SIMPLIFIED STUDY MATERIAL (2024-25)

CLASS 11 CHEMISTRY (PM SHRI SCHOOLS)

SR. N	CONTENT	PAGE NO.
1.	**CH-1 SOME BASIC CONCEPTS OF CHEMISTRY**	**2-4**
2.	**CH-2 STRUCTURE OF ATOM**	**5-6**
3.	**CH-3 CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES**	**7-9**
4.	**CH-4 CHEMICAL BONDING AND MOLECULAR STRUCTURE**	**9-12**
5.	**CH-5 (CHEMICAL THERMODYNAMICS)**	**12-14**
6.	**CH-6 EQUILIBRIUM**	**14-19**
7.	**CH-7 REDOX REACTION**	**20-21**
8.	**CH-8 ORGANIC COMPOUNDS**	**22-24**
9.	**CH-9 HYDROCARBONS**	**25-26**
10.	**SAMPLE PAPER**	**27-29**

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Chapter 1 - Some Basic Concepts of Chemistry

Important Definitions:

Scientific Notation: A way of expressing very large or very small numbers using powers of ten.

Significant Figures: The digits in a measurement that carry meaningful information, including all non-zero digits, any zeros between them, and any trailing zeros in the decimal portion. For example, in 0.00456, the significant figures are 4, 5, and 6.

Law of Conservation of Mass: It states that mass is neither created nor destroyed in a chemical reaction. The total mass of reactants equals the total mass of products.

Law of Definite Proportions: This law states that a given chemical compound always contains the same proportion of elements by mass, regardless of the amount or source of the compound.

Law of Multiple Proportions: This law states that when two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

Gay-Lussac's Law of Gaseous Volumes: This law states that when gases react together at constant temperature and pressure, the volumes of the reacting gases and the volumes of the products (if gaseous) are in simple whole number ratios.

Avogadro's Law: This law states that equal volumes of gases, at the same temperature and pressure, contain an equal number of molecules, regardless of the type of gas.

Stoichiometry: The branch of chemistry that involves the calculation of reactants and products in chemical reactions, based on the conservation of mass and the law of definite proportions.

Limiting Reagent: The reactant in a chemical reaction that is completely consumed first, limiting the amount of product formed. It determines the maximum amount of product that can be produced.

Mole: One mole contains exactly 6.022×10^{23} particles (atoms, molecules, ions, etc.), a number known as Avogadro’s number.

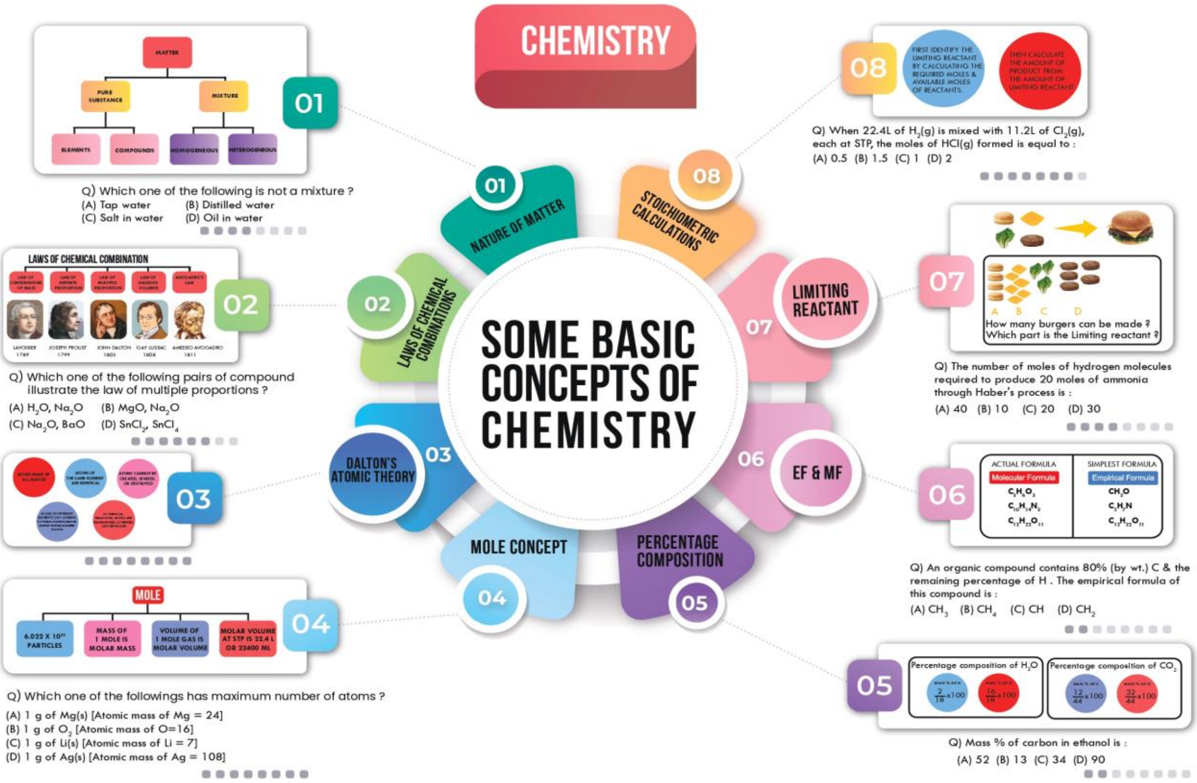
Mole Fraction: The ratio of the number of moles of a component to the total number of moles of all components in a mixture. It is a way of expressing concentration in a solution.

Molarity: A unit of concentration, defined as the number of moles of solute per liter of solution. It is commonly used in laboratory work to express solution concentration.

Molality: A unit of concentration, defined as the number of moles of solute per kilogram of solvent. Molality is often used in cases where temperature variations might affect the volume of a solution.

Molecular Mass: The mass of a molecule, calculated by adding the atomic masses of all the atoms in the molecule.

Formula Mass: The mass of a compound or ionic compound, calculated by adding the atomic masses of the atoms in its empirical formula.



01

Q) Which one of the following is not a mixture ?
(A) Tap water (B) Distilled water
(C) Salt in water (D) Oil in water

02

Q) Which one of the following pairs of compound illustrate the law of multiple proportions ?
(A) H_2O, Na_2O (B) MgO, Na_2O
(C) Na_2O, BaO (D) $SnCl_2$, $SnCl_4$

03

04

Q) Which one of the followings has maximum number of atoms ?
(A) 1 g of Mg(s) [Atomic mass of Mg = 24]
(B) 1 g of O_2 [Atomic mass of O=16]
(C) 1 g of Li(s) [Atomic mass of Li = 7]
(D) 1 g of Ag(s) [Atomic mass of Ag = 108]

05

$\frac{2}{18}\times100$ $\frac{16}{18}\times100$ $\frac{12}{44}\times100$ $\frac{32}{44}\times100$

Q) Mass % of carbon in ethanol is :
(A) 52 (B) 13 (C) 34 (D) 90

06

ACTUAL FORMULA Molecular Formula	SIMPLEST FORMULA Empirical Formula
$C_3H_6O_3$	CH_2O
$C_{10}H_{14}N_2$	C_5H_7N
$C_{12}H_{22}O_{11}$	$C_{12}H_{22}O_{11}$

Q) An organic compound contains 80% (by wt.) C & the remaining percentage of H . The empirical formula of this compound is :
(A) CH_3 (B) CH_4 (C) CH (D) CH_2

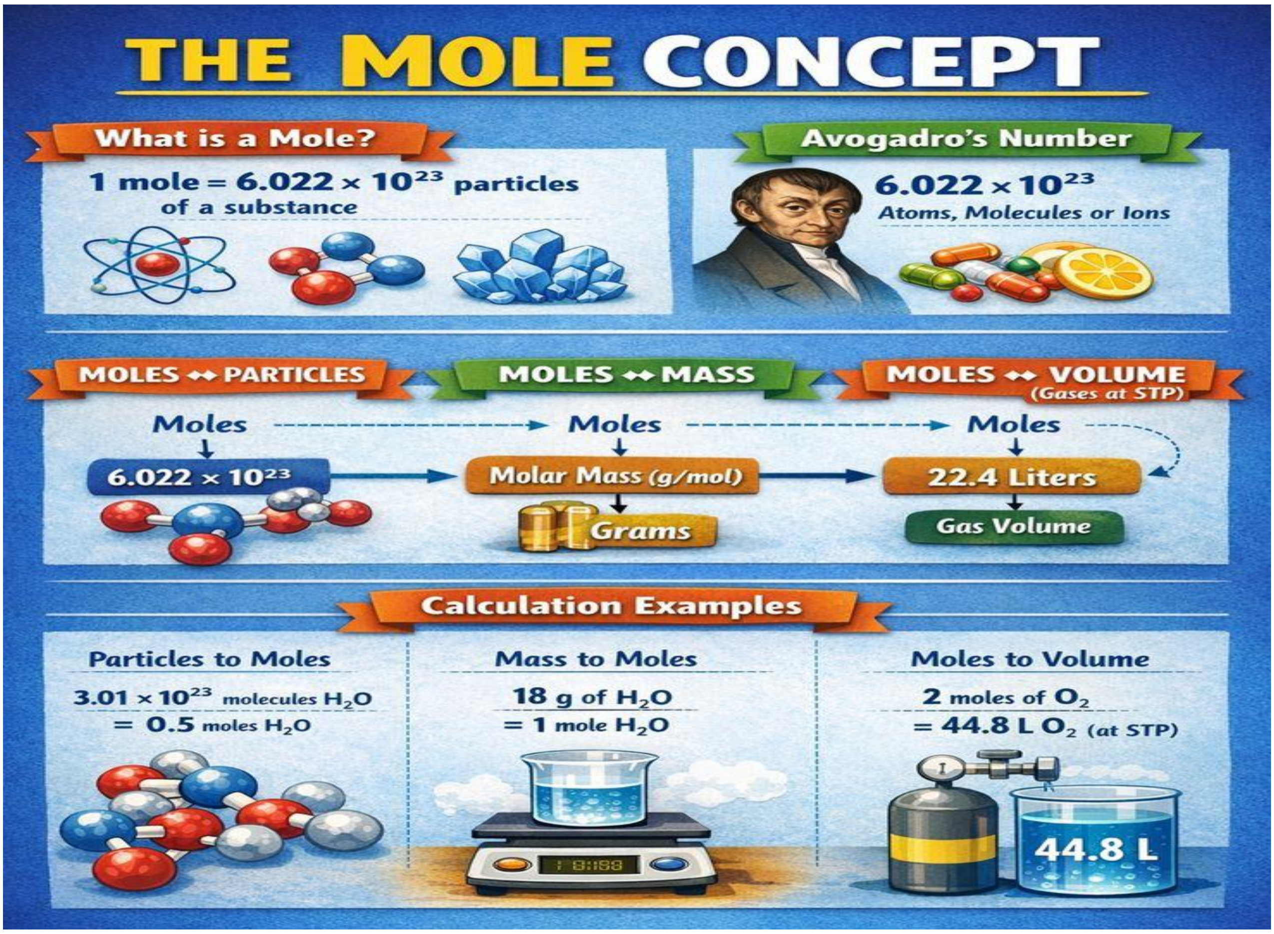
07

How many burgers can be made ?
Which part is the Limiting reactant ?

Q) The number of moles of hydrogen molecules required to produce 20 moles of ammonia through Haber's process is :
(A) 40 (B) 10 (C) 20 (D) 30

08

Q) When 22.4L of $H_2(g)$ is mixed with 11.2L of $Cl_2(g)$, each at STP, the moles of HCl(g) formed is equal to :
(A) 0.5 (B) 1.5 (C) 1 (D) 2



THE MOLE CONCEPT

What is a Mole?

1 mole = 6.022×10^{23} particles of a substance

Avogadro's Number

6.022×10^{23}
Atoms, Molecules or Ions

MOLES ↔ PARTICLES

MOLES ↔ MASS

MOLES ↔ VOLUME (Gases at STP)

Calculation Examples

Particles to Moles

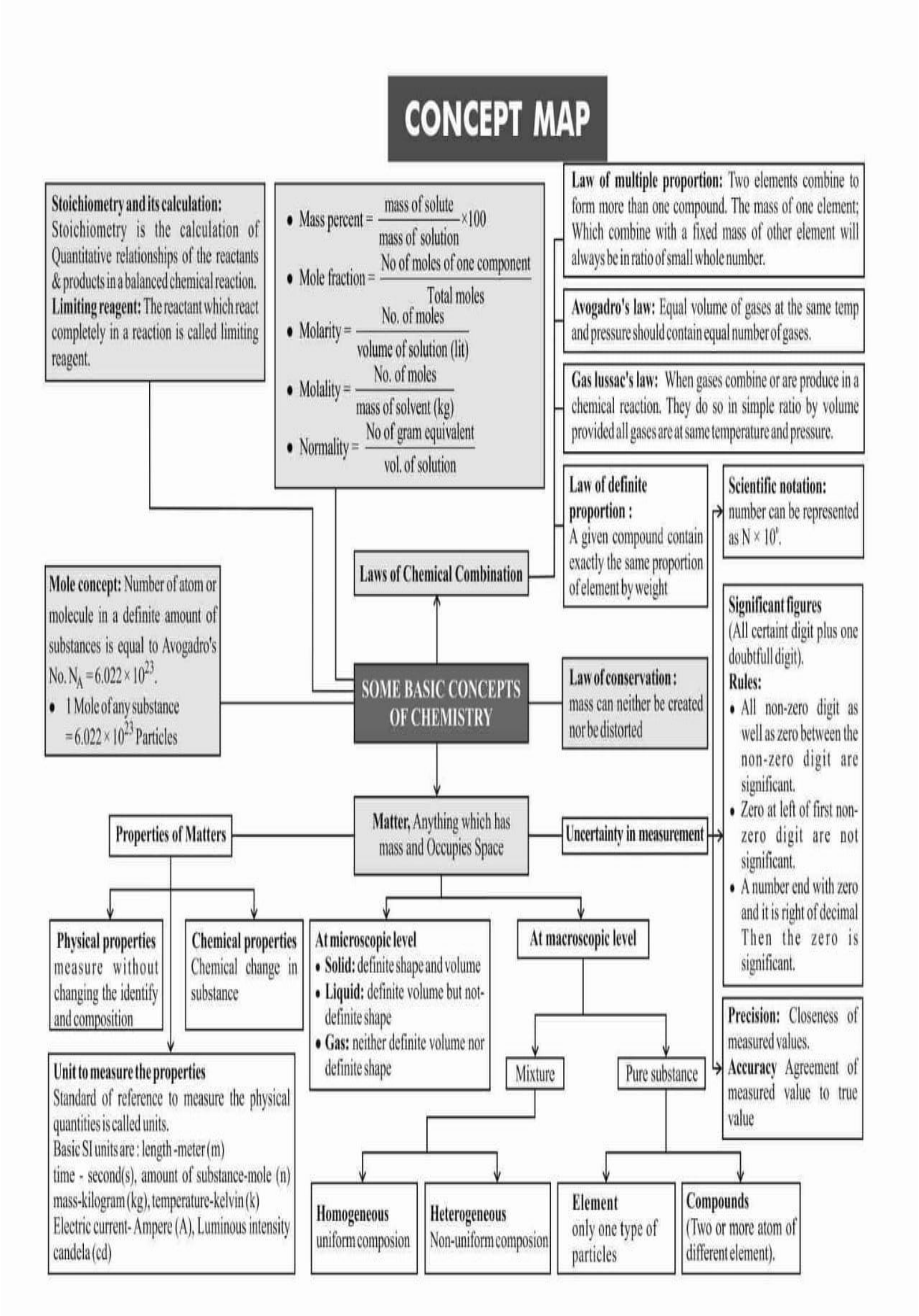
3.01×10^{23} molecules H_2O
= 0.5 moles H_2O

Mass to Moles

18 g of H_2O
= 1 mole H_2O

Moles to Volume

2 moles of O_2
= 44.8 L O_2 (at STP)



CONCEPT MAP

Stoichiometry and its calculation: Stoichiometry is the calculation of Quantitative relationships of the reactants & products in a balanced chemical reaction. **Limiting reagent:** The reactant which react completely in a reaction is called limiting reagent.

- Mass percent = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
- Mole fraction = $\frac{\text{No of moles of one component}}{\text{Total moles}}$
- Molarity = $\frac{\text{No. of moles}}{\text{volume of solution (lit)}}$
- Molality = $\frac{\text{No. of moles}}{\text{mass of solvent (kg)}}$
- Normality = $\frac{\text{No of gram equivalent}}{\text{vol. of solution}}$

Law of multiple proportion: Two elements combine to form more than one compound. The mass of one element; Which combine with a fixed mass of other element will always be in ratio of small whole number.

Avogadro's law: Equal volume of gases at the same temp and pressure should contain equal number of gases.

Gas lussac's law: When gases combine or are produce in a chemical reaction. They do so in simple ratio by volume provided all gases are at same temperature and pressure.

Law of definite proportion: A given compound contain exactly the same proportion of element by weight

Scientific notation: number can be represented as $N \times 10^n$.

Mole concept: Number of atom or molecule in a definite amount of substances is equal to Avogadro's No. $N_A = 6.022 \times 10^{23}$.

- 1 Mole of any substance $= 6.022 \times 10^{23}$ Particles

Law of conservation: mass can neither be created nor be distorted

Significant figures (All certaint digit plus one doubtfull digit).

Rules:

- All non-zero digit as well as zero between the non-zero digit are significant.
- Zero at left of first non-zero digit are not significant.
- A number end with zero and it is right of decimal Then the zero is significant.

Matter, Anything which has mass and Occupies Space

Physical properties measure without changing the identify and composition

Chemical properties Chemical change in substance

At microscopic level

- **Solid:** definite shape and volume
- **Liquid:** definite volume but not-definite shape
- **Gas:** neither definite volume nor definite shape

Precision: Closeness of measured values.

Accuracy Agreement of measured value to true value

Unit to measure the properties
Standard of reference to measure the physical quantities is called units.
Basic SI units are : length-meter (m)
time - second(s), amount of substance-mole (n)
mass-kilogram (kg), temperature-kelvin (k)
Electric current- Ampere (A), Luminous intensity candela (cd)

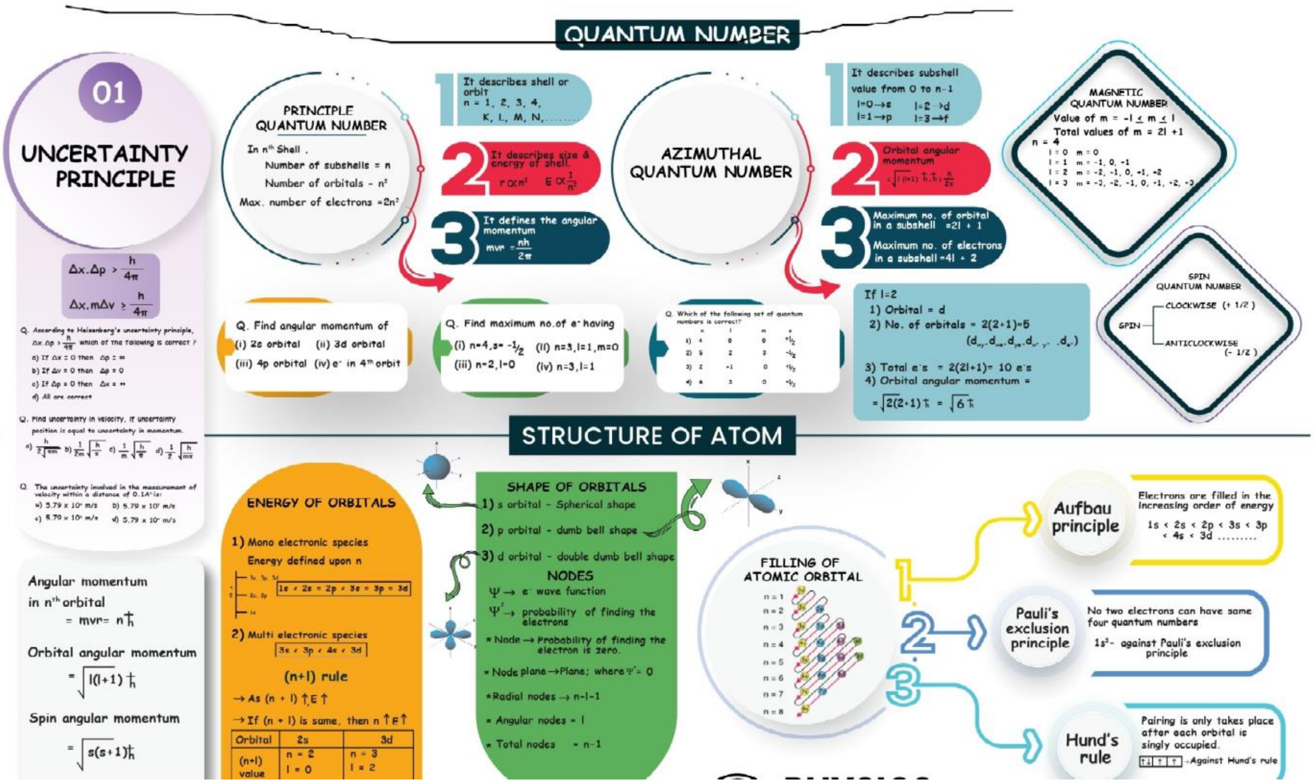
Homogeneous uniform composion

Heterogeneous Non-uniform composion

Element only one type of particles

Compounds (Two or more atom of different element).

Chapter 2:- Structure of Atom



Important Topics:

Discovery of Electron, Proton, Neutron:

Proton	**E. Rutherford**	1909
Neutron	James Chadwick	1932
Electron	J.J. Thomson	1897

Isotopes & Isobars:

Isobars are those elements that have a different atomic number but the same mass number. In contrast, Isotopes are those elements having the same atomic number

Thomson’s Model:

The Thomson model of the atom, also known as the "Plum Pudding Model", describes an atom as a positively charged sphere with negatively charged electrons scattered throughout

Rutherford’s Model of Atom :

Rutherford's model proposed that the negatively charged electrons surround the nucleus of an atom. He also claimed that the electrons surrounding the nucleus revolve around it with very high speed in circular paths. He named these circular paths as orbits.

LIMITATIONS OF RUTHERFORD’S NUCLEAR ATOMIC MODEL:-

- Rutherford’s model cannot explain the stability of an atom.
- It doesn’t given any idea about the distribution of electrons

around the nucleus and about their energies.

3. BOHR’S MODEL FOR HYDROGEN ATOM

Main Postulates of Bohr’s Model for Hydrogen Atom :-

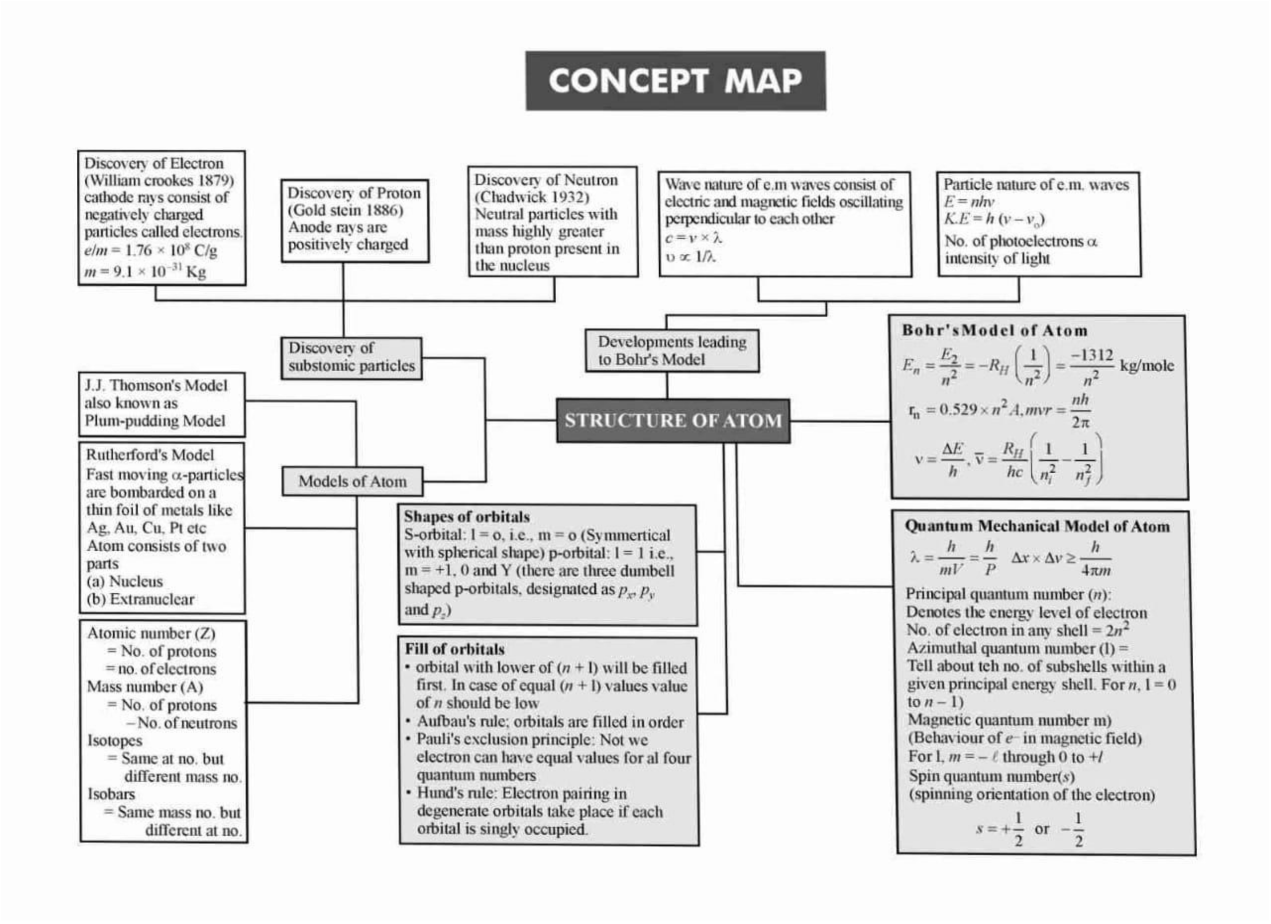
- In the hydrogen atom electrons are moving around the nucleus in a circular path of fixed radius and energy. These paths are called orbits.
- The energy of an electron in the orbit remains constant with time on absorbing sufficient amount of energy, electrons move from lower stationary state to a higher stationary state.
- The frequency of radiation absorbed or emulated when transition occurs between two stationary state differ by ΔE.

$V = \Delta E /h = E_2 - E_1/h$

Where E1 and E2 are the energies of the lower and higher energy states respectively.

Electronic configuration of Atom:

electronic configuration, the arrangement of electrons in orbitals around an atomic nucleus. The electronic configuration of an atom in the quantum-mechanical model is stated by listing the occupied orbitals, in order of filling, with the number of electrons in each orbital indicated by superscript



CONCEPT MAP

Discovery of Electron (William crookes 1879) cathode rays consist of negatively charged particles called electrons.
$e/m = 1.76 \times 10^8$ C/g
$m = 9.1 \times 10^{-31}$ Kg

Discovery of Proton (Gold stein 1886) Anode rays are positively charged

Discovery of Neutron (Chadwick 1932) Neutral particles with mass highly greater than proton present in the nucleus

Wave nature of e.m waves consist of electric and magnetic fields oscillating perpendicular to each other
$c = \nu \times \lambda$
$\upsilon \propto 1/\lambda$.

Particle nature of e.m. waves
$E = nh\nu$
$K.E = h(\nu - \nu_o)$
No. of photoelectrons α intensity of light

Discovery of substomic particles

Developments leading to Bohr's Model

J.J. Thomson's Model also known as Plum-pudding Model

Rutherford's Model
Fast moving α-particles are bombarded on a thin foil of metals like Ag, Au, Cu, Pt etc
Atom consists of two parts
(a) Nucleus
(b) Extranuclear

Models of Atom

STRUCTURE OF ATOM

Bohr'sModel of Atom

$$E_n = \frac{E_2}{n^2} = -R_H\left(\frac{1}{n^2}\right) = \frac{-1312}{n^2} \text{ kg/mole}$$

$$r_n = 0.529 \times n^2 A, mvr = \frac{nh}{2\pi}$$

$$\nu = \frac{\Delta E}{h}, \bar{\nu} = \frac{R_H}{hc}\left(\frac{1}{n_i^2} - \frac{1}{n_j^2}\right)$$

Atomic number (Z)
= No. of protons
= no. of electrons
Mass number (A)
= No. of protons
– No. of neutrons
Isotopes
= Same at no. but different mass no.
Isobars
= Same mass no. but different at no.

Shapes of orbitals
S-orbital: l = o, i.e., m = o (Symmertical with spherical shape) p-orbital: l = 1 i.e., m = +1, 0 and Y (there are three dumbell shaped p-orbitals, designated as p_x, p_y and p_z)

Fill of orbitals
- orbital with lower of $(n + l)$ will be filled first. In case of equal $(n + l)$ values value of n should be low
- Aufbau's rule; orbitals are filled in order
- Pauli's exclusion principle: Not we electron can have equal values for al four quantum numbers
- Hund's rule: Electron pairing in degenerate orbitals take place if each orbital is singly occupied.

Quantum Mechanical Model of Atom

$$\lambda = \frac{h}{mV} = \frac{h}{P} \quad \Delta x \times \Delta v \geq \frac{h}{4\pi m}$$

Principal quantum number (n):
Denotes the energy level of electron
No. of electron in any shell = $2n^2$
Azimuthal quantum number (l) =
Tell about teh no. of subshells within a given principal energy shell. For n, l = 0 to $n - 1$)
Magnetic quantum number m)
(Behaviour of e^- in magnetic field)
For l, $m = -\ell$ through 0 to $+l$
Spin quantum number(s)
(spinning orientation of the electron)

$$s = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

Chapter3- Classification Of Elements And Periodicity In Properties

1. The Concept of Grouping Elements in Accordance to Their Properties Led To the Development of Periodic Table

2. The Periodic Law;

3. The Significance of Atomic Number and Electronic Configuration as the Basis for Periodic Classification;

4. The elements with z >100 according to IUPAC nomenclature;

5. Elements into *S, P, D, F* Blocks and Learn Their Main Characteristics;

6. The Periodic Trends in Physical and Chemical Properties of Elements like atomic and ionic radius, ionization enthalpies, electron gain enthalpy, electronegativity .

7. The Reactivity of Elements and Correlate It with Their Occurrence in Nature;

8. The Relationship between Ionization Enthalpy and Metallic Character;

9. The periodicity and cause of periodicity.

MENDELEEV'S PERIODIC TABLE (1869)

(a) Mendeleev's periodic law:

The physical and chemical properties of elements are the periodic function of their atomic weight.

(b)Characteristic of Mendeleev's periodic table:

(i) It is based on atomic weight. (ii) 63 elements were known, noble

gases were not discovered.

(iii) He was the first scientist to classify the elements in a systematic manner i.e., in horizontal rows and in vertical columns.

(iv) Horizontal rows are called periods and there are 7 periods in Mendeleev's Periodic Table.

(v)Vertical columns are called groups and there are 8 groups in Mendeleev's Periodic Table.

(vi) Each group up to VII is divided into A & B subgroups. 'A' subgroups elements are called normal elements and 'B' sub group elements are called transition elements.

(vii) The VIII group consists of 9 elements in three rows (Transitional metals group).

(viii) The elements belonging to same group exhibit similar properties.

(c) Merits or advantages of Mendeleev's periodic table:

(i) Study of elements: First time all known elements were

classified in groups according to their similar properties. So, study of the properties of elements become easier.

(ii) Predication of new elements: It gave encouragement to the discovery of new elements as some gaps were left in it.

Sc (Scandium) Ga(Gallium) Ge(Germanium) Tc(Technetium)

These were the elements for whom position and properties were well defined by

Mendeleev even before their discoveries and he left the blank spaces for them in his table.Ex. Blank space at atomic weight 72 in silicon group was called Eka silicon (means properties like silicon) and element discovered later was named Germanium.

Similarly other elements discovered after Mendeleev periodic table were:

i) Correction of doubtful atomic weights:

Corrections were done in atomic weight of some elements.

Eka aluminium- Gallium (Ga) ,Eka Silicon- Germanium (Ge)

Eka Boron- Scandium (SC), Eka Mangense- Technetium(Tc)

Atomic weight = Valency X Equivalent weight

Initially, it was found that equivalent weight of Be is 4.5 and it is trivalent (V=3), so the weight of Be was 13.5 and there is no space in Mendeleev's table for the element. So, after correction, it was found that Be is actually divalent (V=2). So, the weight of Be became 2 X 4.5=9 and there was a space between Li and B for this element in Mendeleev's table.

Defects of Mendeleev's Periodic Table:

(i) Position of hydrogen is uncertain. It has been placed in IA and VII A group because of its resemblance with both the groups.

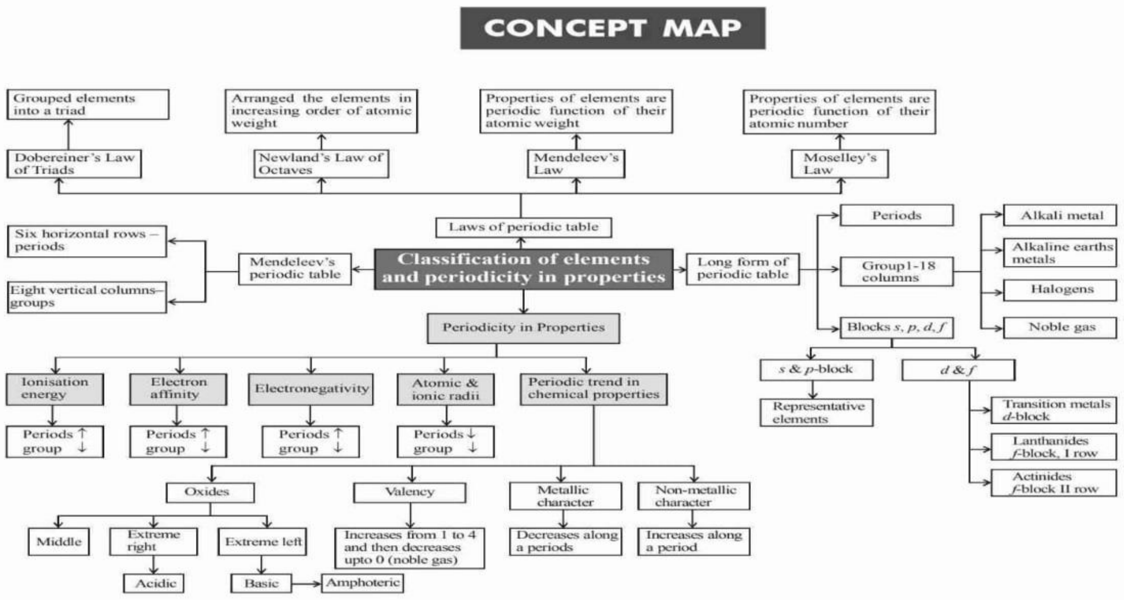
(ii) No separate positions were given to isotopes.

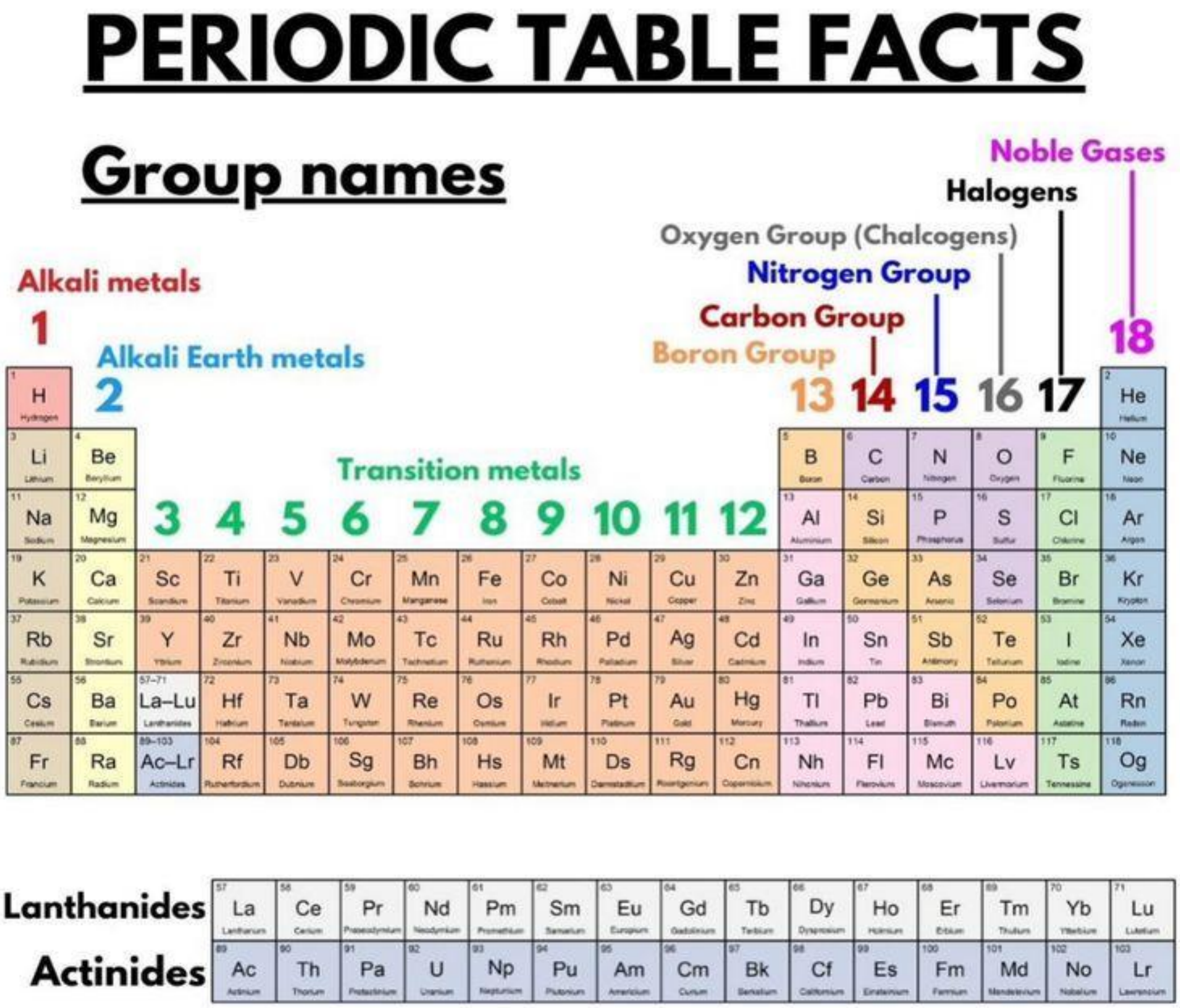
(iii) It is not clear whether the lanthanides and actinides are related to IIA or IIB group.

(iv) Although there is no resemblance except valency of subgroups A and B, they have been put in the same group.

(v) Order of increasing atomic weights is not strictly followed in the arrangement of elements in the periodic table. For

e.g.- Co (At.wt. 58.9) is placed before Ni (127) and Ar (39.9) before K (39).





PERIODIC TABLE FACTS

Group names

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La–Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac–Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

Lanthanides	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinides	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

CH 4: CHEMICAL BONDING AND MOLECULAR STRUCTURE

INTRODUCTION : The attractive force which holds various constituent (atoms, ions , etc) together in different chemical species is called chemical bond . The chemical bonds can be classified into various types likes ionic bond , covalent bond , coordinate and metallic bond , hydrogen bond and vander waal forces. In this chapter , we will study about :

1. Kossel-Lewis approach to chemical bonding :Kossel – Lewis explain about cause of combination between atoms based upon understanding of electronic configuration of noble gases.

2. Octet rule : The principle of attaining maximum of eight electrons in the valence shell of atoms is called octet rule .

3.Formation of different types of bonds : There are two main types of bonds (i) Ionic bond : The electrostatic force of attraction which holds the oppositely charged ions together is known as ionic bond . (ii) Covalent bond : The bond formed by mutual sharing of electrons between combining atoms of same or different elements is called covalent bond .

4. VSEPR theory: The basic idea of this theory is that bonded atoms in molecule adopt that particular arrangement in space around the central atom which keeps them on average as far apart as possible .

5.Valence bond theory: It is based on the knowledge of atomic orbital and electronic configuration of elements , overlap criteria of atomic orbitals and stability of molecules

6. Molecular orbital theory : The basic idea of molecular orbital theory is that atomic orbitals of individual atoms combine to form molecular orbitals .

7.Different types of hybridization involving s, p and d-orbitals: sp , sp2 , sp3 and draw shapes of simple covalent molecules.

8.Concept of hydrogen bond and metallic bond.

Covalent bond may be defined as the sharing of one or more electron pairs with similar or dissimilar atoms. In this type of chemical bonding; each atom contributes equal number of electrons for sharing. In such a way, the participating covalent bonding atoms complete their octet structures. The compound so formed are called covalent compounds.

Octet Rule : When two atoms form covalent compounds, they attain a noble gas configuration in their valance shell. This is known as Octet Rule.

Lewis Dot Structure for simple molecules: Lewis structure is simplified representation of the valence shell electrons in a molecule. Electrons are shown as “dots” ; bonding electrons are as a line between the two atoms.

Formal charge:

It is the charge assigned to an atom in molecule assuming that electrons in all chemical bonds are equally shared equally between the atoms regardless of electronegativity.

$FC = V - L - \frac{1}{2} S$

FC= formal charge of an atom

V= total number of valence electrons

L= total number of non-bonding electrons (lone pair)

S= total number of bonding electrons (shared pair)

Ozone Molecule:

‘O’ atom marked as ‘1’ = $6 - 2 - \frac{1}{2} \times 6 = +1$

‘O’ atom marked as ‘2’ = $6 - 4 - \frac{1}{2} \times 4 = 0$

‘O’ atom marked as ‘3’ = $6 - 6 - \frac{1}{2} \times 2 = -1$

CHEMICAL BONDING

[redacted]

01

VSEPR

2,0- Linear - 180°
3,0- Trigonal planar - 120 °
2,1- Bent shape
4,0- Tetrahedral - 109°28'
3,1- Pyramidal - 107°
2,2- Bent - 104.5°

5,0- Trigonal Bipyramidal
4,1- See saw
3,2- T-shape
2,3- Linear - 180°

6,0- Octahedral - 90°
5,1- Square pyramidal
4,2- Square planar

7,0- Pentagonal bipyramidal
6,1- Distorted octahedral

Q1 In which of the following pairs are the two species trigonal pyramidal ?

1) BrO_3^- and XeO_3 2) SF_4 and XeF_4
2) SO_3^{2-} and NO_3^- 4) BF_3 and NF_3

02

EP $= \frac{CVE + \text{no. of MV}}{2}$

MV ↓ BP Remaining ↓ LP Monoatomic surrounding atoms

EP $= \frac{CVE + 0}{2}$

DV ↓ BP Remaining ↓ LP Diatomic surrounding atoms

EP $= \frac{CVE + \text{negative charge} - \text{positive charge}}{2}$

MV/DV ↓ BP Remaining ↓ LP Charged Compound

LP-LP > LP-BP > BP-BP
(Repulsion)

Q2 Total no. of Lone pairs in I_3^- is

1) 3 2) 6
3) 9 4) 12

03

HYBRIDISATION

2- sp - Linear
3- sp^2 - Trigonal planar
4- sp^3 - Tetrahedral
5- sp^3d - Trigonal pyramidal
6- sp^3d^2 - Octahedral
7- sp^3d^3 - Pentagonal bipyramidal

CH_3-CH_3 - sp^3
$CH_2=CH_2$ - sp^2
$CH\equiv CH$ - sp
=C= - sp

Q3 In XeF_6, oxidation state and state of hybridisation of Xe and shape of the molecule are, respectively

1) +6, sp^3d^3, distorted octahedral
2) +4, sp^3d^2, square planar
3) +6, sp^3, pyramidal
4) +6, sp^3d^2 square pyramidal

04

VALENCE BOND THEORY

$C-C = 1\sigma$ bond
$C=C = 1\sigma + 1\pi$
$C\equiv C = 1\sigma + 2\pi$

MOLECULAR ORBITAL THEORY

Up to N_2 the energy order is,
$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x = \pi 2p_y) < \sigma 2p_z < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$

After N_2 the energy order is, (for O_2, F_2, Ne_2)
$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \sigma 2p_z < (\pi 2p_x = \pi 2p_y) < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$

B_2-10 Paramagnetic $\pi 2p_x^1 = \pi 2p_y^1$
O_2-16 Paramagnetic $\pi^* 2p_x^1 = \pi^* 2p_y^1$

Q4 The no. of unpaired electrons in O_2 molecule is:

1) 0 2) 1
3) 2 4) 3

05

BOND ORDER

Total no.of e^-

2 - 1 - Diamagnetic
3 - 0.5 - Paramagnetic
4 - 0 - Diamagnetic
5 - 0.5 - Paramagnetic
6 - 1 - Diamagnetic
7- 0.5- Paramagnetic
8 - 0 - Diamagnetic
9 - 0.5 - Paramagnetic
10 - 1 - Paramagnetic
11 - 1.5 - Paramagnetic
12 - 2 - Diamagnetic
13 - 2.5 - Paramagnetic
14 - 3 - Diamagnetic
15 - 2.5 - Paramagnetic
16 - 2 - Paramagnetic
17 - 1.5 - Paramagnetic
18 - 1 - Diamagnetic
19 - 0.5 - Paramagnetic
20 - 0 - Diamagnetic

Odd number + 10&16 - Paramagnetic
Even number except 10&16 - Diamagnetic

BOND PARAMETERS

BOND ORDER

B.O $\propto \frac{1}{B.L} \propto$ Stability $\propto$ Bond strength

B.O-$\underset{3}{O_2^{2+}} > \underset{2.5}{O_2^{+}} > \underset{2}{O_2}$

B.L-$O_2 > O_2^+ > O_2^{2+}$

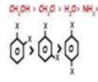
BOND ENERGY

1) Hybridisation
$sp > sp^2 > sp^3$
2) B.E. $\propto$ B.O. (No. of Bonds)
$N_2 > O_2 > F_2$
3) B.E. $\propto$ E.N. difference
HF > HCl > HBr > HI
4) Halogens (B.E.)
$Cl_2 > Br_2 > F_2 > I_2$

DIPOLE MOMENT (μ)

1) Polar molecule, $\mu \neq 0$
Irregular geometry.
a) Different bonds
b) Lone pairs
c) Different surrounding atoms
eg: SF_4, CH_3Cl
2) Non polar molecule, $\mu = 0$
Regular geometry
a) Same bonds
b) Zero lone pairs
c) Same surrounding atoms
eg: CH_4, CO_2

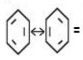
$CH_3OH > CH_3Cl > H_2O > NH_3 > NF_3 > BF_3$
$\mu=0$



RESONANCE

Resonance Changes B.O.

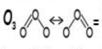
1) Benzene



$= \frac{3}{2} = 1.5$

B.L.: $C_2H_6 > C_6H_6 > C_2H_4 > C_2H_2$
B.O.: 1 1.5 2 3

2) O_3



$= \frac{3}{2} = 1.5$

B.L.: $H_2O_2 > O_3 > O_2$
B.O.: 1 1.5 2

BOND ANGLE

1) Max B.A-180°[sp]
2) Hybridisation
$sp > sp^2 > sp^3$
$\underset{sp}{NO_2^+} > NO_2 > \underset{sp^2}{NO_2^-}$
3) if LP↑ B.A↓
$CH_4 > NH_3 > H_2O$ [sp^3]
0 LP 1 LP 2 LP
4) If electronegativity of C.A↑ B.A↑
$NH_3 > PH_3 > AsH_3 > SbH_3$
5) If electronegativity oF S.A↑ B.A↓
$NI_3 > NBr_3 > NCl_3 > NF_3$

BOND LENGTH

1) B.L ↑ with Size of atom
HI > HBr > HCl > HF
2) B.L decreases with multiplicity
$C-C > C=C > C\equiv C$

Q5 Which of the following molecules has the maximum bond enthalpy?

1) N_2 2) CO
3) F_2 4) HF

Q6 Which of the following will have maximum dipole moment?

1) NF_3 2) NH_3
3) CH_4 4) PCl_3

Q7 The correct order in which the O-O bond length decreases in the following is

1) $O_3 > H_2O_2 > O_2$ 2) $O_2 > O_3 > H_2O_2$
3) $O_2 > H_2O_2 > O_3$ 4) $H_2O_2 > O_3 > O_2$

Q8 The correct order of bond angle is:

1) $H_2O > NH_3 > CH_4 > CO_2$
2) $H_2O < NH_3 < CO_2 > CH_4$
3) $H_2O < NH_3 > CO_2 > CH_4$
4) $CO_2 > CH_4 > NH_3 > H_2O$

Q9 The correct order of increasing bond length of C-H, C-O, C-C & C=C is:

1) C-C < C=C < C-O < C-H
2) C-O < C-H < C-C < C=C
3) C-H < C-O < C-C < C=C
4) C-H < C=C < C-O < C-C

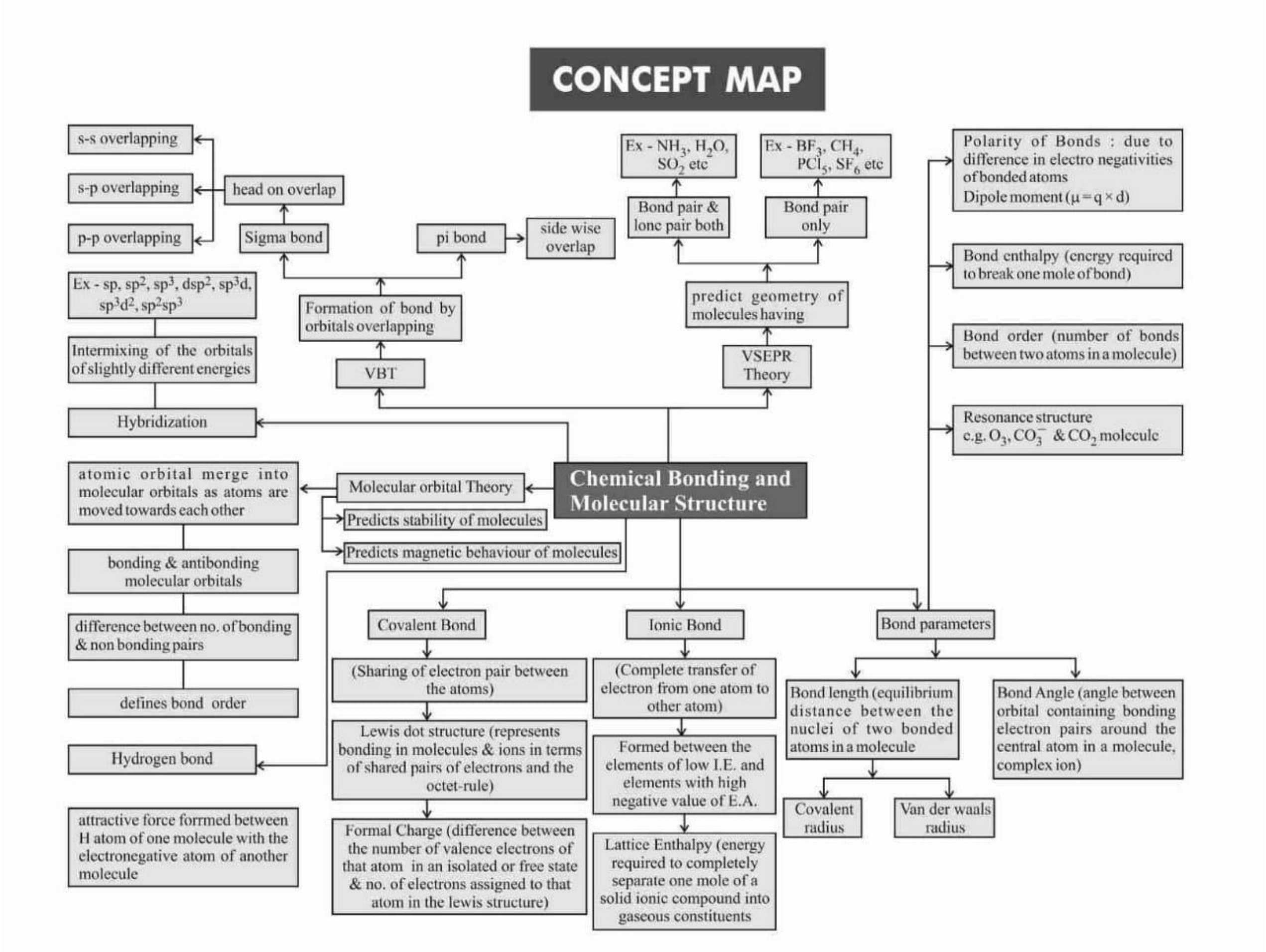
Hydrogen bond- Formation of bond between hydrogen & most electronegetive elements like F,O,N

Intramolecular - H - Bonding occur within one single molecule.

Intermolecular - H Bonding between two different molecules of same or different compounds.

Intramolecular H bonding increases the boiling point. eg: Intermolecular H bond in p-nitrophenol increases the boiling point. HF(HB) > HI > HBr > HCl. H_2O(4HB) > H_2Te > H_2Se > H_2S

Fajan's Rule - Indicates Covalent character in an ionic bond.
1) Size of the cation- Smaller the cation higher is the polarisation, so covalent character increases
LiCl > NaCl > KCl > RbCl > CsCl
2) Size of anion- As the size of anion increases, polarisation increases, covalent character increases
AgF < AgCl < AgBr < AgI
3) Charge on the cation- As the charge increases, Covalent character also increases
LiCl < $BeCl_2$ < BCl_3 < CCl_4



CHAPTER 5: CHEMICAL THERMODYNAMICS

DEFINITION: Thermodynamics is branch of science which deals with study of different forms of energy and their interconversion.

BASICS TERMS:

●SYSTEM : A part of universe which is under observation.A system which can exchange matter as well energy with surroundings is open system eg. chemical reaction carried out in beaker.A system which can exchange energy but not matter with surroundings is closed system eg. Chemical reaction carried in sealed test tube. A system which can neither exchange energy nor matter with surroundings is isolated system eg. Chemical reactions carried out in closed containers insulated from surroundings.

● ISOTHERMAL PROCESS: A process during which temperature of system remains constant.

● ADIABATIC PROCESS: A process in which no heat enters or leaves the system.

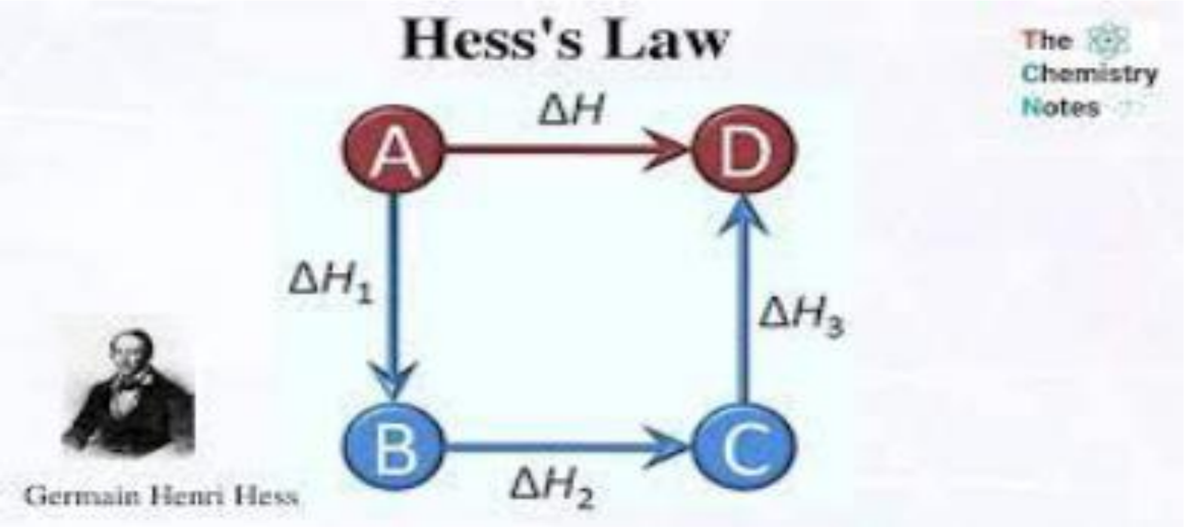
●REVERSIBLE PROCESS:A process in which direction may be reversed at any stage by small change in variable like temperature, pressure.

EXTENSIVE PROPERTIES : Properties of system whose value depends upon the amount of substance present in system like enthalpy,internal energy,entropy,free energy.These properties are state functions because their values depend upon the state of the system and is independent of path or manner by which state is reached.

ENTHALPY:Enthalpy change is heat absorbed or evolved at constant temperature and constant pressure and we always consider enthalpy change(-ve during exothermic and +ve during endothermic reaction) because atmospheric pressure is considered constant. There are different forms of enthalpies depending upon types of reaction(combustion,hydration,neutralisation,sublimation).

BOMB CALORIMETER:Bomb Calorimeter is used to calculate internal energy

HESS LAW: This law is also known as law of constant heat summation.It states that heat absorbed or evolved in a particular reaction is same whether reaction takes place in one step or in a number of steps.



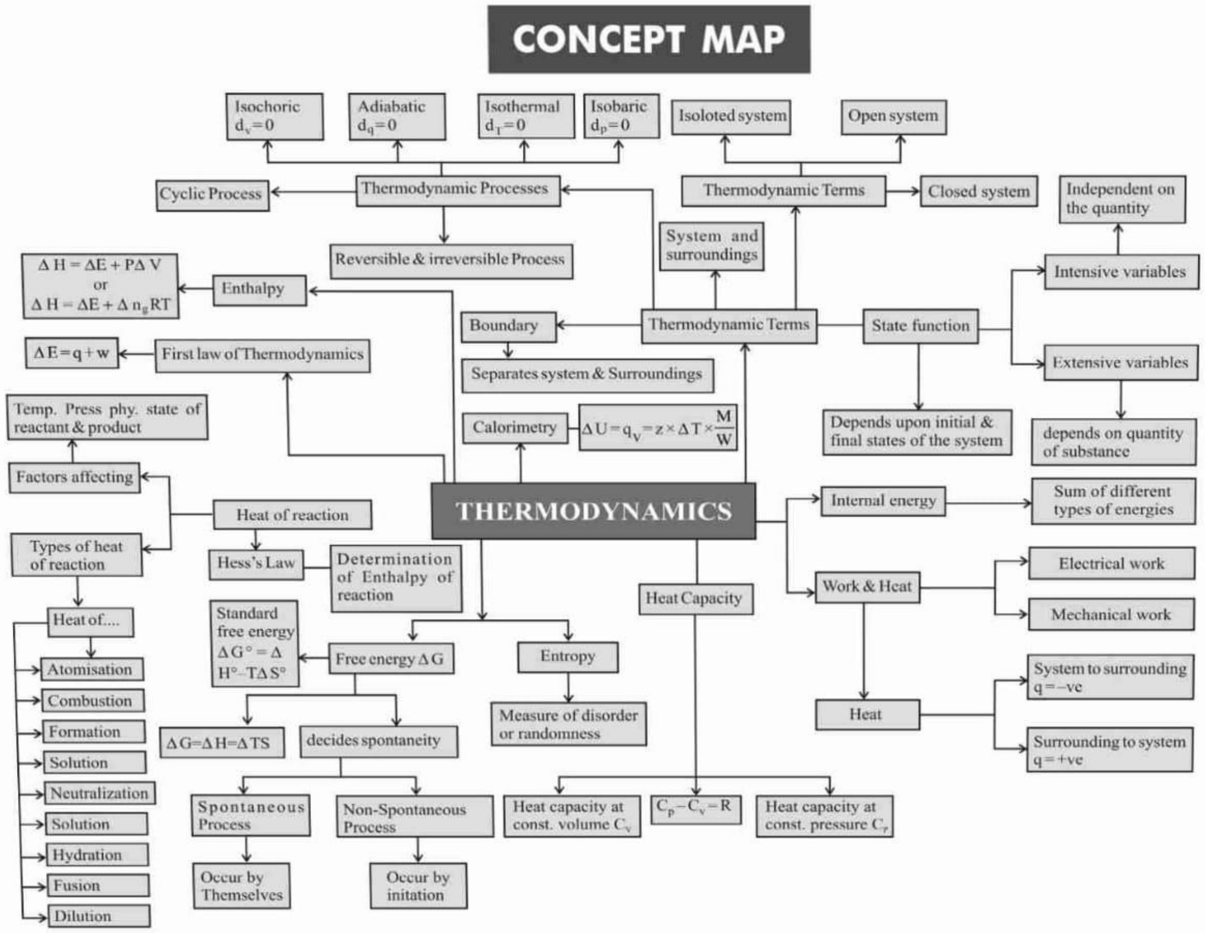
BORN HABER CYCLE:Lattice enthalpy is heat evolved when one mole of an ionic compound is formed from its ions is sum of enthalpy of sublimation,enthalpy of atomisation,ionisation enthalpy,electron gain enthalpy .

FIRST LAW OF THERMODYNAMICS:First law of thermodynamics (law of conservation of energy) does not provide any information about spontaneous nature of reaction (why reaction can take place by itself?).Negative enthalpy and positive entropy (degree of randomness) are driving forces which make process spontaneous.

GIBBS HELMHOLTZ EQUATION:▲G=▲H-T▲S which combines enthalpy and entropy in single equation.For spontaneous process ▲G (change in free energy which is a measure to do useful work,must be negative).That’s why endothermic processes are favourable at high temperature and exothermic processes are favourable at low temperature.

SECOND LAW OF THERMODYNAMICS :Entropy of the universe always increase during spontaneous process.

THIRD LAW OF THERMODYNAMICS: At absolute zero entropy of a perfectly crystalline substance is taken as zero.



CHAPTER 6: Equilibrium

Chemical Equilibrium

In a chemical reaction, chemical equilibrium is defined as the state at which there is no further change in concentration of reactants and products.

For example: A + B C + D

- At equilibrium the rate of forward reaction is equal to the rate of backward reaction.
- The reactions in which the concentrations of both the reactants and products are comparable when the system is in equilibrium.

Equilibrium in Physical Processes

1. Solid-Liquid Equilibrium:

Rate of melting of ice = Rate of freezing of water.

The system here is in dynamic equilibriums and following can be inferred.

✓ Both the opposing processes occur simultaneously

✓ Both the processes occur at the same rate so that the amount of ice and water remains constant.

2. Liquid-Vapour Equilibrium

Rate of evaporation = Rate of condensation

3. Equilibrium involving Dissolution of Solid in Liquid solution:

When a limited amount of salt or sugar or any solute dissolves in a given amount of water solution is formed. At a given temperature state is reached when no more solute can be dissolved then the solution is called saturated solution.The equilibrium between a solid and its solution is indicated by the saturated solution and may be represented as Equilibrium involving Dissolution of Solid in Liquid.

5. Equilibrium between a Gas and its Solution in Liquid

Henry's law: It states that the mass of a gas that dissolves in a given mass of a solvent at any temperature is proportional to the pressure of the gas above the surface of the solvent.

M = kHp ; here kH = Henry's constant

Characteristics of Equilibria Involving Physical Processes

✓ The equilibrium can be attained only in closed systems at a given temperature.

✓ At the equilibrium the measurable properties of the system remain constant.

✓ The equilibrium is dynamic since both the forward and backward processes occur at same rate.

✓ At equilibrium, the concentrations of substances become constant at constant temperature.

✓ The value of equilibrium constant represents the extent to which the process proceeds before equilibrium is achieved.

At equilibrium, $aA+bB \rightleftharpoons cC+dD$

Rate of forward reaction = Rate of backward reaction

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

equilibrium Constant in terms of Partial Pressure

$$K_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

Relation between Kp and Kc

$K_p = K_c (RT)^{\Delta n}$

where R is gas constant 0.0821 liter.atm/°C/mole and Δn is the total number of molecules of the product – total number of molecules of the reactants.

Relation between Equilibrium Constant and Standard Free Energy

$\Delta G^o = - RT \ln K_{eq}$,

$\Delta G^o = - 2.303RT \log K_{eq}$

Le Chatelier's Principle : A change in one of the variable s that describe a system at equilibrium produces a shift in the position of the equilibrium that counteracts the effect of this change.

pH Calculation

pH = -log[H+]

pOH = -log[OH-]

pH + pOH = 14

BUFFER Solution : calculating the pH of buffer solutions.

For an acid (HA) and its conjugate base (A^-), the Henderson-Hasselbalch equation is given by:

$pH = pK_a + \log([A-]/[HA])$

For a base (B) and its conjugate acid (BH^+), the equation is modified as follows:

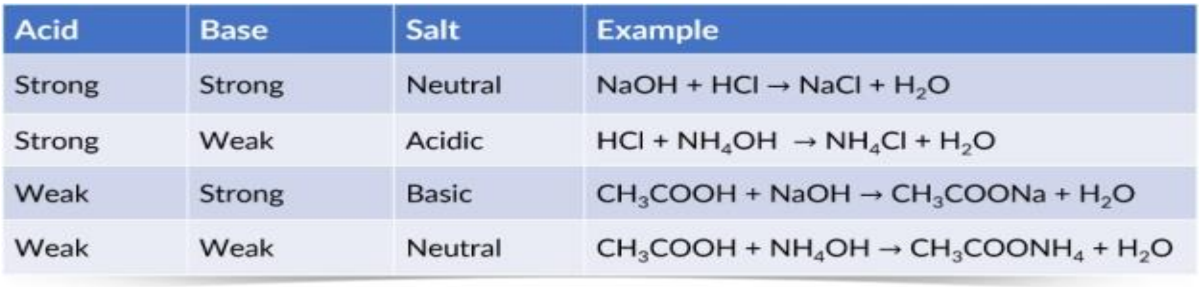
$pOH = pK_b + \log([BH+]/[B])$

The solubility product constant Ksp is the equilibrium constant for the equilibrium established between a slightly soluble ionic compound and its ions in a saturated aqueous solution.

$$K_{SP} = (xs)^x (ys)^y = x^x.y^y.(s)^{x+y}$$

.

Salt hydrolysis



Acid	Base	Salt	Example
Strong	Strong	Neutral	$NaOH + HCl \rightarrow NaCl + H_2O$
Strong	Weak	Acidic	$HCl + NH_4OH \rightarrow NH_4Cl + H_2O$
Weak	Strong	Basic	$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
Weak	Weak	Neutral	$CH_3COOH + NH_4OH \rightarrow CH_3COONH_4 + H_2O$

STRONG ACID IN WATER : A strong acid is one which dissociates 100% in solution. Hydrogen chloride is a strong acid in water. That means it reacts with water completely.

HCl (g) + H_2O (l) ⇄ H_3O^+ (aq) + Cl^- (aq) Complete reaction (100%)

WEAK ACID IN WATER : A weak acid is one which forms equilibrium in solution, rather than dissociate 100%. Hydrogen fluoride doesn't react completely with water. It sets up an equilibrium with water. Some HF remains in solution undissociated. HF is a weak acid in water.

HF (aq) + H_2O (l) ⇄ H_3O^+(aq) + F^- (aq) not complete reaction (equilibrium).

STRONG AND WEAK BASES : The issue is similar with bases, a strong base is a base that is 100% ionized in solution. If it is less than 100% ionized in solution, it isa weak base. There are very few strong bases. All strong bases are OH compounds. So a base based on some other mechanism, such as NH_3 (which does not contain OH^- ions as part of its formula), will be a weak base.

HOW ACID IS FORMED : An acid is formed when a compound has a hydrogen bonded to an atom that is moderately stable as a negative ion after the hydrogen ion is stripped off by a base. Inorganic acids generally are formed when hydrogens are attached to electronegative species like a halogen or complex ionslike sulfates, which can exist as negative ions Organic molecules can have acidic properties. The most common type of organic acid is a carboxylic acid, as we seein the acetic acid molecule.

CH_3COOH (aq) + H_2O (l) ⇄ H_3O^+ + CH_3COO^- (aq)

Acetic acid is also a weak acid. The hydrogen on the oxygen is the acidic hydrogen. It leaves behind a carboxylate ion (called the acetate ion in this case).

ACID BASE EQUILIBRIUM : Acid-base reactions, in which protons are exchanged between donor molecules (acids) and acceptors (bases), form the basis of the most common kinds of equilibrium problems which you will encounter in almost any application of chemistry. Following theories explain acid base equilibrium.

- The Arrhenius concept of acids and bases
- The Bronsted-Lowry concept, conjugate acids and bases

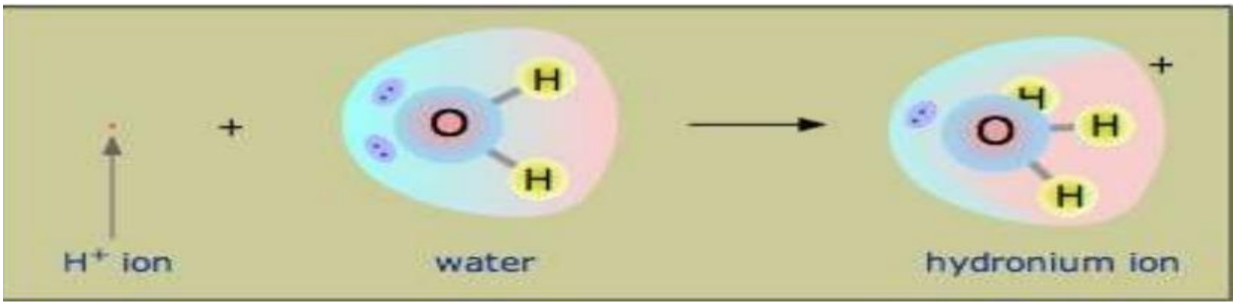
Arrhenius Theory : An Arrhenius acid is any species that increases the concentration of H^+ in aqueous solution. An Arrhenius base is any species that increases the concentration of OH^-in aqueous solution.

Bronsted-Lowry Theory : In the Bronsted-Lowry definition, acids are proton donors, and bases are proton acceptors. Note that these definitions are interrelated. Defining a base as a proton acceptor means an acid must be available to provide the proton.

For example, in reaction below acetic acid, CH_3COOH, donates a proton to ammonia, NH_3, which serves as the base.

CH_3COOH(aq) + NH_3(aq) ⇌ CH_3COO^- (aq) + NH_4^+(aq)

The acetate ion considered as a conjugate base (which is strong because weaker the acid means stronger the conjugate base), the ammonium ion is the conjugate acid for the ammonia (the base).



pH DEFINITION : pH is a measure of hydrogen ion concentration, a measure of the acidity or alkalinity of a solution. The pH scale usually ranges from 0 to 14. Aqueous solutions at 25°C with a pH less than 7 are acidic, while those with a pH greater than 7 are basic or alkaline. A pH level of 7.0 at 25°C is defined as "neutral" because the concentration of $H3O^+$ equals the concentration of $OH^−$ in pure water. Very strong acids might have a negative pH, while very strong bases might have a pH greater than 14.5. The equation for calculating pH was proposed in 1909 by Danish biochemist Soren Peter Lauritz Sorensen:

$$pH = -\log[H^+]$$

SALTS : If the solution containing the salt alone we must recognize the type of this salt,there are three major types:

NEUTRAL SALTS : Salts of strong acid and base (NaCl) this type has neutral acidity (pH = 7).

BASIC SALTS : Salts of weak acid and strong base (CH3COONa) this type of salt has pH > 7. $[H^+] = \sqrt{KwKa/Salt}$

pH = (pKw + pKa + log [salt])

ACIDIC SALTS : Salts of strong acid and weak base (NH4Cl) this type of salt has pH < 7 $[H^+] = \sqrt{Kw}$ ~~$[Salt]/Kb$~~

pH = (pKw - pKb - log [salt])

We can manipulate the equations above by using: **Kw = Ka x Kb**

COMMEN ION EFFECT : The common ion effect is an effect that suppresses the ionization of an electrolyte when another electrolyte (which contains an ion which is also present in the first electrolyte, i.e. a common ion) is added. It is considered to be a consequence of Le Chatlier's principle (or the Equilibrium Law). An example of the common ion effect can be observed when gaseous hydrogen chloride is passed through a sodium chloride solution, leading to the precipitation of the NaCl due to the excess of chloride ions in the solution (brought on by the dissociation of HCl). This effect cannot be observed in the compounds of transition metals. This is because the d-block elements have a tendency to form complex ions. This can be observed in the compound cuprous chloride, which is insoluble in water. This compound can be dissolved in water by the addition of chloride ions leading to the formation of the $CuCl2^‾$ complex ion, which is soluble in water.

EFFECT OF COMMEN-ION ON SOLUBILITY: The common ion effect can be used to obtain drinking water from aquifers (underground layer of water mixed with permeable rocks or other unconsolidated materials) containing chalk or limestone.

Sodium carbonate (chemical formula Na_2CO_3) is added to the water in order todecrease the hardness of the water.

EFFECT OF COMMEN-ION ON pH : When the conjugate ion of a buffer solution (solution containing a base and its conjugate acid, or acid and its conjugate base) is added to it, the pH of the buffer solution changes due to the common ion effect. An example of such an effect can be observed when acetic acid and sodium acetate are both dissolved in a given solution, generating acetate ions. However, sodium acetate completely dissociates but the acetic acid only partly ionizes. This is because acetic acid is a weak acid whereas sodium acetate is a strong electrolyte.

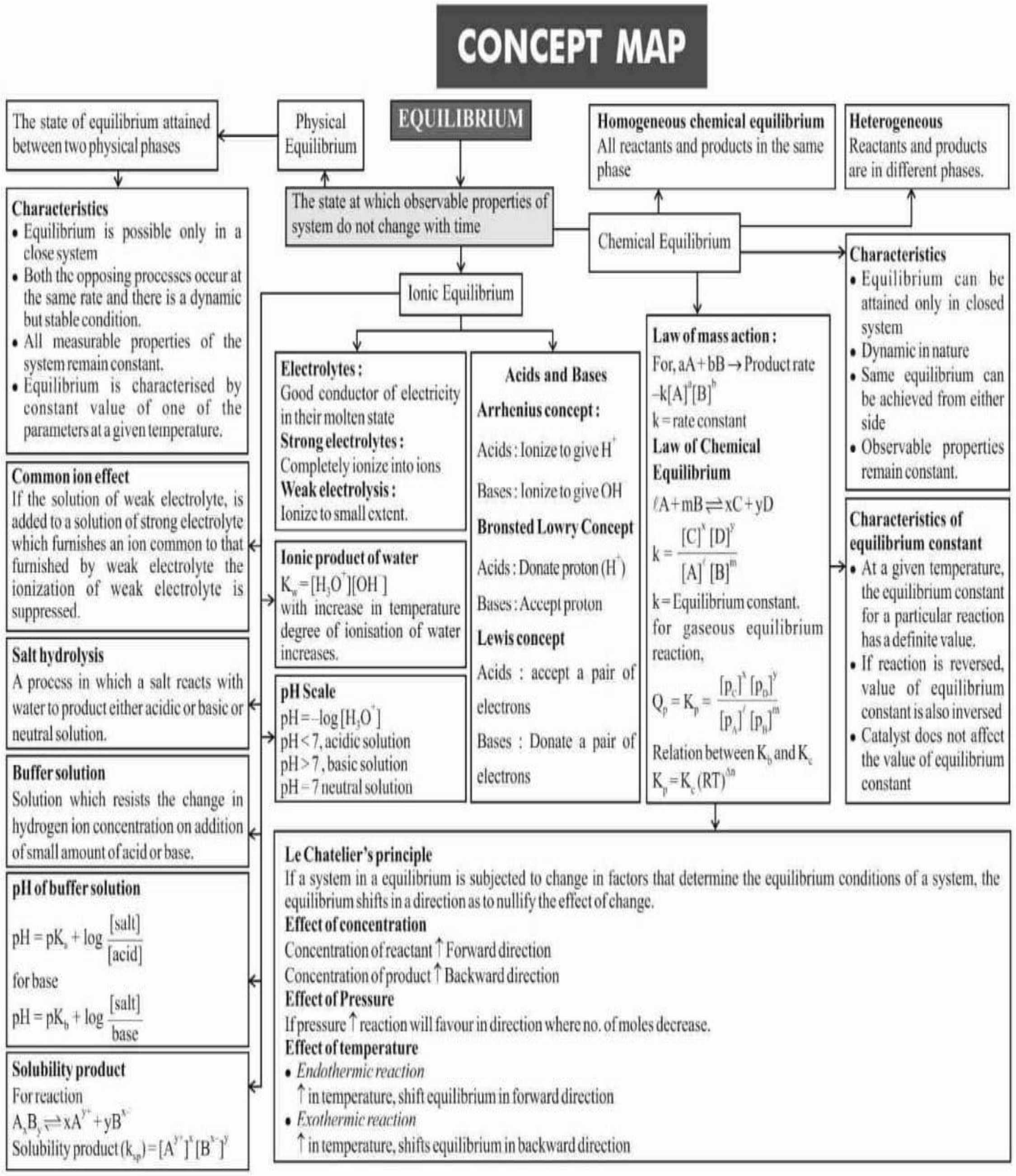
SOLUBILITY PRODUCT Ksp : The solubility product constant is the equilibrium constant for the dissolution of a solid substance into an aqueous solution. It is denoted by the symbol Ksp. The solubility product is a kind of equilibrium constant and its value depends on temperature. Ksp usually increases with an increase in temperature. Suppose barium sulphate along with its saturated aqueous solution is taken. The following equation represents the equilibrium set up between the undissolved solids and ions:

$$BaSO4 \rightleftarrows Ba^{+2}(aq) + SO4^{-}(aq)$$

The equilibrium constant in the above case is: K = $\frac{[Ba^{+2}][SO_4]-}{[BaSO4]}$

$Ksp = K[BaSO4] = [Ba^{+2}][SO4^{-}]$ Here Ksp is known as the solubility product constant.

SIGNIFICANCE OF SOLUBILITY PRODUCT : Solubility depends on a number of parameters amongst which lattice enthalpy of salt and solvation enthalpy of ions in the solution are of most importance.



CONCEPT MAP

EQUILIBRIUM

The state at which observable properties of system do not change with time

Physical Equilibrium

The state of equilibrium attained between two physical phases

Characteristics
- Equilibrium is possible only in a close system
- Both the opposing processes occur at the same rate and there is a dynamic but stable condition.
- All measurable properties of the system remain constant.
- Equilibrium is characterised by constant value of one of the parameters at a given temperature.

Ionic Equilibrium

Common ion effect
If the solution of weak electrolyte, is added to a solution of strong electrolyte which furnishes an ion common to that furnished by weak electrolyte the ionization of weak electrolyte is suppressed.

Salt hydrolysis
A process in which a salt reacts with water to product either acidic or basic or neutral solution.

Buffer solution
Solution which resists the change in hydrogen ion concentration on addition of small amount of acid or base.

pH of buffer solution

$pH = pK_a + \log \frac{[salt]}{[acid]}$

for base

$pH = pK_b + \log \frac{[salt]}{base}$

Solubility product
For reaction
$A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}$
Solubility product $(k_{sp}) = [A^{y+}]^x[B^{x-}]^y$

Electrolytes :
Good conductor of electricity in their molten state
Strong electrolytes :
Completely ionize into ions
Weak electrolysis :
Ionize to small extent.

Ionic product of water
$K_w = [H_3O^+][OH^-]$
with increase in temperature degree of ionisation of water increases.

pH Scale
$pH = -\log[H_3O^+]$
pH < 7, acidic solution
pH > 7, basic solution
pH = 7 neutral solution

Acids and Bases
Arrhenius concept :
Acids : Ionize to give H^+
Bases : Ionize to give OH^-
Bronsted Lowry Concept
Acids : Donate proton (H^+)
Bases : Accept proton
Lewis concept
Acids : accept a pair of electrons
Bases : Donate a pair of electrons

Chemical Equilibrium

Homogeneous chemical equilibrium
All reactants and products in the same phase

Heterogeneous
Reactants and products are in different phases.

Characteristics
- Equilibrium can be attained only in closed system
- Dynamic in nature
- Same equilibrium can be achieved from either side
- Observable properties remain constant.

Law of mass action :
For, $aA + bB \rightarrow$ Product rate $-k[A]^a[B]^b$
k = rate constant
Law of Chemical Equilibrium
$\ell A + mB \rightleftharpoons xC + yD$

$k = \frac{[C]^x[D]^y}{[A]^\ell[B]^m}$

k = Equilibrium constant.
for gaseous equilibrium reaction,

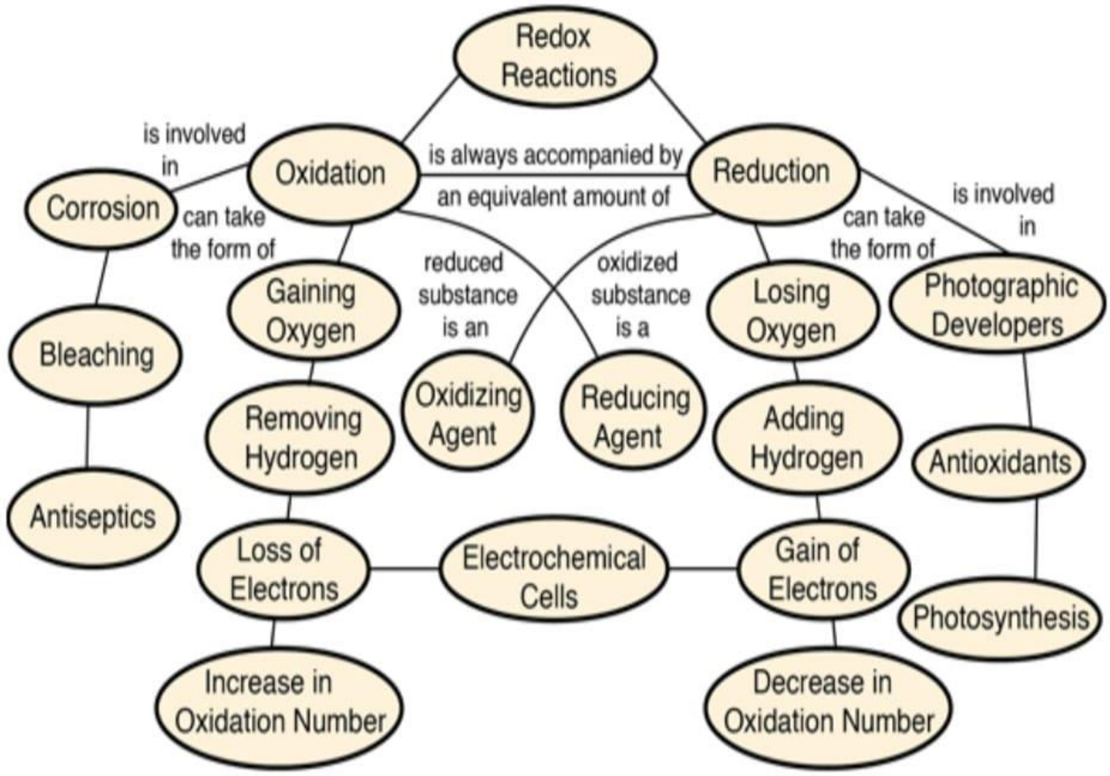
$Q_p = K_p = \frac{[p_C]^x[p_D]^y}{[p_A]^\ell[p_B]^m}$

Relation between K_b and K_c
$K_p = K_c(RT)^{\Delta n}$

Characteristics of equilibrium constant
- At a given temperature, the equilibrium constant for a particular reaction has a definite value.
- If reaction is reversed, value of equilibrium constant is also inversed
- Catalyst does not affect the value of equilibrium constant

Le Chatelier's principle
If a system in a equilibrium is subjected to change in factors that determine the equilibrium conditions of a system, the equilibrium shifts in a direction as to nullify the effect of change.
Effect of concentration
Concentration of reactant ↑ Forward direction
Concentration of product ↑ Backward direction
Effect of Pressure
If pressure ↑ reaction will favour in direction where no. of moles decrease.
Effect of temperature
- *Endothermic reaction*
  ↑ in temperature, shift equilibrium in forward direction
- *Exothermic reaction*
  ↑ in temperature, shifts equilibrium in backward direction

CH:7 REDOX REACTIONS



SOME KEY POINTS

SOME KEY POINTS

1. (Oxidation number: Charge on atom which appears on it when it is present in the combined state.

2. Sum of the oxidation states in a compound/ion should be equal to the zero or to the net charge on the ion.

3. Some elements show variable oxidation states.

4. Oxidation: The process in which electrons are lost.

5. Reduction: The process in which electrons are gained.6. Oxidizing agent: A substance which oxidize the other.7. Reducing agent: A substance which reduces the other.

8. Redox reaction: When oxidation and reduction take place together is known as redox reaction.

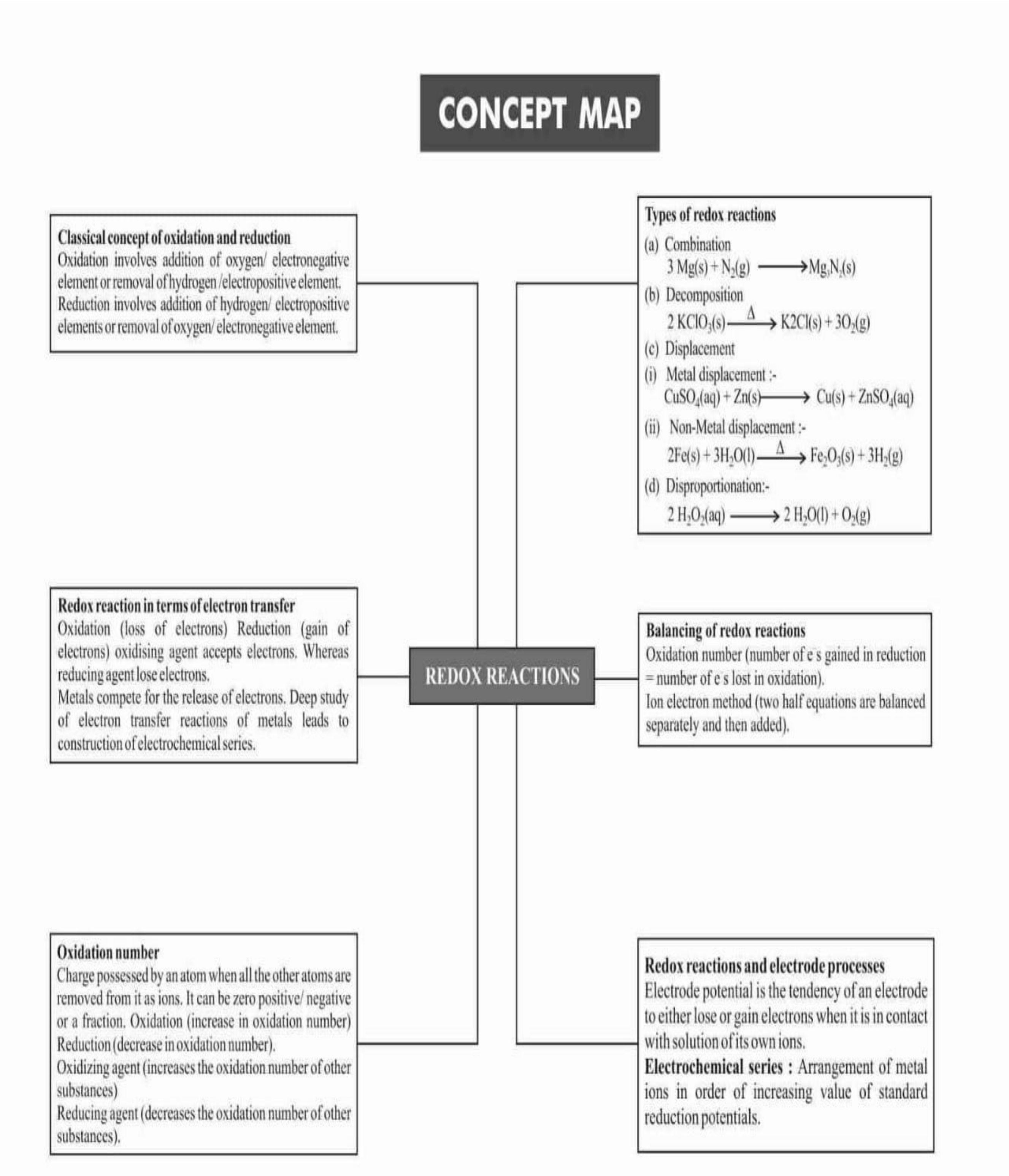
9. Disproportionation reaction: The reaction in which same species is simultaneously oxidised as well as reduced.

10. The change in oxidation state of any element in a compound is useful in calculating the equivalent weight.

11. Electrochemical series: Arrangement of E° red of different electrodes in increasing order of electrode potential.

12. Electrode Potential: The tendency of an electrode to lose or gain electrons is called electrode potential.

The standard electrode potentials of a large number of electrodes have been determined using standard hydrogen electrode as the reference electrode.



CONCEPT MAP

REDOX REACTIONS

Classical concept of oxidation and reduction
Oxidation involves addition of oxygen/ electronegative element or removal of hydrogen /electropositive element. Reduction involves addition of hydrogen/ electropositive elements or removal of oxygen/ electronegative element.

Types of redox reactions

(a) Combination
$3\,Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$

(b) Decomposition
$2\,KClO_3(s) \xrightarrow{\Delta} K2Cl(s) + 3O_2(g)$

(c) Displacement

(i) Metal displacement :-
$CuSO_4(aq) + Zn(s) \longrightarrow Cu(s) + ZnSO_4(aq)$

(ii) Non-Metal displacement :-
$2Fe(s) + 3H_2O(l) \xrightarrow{\Delta} Fe_2O_3(s) + 3H_2(g)$

(d) Disproportionation:-
$2\,H_2O_2(aq) \longrightarrow 2\,H_2O(l) + O_2(g)$

Redox reaction in terms of electron transfer
Oxidation (loss of electrons) Reduction (gain of electrons) oxidising agent accepts electrons. Whereas reducing agent lose electrons.
Metals compete for the release of electrons. Deep study of electron transfer reactions of metals leads to construction of electrochemical series.

Balancing of redox reactions
Oxidation number (number of e s gained in reduction = number of e s lost in oxidation).
Ion electron method (two half equations are balanced separately and then added).

Oxidation number
Charge possessed by an atom when all the other atoms are removed from it as ions. It can be zero positive/ negative or a fraction. Oxidation (increase in oxidation number) Reduction (decrease in oxidation number).
Oxidizing agent (increases the oxidation number of other substances)
Reducing agent (decreases the oxidation number of other substances).

Redox reactions and electrode processes
Electrode potential is the tendency of an electrode to either lose or gain electrons when it is in contact with solution of its own ions.
Electrochemical series : Arrangement of metal ions in order of increasing value of standard reduction potentials.

Chapter 8 some basic concept of organic chemistry

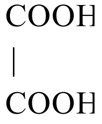
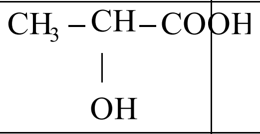
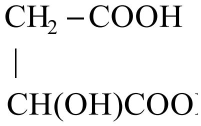
Nomenclature of Organic Compounds

Mainly three systems are adopted for naming an organic compound :

1. Common Name or Trivial System
2. Derived System
3. IUPAC system or Geneva system

Common or Trivial System

(*i*) On the basis of source from which they were obtained.

S.No.	Organic Compound	Trivial Name	Source
1.	CH_3OH	Wood spirit or Methyl spirit	Obtained by destructive distillation of wood.
2.	NH_2CONH_2	Urea	Obtained from urine.
3.	CH_4	Marsh gas (fire damp)	It was produced in marsh places.
4.	CH_3COOH	Vinegar	Obtained from Acetum – *i.e.,* Vinegar.
5.	 $\begin{array}{c}COOH\\ \vert\\ COOH\end{array}$	Oxalic acid	Obtained from oxalis plant.
6.	HCOOH	Formic acid	Obtained from Formicas [Red ant]
7.	 $CH_3-\underset{\underset{OH}{\vert}}{CH}-COOH$	Lactic acid	Obtained from lacteous (milk)
8.	 $\begin{array}{l}CH_2-COOH\\ \vert\\ CH(OH)COO\end{array}$	Malic acid	Obtain from Apple.
9.	$CH_3CH_2CH_2COOH$	Butyric acid	Obtained from butter.
10.	$CH_3(CH_2)_4COOH$	Caproic acid	Obtained from goats.
11.	C_2H_5OH	Grain alcohol	Obtained from barley.

(*ii*) On the basis of property

1. Glucose – Sweet in taste 2. Glycol – Sweet poisonous
3. Glycerol – Sweet

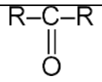
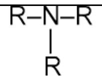
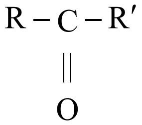
(*iii*) On the basis of discovery

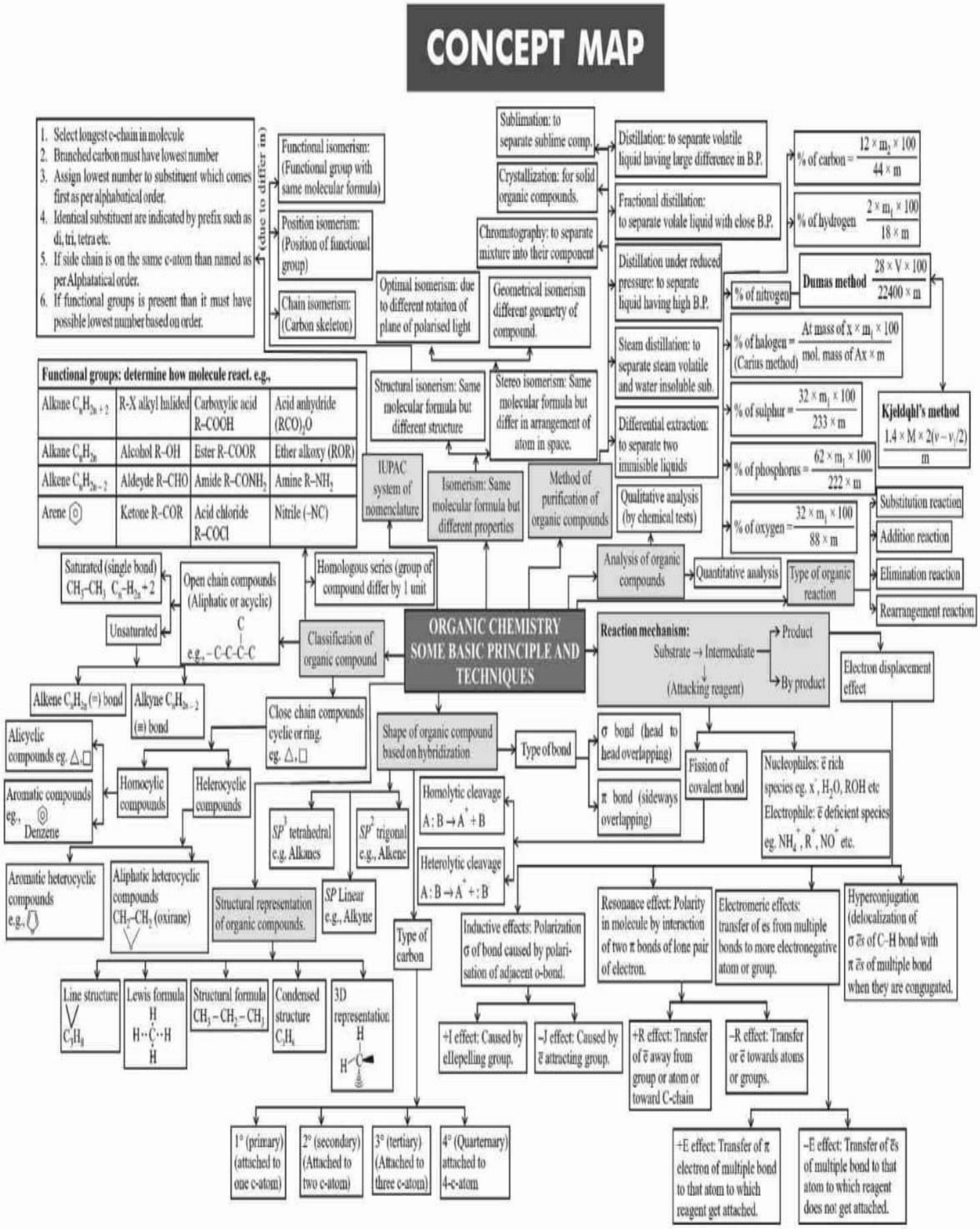
1. RMgx (Grignard Reagent)

(*iv*) On the basis of structure

S.No.	No. of Carbon atom	Word Root
(*i*)	1C	Meth
(*ii*)	2C	Eth
(*iii*)	3C	Prop
(*iv*)	4C	But
(*v*)	5C	Pent
(*vi*)	6C	Hex
(*vii*)	7C	Hept
(*viii*)	8C	Oct
(*ix*)	9C	Non
(*x*)	10C	Dec

Common Names for Hydrocarbon Derivatives

S.No.	Compound	Name
1.	R – X	Alkyl halide
2.	R – OH	Alkyl alcohol
3.	R – SH	Alkyl thio alcohol
4.	R – NH2	Alkyl amine
5.	R–O–R	Dialkyl ether
6.	 $R-\underset{\underset{O}{\|}}{C}-R$	Dialkyl ketone
7.	R–NH–R	Dialkyl amine
8.	 $R-\underset{\underset{R}{\|}}{N}-R$	Trialkyl amine
9.	R–O–R’	Alkyl alkyl’ ether
10.	 $R-\underset{\underset{O}{\|}}{C}-R'$	Alkyl alkyl’ ketone
11.	R–NH–R’	Alkyl alkyl’ amine


CONCEPT MAP
ORGANIC CHEMISTRY SOME BASIC PRINCIPLE AND TECHNIQUES
1. Select longest c-chain in molecule 2. Branched carbon must have lowest number 3. Assign lowest number to substituent which comes first as per alphabatical order. 4. Identical substituent are indicated by prefix such as di, tri, tetra etc. 5. If side chain is on the same c-atom than named as per Alphatatical order. 6. If functional groups is present than it must have possible lowest number based on order.
(due to differ in)
Functional isomerism: (Functional group with same molecular formula)
Position isomerism: (Position of functional group)
Chain isomerism: (Carbon skeleton)
Optimal isomerism: due to different rotaiton of plane of polarised light
Geometrical isomerism different geometry of compound.
Structural isonerism: Same molecular formula but different structure
Stereo isomerism: Same molecular formula but differ in arrangement of atom in space.
Isomerism: Same molecular formula but different properties
IUPAC system of nomenclature
Method of purification of organic compounds
Sublimation: to separate sublime comp.
Crystallization: for solid organic compounds.
Chromatography: to separate mixture into their component
Distillation: to separate volatile liquid having large difference in B.P.
Fractional distillation: to separate volale liquid with close B.P.
Distillation under reduced pressure: to separate liquid having high B.P.
Steam distillation: to separate steam volatile and water insoluble sub.
Differential extraction: to separate two immisible liquids
Qualitative analysis (by chemical tests)
Analysis of organic compounds
Quantitative analysis
% of carbon = 12 × m₂ × 100 / 44 × m
% of hydrogen 2 × m₁ × 100 / 18 × m
% of nitrogen
Dumas method 28 × V × 100 / 22400 × m
Kjeldqhl's method 1.4 × M × 2(v − v₁/2) / m
% of halogen = At mass of x × m₁ × 100 / mol. mass of Ax × m (Carius method)
% of sulphur = 32 × m₁ × 100 / 233 × m
% of phosphorus = 62 × m₁ × 100 / 222 × m
% of oxygen = 32 × m₁ × 100 / 88 × m
Type of organic reaction
Substitution reaction
Addition reaction
Elimination reaction
Rearrangement reaction
Functional groups: determine how molecule react. e.g.,
Alkane CₙH₂ₙ₊₂ | R-X alkyl halided | Carboxylic acid R-COOH | Acid anhydride (RCO)₂O
Alkane CₙH₂ₙ | Alcohol R-OH | Ester R-COOR | Ether alkoxy (ROR)
Alkene CₙH₂ₙ₋₂ | Aldeyde R-CHO | Amide R-CONH₂ | Amine R-NH₂
Arene | Ketone R-COR | Acid chloride R-COCl | Nitrile (-NC)
Homologous series (group of compound differ by 1 unit
Saturated (single bond) CH₃-CH₃ CₙH₂ₙ+2
Unsaturated
Open chain compounds (Aliphatic or acyclic) e.g., – C-C-C-C
Classification of organic compound
Alkene CₙH₂ₙ (=) bond
Alkyne CₙH₂ₙ₋₂ (≡) bond
Close chain compounds cyclic or ring. eg. △,□
Alicyclic compounds eg. △,□
Aromatic compounds eg., Denzene
Homocylic compounds
Helerocyclic compounds
Aromatic heterocyclic compounds e.g.,
Aliphatic heterocyclic compounds CH₂-CH₂ (oxirane)
Structural representation of organic compounds.
Line structure C₃H₈
Lewis formula H H··C··H H
Structural formula CH₃ – CH₂ – CH₃
Condensed structure C₃H₈
3D representation H H C
Shape of organic compound based on hybridization
SP³ tetrahedral e.g. Alkanes
SP² trigonal e.g., Alkene
SP Linear e.g., Alkyne
Type of carbon
1° (primary) (attached to one c-atom)
2° (secondary) (Attached to two c-atom)
3° (tertiary) (Attached to three c-atom)
4° (Quarternary) attached to 4-c-atom
Type of bond
σ bond (head to head overlapping)
π bond (sideways overlapping)
Homolytic cleavage A : B → A˙ + B˙
Heterolytic cleavage A : B → A⁺ + : B⁻
Reaction mechanism: Substrate → Intermediate → Product → By product (Attacking reagent)
Fission of covalent bond
Nucleophiles: ē rich species eg. X⁻, H₂O, ROH etc Electrophile: ē deficient species eg. NH₄⁺, R⁺, NO⁺ etc.
Electron displacement effect
Inductive effects: Polarization σ of bond caused by polari-sation of adjacent σ-bond.
+I effect: Caused by ellepelling group.
–J effect: Caused by ē attracting group.
Resonance effect: Polarity in molecule by interaction of two π bonds of lone pair of electron.
+R effect: Transfer of ē away from group or atom or toward C-chain
–R effect: Transfer or ē towards atoms or groups.
Electromeric effects: transfer of es from multiple bonds to more electronegative atom or group.
+E effect: Transfer of π electron of multiple bond to that atom to which reagent get attached.
–E effect: Transfer of ēs of multiple bond to that atom to which reagent does not get attached.
Hyperconjugation (delocalization of σ ēs of C–H bond with π ēs of multiple bond when they are congugated.

CHAPTER 9: HYDROCARBONS

Classification of Hydrocarbons

- Saturated (Alkanes): Single C-C bonds (e.g., Methane, Ethane).
- Unsaturated (Alkenes/Alkynes): Double or triple C-C bonds (e.g., Ethene, Ethyne).
- Aromatic (Arenes): Cyclic, stable with alternating double bonds (e.g., Benzene).

2. Alkanes (Saturated)

- General Formula: CnH_{2n+2}
- Isomerism: Chain, position, conformational (eclipsed/staggered).
- Preparation: Hydrogenation, Wurtz Reaction, Decarboxylation.
- Properties: Relatively unreactive, undergo free radical substitution (halogenation), combustion.

3. Alkenes (Unsaturated)

- General Formula: C_nH_{2n} sp^2 ,hybridization, trigonal planar geometry.
- Isomerism: Geometrical (cis-trans).
- Reactions: Electrophilic addition (Markownikoff's Rule, Anti-Markownikoff), hydration, polymerization.

4. Alkynes (Unsaturated)

- General Formula: CnH_{2n-2}

.sphybridization, linear geometry.

- Reactions: Electrophilic addition (less reactive than alkenes), hydration, polymerization.

5. Aromatic Hydrocarbons (Arenes)

- Benzene C_6H_6

: Cyclic, planar, resonance, aromaticity (Hückel's Rule).

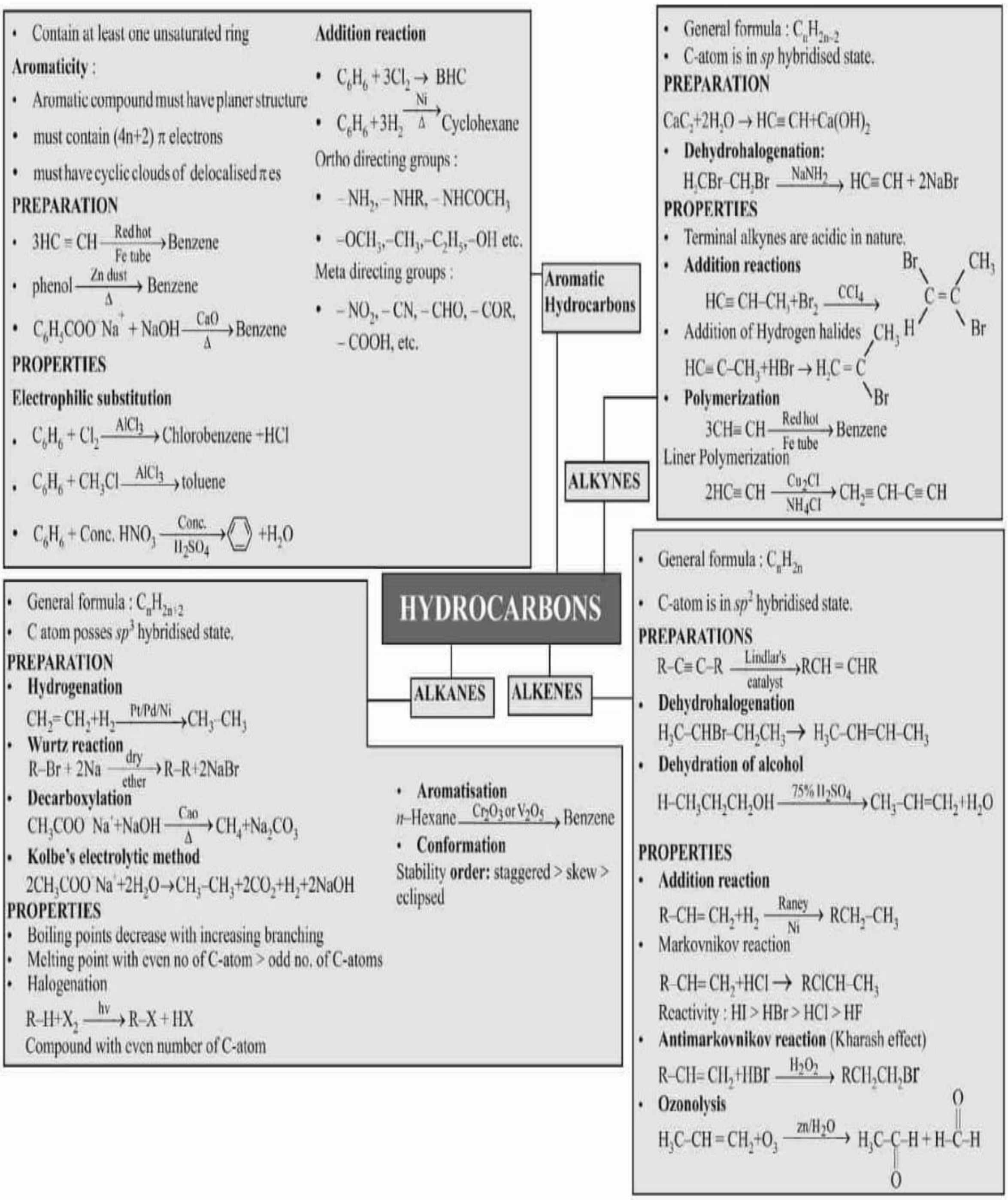
- Reactions: Electrophilic Substitution (Nitration, Halogenation, Sulphonation, Friedel-Crafts Alkylation/Acylation).
- Directive Influence: Ortho-, Para-, Meta-directors.

6. Important Concepts & Uses

- Isomerism: Crucial for distinguishing isomers (e.g., N-pentane vs. Neopentane M.P./B.P.).
- Conformational Isomerism: Eclipsed vs. Staggered (Ethane).
- Uses: Fuels (LPG, CNG, Petrol), solvents, polymers (Polythene).

Toxicity/Environmental Impact: Mentioned in context of fuels.

CONCEPT MAP



HYDROCARBONS

Aromatic Hydrocarbons

- Contain at least one unsaturated ring

Aromaticity :

- Aromatic compound must have planer structure
- must contain (4n+2) π electrons
- must have cyclic clouds of delocalised π es

PREPARATION

- $3HC\equiv CH \xrightarrow[\text{Fe tube}]{\text{Red hot}} \text{Benzene}$
- $\text{phenol} \xrightarrow[\Delta]{\text{Zn dust}} \text{Benzene}$
- $C_6H_5COO^-Na^+ + NaOH \xrightarrow[\Delta]{CaO} \text{Benzene}$

PROPERTIES

Electrophilic substitution

- $C_6H_6 + Cl_2 \xrightarrow{AlCl_3} \text{Chlorobenzene} + HCl$
- $C_6H_6 + CH_3Cl \xrightarrow{AlCl_3} \text{toluene}$
- $C_6H_6 + \text{Conc. } HNO_3 \xrightarrow[H_2SO_4]{\text{Conc.}}$ ⌬ $+H_2O$

Addition reaction

- $C_6H_6 + 3Cl_2 \rightarrow BHC$
- $C_6H_6 + 3H_2 \xrightarrow[\Delta]{Ni} \text{Cyclohexane}$

Ortho directing groups :

- $-NH_2, -NHR, -NHCOCH_3$
- $-OCH_3, -CH_3, -C_2H_5, -OH$ etc.

Meta directing groups :

- $-NO_2, -CN, -CHO, -COR, -COOH$, etc.

ALKYNES

- General formula : C_nH_{2n-2}
- C-atom is in *sp* hybridised state.

PREPARATION

$CaC_2 + 2H_2O \rightarrow HC\equiv CH + Ca(OH)_2$

- **Dehydrohalogenation:**

$H_2CBr-CH_2Br \xrightarrow{NaNH_2} HC\equiv CH + 2NaBr$

PROPERTIES

- Terminal alkynes are acidic in nature.
- **Addition reactions**

$HC\equiv CH-CH_3 + Br_2 \xrightarrow{CCl_4} \underset{H}{\overset{Br}{>}}C=C\underset{Br}{\overset{CH_3}{<}}$

- Addition of Hydrogen halides

$HC\equiv C-CH_3 + HBr \rightarrow H_2C=C\underset{Br}{\overset{CH_3}{<}}$

- **Polymerization**

$3CH\equiv CH \xrightarrow[\text{Fe tube}]{\text{Red hot}} \text{Benzene}$

Liner Polymerization

$2HC\equiv CH \xrightarrow[NH_4Cl]{Cu_2Cl} CH_2\equiv CH-C\equiv CH$

ALKANES

- General formula : C_nH_{2n+2}
- C atom posses sp^3 hybridised state.

PREPARATION

- **Hydrogenation**

$CH_2=CH_2 + H_2 \xrightarrow{Pt/Pd/Ni} CH_3-CH_3$

- **Wurtz reaction**

$R-Br + 2Na \xrightarrow[\text{ether}]{\text{dry}} R-R + 2NaBr$

- **Decarboxylation**

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

- **Kolbe's electrolytic method**

$2CH_3COO^-Na^+ + 2H_2O \rightarrow CH_3-CH_3 + 2CO_2 + H_2 + 2NaOH$

PROPERTIES

- Boiling points decrease with increasing branching
- Melting point with even no of C-atom > odd no. of C-atoms
- Halogenation

$R-H + X_2 \xrightarrow{h\nu} R-X + HX$

Compound with even number of C-atom

- **Aromatisation**

$n\text{-Hexane} \xrightarrow{Cr_2O_3 \text{ or } V_2O_5} \text{Benzene}$

- **Conformation**

Stability **order:** staggered > skew > eclipsed

ALKENES

- General formula : C_nH_{2n}
- C-atom is in sp^2 hybridised state.

PREPARATIONS

$R-C\equiv C-R \xrightarrow[\text{catalyst}]{\text{Lindlar's}} RCH=CHR$

- **Dehydrohalogenation**

$H_3C-CHBr-CH_2CH_3 \rightarrow H_3C-CH=CH-CH_3$

- **Dehydration of alcohol**

$H-CH_2CH_2CH_2OH \xrightarrow{75\%\ H_2SO_4} CH_3-CH=CH_2 + H_2O$

PROPERTIES

- **Addition reaction**

$R-CH=CH_2 + H_2 \xrightarrow[Ni]{Raney} RCH_2-CH_3$

- Markovnikov reaction

$R-CH=CH_2 + HCl \rightarrow RClCH-CH_3$

Reactivity : HI > HBr > HCl > HF

- **Antimarkovnikov reaction** (Kharash effect)

$R-CH=CH_2 + HBr \xrightarrow{H_2O_2} RCH_2CH_2Br$

- **Ozonolysis**

$H_3C-CH=CH_2 + O_3 \xrightarrow{zn/H_2O} H_3C-\underset{\underset{O}{\|}}{C}-H + H-\overset{\overset{O}{\|}}{C}-H$

SAMPLE PAPER

Time: 3hrs **Examinations march, 2026 Class 10+1 (CHEMISTRY)** **M. M:70**

NOTE: Section-A has 20 parts carrying 1 mark each.
Section-B has 14 questions carry 2 marks each.

Section-C has 4 questions carry 3 marks each.
Section-D has 2 questions carry 5 marks each.
1. SECTION-A (each question carry one mark)

i. Significant figures in 0.00051 are
(a) 5 (b) 3 (c) 2 (d) 6
ii. Which of the following is independent of temperature?
(a) Molarity (b) Molality (c) Both (a) & (b) (d) None of these
iii. The number of moles present in 6 gm of carbon is:
(a) 27 (b) 0.5 (c) 5 (d) 15
iv. The state of the system is described by its:
(a) Pressure(P) (b) Volume(V)
(c) Pressure and Volume (d) Pressure(P) Volume(V) and temp(T)
v. Acidic solutions possess a concentration of hydrogen ion:
(a) $[H^+] = 10^{-7}$M (b) $[H^+] < 10^{-7}$M (c) $[H^+] > 10^{-}7$M (d) None of these
vi. What is the oxidation state of K in KO_2?
(a) +1 (b) +2 (c) 0 (d) -2
vii. Which of the following have sp2-hybridisation.
(a) C_2H_4 (b) CH_4 (c) NH_3 (d) C_2H_2
viii. Name the structure of PCl_5.
ix. Calculate the formal charge on oxygen in the structure of O_3 (ozone).
x. State Le Chatelier's Principle.
True/ False
xi. Enthalpy change is always negative for an endothermic process.
xii. Entropy always increases in a spontaneous process.
xiii. A reducing agent is itself oxidized during the reaction.
xiv. In organic compounds, the carbon atom can form bonds with a maximum of four other atoms.

xv. Aldehydes are more reactive than ketones due to the presence of a hydrogen atom attached to the carbonyl group.

Comprehension Passage
Hydrogen has resemblance to alkali metals, which lose one electron to form unipositive ions, as well as with halogens, which gain one electron to form uni-negative ion. Like alkali metals, hydrogen form oxides, halides and sulphides. However, unlike metals, it has a very high ionization enthalpy and does not possess metallic characteristics under normal conditions. In fact, in terms of ionization enthalpy, hydrogen ΔiH of Li is 520 $kJmol^{-1}$. F is 1689 $kJmol^{-1}$ and that of H is 1312 $kJmol^{-1}$. Like halogens, it forms a diatomic molecules. However, in terms of reactivity, it is very low as compared to halogens. In spite of the fact that hydrogen, to a certain extent resembles both with alkali metals and halogens.
xvi. In what way hydrogen resembles to alkali metals?
xvii. How hydrogen form unipositive as well as uni-negative ions?
xviii. Under what conditions hydrogen resembles with halogens?

xix. Why hydrogen is placed separately in the periodic table?

xx. Hydrogen, to a certain extent resembles both with alkali metals and halogens. Explain.

SECTION- B (each question carry two marks)

2. Calculate the molar mass of calcium carbonate ($CaCO_3$) and HNO_3

OR

How many moles are present in 5.6 litres of oxygen gas at STP?

3. Calculate the total number of electrons present in 1.4g of dinitrogen gas.
4. What are the main postulates of Bohr's atomic model?
5. What do you mean by the diagonal relationship of elements in the periodic table? Explain with example.
6. Write the resonance structures of CO_3^{2-} And NO_2.
7. Derive the relationship between Cp and Cv for an Ideal gas.
8. What is Gibbs free energy? Write its formula.
9. Define Law of Chemical Equilibrium with example.
10. Is decrease in enthalpy a criterion for Spontaneity of a reaction?
11. Given: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, Δ r H- = -92.4KJ mol-1. What is standard enthalpy of formation of NH_3 gas?

OR

If the equilibrium constant for the reaction $H_2(g)+I_2(g)\rightleftharpoons 2HI(g)$ is 50 at a certain temperature, and the initial concentrations of H_2, I_2, and HI are 0.1 M, 0.1 M, and 0.2 M, respectively, calculate the equilibrium concentrations.

12. What is meant by an oxidizing agent and a reducing agent? Explain with an example.

OR

Explain why chlorine is a stronger oxidizing agent than iodine.

13. Why Benzene is extraordinarily stable though it contains three double bonds?

OR

What is the difference between a structural isomer and a stereoisomer? Explain with examples.

14. Differentiate between saturated and unsaturated hydrocarbons with examples.
15. Write the IUPAC names of the following compounds:

a) 2,3-dimethylbutane b) 1-bromo-3-chlorocyclohexane

SECTION- C (each question carry three marks)

16. Calculate the mass percent of calcium, phosphorus and oxygen in calcium phosphate $Ca_3(PO_4)_2$.
17. Calculate the frequency and energy of light associated with a photon of wavelength 500 nm.

OR

Calculate the wavelength, frequency and wave number of a wave whose time period is $2x10^{-10}$s.

18. Balance the following redox reactions:

a) $MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$ (in acidic solution)

b) $Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$

OR

What are the characteristics of a redox reaction? Explain with a suitable example how a reaction can be both oxidation and reduction.

19. Write structure of

I) 2-Methylbuta-1,3-diene.

II) But-4-ol-1yne.

SECTION- D (each question carry five marks)

20. Explain the concept of bond order. Calculate the bond order, stability, magnetic property and electronic configuration of O_2 molecule on the basis of Molecular Orbital Theory.

OR

What is the VSEPR theory? Explain how it helps predict the shapes of molecules with examples like NH_3, H_2O and CH_4.

21. Write the following Reactions:

I) Ethyne into Benzene.
II) Acetylene into Benzene.
III) Propyne into propane IV) Wurtz Reaction
V) Kolbe's Reaction (1x5=5)

OR

(I) Why do alkenes undergo Electrophilic Addition Reaction?
(II) Explain Aromatization of alkanes containing six to eight carbon atoms with example.
(III) Define Polymerization with an example of Ethene. (2+2+1)

STATE RESOURCE TEAM CHEMISTRY