chemicals.

Eutrophication: Process in which nutrient enriched water bodies support a dense plant population which kills animals life and loss of biodiversity.

·International standards for Drinking water
(i)Fluoride: Excess fluoride (over 10 ppm) causes harmful effects to bones and teeth.
(ii)Lead: Prescribed upper limit in drinking Water is about 50ppm.
(iii)Sulphate: >500ppm causes laxative effect.
(iv)Nitrate: Maximum limit in water is 50ppm.
(v)Other metals: Fe, Mn, Al, Cu, Zn, Cd.

Atmospheric Pollution

Stratospheric Pollution:

Ozone present in upper stratosphere protects us from cancer causing harmful uv radiations.
Depletion of Ozone layer:

$$Cl\dot{O}_{(g)} + NO_{2(g)} \longrightarrow ClONO_{2(g)}$$

$$\dot{C}l_{(g)} + CH_{4(g)} \longrightarrow \dot{C}H_{3(g)} + HCl_{(g)}$$

$$ClONO_{2(g)} + H_2O_{(g)} \longrightarrow HOCl_{(g)} + HNO_{3(g)}$$

$$ClONO_{2(g)} + HCl_{(g)} \longrightarrow Cl_{2(g)} + HNO_{3(g)}$$

$$HOCl_{(g)} \xrightarrow{hv} \dot{O}H_{(g)} + \dot{C}l_{(g)}$$

$$Cl_{2(g)} \xrightarrow{hv} 2\dot{C}l_{(g)}$$

Tropospheric Pollution

Occurs due to the presence of undesirable solid or gaseous particles in the air.

Pollutants presents in Troposphere:
(1) Gaseous air pollutants
(a)Oxides of sulphur produced when sulphur containing fossil fuel is burnt. It causes respiratory diseases.
(b)Oxides of Nitrogen: At high altitudes when lightning strikes, dinitrogen and dioxygen combine to form oxides of hydrogen. In an automobile engine, when fossil fuel is burnt, dinitrogen and dioxygen combine to form NO and NO_2
(c)Hydrocarbons: Formed by incomplete combustion of fuel used in automobiles. They are carcinogenic and harm plants also.
(d)Oxides of Carbon: CO, CO_2
(2) Particulate pollutants: (Minute solid particles or liquid droplets in air present in vehicle emissions, smoke particles from fires, dust particles and ash from industries).
(a) Smog: Smoke + Fog

Types of smog:
(1)Classical smog occurs in cool humid climate. It is a mixture of smoke, fog and sulphur dioxide.
(2)Photochemical smog occurs in warm, dry and sunny climate. It has high concentration of oxidising agents. It contains ozone, nitric oxide, acrolain, formaldehyde and PAN. It causes serious health problems.