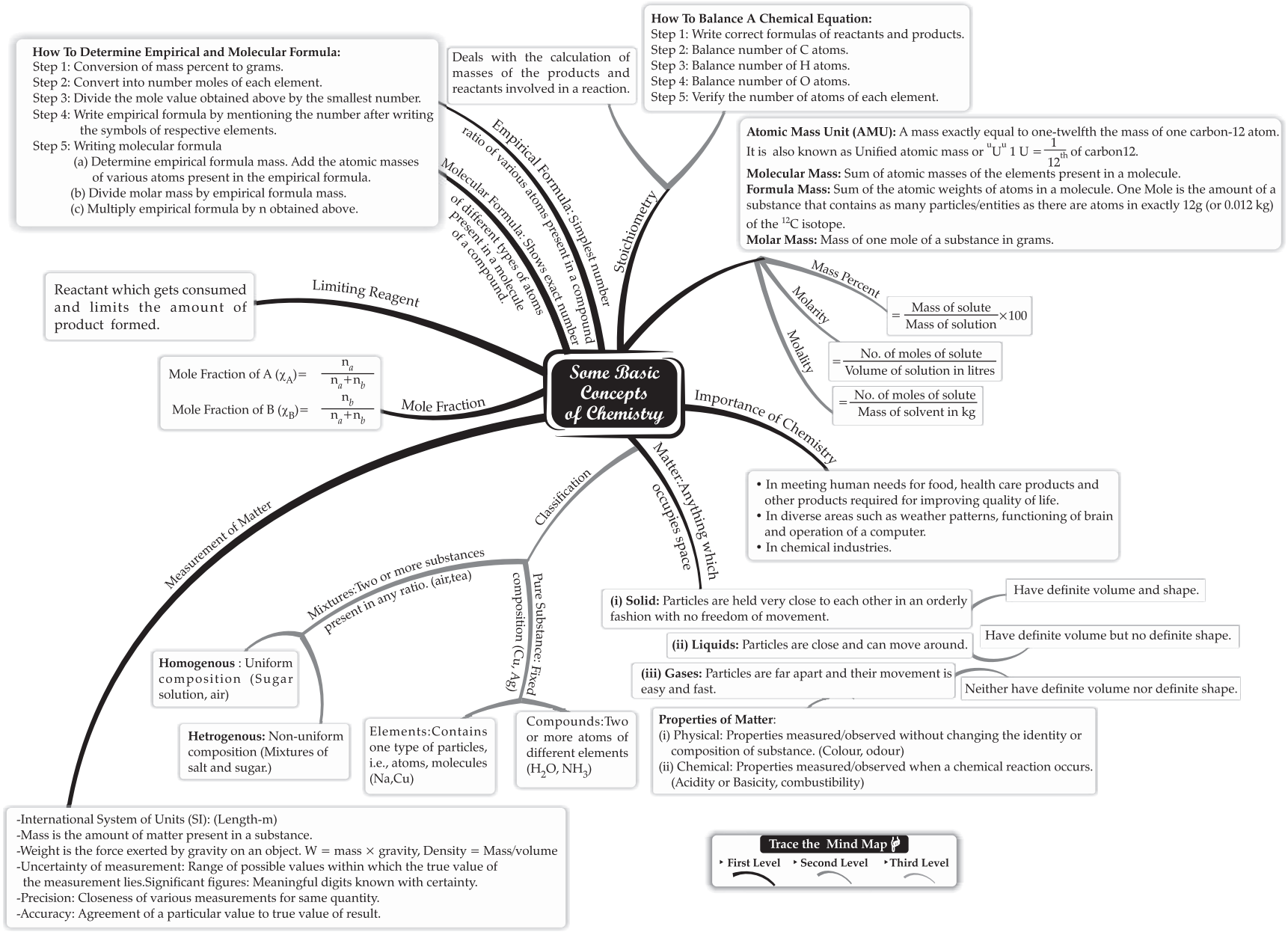

Some Basic Concepts of Chemistry
Empirical Formula: Simplest number ratio of various atoms present in a compound
Molecular Formula: Shows exact number of different types of atoms present in a molecule of a compound.
Stoichiometry
Deals with the calculation of masses of the products and reactants involved in a reaction.
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.
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 (AMU): A mass exactly equal to one-twelfth the mass of one carbon-12 atom. It is also known as Unified atomic mass or uU 1 U = 1/12th of carbon12.
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 12C isotope.
Molar Mass: Mass of one mole of a substance in grams.
Mass Percent
= Mass of solute / Mass of solution ×100
Molarity
= No. of moles of solute / Volume of solution in litres
Molality
= No. of moles of solute / Mass of solvent in kg
Limiting Reagent
Reactant which gets consumed and limits the amount of product formed.
Mole Fraction
Mole Fraction of A (χA)= nA / (nA+nb)
Mole Fraction of B (χB)= nb / (nA+nb)
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)
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 (H2O, NH3)
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. W = mass × gravity, Density = Mass/volume
-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.
Trace the Mind Map
First Level
Second Level
Third Level

Structure of Atom

Mass Number (A) — Number of protons (Z) + number of neutrons (*n*)

Atomic Number (Z) — Number of protons in nucleus of an atom or Number of electrons in a neutral atom

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)

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

Discovery of Electron: — **By J.J.Thomson Charge to mass ratio of electron** = 1.758820×10^{11} C kg^{-1}; Charge an electron = 1.6022×10^{-19}C; Mass of electron = 9.1094×10^{-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 Chadwick** 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 i.e., 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.

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.
- $\nu = \frac{\Delta E}{n} = \frac{E_2 - E_1}{n}$ Bohr's frequency rule.

Angular momentum of electron: $m_e v_r = n\frac{h}{2p}$ 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.

Developments Leading To Bohr's Model of Atom:

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.

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

Dual nature i.e., 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 × Wavelength

Wave number ($\bar{\nu}$): Number of wavelengths per unit length. Unit – m^{-1}

- **Particle nature of electromagnetic radiation:** Planck's quantum theory: E= hn; Planck's constant (h) = 6.626×10^{-34} Js

Quantum Mechanical Model of Atom:

Schrodinger Equation

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

$\Delta x.\ \Delta P_x \geq = \frac{h}{4p}$

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

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

Quantum Numbers

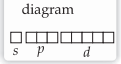
(i) Principal quantum number (n) :n = 1,2,3,4…….. Shell = K, L, M, N.....
(ii) Azimuthal Quantum number : For given value of n , *l* = 0 to n-1
(iii) Magnetic Quantum number (m) : for subshell with '*l*' value m_l=2l+1
(iv) Spin Quantum Number (m_s): $+1/2$ (↑), $-1/2$ (↓)

Filling of orbitals in atom

- **Aufbau Principle:** In the ground state of atoms, the orbital's are filled according to increasing energies.
- **Pauli Exclusion Principle:** No two electrons in an atom can have same set of four quantum numbers
- **Hund's rule**: Pairing of electrons in the orbital belonging to same subshell does not take place until each electron belonging to the subshell is singly occupied.

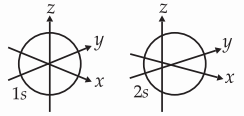
Electronic Configuration of atoms

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

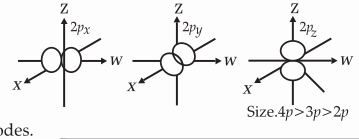


Shapes of orbitals:

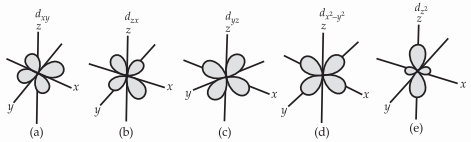
***s* Orbitals:** Spherically symmetric Size increases with increase in size in n, i.e., 4s>3s>2s>1s



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



***d* Orbitals** – ''Clover leaf'' distribution. – Two angular nodes.



Energy of orbitals: Lower the value of $(n+l)$ for an orbital, lower is its energy.

Trace the Mind Map

First Level · Second Level · Third Level

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 – are numbered from 1 to 7 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.

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 gaseous 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 outermost orbit of an isolated atom in gaseous state.	Decreases	Increases
(d) Electronegativity : Tendency of an atom to attract the shared pair of electrons towards itself.	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

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

f -Block: Elements

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

d- Block: Elements

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

p- Block elements

- 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.

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.

Trace the Mind Map

First Level · Second Level · Third Level

Chemical Bonding and Molecular Structure

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 transfer of electrons from one atom to another acquire the stable nearest noble gas configuration.

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

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

Lattice Enthalpy

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), p (Pi), d (Delta)

Molecular orbital theory

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 attains noble gas electronic configuration.
- 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 i.e. 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 substances.
(ii) Intramolecular : H atom is between two highly electronegative atom of same molecule..

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:**

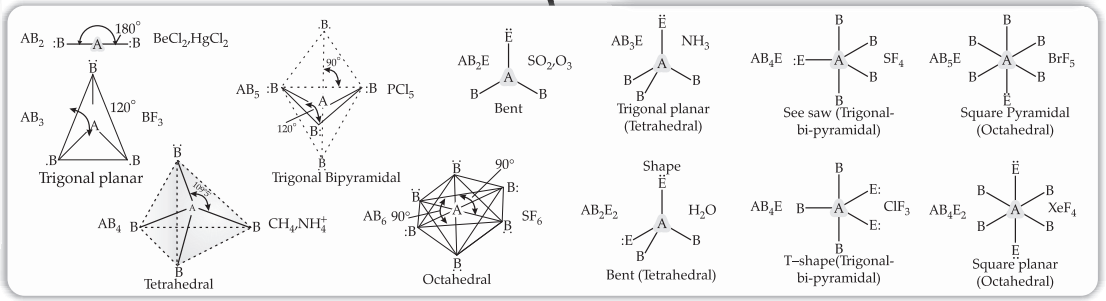
lp – lp > lp – bp > bp – bp

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 (π) 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: 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. μ= Q x r

Hybridization



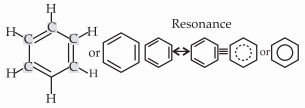
Trace the Mind Map
First Level Second Level Third Level

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-)
- Structure:



- Aromaticity: Planarity, complete delocalisation of the π–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

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

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

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:

$$\text{C}_6\text{H}_6 + \text{conc. } HNO_3 + \text{conc. } H_2SO_4 \xrightarrow{323\text{-}333K} C_6H_5NO_2 + H_2O$$

$$C_6H_6 + Cl_2 \xrightarrow{\text{Anhyd. } AlCl_3} C_6H_5Cl + HCl \; ; \; C_6H_5Cl + H_2SO_4 \xrightarrow{\Delta} C_6H_4Cl(SO_3H) + H_2O$$

$$C_6H_6 + CH_3Cl \xrightarrow{\text{Anhyd. } AlCl_3} C_6H_5CH_3 + HCl$$

$$C_6H_6 + CH_3COCl \xrightarrow{\text{Anhyd. } AlCl_3} C_6H_5COCH_3 + HCl$$

Classification

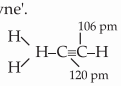
- **Saturated:** Contain C-C and C-H single bonds. (alkanes)
- **Unsaturated:** Contain C-C multiple bonds (alkenes, alkynes)
- **Aromatic**: Contain cyclic compounds

Mechanism of Electrophilic substitution reactions

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

Alkynes (C_nH_{2n-2})

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



Preparation:

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

$$CaO + 3C \longrightarrow CaC_2 + CO$$

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

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

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:

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

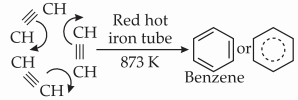
$$HC \equiv CH + H_2 \xrightarrow{Pt/Pd/Ni} [H_2C=CH_2] \xrightarrow{H_2} CH_3-CH_3$$

$$CH_3-C \equiv CH + Br_2 \longrightarrow [CH_3CBr=CHBr] \xrightarrow{Br_2} CH_3-\underset{Br}{\overset{Br}{C}}-\underset{Br}{\overset{Br}{CH}}$$

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

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

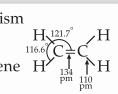
Polymerisation



Alkenes (C_2H_{2n})

- Shows structural and geometrical isomerism
- Preparation:

$$RC \equiv CR' + H_2 \xrightarrow{Pd/C} \underset{H}{\overset{R}{}}C=C\underset{H}{\overset{R'}{}} \quad \textit{cis}\text{-Alkene}$$



$$RC \equiv CR' + H_2 \xrightarrow{Na/\text{Liquid } NH_3} \underset{H}{\overset{R}{}}C=C\underset{R'}{\overset{H}{}} \quad \textit{trans}\text{-Alkene}$$

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

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

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

- 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 increase in molecular size.
- Chemical Properties:

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

$$CH_2=CH_2 + HBr \longrightarrow CH_3CH_2Br$$

$$CH_3-CH=CH_2 + HBr \xrightarrow[\text{rule}]{Markovnikov} \begin{cases} CH_3-\underset{Br}{CH}-CH_3 \\ CH_3-CH_2-CH_2-Br \end{cases}$$

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

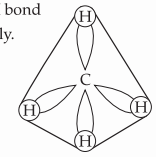
$$CH_3-CH=CH-CH_3 \xrightarrow{KMnO_4/H^+} 2CH_3COOH$$

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

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

Alkane (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 :

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

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

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:

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

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

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

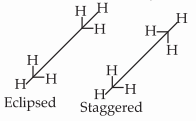
$$CH_4 + O_2 \xrightarrow[\Delta]{Mn_2O_3} HCHO + H_2O$$

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

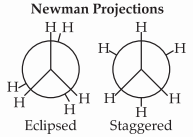
$$CH_4 + H_2O \xrightarrow[\Delta]{Ni} CO + 3H_2$$

$$C_6H_{14} \xrightarrow{733\ K} \begin{cases} C_6H_{12} + H_2 \\ C_4H_8 + C_2H_6 \\ C_3H_6 + C_2H_4 + CH_4 \end{cases}$$

Seahorse Projections

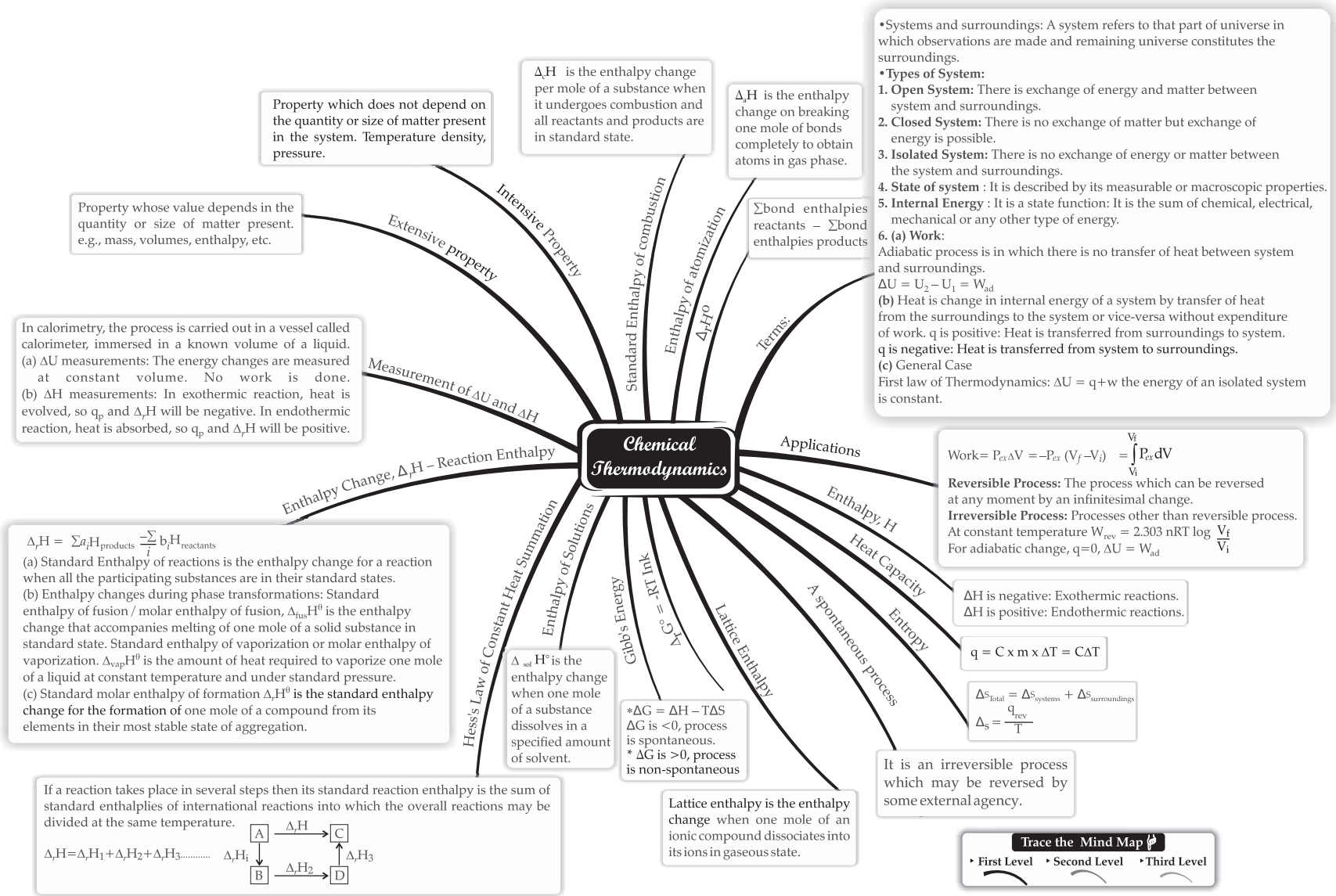


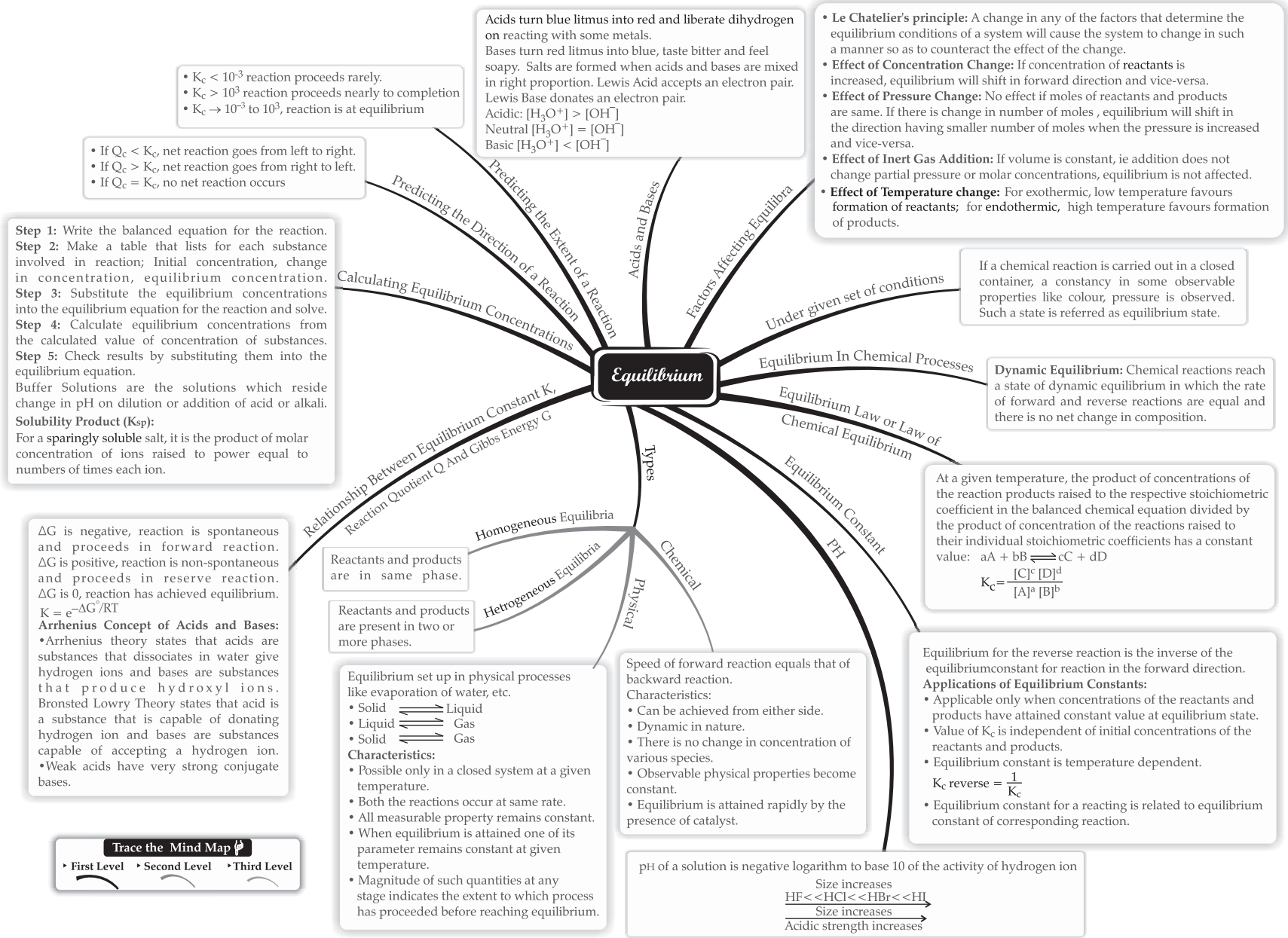
Newman Projections

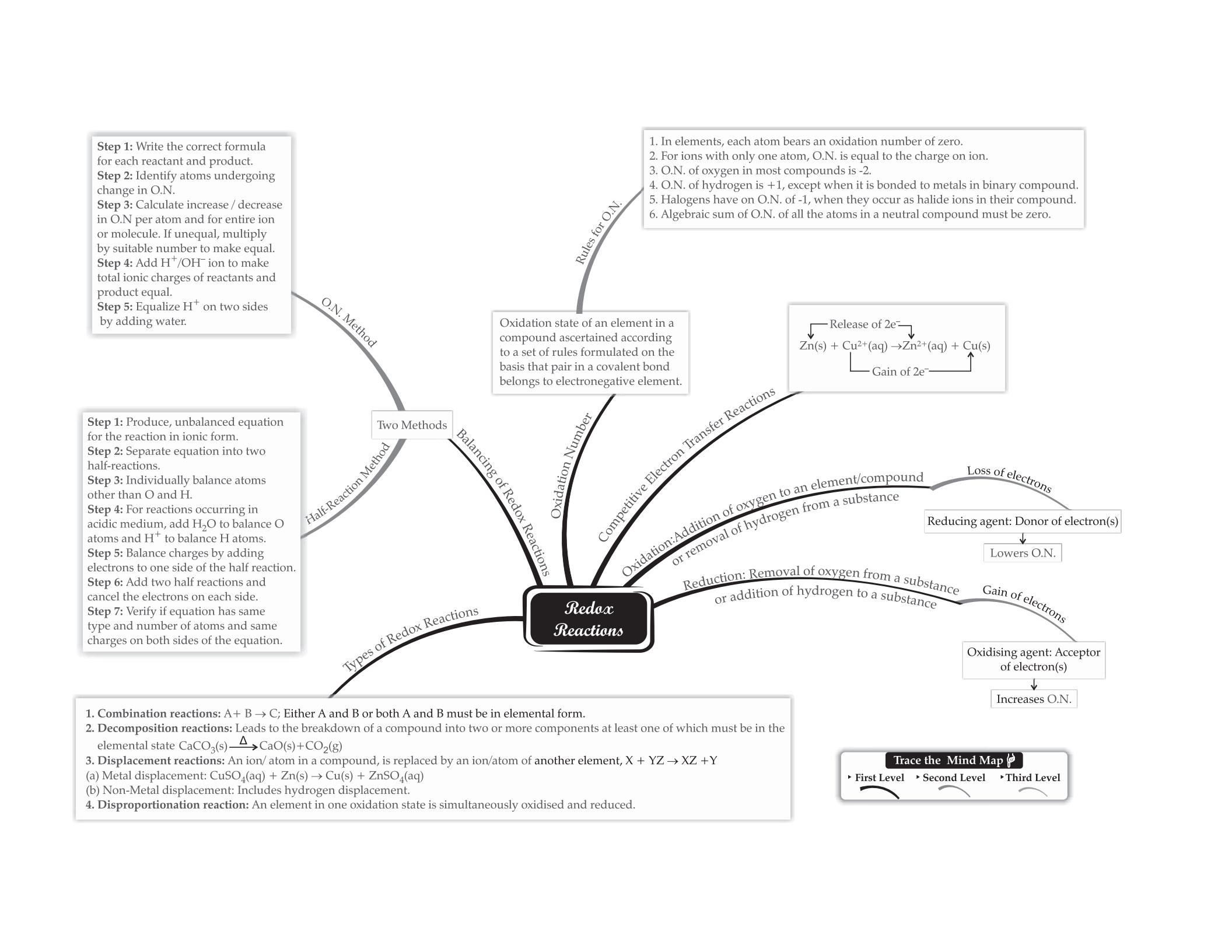


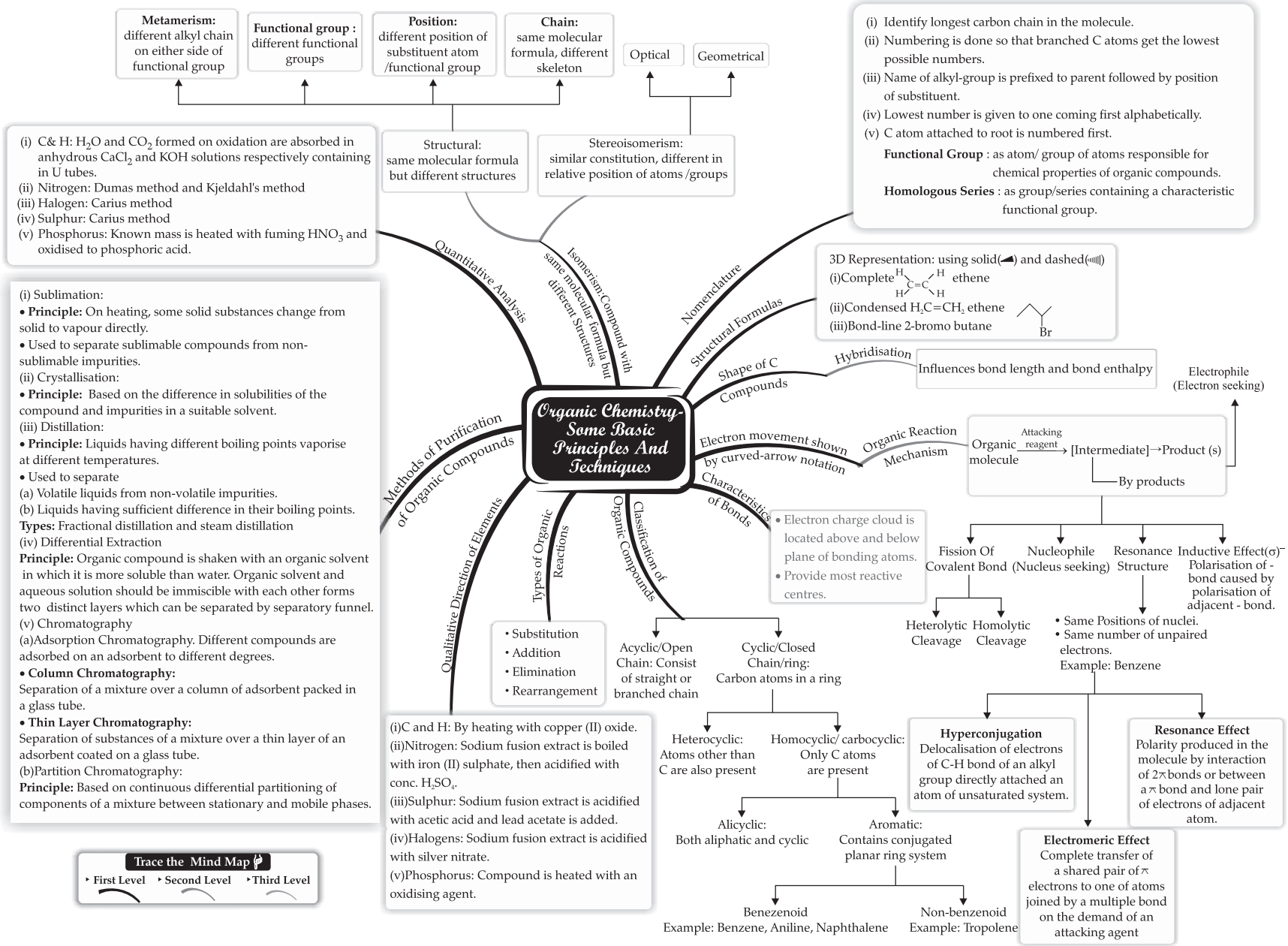
Trace the Mind Map

First Level · Second Level · Third Level


Chemical Thermodynamics
Intensive Property
Property which does not depend on the quantity or size of matter present in the system. Temperature density, pressure.
Extensive property
Property whose value depends in the quantity or size of matter present. e.g., mass, volumes, enthalpy, etc.
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) ΔU measurements: The energy changes are measured at constant volume. No work is done.
(b) ΔH measurements: In exothermic reaction, heat is evolved, so q_p and $\Delta_r H$ will be negative. In endothermic reaction, heat is absorbed, so q_p and $\Delta_r H$ will be positive.
Enthalpy Change, $\Delta_r H$ – Reaction Enthalpy
$\Delta_r H = \sum a_i H_{products} - \sum b_i H_{reactants}$
(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, $\Delta_{fus}H^\theta$ 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. $\Delta_{vap}H^\theta$ 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 $\Delta_f H^\theta$ is the standard enthalpy change for the formation of one mole of a compound from its elements in their most stable state of aggregation.
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.
$\Delta_r H = \Delta_r H_1 + \Delta_r H_2 + \Delta_r H_3 .....$
A, B, C, D, $\Delta_r H$, $\Delta_r H_i$, $\Delta_r H_2$, $\Delta_r H_3$
Enthalpy of Solutions
$\Delta_{sol} H^\circ$ is the enthalpy change when one mole of a substance dissolves in a specified amount of solvent.
Gibb's Energy
$\Delta G^\circ = -RT \ln K$
*ΔG = ΔH – TΔS
ΔG is <0, process is spontaneous.
* ΔG is >0, process is non-spontaneous
Lattice Enthalpy
Lattice enthalpy is the enthalpy change when one mole of an ionic compound dissociates into its ions in gaseous state.
A spontaneous process
It is an irreversible process which may be reversed by some external agency.
Entropy
$\Delta S_{Total} = \Delta S_{systems} + \Delta S_{surroundings}$
$\Delta_S = \frac{q_{rev}}{T}$
Heat Capacity
q = C x m x ΔT = CΔT
Enthalpy, H
ΔH is negative: Exothermic reactions.
ΔH is positive: Endothermic reactions.
Applications
Work= $P_{ex}\Delta V = -P_{ex}(V_f - V_i) = \int_{V_i}^{V_f} P_{ex} dV$
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, ΔU = W_{ad}
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 : It is described by its measurable or macroscopic properties.
5. Internal Energy : It is 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: ΔU = q+w the energy of an isolated system is constant.
Standard Enthalpy of combustion
$\Delta_c H$ 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
$\Delta_a H$ is the enthalpy change on breaking one mole of bonds completely to obtain atoms in gas phase.
$\Delta_r H^o$
∑bond enthalpies reactants – ∑bond enthalpies products
Trace the Mind Map
First Level
Second Level
Third Level


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
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.
Relationship Between Equilibrium Constant K, Reaction Quotient Q And Gibbs Energy G
ΔG is negative, reaction is spontaneous and proceeds in forward reaction.
ΔG is positive, reaction is non-spontaneous and proceeds in reserve reaction.
ΔG 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.
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 ⇌ Liquid
• Liquid ⇌ Gas
• Solid ⇌ 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.
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.
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: $aA + bB \rightleftharpoons cC + dD$
$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$
Equilibrium Constant
Equilibrium for the reverse reaction is the inverse of the equilibriumconstant 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.
PH
pH of a solution is negative logarithm to base 10 of the activity of hydrogen ion
Size increases
HF<<HCl<<HBr<<HI
Size increases
Acidic strength increases
Trace the Mind Map
First Level
Second Level
Third Level


Redox Reactions
Types of Redox Reactions
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 H2O 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.
Oxidation Number
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 neutral compound must be zero.
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.
Competitive Electron Transfer Reactions
Release of 2e−
Zn(s) + Cu2+(aq) →Zn2+(aq) + Cu(s)
Gain of 2e−
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.
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 CaCO3(s) —Δ→ CaO(s)+CO2(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: CuSO4(aq) + Zn(s) → Cu(s) + ZnSO4(aq)
(b) Non-Metal displacement: Includes hydrogen displacement.
4. Disproportionation reaction: An element in one oxidation state is simultaneously oxidised and reduced.
Trace the Mind Map
First Level
Second Level
Third Level


Organic Chemistry- Some Basic Principles And Techniques
Quantitative Analysis
(i) C& H: H_2O and CO_2 formed on oxidation are absorbed in anhydrous $CaCl_2$ and KOH solutions respectively containing in U tubes.
(ii) Nitrogen: Dumas method and Kjeldahl's method
(iii) Halogen: Carius method
(iv) Sulphur: Carius method
(v) Phosphorus: Known mass is heated with fuming HNO_3 and oxidised to phosphoric acid.
Isomerism: Compound with same molecular formula but different Structures
Structural: same molecular formula but different structures
Stereoisomerism: similar constitution, different in relative position of atoms /groups
Metamerism: different alkyl chain on either side of functional group
Functional group : different functional groups
Position: different position of substituent atom /functional group
Chain: same molecular formula, different skeleton
Optical
Geometrical
Nomenclature
(i) Identify longest carbon chain in the molecule.
(ii) Numbering is done so that branched C atoms get the lowest possible numbers.
(iii) Name of alkyl-group is prefixed to parent followed by position of substituent.
(iv) Lowest number is given to one coming first alphabetically.
(v) C atom attached to root is numbered first.
Functional Group : as atom/ group of atoms responsible for chemical properties of organic compounds.
Homologous Series : as group/series containing a characteristic functional group.
Structural Formulas
3D Representation: using solid(▲) and dashed(▲)
(i)Complete ethene
(ii)Condensed $H_2C=CH_2$ ethene
(iii)Bond-line 2-bromo butane
Br
Shape of C Compounds
Hybridisation
Influences bond length and bond enthalpy
Electron movement shown by curved-arrow notation
Organic Reaction Mechanism
Organic molecule → [Intermediate] → Product (s)
Attacking reagent
By products
Electrophile (Electron seeking)
Characteristics of Bonds
• Electron charge cloud is located above and below plane of bonding atoms.
• Provide most reactive centres.
Fission Of Covalent Bond
Heterolytic Cleavage
Homolytic Cleavage
Nucleophile (Nucleus seeking)
Resonance Structure
• Same Positions of nuclei.
• Same number of unpaired electrons.
Example: Benzene
Inductive Effect(σ)
Polarisation of σ-bond caused by polarisation of adjacent σ-bond.
Hyperconjugation
Delocalisation of electrons of C-H bond of an alkyl group directly attached an atom of unsaturated system.
Resonance Effect
Polarity produced in the molecule by interaction of 2π bonds or between a π bond and lone pair of electrons of adjacent atom.
Electromeric Effect
Complete transfer of a shared pair of π electrons to one of atoms joined by a multiple bond on the demand of an attacking agent
Methods of Purification of Organic Compounds
(i) Sublimation:
• Principle: On heating, some solid substances change from solid to vapour directly.
• Used to separate sublimable compounds from non-sublimable impurities.
(ii) Crystallisation:
• Principle: Based on the difference in solubilities of the compound and impurities in a suitable solvent.
(iii) Distillation:
• Principle: Liquids having different boiling points vaporise at different temperatures.
• Used to separate
(a) Volatile liquids from non-volatile impurities.
(b) Liquids having sufficient difference in their boiling points.
Types: Fractional distillation and steam distillation
(iv) Differential Extraction
Principle: Organic compound is shaken with an organic solvent in which it is more soluble than water. Organic solvent and aqueous solution should be immiscible with each other forms two distinct layers which can be separated by separatory funnel.
(v) Chromatography
(a)Adsorption Chromatography. Different compounds are adsorbed on an adsorbent to different degrees.
• Column Chromatography:
Separation of a mixture over a column of adsorbent packed in a glass tube.
• Thin Layer Chromatography:
Separation of substances of a mixture over a thin layer of an adsorbent coated on a glass tube.
(b)Partition Chromatography:
Principle: Based on continuous differential partitioning of components of a mixture between stationary and mobile phases.
Qualitative Direction of Elements
(i)C and H: By heating with copper (II) oxide.
(ii)Nitrogen: Sodium fusion extract is boiled with iron (II) sulphate, then acidified with conc. H_2SO_4.
(iii)Sulphur: Sodium fusion extract is acidified with acetic acid and lead acetate is added.
(iv)Halogens: Sodium fusion extract is acidified with silver nitrate.
(v)Phosphorus: Compound is heated with an oxidising agent.
Types of Organic Reactions
• Substitution
• Addition
• Elimination
• Rearrangement
Organic Compounds
Classification of
Acyclic/Open Chain: Consist of straight or branched chain
Cyclic/Closed Chain/ring: Carbon atoms in a ring
Heterocyclic: Atoms other than C are also present
Homocyclic/ carbocyclic: Only C atoms are present
Alicyclic: Both aliphatic and cyclic
Aromatic: Contains conjugated planar ring system
Benezenoid
Example: Benzene, Aniline, Naphthalene
Non-benzenoid
Example: Tropolene
Trace the Mind Map
First Level
Second Level
Third Level