CAREERS 360
PREPARATION Series

Class 11-12

Chemistry
Formula Book

Content

Chapter 1- Some Basic Concepts in Chemistry

Important Formulas
1. Molar Mass (M):

$\text{Molar mass } (\mathrm{g/mol}) = \text{Mass of 1 mole of a substance}$
2. Number of Moles (n):

$$n = \frac{\text{Given mass (g)}}{\text{Molar mass (g/mol)}}$$

3. Number of Particles (atoms/molecules/ions):

$$\text{Number of particles } = n \times N_A = \frac{\text{Given mass}}{\text{Molar mass}} \times 6.022 \times 10^{23}$$

4. Volume of Gas at STP (Standard Temperature and Pressure):

1 mole of gas at STP $= 22.4$ L

$$\text{Volume (L)} = n \times 22.4$$

5. Percentage Composition of Element in a Compound:

$$\% \text{ of element } = \frac{\text{Mass of element in 1 mole}}{\text{Molar mass of compound}} \times 100$$

6. Empirical Formula Calculation
Step 1- Divide % by atomic mass
Step 2- Divide all values by the smallest result

Step 3- Multiply to get whole numbers if needed

7. Molecular Formula:

$$\text{Molecular formula } = \mathrm{n} \times \text{ Empirical formula}$$

where

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}}$$

8. Limiting Reagent Concept:
Determine moles of each reactant
Use balanced equation to find the one that gets used up first

9. Concentration (Molarity):

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Volume of solution in L}}$$

10. Dilution Formula:

$$M_1V_1 = M_2V_2$$

Significant figures are meaningful digits which are known with certainty plus one which is estimated or uncertain.

Rules for Significant Figures

1. **All non-zero digits** are significant.
   Example: 123 → 3 significant figures

2. **Zeros between non-zero digits** are significant.
   Example: 1003 → 4 significant figures

3. **Leading zeros** (before non-zero digits) are **not** significant.
   Example: 0.0045 → 2 significant figures

4. **Trailing zeros** after decimal point are significant.
   Example: 12.300 → 5 significant figures

5. **Trailing zeros in a whole number without decimal** are **not** significant.
   Example: 1500 → 2 significant figures

Important Points to Remember

Law of Conservation of Mass-
Mass is neither created nor destroyed in a chemical reaction.

Law of Definite Proportions-
A given compound always contains exactly the same proportion of elements by mass.

Law of Multiple Proportions-
When two elements form more than one compound, the masses of one element combining with fixed mass of the other are in whole-number ratios.

Atomic Mass Unit (amu or u):
$1 \mathrm{amu} = \frac{1}{12}$ th of the mass of a carbon-12 atom

Avogadro's Number (NA):

6.022×10^{23} particles $/ \mathrm{mol}$
Stoichiometry:

Use balanced equations to calculate masses, volumes and particles involved.

Significant Figures & Scientific Notation:

- Be careful with rounding
- Use appropriate significant figures in final answers

Chapter 2- Structure of Atoms

Important Formulas
1. Wavelength-Frequency Relationship:

$$c = \lambda \nu$$

where:
$c =$ speed of light $= 3 \times 10^{8} \mathrm{~m} / \mathrm{s}$
$\lambda =$ wavelength (m)
$\nu =$ frequency (Hz)

2. Energy of a Photon (Planck's Equation):

$$E = h\nu = \frac{hc}{\lambda}$$

$h =$ Planck's constant $= 6.626 \times 10^{-34}$Js

3. Bohr's Radius (first orbit of H -atom):

$r_n = 0.529 \times \frac{n^2}{Z} A$

4. Energy of Electron in nth Orbit (Hydrogen-like atom):

$E_n = -13.6 \times \frac{Z^2}{n^2} \mathrm{eV}$

5. Wave Number (for Hydrogen spectral lines):

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

$R = 1.097 \times 10^7 \mathrm{~m}^{-1}$ (Rydberg constant)
6. de Broglie Wavelength:

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

(Wave nature of particles)

Key Concepts

Subatomic Particles

- Proton (p): +1 charge, mass ≈ 1 amu
- Neutron (n): 0 charge, mass ≈ 1 amu
- Electron (e): –1 charge, negligible mass

Isotopes, Isobars and Isotones

- **Isotopes:** Same atomic number, different mass numbers
- **Isobars:** Same mass number, different atomic numbers
- **Isotones:** Same number of neutrons

Quantum Numbers

- **n (principal):** shell, energy level
- **l (azimuthal):** subshell, shape (0 = s, 1 = p...)
- **m (magnetic):** orbital orientation
- **s (spin):** +½ or –½

Aufbau Principle: Fill lower energy orbitals first

Pauli's Exclusion Principle: No two electrons can have same 4 quantum numbers

Hund's Rule: Pairing happens after all orbitals in a subshell are singly filled

Electronic configuration is the arrangement of electrons in the orbitals (energy levels and sublevels) of an atom in a specific order, following the principles of quantum mechanics.

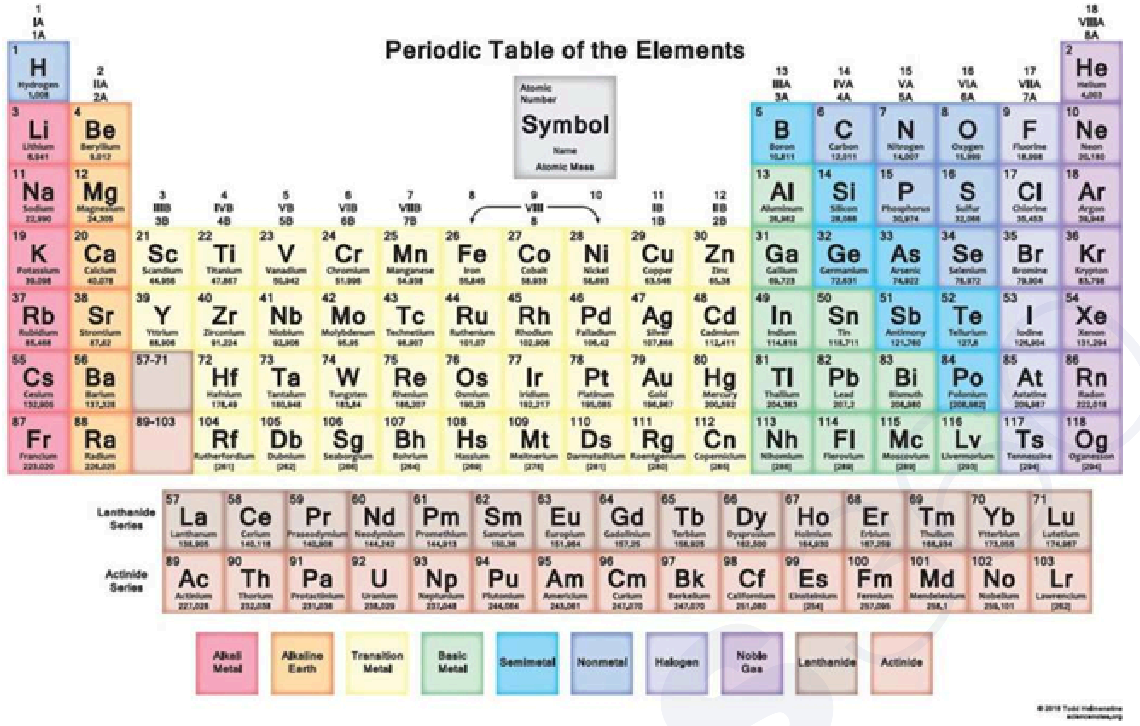
- e.g., Carbon = $1s^2\ 2s^2\ 2p^2$

Chapter 3- Classification of Elements and periodicity in properties

Historical Periodic Laws

1. **Dobereiner's Triads** (1817): The atomic mass of the middle element was roughly the average of the other two.
    - Example: Li (7), Na (23), K (39)
2. **Newlands' Law of Octaves** (1864):
    - Every eighth element had properties similar to the first when arranged by atomic mass.
3. **Mendeleev's Periodic Law** (1869):
    - **"Properties of elements are a periodic function of their atomic masses."**
    - Limitations: Position of isotopes, hydrogen, and no place for noble gases.
4. **Modern Periodic Law** (Moseley, 1913): "*Properties of elements are a periodic function of their atomic numbers (Z).*"

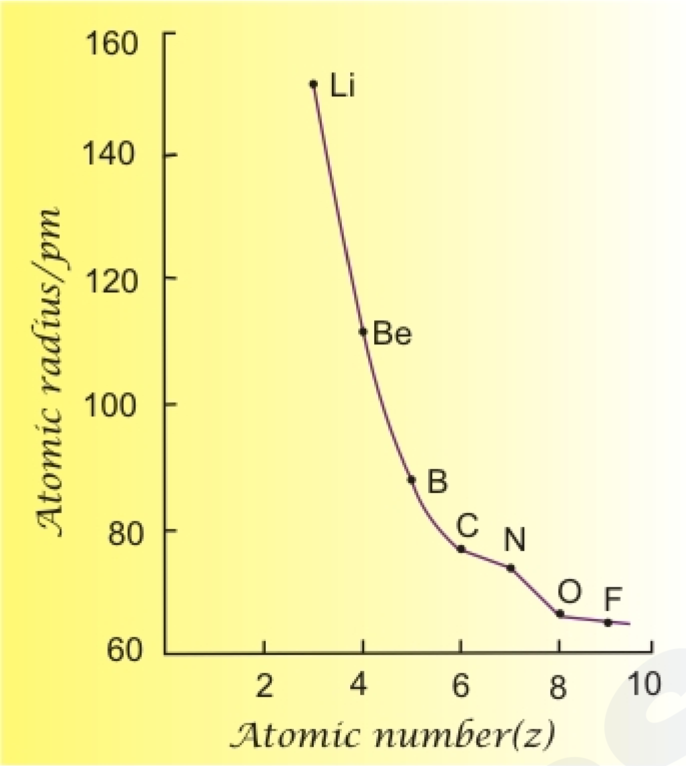
- **Long Form of Periodic Table:**
    - **Groups-** 18 vertical columns, **Periods-** 7 horizontal rows
    - Elements are arranged in order of increasing atomic number.
- **Types of Elements:**
    - **s-block:** Groups 1 and 2
    - **p-block:** Groups 13 to 18
    - **d-block:** Transition metals (Groups 3 to 12)
    - **f-block:** Inner transition metals (Lanthanides and Actinides)



Comparison of the ionic radii and atomic radii

Variation in a Period

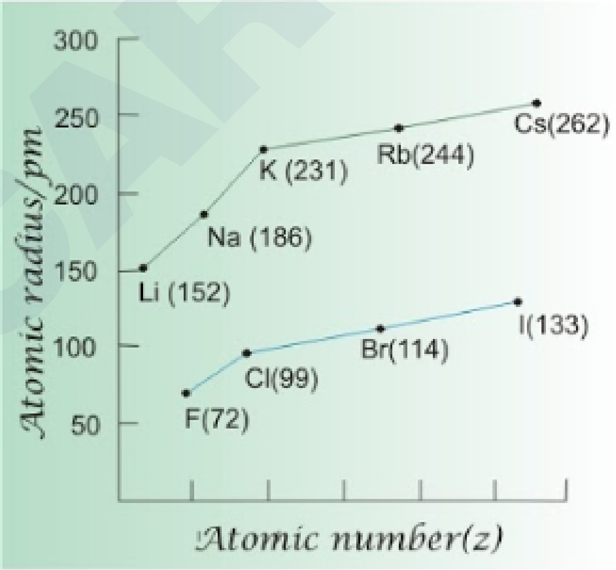
In moving from left to right in a period, the effective nuclear charge increase and the atomic size decreases in the period.



Variation of atomic radius with atomic number across the second period

Variation in a Group

In moving from top to bottom in a group, the number of shells increases due to which the atomic size increases.



Variation of atomic radius with atomic number for alkali metals and halogens

- $M^{+3} < M^{+2} < M^{+} < M$
- $M^{-3} > M^{-2} > M^{-} > M$

The Screening effect or Shielding effect

The decrease in the force of attraction between the outer electrons and the nucleus due to the presence of inner electrons is called the screening effect or shielding effect.

Calculation of the screening effect

- **For ns or np orbital electrons**
  - All electrons in the (ns, np) group contribute to 0.35 each to the screening effect constant. Except for 1s electrons which contribute by 0.30.
  - All electrons in (n-1) shell contribute by 0.85 each to the screening effect constant.
  - All electrons in (n-2) shell or lower contribute by 1.0 each to the screening effect constant.
- **For d- or f-electrons**
  - All electrons in the (ns, np) group contribute to 0.35 each to the screening effect constant.
  - All the electrons in groups lower than (nd, nf) contribute by 1.0 each to the screening effect.

Effective Nuclear Charge

$Z^{\star} = (Z - \sigma)$, where σ is the screening effect constant.

Isoelectronic species -

A series of atoms, ions and molecules in which each species contains the same number of electrons but a different nuclear charge.

e.g. N^{3-}, O^{2-}, F^{-}, Ne, Na^{+}, Mg^{2+}, Al^{3+}

Ionisation Enthalpy

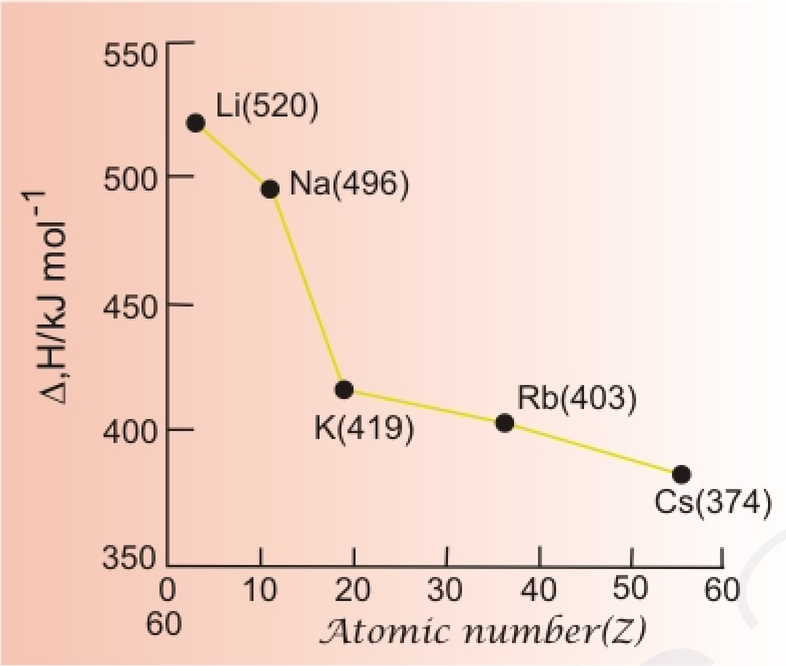
The minimum energy required to remove the most loosely bound electron from an isolated gaseous atom to convert it into a gaseous monovalent positive ion.

$$M(g) \xrightarrow{IE_1} M^{+}(g) + e$$

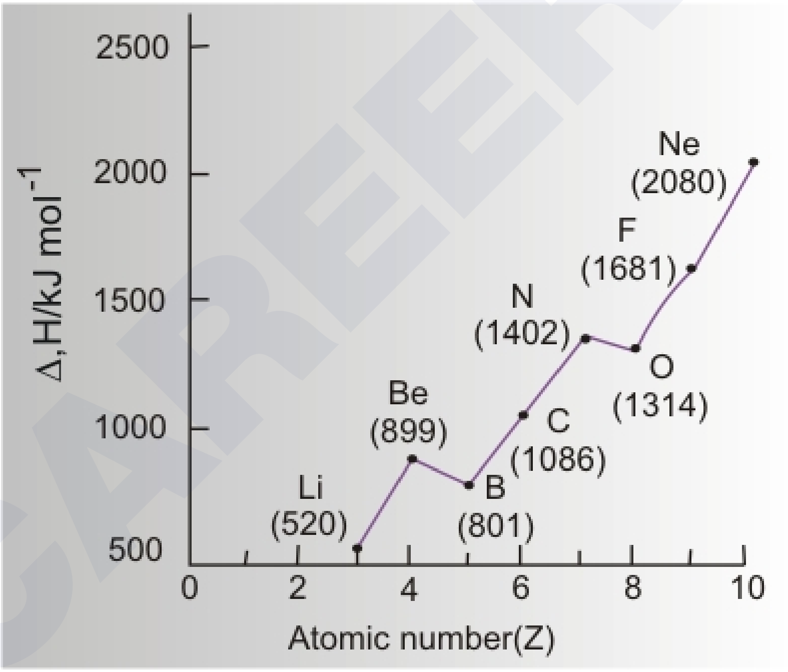
IE_1 is ionisation enthalpy or also known as first ionisation enthalpy.

Variation of Ionisation Enthalpy

- In moving down the group, the ionisation enthalpy decreases
- The elements with atomic number from 73 to 82 have higher ionisation enthalpy than the earlier elements in their respective group. This deviation of behaviour is because of the lanthanide contraction.



- In moving from left to right in a period, the ionisation enthalpy increases.
- For every element, the successive ionisation energy increases. This is because of the increase in the nuclear charge due to the successive removal of electrons.



Electron Gain Enthalpy $(\Delta_{eg}H)$

The energy change that occurs when an electron is added to a neutral gaseous atom to form a negative ion.

$A(g) + e \rightarrow A^-(g) + \Delta_{eg}H$

Variation of Electron Gain Enthalpy

- Chlorine has the most negative electron gain enthalpy value than fluorine. Because fluorine is very small in size due to which there is a very strong inter-electronic repulsion for the incoming electron, thus its electron gain enthalpy is less than chlorine.

Chapter 4 - Chemical Bonding and Molecular Structure

Why do Atoms Combine?

- To attain **octet configuration** (noble gas-like stability).
- Lower potential energy and increased stability.

Type	**Description**	**Example**
Ionic Bond	Complete transfer of electrons (metal + non-metal)	NaCl, MgO
Covalent Bond	Sharing of electrons between atoms (usually non-metals)	H_2O, CH_4, O_2
Coordinate Bond	One atom donates both electrons to the bond	NH_4^+, $[BF_4]^-$

Octet Rule

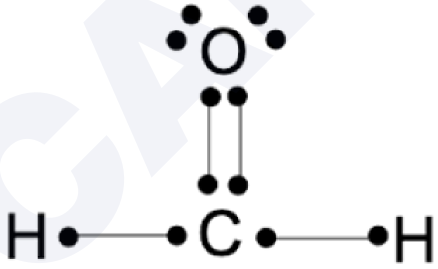
Atoms tend to have 8 electrons in their valence shell after bonding. Although there are some exceptions-

- Expanded octet (e.g. PCl_5, SF_6),
- Odd-electron molecules (e.g. NO)

Lewis Structures

Visual representations using dots for electrons.
Helps determine bonding and lone pairs.

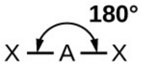
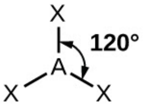
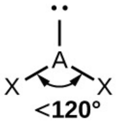
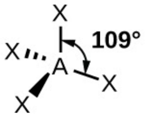
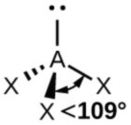
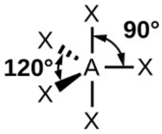
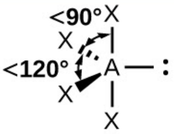
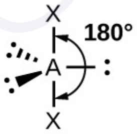
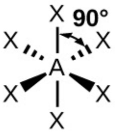
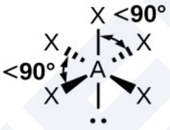
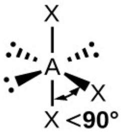
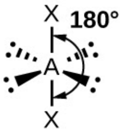
Example-



Formal Charge

$$\text{Formal Charge} = (\text{ Valence electrons }) - (\text{ Lone pair electrons }) - \frac{1}{2}\,(\text{Bonding electrons})$$

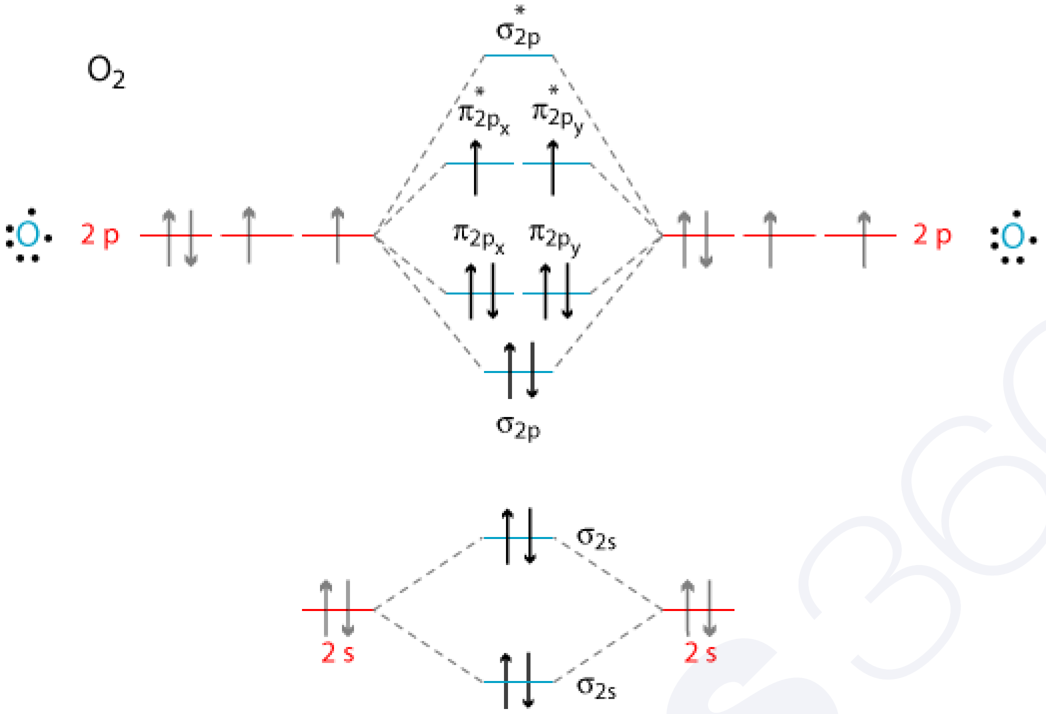
VSEPR Theory

Number of electron pairs	Electron pair geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramid	 Bent or angular		
5	 Trigonal bipyramid	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square pyramid	 Square planar	 T-shape	 Linear

Molecular orbital theory

Molecular orbital theory describes the distribution of electrons in molecules in much the same way that the distribution of electrons in atoms is described using atomic orbitals.

For example, the electronic configuration of O_2 molecule is given below:



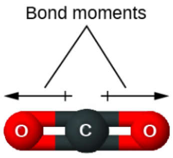
Thus, the electronic configuration of O_2 molecule can be written as:

$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^1 \pi^* 2p_y^1$

Dipole moment (μ)

$$\mu = Qr$$

where Q is the magnitude of the partial charges (determined by the electronegativity difference) and r is the distance between the charges



Overall dipole moment = 0
(a)

Hydrogen Bonding

Greater the electronegativity of the atom will result in an increase in hydrogen-bond strength.

Hydrogen bonds are responsible for holding together DNA, proteins, and other macromolecules.

Chapter 5 - Thermodynamics

Thermodynamics deals with the **energy changes** in chemical reactions and physical processes.

Key terms

Term	Definition
System	Part of the universe under study
Surroundings	Everything outside the system
Open system	Can exchange both matter and energy
Closed system	Exchanges energy but not matter
Isolated system	No exchange of energy or matter
State function	It depends only on the initial and final state (e.g., internal energy, enthalpy)
Path function	Depends on the path taken (e.g., heat q, work w)

Internal energy

Total energy (kinetic + potential) of a system.
Change in internal energy:

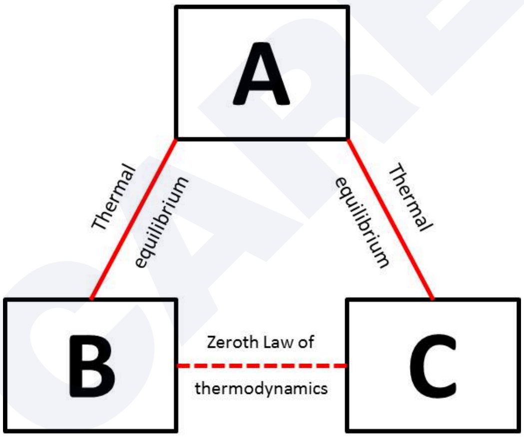
$$\Delta U = q + w$$

Work (w)

Work done by/on the system during expansion or compression:

$$w = -P\Delta V$$

The Zeroth Law of Thermodynamics-



First Law of Thermodynamics

Energy can neither be created nor destroyed; it can only be transformed. Mathematical form:

$$\Delta U = q + w$$

Enthalpy (H) Heat content of a system at constant pressure:

$$H = U + PV$$

Change in enthalpy:

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

Types of Process

Type	Description
Isothermal	Temperature remains constant (ΔT = 0)
Adiabatic	No heat exchange (q = 0)
Isobaric	Constant pressure
Isochoric	Constant volume
Reversible	Happens infinitely slowly
Irreversible	Fast and not easily reversed

Heat capacity (C):

$$C = \frac{q}{\Delta T}$$

Molar heat capacity: Heat required to raise temp. of 1 mol by $1^\circ\mathrm{C}$
Specific heat (c): Per gram basis

$$q = mc\Delta T$$

Spontaneous and Non-spontaneous Processes

- **Spontaneous:** Happens on its own (e.g. ice melting at room temp)
- **Non-spontaneous:** Needs external help

Second Law of Thermodynamics
Entropy (randomness) of the universe increases for spontaneous processes.

Represented as:

$$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$$

$$\Delta G = \Delta H - T\Delta S$$

Gibbs Free Energy (G)

$$\Delta G = \Delta H - T\Delta S$$

ΔG value	Interpretation

ΔG < 0	Spontaneous
ΔG > 0	Non-spontaneous
ΔG = 0	Equilibrium

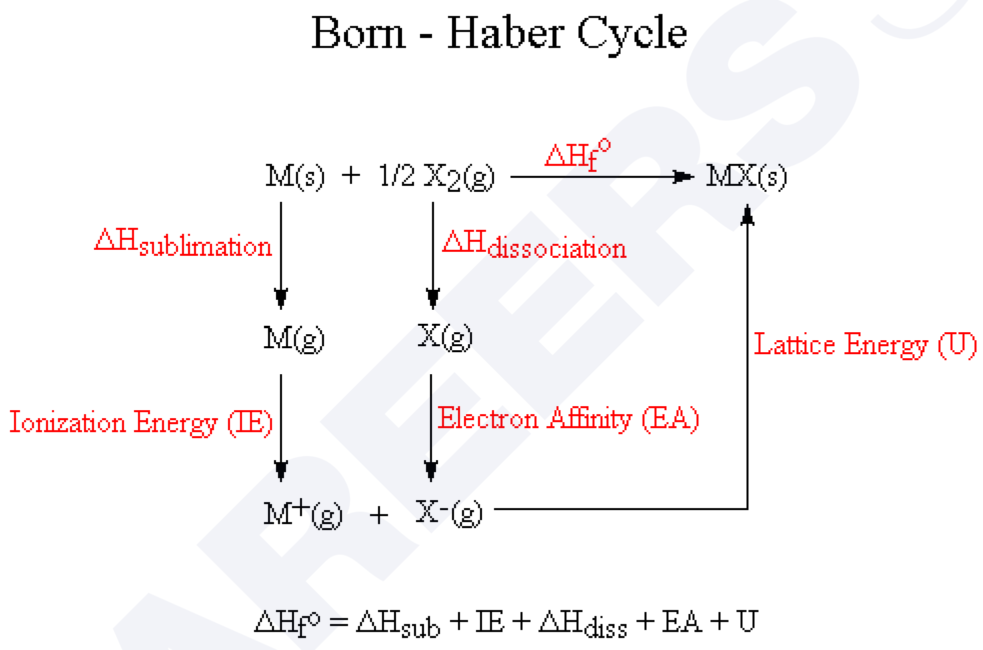
Important Relationships

At constant pressure:

$$q_p = \Delta H$$

At constant volume:

$$q_v = \Delta U$$



Born - Haber Cycle

$$\Delta H_f° = \Delta H_{sub} + IE + \Delta H_{diss} + EA + U$$

Hess's Law

According to Hess's Law, "The enthalpy change for a process is independent of path or way of a process.

$$\begin{tikzcd}
A \arrow[rr, "\Delta H_o"] \arrow[d, "\Delta H_1"'] & & E \\
B \arrow[r, "\Delta H_2"'] & C \arrow[r, "\Delta H_3"'] & D \arrow[u, "\Delta H_4"']
\end{tikzcd}$$

$$\Delta H_o = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4$$

Third Law of Thermodynamics

The entropy of any pure crystalline substance approaches zero as the temperature approaches absolute zero.

Isothermal reversible and irreversible

When the volume of the gas changes from $\mathrm{V}_1 - \mathrm{V}_2$, the total work done (W) can be given as $\mathrm{W} = \mathrm{P}. \int \cdot \mathrm{dV}$ If we consider the external pressure (P) to be constant than

$$\mathrm{W} = \mathrm{P} \int^{\mathrm{d}} \mathrm{V} = \mathrm{P}\left(\mathrm{V}_2 - \mathrm{V}_1\right) = \mathrm{P} \cdot \Delta \mathrm{V}$$
$$\mathrm{W} = \mathrm{P} \cdot \Delta \mathrm{V}$$

Isothermal irreversible expansion of an ideal gas

$$\mathrm{W} = -\int_{\mathrm{V}_1}^{\mathrm{V}_2} \mathrm{P}_{\mathrm{ext}} \mathrm{dV}$$

$$\begin{aligned}
&= -P_{\text{ext}} \int_{V_1}^{V_2} \\
&= -P_{\text{ext}}\left(V_2 - V_1\right) \\
W &= -P_{\text{ext}} \cdot \Delta V
\end{aligned}$$

Work done in Isothermal reversible expansion of an ideal gas

$$dW = -PdV$$

So the total work done when the gas expands from initial volume V_1 to final volume V_2 is given a

$$\mathrm{W}_{\text{rev}} = -2.303 \mathrm{nRT} \log_{10} \frac{\mathrm{V}_2}{\mathrm{V}_1}$$
$$\mathrm{W}_{\text{rev}} = -2.303 \mathrm{nRT} \log_{10} \frac{\mathrm{P}_1}{\mathrm{P}_2}$$

Here negative sign indicates work of expansion and it is generally greater than work in the irreversible process.

Here $\Delta \mathrm{H}$ can be found out as follows:

$$\Delta \mathrm{H} = \Delta \mathrm{E} + \Delta \mathrm{n}_{\mathrm{g}} \mathrm{RT}$$

As, for isothermal process, $\Delta E = 0, \Delta T = 0$ So $\Delta H = 0$

Adiabatic Reversible Expansion of An Ideal Gas

(1) Process Equations for Reversible Adiabatic Process

$$PV^{\gamma}=\text{constant}$$

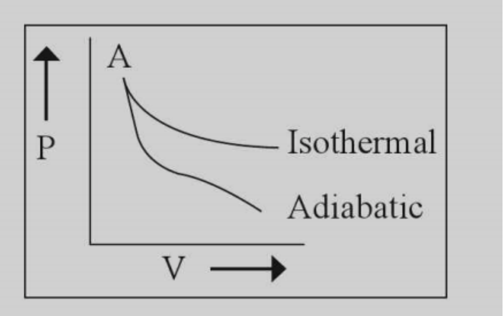
$$TV^{\gamma-1}=\text{constant}$$

$$T^{\gamma}P^{1-\gamma}=\text{constant}$$

(2)Irreversible Adiabatic Free Expansion

$\Delta E=W=0$
$\Delta\ T=0, \Delta H=0$

Comparison between Isothermal and Adiabatic Curves



Heat of Combustion

calorific value $=\frac{\text{Heat of combustion}}{\text{Molecular wt.}}$

$\Delta H(\text{ heat of reaction })=-\Sigma\Delta H_{P}^{\circ}-\Sigma H_{R}^{\circ}$

Enthalpy of Dissociation or Ionization

$H_2O(l)\rightarrow H^{+}+OH^{-}\quad \Delta H=13.7\text{Kcal}$

Heat of Atomization

It is denoted by ΔH_a or ΔH°.
Example,

$H_2(\ g)\rightarrow 2H(\ g)-435\ kJ$
$\Delta H=+435\ kJ/mol$
$CH_4(\ g)\rightarrow C(\ g)+4H(\ g)+1665\ kJ$
$\Delta H=+1665\ kJ/mol$

Phase Transition and Transition Energy

Example,
$C\ (\text{diamond})\ \rightarrow C\ (\text{amorphous})$
$\Delta H=3.3\text{Kcal}$

Chapter 6 - Equilibrium

A state in which forward and backward reactions occur at the same rate and concentrations remain constant is called Equilibrium.

Type	Description	Example
Physical	Involves physical processes	Ice ⇌ Water ⇌ Vapour
Chemical	Involves chemical reactions	$N_2 + 3H_2 \rightleftharpoons 2NH_3$

Law of Chemical Equilibrium (Law of Mass Action)
For a reaction:

$$aA + bB \rightleftharpoons cC + dD$$

ilibrium constant (Kc):

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

Only gases and aqueous species are included in the expression.
Pure solids and liquids are excluded.

Equilibrium Constant (Kp) for Gaseous Reactions

$K_p = K_c(RT)^{\Delta n}$
Where:
$\Delta n =$ moles of gaseous products - reactants
R $=$ gas constant
$T =$ temperature (in Kelvin)

Reaction Quotient (Q)

- Similar to Kc, but for **non-equilibrium conditions**.

Condition	**Interpretation**
Q = K	System is at equilibrium
Q < K	Forward reaction favoured
Q > K	Backward reaction favoured

Le Chatelier’s Principle

If a system at equilibrium is disturbed, it shifts to minimize the disturbance.

Change Applied	Direction of Shift
Increase in Concentration	Opposite side
Increase in Pressure	Toward fewer gas molecules
Increase in Temperature	Depends on exo/endo reaction
Adding Catalyst	No effect on equilibrium position, only speeds up reaching equilibrium

Factors Affecting Equilibrium

- Concentration
- Pressure/Volume
- Temperature
- Inert gases
- Catalyst

Ionic Equilibrium in Aqueous Solutions

- Involves weak and strong electrolytes.
- Acids, bases, and salts in water dissociate to form ions

Ionisation of Acids and Bases

- **Strong acid/base** → completely ionised
- **Weak acid/base** → partially ionised

Acid-Base Equilibrium (Arrhenius, Bronsted-Lowry)

- **Arrhenius:** Acid → H^+ donor; Base → OH^- donor
- **Bronsted-Lowry:** Acid → proton donor; Base → proton acceptor.

Ostwald's Dilution Law

For weak electrolytes:

$$K_a = \frac{C\alpha^2}{1-\alpha}$$

(Where $\alpha =$ degree of dissociation, $\mathrm{C} =$ concentration)

Buffer Solutions
Maintain constant pH
Made of a weak acid + salt of its conjugate base (or vice versa)
Use Henderson-Hasselbalch equation:

$$\mathrm{pH} = \mathrm{p}K_a + \log \frac{[\text{ Salt }]}{[\text{ Acid }]}$$

Solubility Product (Ksp)

For a salt $AB \rightleftharpoons A^{+} + B^{-}$

$K_{sp} = [A^{+}][B^{-}]$

pH and pOH

$\mathrm{pH} = -\log[\mathrm{H}^{+}] \quad ; \quad \mathrm{pOH} = -\log\left[\mathrm{OH}^{-}\right] \quad ; \quad \mathrm{pH} + \mathrm{pOH} = 14$

Chapter 7 - Redox reactions

- **Redox** = **Reduction + Oxidation**
- Occurs **simultaneously** in a chemical reaction.

Oxidation:

- Loss of electrons
- Increase in oxidation number
- Addition of oxygen / removal of hydrogen

Reduction:

- Gain of electrons
- Decrease in oxidation number
- Addition of hydrogen / removal of oxygen

Redox Reaction

A chemical reaction where one species is oxidised and another is reduced. Example:

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

Zn is oxidised, Cu^{2+} is reduced.

Oxidising and Reducing agent

An oxidising agent is a substance that gains electrons and gets reduced itself while causing another substance to be oxidised.
A reducing agent is a substance that loses electrons and gets oxidised itself while causing another substance to be reduced.

Example:
In the reaction:

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

Zinc (Zn) loses electrons and is oxidised, so it is the reducing agent.
Copper ion $\left(\mathrm{Cu}^{2+}\right)$ gains electrons and is reduced, so it is the oxidising agent.

Oxidation Number (O.N.)

Hypothetical charge assigned to atoms in a compound assuming electrons are completely transferred.

Common rules:

- O.N. of free element = 0
- O.N. of oxygen = –2 (except in peroxides = –1)
- O.N. of hydrogen = +1 (–1 in metal hydrides)
- Sum of O.N. in a neutral compound = 0

Balancing Redox Reactions

1. Oxidation number method
2. Ion-electron (half-reaction) method

Steps in Ion-Electron Method (in acidic medium)

1. Split into oxidation and reduction half-reactions
2. Balance all atoms except H and O
3. Balance O using H_2O, and H using H^+
4. Balance charges using electrons
5. Equalise electrons and add half-reactions
6. Cancel common terms

Disproportionation Reaction
Same element is oxidised and reduced simultaneously.
Example:

$2H_2O_2 \rightarrow 2H_2O + O_2$

Characteristics of Electrochemical Series

Li^+/Li	$Li^+(aq.) + e^- \longrightarrow Li\ (s)$	-3.04
K^+/K	$K^+(aq.) + e^- \longrightarrow K(s)$	-2.93
Ca^{2+}/Ca	$Ca^{2+}\ (aq.) + 2e^- \longrightarrow Ca(s)$	-2.87
Na^+/Na	$Na^+(aq.) + e^- \longrightarrow Na\ (s)$	-2.71
Mg^{2+}/Mg	$Mg^{2+}\ (aq.) + 2e^- \longrightarrow Mg(s)$	-2.37
$Pt, H_2/H^-$	$H_2(g) + 2e^- \longrightarrow 2H^-(aq.)$	-2.25
Al^{3+}/Al	$Al^{3+}\ (aq.) + 3e^- \longrightarrow Al\ (s)$	-1.66
Mn^{2+}/Mn	$Mn^{2+}\ (aq.) + 2e^- \longrightarrow Mn(s)$	-0.91
$OH^-/H_2, Pt$	$2H_2O(\ell) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq.)$	-0.83
Zn^{2+}/Zn	$Zn^{2+}\ (aq.) + 2e^- \longrightarrow Zn\ (s)$	-0.76
Cr^{3+}/Cr	$Cr^{3+}\ (aq.) + 3e^- \longrightarrow Cr(s)$	-0.74
Fe^{2+}/Fe	$Fe^{2+}\ (aq.) + 2e^- \longrightarrow Fe\ (s)$	-0.44
$Cr^{3+}/Cr^{2+}, Pt$	$Cr^{3+}\ (aq.) + e^- \longrightarrow Cr^{2+}\ (aq.)$	-0.41

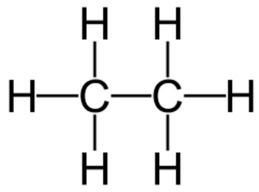
Cd^{2+}/Cd	$Cd^{2+}(aq.) + 2e^{-} \longrightarrow Cd(s)$	-0.40
Co^{2+}/Co	$Co^{2+}(aq.) + 2e^{-} \longrightarrow Co(s)$	-0.28
Ni^{2+}/Ni	$Ni^{2+}(aq.) + 2e^{-} \longrightarrow Ni(s)$	-0.25
$I^{-}/AgI/Ag$	$AgI(s) + e^{-} \longrightarrow Ag(s) + I^{-}(aq.$	-0.15
Sn^{2+}/Sn	$Sn^{2+}(aq.) + 2e^{-} \longrightarrow Sn(s)$	-0.14
Pb^{2+}/Pb	$Pb^{2+}(aq.) + 2e^{-} \longrightarrow Pb(s)$	-0.13
Fe^{3+}/Fe	$Fe^{3+}(aq.) + 3e^{-} \longrightarrow Fe(s)$	-0.04
$H^{+}/H_2, Pt$	$2H^{+}(aq.) + 2e^{-} \longrightarrow H_2(g)$	0.00
$Br^{-}/AgBr/Ag$	$AgBr(s) + e^{-} \longrightarrow Ag(s) + Br^{-}(aq.$	0.10
$Cu^{2+}/Cu^{+}, Pt$	$Cu^{2+}(aq.) + e^{-} \longrightarrow Cu^{+}(aq.)$	0.16
$Sn^{4+}/Sn^{2+}, Pt$	$Sn^{4+}(aq.) + 2e^{-} \longrightarrow Sn^{2+}(aq.)$	0.15
$SO_4^{2-} + H_2SO_3$	$SO_4^{2-}(aq) + .4H^{+} + 2e^{-} \longrightarrow H_2SO_3(aq.) + H_2O(\ell)$	0.17
$Cl^{-}/AgCl/Ag$	$AgCl(s) + e^{-} \longrightarrow Ag(s) + Cl^{-}(aq.)$	0.22
$Cl^{-}/Hg_2Cl_2/Hg(Pt)$	$Hg_2Cl_2(s) + 2e^{-} \longrightarrow 2Hg(\ell) + 2Cl^{-}(aq.)$	0.27
Cu^{2+}/Cu	$Cu^{2+}(aq.) + 2e^{-} \longrightarrow Cu(s)$	0.34
$Pt, O_2/OH^{-}$	$O_2(g) + 2H^{+}(aq.) + 2e^{-} \longrightarrow H_2O_2(aq.)$	0.40
Cu^{+}/Cu	$Cu^{+}(aq.) + e^{-} \longrightarrow Cu(s)$	0.52
$I_2/I^{-}, Pt$	$1/2I_2(s) + e^{-} \longrightarrow I^{-}(aq.)$	0.54
$Pt, O_2/H_2O_2$	$O_2(g) + 2H^{+}(aq.) + 2e^{-} \longrightarrow H_2O_2(aq.)$	0.68
$Fe^{3+}/Fe^{2+}, Pt$	$Fe^{3+}(aq.) + e^{-} \longrightarrow Fe^{2+}(aq.)$	0.77
$Hg_2^{2+}/Hg(Pt)$	$1/2Hg_2^{2+}(aq.) + e^{-} \longrightarrow Hg(s)$	0.79
Ag^{+}/Ag	$Ag^{+}(aq.) + e^{-} \longrightarrow Ag(s)$	0.80
Hg^{2+}/Hg_2^{2+}	$2Hg^{2+}(aq.) + 2e^{-} \longrightarrow Hg_2^{2+}(aq.)$	0.92
$NO_3^{-}/NO, Pt$	$NO_3^{-} + 4H(aq.) + 3e^{-} \longrightarrow NO(g) + 2H_2O(\ell)$	0.97
$Pt, Br_2/Br^{-}$	$Br_2(\ell) + 2e^{-} \longrightarrow 2Br^{-}(aq.)$	1.09
MnO_2/Mn^{2+}	$MnO_2(s) + 4H^{+}(aq.) + 2e^{-} \longrightarrow Mn^{2+}(aq.) + 2H_2O(\ell)$	1.23
$H^{+}/O_2/Pt$	$O_2(g) + 4H^{+}(aq.) + 4e^{-} \longrightarrow 2H_2O(\ell)$	1.23
$Cr_2O_7^{2-}/Cr^{3+}$	$Cr_2O_7^{2-}(aq.) + 14H^{+}(aq.) + 6e^{-} \longrightarrow 2Cr^{3+}(aq.) + 7H_2O(\ell)$	1.33
Cl_2/Cl^{-}	$1/2Cl_2(g) + e^{-} \longrightarrow Cl^{-}(aq.)$	1.36
Au^{3+}/Au	$Au^{3+}(aq.) + 3e^{-} \longrightarrow Au(s)$	1.40
$MnO_4^{-}/Mn^{2+}, H^{+}/Pt$	$MnO_4^{-}(aq.) + 8H^{+}(aq.) + 5e \longrightarrow Mn^{2+}(aq.) + 4H_2O(\ell)$	1.52
$Ce^{4+}/Ce^{3+}, Pt$	$Ce^{4+} + e^{-} \longrightarrow Ce^{3+}(aq.)$	1.72
H_2O_2/H_2O	$H_2O_2(\ell) + 2H^{+}(aq.) + 2e^{-} \longrightarrow 2H_2O(\ell)$	1.78
$Co^{3+}/Co^{2+}, Pt$	$Co^{3+}(aq.) + e^{-} \longrightarrow Co^{2+}(aq.)$	1.81
O_3/O_2	$O_3(g) + 2H^{+}(aq.) + 2e^{-} \longrightarrow O_2(g) + H_2O(\ell)$	2.07
$Pt, F_2/F$	$F_2(g) + 2e^{-} \longrightarrow 2F^{-}(aq.)$	2.87

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
Here Cu is replaced by Zn due to more oxidation potential or reactivity of Zn, while Zn is coated by Cu. Zn-Cu couple is also coated by Cu.

Chapter 8 - Organic Chemistry - Some Basic Principles and Techniques

Classification of organic compounds

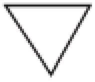
Acyclic or Open-chain compounds

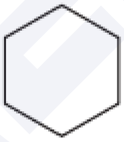


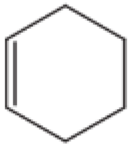
Cyclic or Closed-chain compounds

(a) Homocyclic compounds: These are the compounds having a ring or rings of carbon atoms only in the molecule.

- **Alicyclic compounds:**

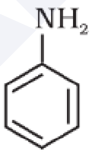

Cyclopropane


Cyclohexane

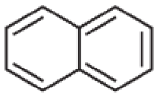

Cyclohexene

- **Aromatic compounds:**

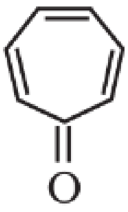

Benzene



Aniline

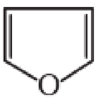

Naphthalene

However, there are aromatic compounds, which have structural units different from benzenoid type and are known as non-benzenoid aromatics.

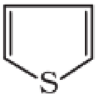


Tropone

(b) Heterocyclic compounds:



Furan



Thiophene

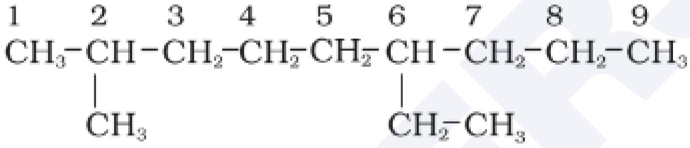
Pyridine

Nomenclature of straight-chain hydrocarbons

Name	Molecular formula	Name	Molecular formula
Methane	CH_4	Heptane	C_7H_{16}
Ethane	C_2H_6	Octane	C_8H_{18}
Propane	C_3H_8	Nonane	C_9H_{20}
Butane	C_4H_{10}	Decane	$C_{10}H_{22}$
Pentane	C_5H_{12}	Icosane	$C_{20}H_{42}$
Hexane	C_6H_{14}	Triacontane	$C_{30}H_{62}$

Nomenclature of branched-chain alkanes:
The rules for naming branched-chain alkanes are as follows:

1. First of all, the longest carbon chain in the molecule is identified.



$$\overset{1}{\mathrm{CH_3}}-\underset{\underset{\mathrm{CH_3}}{|}}{\overset{2}{\mathrm{CH}}}-\overset{3}{\mathrm{CH_2}}-\overset{4}{\mathrm{CH_2}}-\overset{5}{\mathrm{CH_2}}-\underset{\underset{\mathrm{CH_2-CH_3}}{|}}{\overset{6}{\mathrm{CH}}}-\overset{7}{\mathrm{CH_2}}-\overset{8}{\mathrm{CH_2}}-\overset{9}{\mathrm{CH_3}}$$

2. The names of alkyl groups attached as a branch are then prefixed to the name of the parent alkane and position of the substituents is indicated by the appropriate numbers. If different alkyl groups are present, they are listed in alphabetical order.

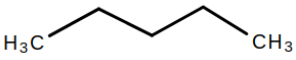
3. If two or more identical substituent groups are present then the numbers are separated by commas. The names of identical substituents are not repeated, instead prefixes such as di (for 2), tri (for 3), tetra (for 4), penta (for 5), hexa (for 6) etc. are used. While writing the name of the substituents in alphabetical order, these prefixes, however, are not considered.

4. If the two substituents are found in equivalent positions, the lower number is given to the one coming first in the alphabetical listing.

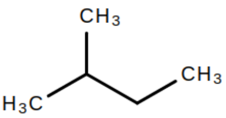
Organic compounds having Functional Groups
When there are more functional groups then a priority order is followed as:
-COOH, –SO_3H, -COOR (R=alkyl group), COCl, -$CONH_2$, -CN,-HC=O, >C=O, -OH, -NH_2, >C=C<, -C≡C-

Structural isomerism
Compounds having the same molecular formula but different structures (manners in which atoms are linked) are classified as structural isomers. Some typical examples of different types of structural isomerism are given below:

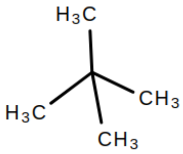
- **Chain isomerism:**



n- Pentane



Isopentane

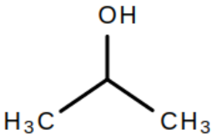


Neopentane

- .

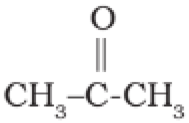
- **Position isomerism:**

Propan-1-ol



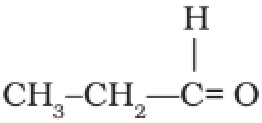
Propan-2-ol

- **Functional group isomerism:**



$CH_3-\overset{\overset{O}{\|}}{C}-CH_3$

Propanone



$CH_3-CH_2-\overset{\overset{H}{|}}{C}=O$

Propanal

- **Metamerism:** It arises due to different alkyl chains on either side of the functional group in the molecule. For example, $C_4H_{10}O$ represents methoxypropane ($CH_3OC_3H_7$) and ethoxyethane ($C_2H_5OC_2H_5$).

Stereoisomerism

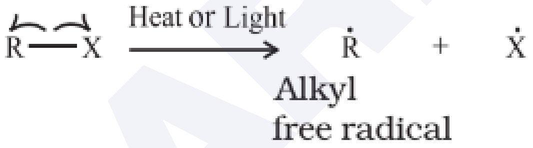
The compounds that have the same constitution and sequence of covalent bonds but differ in relative positions of their atoms or groups in space are called stereoisomers. This special type of isomerism is called as stereoisomerism and can be classified as geometrical and optical isomerism.

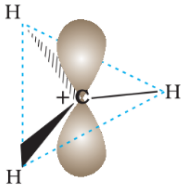
A covalent bond can get cleaved either by:

- Heterolytic cleavage
- Homolytic cleavage

Heterolytic cleavage

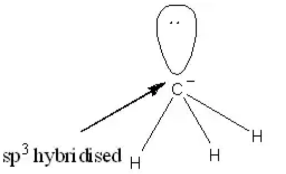
Homolytic cleavage





Carbocation

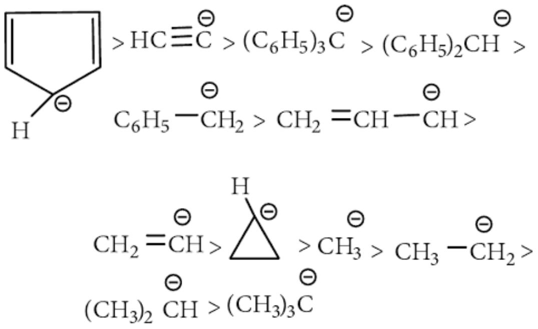
The observed order of carbocation stability is: $^+CH3 < CH_3C^+H_2 < (CH_3)_2C^+H < (CH_3)_3C^+$.



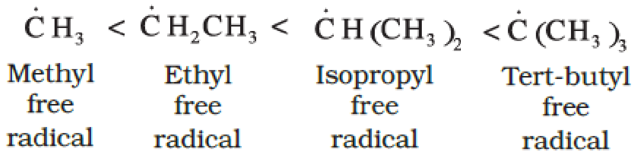
Orbital structure of carbanion

Carbanions

The group having +I effect decreases the stability while groups having -I effect increase the stability of carbanions.



Free radicals



$\dot{C}H_3 < \dot{C}H_2CH_3 < \dot{C}H(CH_3)_2 < \dot{C}(CH_3)_3$

A reagent that brings an electron pair to the reactive site is called a nucleophile (Nu:) i.e., nucleus seeking and the reaction is then called **nucleophilic**.

A reagent that takes away an electron pair from the reactive site is called electrophile (E+) i.e., electron seeking and the reaction is called **electrophilic.**

Some examples of **nucleophiles** are the negatively charged ions with lone pair of electrons such as hydroxide (HO^-), cyanide (NC^-) ions and carbanions ($R_3C:^-$). Neutral molecules such as H_2O, R_3N, R_2NH, etc., can also act as nucleophiles due to the presence of lone pair of electrons.

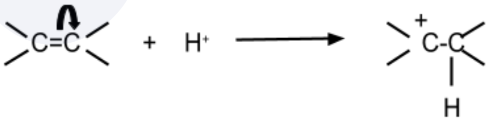
Examples of **electrophiles** include carbocations ($^+CH_3$) and neutral molecules having functional groups like carbonyl group (>C=O) or alkyl halides (R_3C-X, where X is a halogen atom).

Inductive effect

(1) Electron Withdrawing Groups: Halogens and many other groups such as nitro (-NO_2), cyano (-CN), carboxy (-COOH), ester (COOR), aryloxy (-OAr, e.g. – OC_6H_5), etc. are electron-withdrawing groups.

(2) Electron Donating Groups: Alkyl groups like methyl (–CH_3) and ethyl(–CH_2–CH_3) are electron-donating groups.

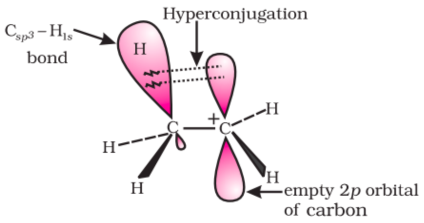
Electromeric effect



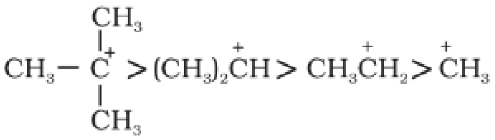
Resonance effect

+R effect: $-\ddot{X}, -OH, -OR, -OCOR, -\overrightarrow{H}_2, -NHR, -NR_2, -NHCOR$
R effect: $-COOH, -CHO, > C = O, -CN, -NO_2, -SO_3H$

Hyperconjugation



The relative stability of carbocations:



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

Sublimation is used for those sublimable substances which are associated with non-volatile impurities. e.g., camphor, naphthalene, anthracene, benzoic acid, iodine, etc.

Crystallisation is the process of formation of solid crystals from solution, melt or by deposition directly from a gas phase. For example, Benzoic acid mixed with naphthalene can be seperated using hot water.

Simple distillation can be used effectively to separate liquids that have some major degrees difference in their boiling points. e.g- Acetone and water.

Fractional distillation is used for the separation of miscible liquids. E.g.,

(i) Acetone (b.pt. 329 K) from methyl alcohol (b.pt. 338 K)

(ii) Crude oil into various useful fractions such as gasoline, kerosene oil, lubricating oil, etc.

Differential extraction method is based on the fact that organic substances are more soluble in organic solvents than in water.

Chromatography is the proces where we apply the mixture to be separated on a stationary phase (solid or liquid) and a pure solvent such as water or any gas is allowed to move slowly over the stationary phase, carrying the components separately as per their solubility in the pure solvent.

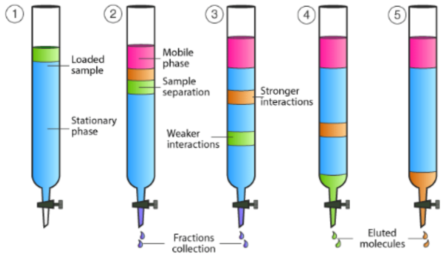
There are four main types of chromatography:

1. Adsorption chromatography
2. Column chromatography
3. Thin layer chromatography
4. Partition chromatography

Adsorption chromatography

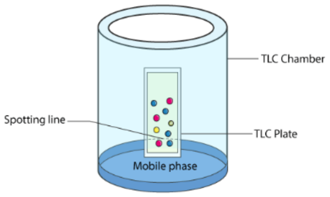
Different compounds are adsorbed on the adsorbent to different degrees based on the absorptivity of the component.

Column chromatography



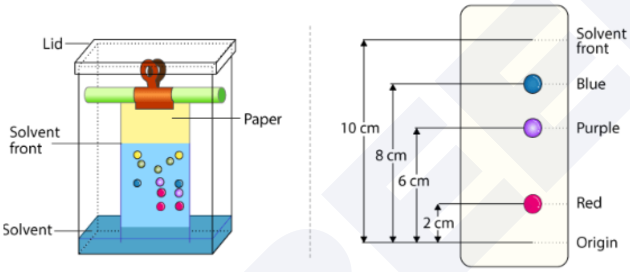
Thin layer chromatography

The solution of the mixture to be separated is applied as a small spot at a distance of 2 cm above one end of the plate.



Partition chromatography

In this process, a continuous differential partitioning of components of a mixture into a stationary phase and mobile phase takes place.Paper chromatography is a type of partition chromatography.



Qualitative analysis

$$C + 2CuO \xrightarrow{\Delta} 2Cu + CO_2$$

$$2H + CuO \xrightarrow{\Delta} Cu + H_2O$$

Test for Nitrogen

$$6CN + Fe^{2+} \rightarrow [Fe(CN)_6]^{4-}$$

$$3[Fe(CN)_6]^{4-} + 4Fe^{3+} \xrightarrow{xH_2O} Fe_4[Fe(CN)_6]_3 \cdot \times H_2O$$

Prussian blue

Test for Sulphur

$$S^{2-} + [Fe(CN)_5NO]^{2-} \longrightarrow [Fe(CN)_5NOS]^{4-}$$

- In case, nitrogen and sulphur both are present in an organic compound, sodium thiocyanate is formed.

$$Na + C + N + S \longrightarrow NaSCN$$
$$Fe^{3+} + SCN^{-} \longrightarrow [Fe(SCN)]^{2+}$$

Halogens

$$X^- + Ag^+ \longrightarrow AgX$$
$$\left(X = Cl^-, Br^-, I^-\right)$$

Test for Phosphorus

$$Na_3PO_4 + 3HNO_3 \longrightarrow H_3PO_4 + 3NaNO_3$$
$$\underset{\text{Ammonium molybdate}}{H_3PO_4 + 12(NH_4)_2MoO_4} + 21HNO_3 \longrightarrow \underset{\text{Ammonium phosphomolybdate}}{(NH_4)_3PO_4 \cdot 12MoO_3} + 21NH_4NO_3 + 12H_2O$$

Quantitative analysis

Estimation of carbon and hydrogen in an organic compound is based on their conversion to CO_2 and H_2O respectively. The percentage of carbon and hydrogen are calculated from the masses of CO_2 and H_2O.

Duma's Method

$$C_xH_yN_z + (2x + y/2)CuO \longrightarrow xCO_2 + y/2H_2O + z/2\,N_2 + (2x + y/2)Cu$$

Volume of nitrogen at STP $= \frac{p_1V_1 \times 273}{760 \times T_1}$

Kjeldahl method

$$\text{Organic compound } + H_2SO_4 \longrightarrow (NH_4)_2SO_4 \xrightarrow{2NaOH} Na_2SO_4 + 2NH_3 + 2H_2O$$

Kjeldahl method is not applicable to compounds containing nitrogen in nitro and azo groups and nitrogen present in the ring (e.g. pyridine) as nitrogen of these compounds does not change to ammonium sulphate under these conditions.

Halogens (Carius method)

$$\text{Mass of halogen in } m_1 \text{ g of AgX} = \frac{\text{atomic mass of X} \times m_1 \text{ g}}{\text{molecular mass of AgX}}$$
$$\text{Percentage of halogen} = \frac{\text{atomic mass of X} \times m_1 \times 100}{\text{molecular mass of AgX} \times m}$$

Sulphur
Let the mass of organic compound taken = m g
and the mass of barium sulphate formed = m_1 g
1 mol of $BaSO_4$ = 233 g $BaSO_4$ = 32 g sulphur

$m_1 g BaSO_4$ contains $\frac{32 \times m_1}{233}$ g sulphur
Percentage of sulphur $= \frac{32 \times m_1 \times 100}{233 \times m}$

Carius Method (oxygen)

$$10C + 5O_2 + 2I_2O_5 \longrightarrow 2I_2 + 10CO_2$$

Let the mass of organic compound taken be mg Mass of carbon dioxide produced be m_1 g $\therefore$ m_1 g carbon dioxide is obtained from $\frac{32 \times m_1}{88}$ gO_2

$$\therefore \text{ Percentage of oxygen } = \frac{32 \times m_1 \times 100}{88 \times m}\%$$

Chapter 9 - Hydrocarbons

They are carbon and hydrogen containing organic compounds that are classified into alkanes, alkenes and alkynes.

Preparation of Alkanes

a) From Grignard reagents

$$CH_3MgBr + C_2H_5OH \longrightarrow CH_4 + (C_2H_5O)MgBr$$

Decarboxylation of Fatty acids

$$RCOONa + NaOH \xrightarrow[\Delta]{CaO} RH + Na_2CO_3$$

Kolbe's Electrolysis

$$2R - COONa \xrightarrow{\text{electrolytic hydrolysis}} R - R + 2CO_2 + 2NaOH + H_2 \uparrow$$

Chemical properties of Alkanes

Halogenation
The rate of reaction of alkanes with halogen follows the following order:
$F_2 > Cl_2 > Br_2 > I_2$

$$C_2H_6 + Cl_2 \xrightarrow{hv} C_2H_5Cl + HCl$$

Nitration

$$R - H + HONO_2 \xrightarrow[\text{Temperature}]{\text{Hgh}} R - NO_2 + H_2O$$

However, when a mixture of vapours of an alkane and nitric acid is heated at 673-773K, nitroalkane is formed readily. This is known as **vapour phase nitration**. By this process, lower, as well as higher alkanes, can be converted into nitroalkanes.

$$CH_3 + HONO_2 \xrightarrow{723\text{ K}} CH_3 - NO_2 + H_2O$$

Sulphonation

$$R - H + HOSO_3H \xrightarrow[\text{Prolonged heating}]{SO_3} R - SO_3H + H_2O$$

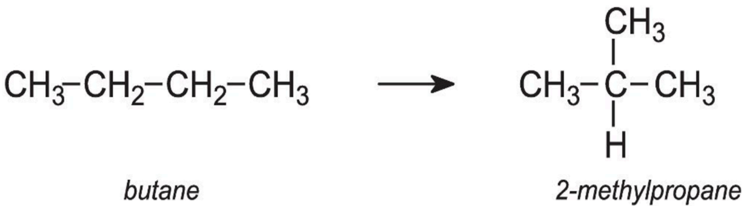
Combustion
$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(1), \Delta_cH^\circ = -890\text{ kJ/mol}$
Due to the production of large amount of heat, alkanes are used as fuels.

Catalytic oxidation

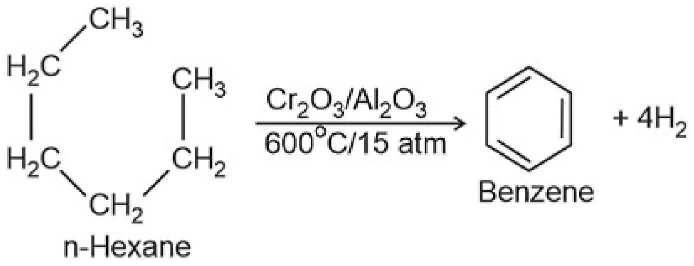
$$2CH_4 + O_2 \xrightarrow{\text{Cu,523 K}} 2CH_3OH$$

$$2CH_3CH_3 + 3O_2 \xrightarrow{(CH_3COO)_2Mn} 2CH_3COOH + 2H_2O$$

Isomerisation



Aromatization



Pyrolysis

$$C_{12}H_{26} \xrightarrow[973\ K]{Pt/Pd/Ni} C_7H_{16} + C_5H_{10} + \text{other products}$$

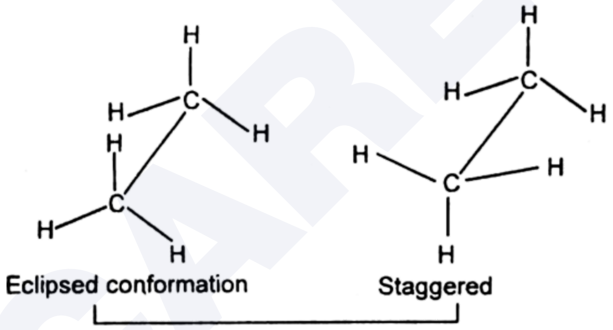
Dodecane Heptane Pentene

Conformations

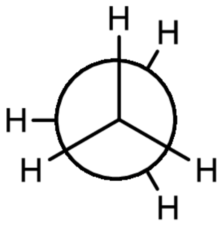
Free rotation about C–C single bond. Alkanes can thus have infinite number of conformations by rotation around C-C single bonds. It is hindered by a small energy barrier of 1-20 kJ mol^{-1} due to weak repulsive interaction between the adjacent bonds. Such a type of repulsive interaction is called torsional strain.

Sawhorse projections

Each carbon has three lines attached to it corresponding to three hydrogen atoms. The lines are inclined at an angle of 120° to each other.



Newman projections



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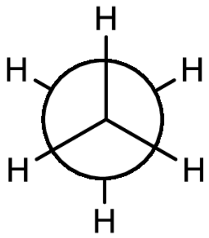
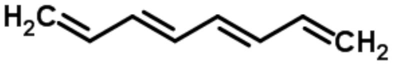
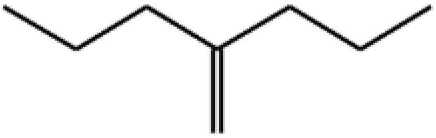


Staggered

Nomenclature of Alkenes
Examples



IUPAC name: Octa-1,3,5,7-tetraene



IUPAC name: 2-n-propylpent-1-ene

Isomerism
Alkenes show two kinds of isomerism i.e, stereoisomerism and geometrical isomerism

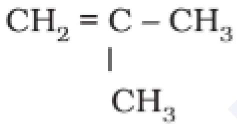
- **Stereoisomerism:** Ethene (C2H4) and propene (C3H6) can have only one structure but alkenes higher than propene have different structures.For example, But-1-ene can have three different structures as follows:

$CH_2 = CH - CH_2 - CH_3$

But-1-ene

$CH_3 - CH = CH - CH_3$

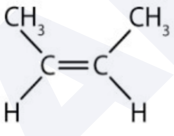
But-2-ene



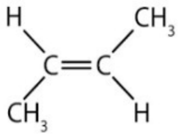
$$\begin{array}{l} CH_2 = C - CH_3 \\ \quad\quad\; | \\ \quad\quad CH_3 \end{array}$$

2-Methyprop-1-ene

- **Geometrical isomerism:** The two structures or isomers are known as *cis* and *trans* isomers.



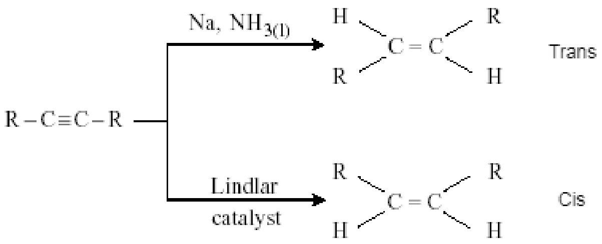
cis-2-butene



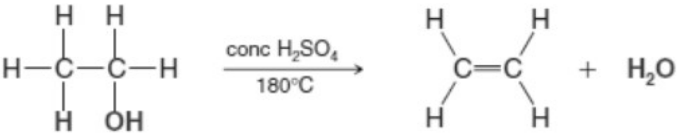
trans-2-butene

Preparation of alkenes

a) From Alkynes



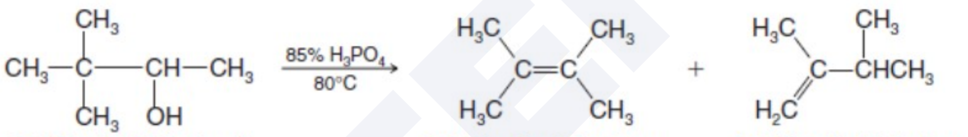
b) Dehydration of Alcohols



This reaction can be used to dehydrate all three types of alcohol viz. Primary, secondary and tertiary alcohols.

It is to be noted that the dehydration usually occurs via the Unimolecular elimination reaction (E_1) and involves a carbocation intermediate which can undergo rearrangement via hydride or alkyl shift and also undergo ring expansion for suitable substrates where the ring strain can be released.

Case of Methyl shift

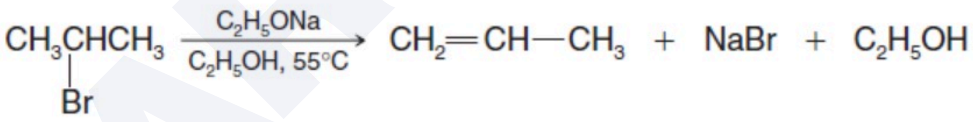


3,3-Dimethyl-2-butanol

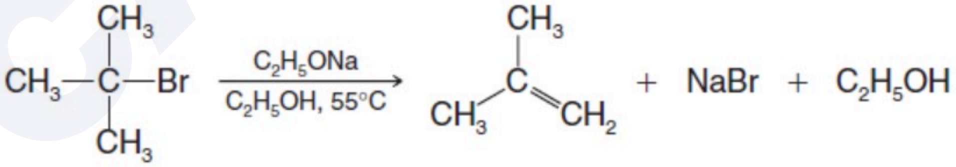
2,3-Dimethyl-2-butene (80%)

2,3-Dimethyl-1-butene (20%)

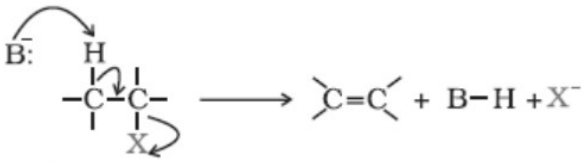
c) Dehydrohalogenation of Haloalkanes with Strong bases



$CH_3CHBrCH_3 \xrightarrow[C_2H_5OH,\ 55^\circ C]{C_2H_5ONa} CH_2{=}CH{-}CH_3 + NaBr + C_2H_5OH$

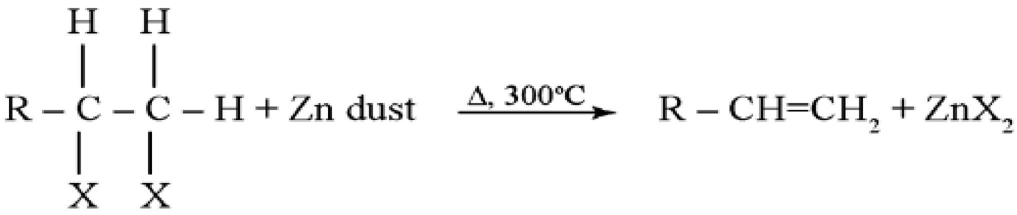


This reaction is an example of β elimination in which a β - hydrogen is eliminated along with a halogen at the α carbon.



B=Base ; X=Leaving group

d) With vicinal dihalides

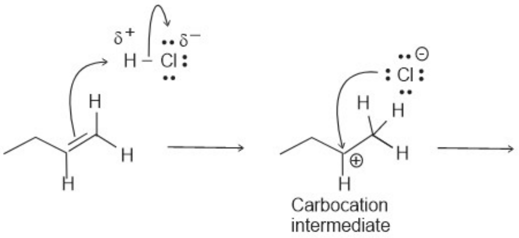


$$R-\underset{X}{\overset{H}{C}}-\underset{X}{\overset{H}{C}}-H + \text{Zn dust} \xrightarrow{\Delta,\ 300^{\circ}C} R-CH{=}CH_2 + ZnX_2$$

Chemical properties of alkenes

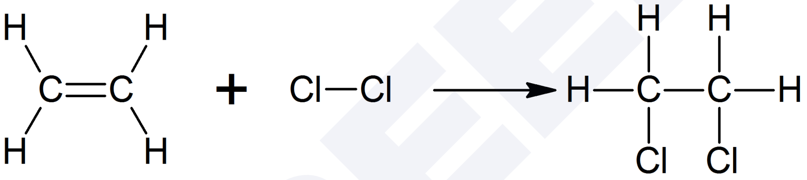
1) All alkenes undergo addition reactions with the hydrogen halides. There is a formation of a carbocation intermediate and rearrangement may occur in cases where there is a possibility of more stability.

The reaction occurs as follows:

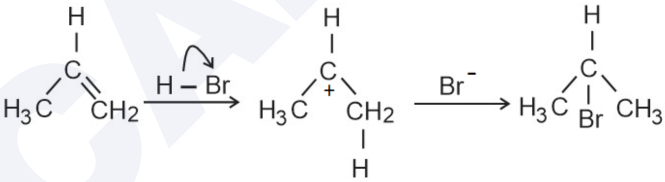


2) Alkenes with Bromine water (Br_2 in CCl_4)

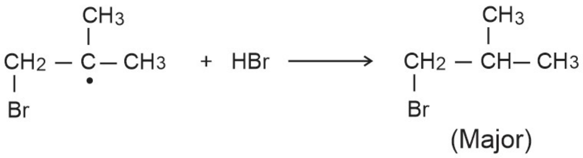
This serves as a test of unsaturation. The addition of halogens to an alkene is an anti-addition .



Markovnikov's rule
This rule states that the acid hydrogen of the protic acid gets attached to the carbon with more hydrogen substituents and negative part adds to the atom with less number of hydrogen atoms.



Anti-Markovnikov's rule
In the presence of peroxide, such as benzoyl peroxide and light, the addition of HBr(not HCl and HI) to unsymmetrical alkenes occurs contrary to Markovnikov's rule.



3) Cold concentrated sulphuric acid adds to alkenes in accordance with Markovnikov rule

For example:

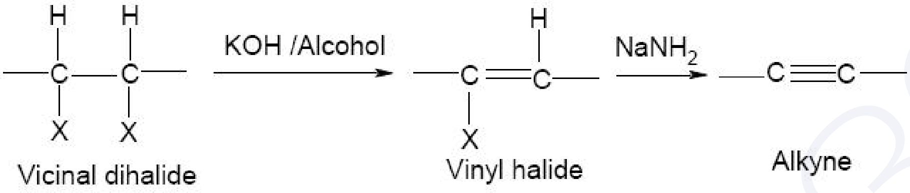
$CH_2 = CH_2 + H_2SO_4 \longrightarrow CH_3CH_2OSO_2OH$

Preparation of Alkynes

a) **Calcium carbide:**

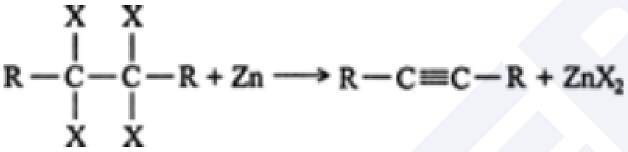
$CaCO_3 \longrightarrow CaO + CO_2$
$CaO + 3C \longrightarrow CaC_2 + CO$
$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H_2$

b) Vicinal dihalides:



Using Zinc:

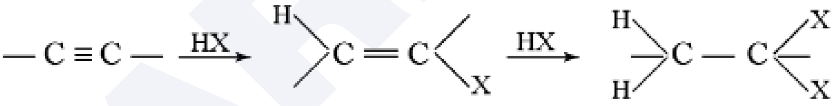
Vicinal tetrahaloalkanes can be dehalogenated with zinc metal in an organometallic reaction to form alkynes.



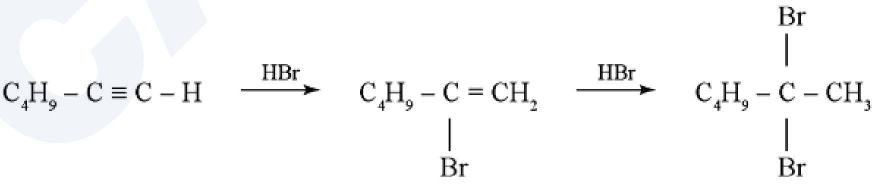
$R-CX_2-CX_2-R + Zn \longrightarrow R-C\equiv C-R + ZnX_2$

Chemical properties of alkynes

Hydrohalogenation
This addition follows Markownikoff's rule.

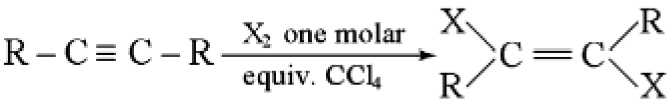


For example:

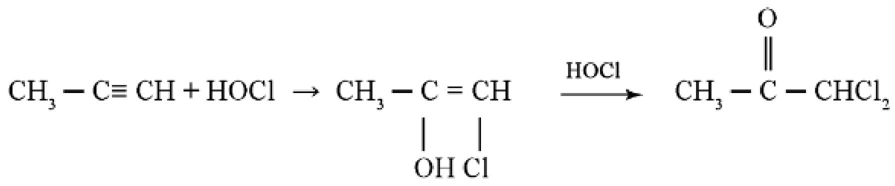


$C_4H_9-C\equiv C-H \xrightarrow{HBr} C_4H_9-C(Br)=CH_2 \xrightarrow{HBr} C_4H_9-CBr_2-CH_3$

Halogenation
Here the addition is Anti Markownikoff's. The reaction occurs as follows:



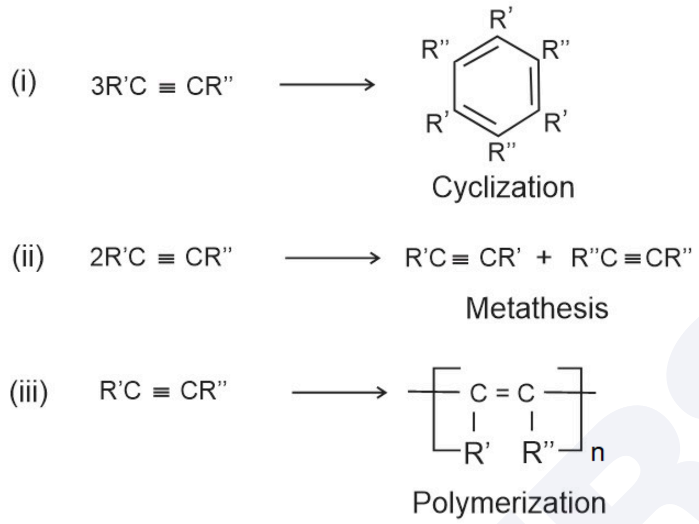
Addition of Hypochlorous acid



$$CH_3-C\equiv CH+HOCl \rightarrow CH_3-\underset{OH}{\underset{|}{C}}=\underset{Cl}{\underset{|}{CH}} \xrightarrow{HOCl} CH_3-\overset{O}{\overset{\|}{C}}-CHCl_2$$

Polymerisation

Alkynes polymerize to give the following compounds. The reactions occur as follows:



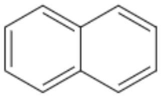
Aromatic hydrocarbons

Aromaticity is defined as "An aromatic compound having a cyclic planar structure with (4n+2)π electrons and have high resonance energy and stability due to delocalization of π electrons." Any compound is aromatic if the following conditions are fulfilled:

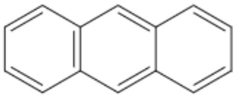
- It has complete delocalization of π electrons.
- Has a high resonance energy.
- Has a conjugate system.
- Has number of π electrons according to 4n +2 or Huckel's rule that is, 2,6,10,14,18. Here, n = 0,1,2 ....
- If number of π electrons 4 "n' i.e., 4, 8, 12, 16, it will be anti-aromatic.
- If any of these conditions is not obeyed it will be non-aromatic.

Some example of aromatic compound include

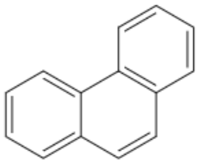
Benzene



Naphthalene

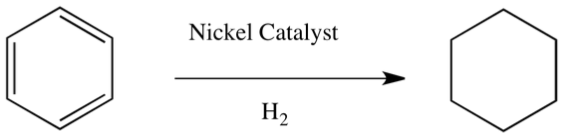


Anthracene



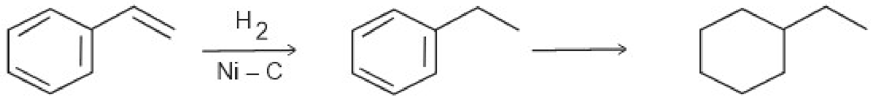
Phenanthrene

Reduction of Aromatic compounds

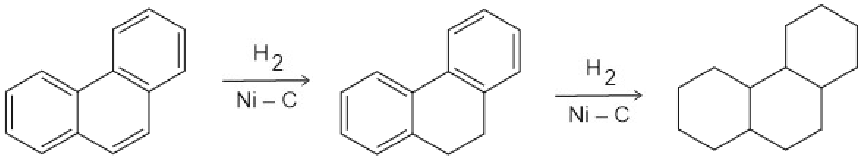


Benzene

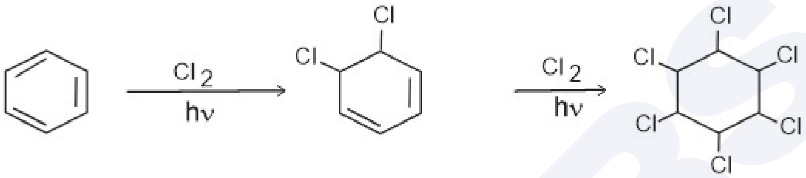
Cyclohexane



In case of Catalytic hydrogenation of Phenanthrene, the middle ring gets reduced first and on subsequent reduction becomes completely saturated.



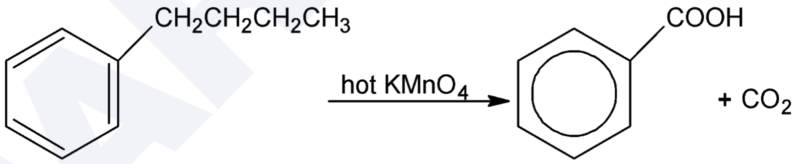
Benzene also shows free radical addition under UV light and adds three molecules of Chlorine to form $C_6H_6Cl_6$ which is also called as Benzene Hexachloride or Gammaxane.



Oxidation

1. **Controlled oxidation with air:** Benzene on oxidation with air at 773 K in presence of V_2O_5 as catalyst gives Maleic anhydride.

2. **Oxidation of Alkyl benzene:** Alkyl groups when attached to the benzene ring, they are easily oxidised by an alkaline solution of potassium manganate.



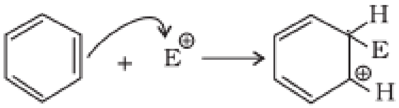
n - butyl benzene

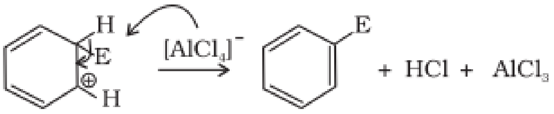
Benzoic acid

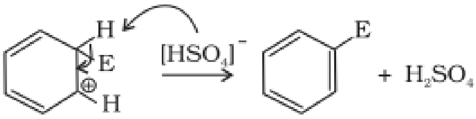
3.

Electrophilic substitution

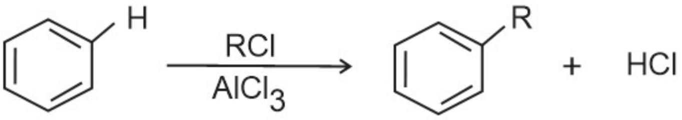
Mechanism



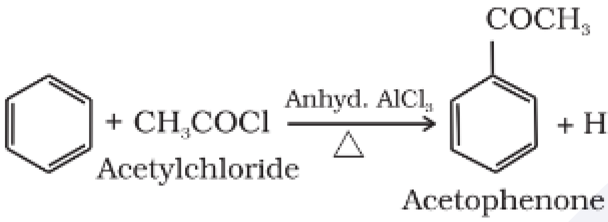




Friedal Crafts Alkylation

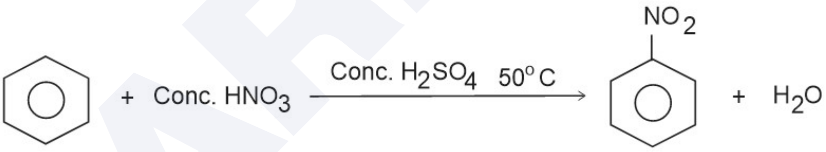


Friedal Crafts Acylation

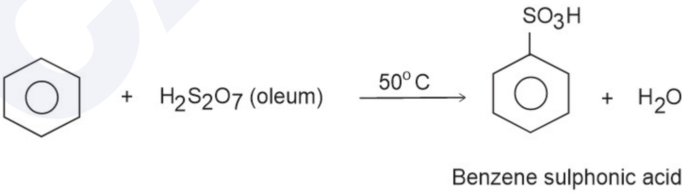


Halogenation

Nitration



Sulphonation



Class 12

Chapter 1. Solutions

Concentration of Solutions

1. Mass percentage (w/w) $= \frac{\text{Mass of the component in the solution}}{\text{Total mass of the solution}} \times 100$

2. Volume percentage (V/V):

Volume % of a component $= \frac{\text{Volume of the component}}{\text{Total volume of solution}} \times 100$

3. Mass by volume percentage (w/V):

Mass by Volume $\%(\mathrm{w}/\mathrm{V}) = \left(\frac{\text{Mass of solute (g)}}{\text{Volume of solution (mL)}}\right) \times 100$

4. Parts per million:

$$\text{Parts per million } = \frac{\text{Number of parts of the component}}{\text{Total number of parts of all components of the solution}} \times 10^6$$

5. Mole fraction:

$= \frac{\text{Number of moles of the component}}{\text{Total number of moles of all the components}}$

For example, in a binary mixture, if the number of moles of A and B are n_{A} and n_{B} respectively, the mole fraction of A will be

$x_{\mathrm{A}} = \frac{n_{\mathrm{A}}}{n_{\mathrm{A}}+n_{\mathrm{B}}}$

For a solution containing i number of components, we have:

$x_i = \frac{n_i}{n_1+n_2+\ldots\ldots+n_i} = \frac{n_i}{\sum n_i}$

6. Molarity:

Molarity $= \frac{\text{Moles of solute}}{\text{Volume of solution in litre}}$

7. Molality:

Molality $(\mathrm{m}) = \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}}$

Solubility of a substance is its maximum amount that can de cussorved in a specined amount of solvent? specified temperature

Raoult's law

$P_{\text{solution}} = (\text{ Mole fraction of solvent }) \times (\text{ Vapour pressure of pure solvent })$

Dalton's law of partial pressures

$P_{\text{total}} = P_1 + P_2 + P_3 + \ldots$

Ideal and Non-Ideal Solutions

Ideal Solutions	Non-Ideal solutions
Obey Raoult's law at all temperatures and concentrations.	Do not obey Raoul's law at all temperatures and concentrations.
$p_1 = x_1 p_1^o ; p_2 = x_2 p_2^o$	$p_1 \neq x_1 p_1^{\circ} ; p_2 \neq x_2 p_2^{\circ}$
$\Delta H_{\text{mix}} = 0, \Delta V_{\text{mix}} = 0$	$\Delta H_{\text{mix}} \neq 0, \Delta V_{\text{mix}} \neq 0$
$A-B$ interactions $\approx A-A$ and $B-B$ interactions.	$A-B$ interactions $\neq A-A$ and $B-B$ interactions.
Do not form azeotropes	Form azeotropes.

Non-ideal solutions snowing positive and negative deviations from Raoult's law :

Solutions showing positive deviation	**Solutions showing negative deviation**

$A-B \ll A-A$ or $B-B$ interactions.	$A-B \gg A-A$ or $B-B$ interactions.
$\Delta H_{\text{mix}}>0, \Delta V_{\text{mix}}>0$	$\Delta H_{\text{mix}}<0, \Delta V_{\text{mix}}<0$
$p_i>p_i^{\circ} x_i$	$p_i<p_i^{\circ} x_i$
Form minimum boiling azeotropes.	Form maximum boiling azeotropes.

Colligative properties

Relative lowering of vapour pressure :
By Raoult's law,

$\frac{P^0-P}{P^0}=X_{\text{solute}}$

Elevation of boiling point :
$\Delta T_b=K_b \cdot m$
where m is molality of solution and K_b is called boiling point elevation

Depression of freezing point :
$\Delta T_f=K_f \cdot m$

$K_f=$ Molal depression constant

Osmosis and osmotic pressure :

$\Pi=CRT$

van't Hoff factor (i)

$i=\frac{\text{Observed colligative property}}{\text{Calculated colligative property}}$

Relation between van't Hoff factor and degree of dissociation, $\alpha=\frac{i-1}{n-1}$
Relation between van't Hoff factor and degree of association, $\alpha=\frac{1-i}{1-1/n}$

Chapter-2 Electrochemistry

Nernst equation for the electrode reaction :

$M_{(aq)}^{n+\infty}+ne^{-} \longrightarrow M_{(s)}$

$$E_{M^{n+}/M}=E_{M^{n+}/M}^{\circ}-\frac{2.303RT}{nF} \log \frac{1}{\left[M^{n+}\right]}=E_{M^{n+}/M}^{\circ}-\frac{0.0591}{n} \log \frac{1}{\left[M^{n+}\right]} \text{ at 298 K}$$

Nernst equation for general electrochemical reaction :

$aA+bB \longrightarrow cC+dD$

$E_{\text{cell}}=E_{\text{cell}}^{\circ}-\frac{2.303RT}{nF} \log \frac{[C]^c[D]^d}{[A]^a[B]^b}$ where, $E_{\text{cell}}^{\circ}=E_{\text{cathode}}^{\circ}-E_{\text{anode}}^{\circ}$ and concentrations of pure solids are taken as unity.

Equilibrium constant from Nernst equation :
At equilibrium, $E_{\text{cell}}=0$ thus, $E_{\text{cell}}^{\circ}=\frac{2.303RT}{nF} \log K_c$
For concentration cell, EMF at 298 K is given by $E_{\text{cell}}=\frac{2.303RT}{nF} \log \frac{C_2}{C_1}$ where $C_2>C_1$

Electrochemical cell and Gibbs energy of the reaction : $\Delta_r G=-nFE_{\text{cell}}$ or $\Delta_r G^{\circ}=-nFE_{\text{cell}}^{\circ}$ or $\Delta_r G^{\circ}=-RT \ln K=-2.303RT \log K$

Conductance of electrolytic solutions:
$R \propto \frac{l}{A}$ or $R=\rho \frac{l}{A}$ where, ρ is called resistivity or specific resistance, having SI units Ωm.
The inverse of resistance, R is called conductance, G,
$G=\frac{1}{R}=\frac{A}{\rho l}=\kappa \frac{A}{l}$ and the SI unit of conductance is S or ohm $^{-1}$ (mho).
$\kappa=\frac{1}{\rho}$ where κ is called conductivity or specific conductance, having SI unit Sm^{-1}.

Molar conductivity (Λ_m **)**

It is the conductance of a solution having SI unit Sm^{-1}.
$\Lambda_m=K \times \frac{1000}{\text{Molarity}}$ where molarity is in molL^{-1}.

Equivalent conductivity (Λ_{eq} **)**

$A_{cq} = \kappa \times \frac{1000}{\text{Normality}}$ where normality is in geq L^{-1}.

Kohlrausch law of independent migration of ions:
$\Lambda_m^{\circ} = v_{+}\lambda_{+}^{\circ} + v_{-}\lambda_{-}^{\circ}$ where λ_{+}° and λ_{-}° are the limiting molar conductivities of the cation and anion respectively and $v_{+}$and $v_{-}$are stoichiometric no. of cations and anions respectively in one formula unit of the electrolyte.
For weak electrolyte, $\alpha = \frac{\Lambda_m}{\Lambda_m^{\circ}}$ and $K_a = \frac{C\alpha^2}{(1-\alpha)} = \frac{C\Lambda_m^2}{\Lambda_m^{\circ}(\Lambda_m^{\circ}-\Lambda_m)}$

Electrolysis
Faraday's first law of electrolysis: $W \propto Q$ or $W = ZQ = Z \times I \times t$

where Z is electrochemical equivalent of the substance deposited and

$Z = \frac{\text{Eq. wt. of substance}}{96500}$

CHAPTER - 3 CHEMICAL KINETICS

Rate of reaction
$R \longrightarrow P$. Rate $= -\frac{\Delta[R]}{\Delta t}$ or $+\frac{\Delta[P]}{\Delta t}$
Average rate of reaction

$$r_{av} = \frac{-\Delta[R]}{\Delta t} = \frac{\Delta[P]}{\Delta t}$$

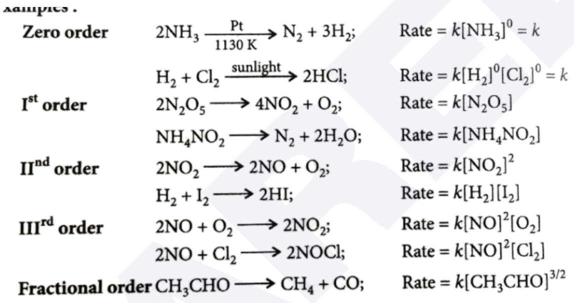
Instantaneous rate of reaction

$$r_{\text{inst}} = \frac{-d[R]}{dt} = \frac{d[P]}{dt}$$

Rate law

$$\text{Rate } = k[A]^x[B]^y$$

$x + y$ gives the overall order of a reaction.



xamples :

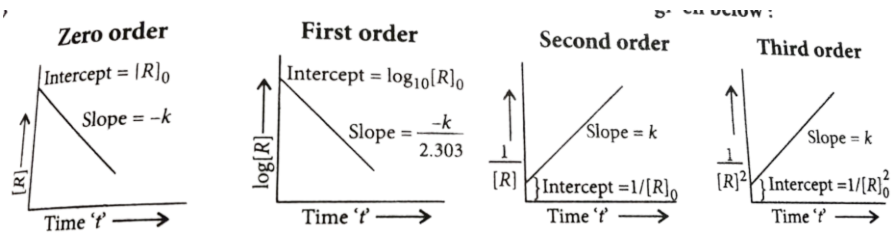
Zero order	$2NH_3 \xrightarrow[1130\text{ K}]{Pt} N_2 + 3H_2$;	Rate = $k[NH_3]^0 = k$
	$H_2 + Cl_2 \xrightarrow{\text{sunlight}} 2HCl$;	Rate = $k[H_2]^0[Cl_2]^0 = k$
I^st order	$2N_2O_5 \longrightarrow 4NO_2 + O_2$;	Rate = $k[N_2O_5]$
	$NH_4NO_2 \longrightarrow N_2 + 2H_2O$;	Rate = $k[NH_4NO_2]$
II^nd order	$2NO_2 \longrightarrow 2NO + O_2$;	Rate = $k[NO_2]^2$
	$H_2 + I_2 \longrightarrow 2HI$;	Rate = $k[H_2][I_2]$
III^rd order	$2NO + O_2 \longrightarrow 2NO_2$;	Rate = $k[NO]^2[O_2]$
	$2NO + Cl_2 \longrightarrow 2NOCl$;	Rate = $k[NO]^2[Cl_2]$
Fractional order	$CH_3CHO \longrightarrow CH_4 + CO$;	Rate = $k[CH_3CHO]^{3/2}$

Units of rate constant :

$k = \frac{\text{Rate}}{[A]^x[B]^y} = \frac{\text{concentration}}{\text{time}} \times \frac{1}{(\text{ concentration })^n}$; $\quad$ (where $[A] = [B]$)
Integrated rate equation for reactions of different orders

Reaction	Order	Rate Law	Rate constant
$R \longrightarrow$ Products	0	Rate $= k[R]^0$	$k = \frac{[R]_0-[R]}{t}$
$R \longrightarrow$ Products	1	Rate $= k[R]$	$k = \frac{2.303}{t}\log\frac{[\mathrm{R}]_0}{[\mathrm{R}]}$
$2R \longrightarrow$ Products	2	Rate $= k[R]^2$	$k = \frac{1}{t}\left(\frac{1}{[R]} - \frac{1}{[R]_0}\right)$
$A + B \longrightarrow$ Products	2	Rate $= k[A][B]$	$k = \frac{1}{t}\frac{2.303}{([\mathrm{A}]_0-[\mathrm{B}]_0)}\log\frac{[\mathrm{B}]_0[\mathrm{A}]}{[\mathrm{A}]_0[\mathrm{B}]}$
$3R \longrightarrow$ Products	3	Rate $= k[R]^3$	$k = \frac{1}{2t}\left[\frac{1}{[R]^2} - \frac{1}{[R]_0^2}\right]$

Linear plots for different order reactions



Half life of a reaction

For zero order reaction: $t_{1/2}=\frac{[R]_0}{2k}$ i.e., $t_{1/2} \propto [R]_0$
For first order reaction: $t_{1/2}=\frac{\ln 2}{k}=\frac{0.693}{k}$ i.e., $t_{1/2}$ is independent of $[R]_0$.

Pseudo first order reactions

Examples:
Acid hydrolysis of ethyl acetate :

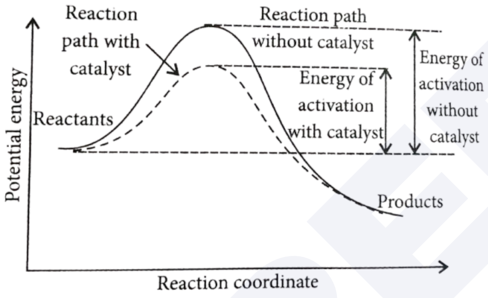
$$CH_3COOC_2H_5 + H_2O \xrightarrow{H^+} CH_3COOH + C_2H_5OH \quad \text{where } k = k'[H_2O]$$
$$\text{Rate } = k'[CH_3COOC_2H_5][H_2O] = k[CH_3COOC_2H_5]$$

Temperature dependence of the rate of a reaction:

For a chemical reaction with a rise in temperature by 10°C, the rate constant is nearly doubled.
Arrhenius equation:

$$k = Ae^{-\frac{E_a}{RT}}$$

Effect of the catalyst on Rate of Reaction



Effect of catalyst on activation energy

Chapter-4 The d and f block elements

Transition elements (or d-block elements) :

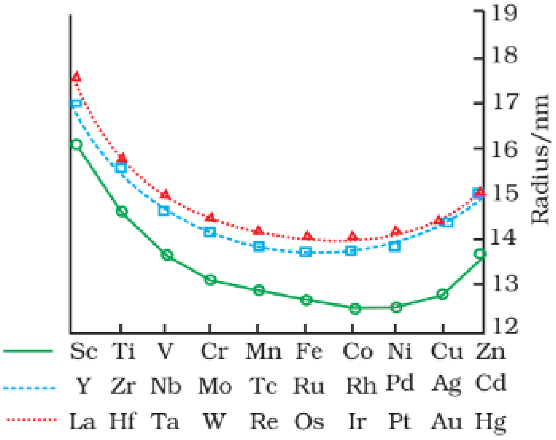
- They are called $\boldsymbol{d}$-block elements as the last electron enters in the d-orbital.
- General electronic configuration : $(n-1)d^{1-10}ns^{0-2}$

Transition series : d-block consists of four transition series,

1^{st} Transition series or $3d$ series ${}_{21}Sc - {}_{30}Zn$
2^{nd} Transition series or $4d$ series ${}_{39}Y - {}_{48}Cd$
3^{rd} Transition series or $5d$ series ${}_{57}La, {}_{72}Hf - {}_{80}Hg$
4^{th} Transition series or $6d$ series ${}_{89}Ac, {}_{104}Rf - {}_{112}Cn$

Properties

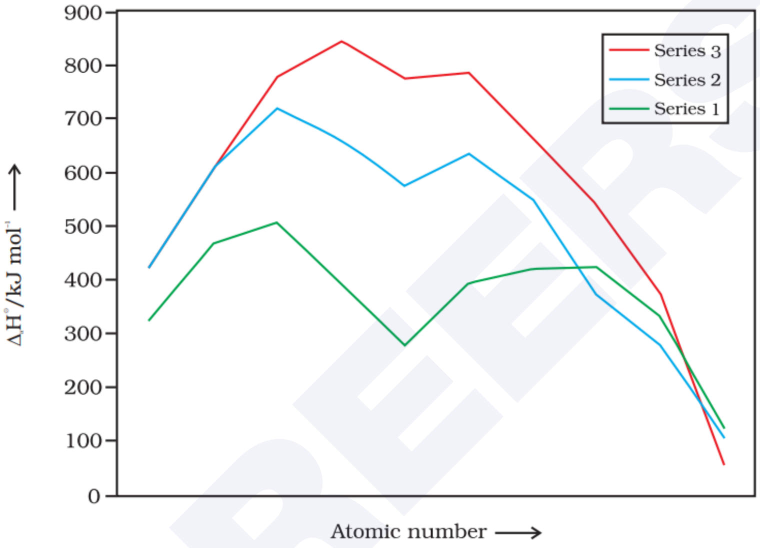
Atomic size



Metallic character

- All the transition elements or *d*-block elements are metals, since the number of electrons in the outermost shell is very small, i.e, either 1 or 2.

Enthalpy of atomisation



Ionisation enthalpy

Ionisation enthalpy in d-block elements shows an irregular trend across a period and decreases slightly down a group.

Oxidation State

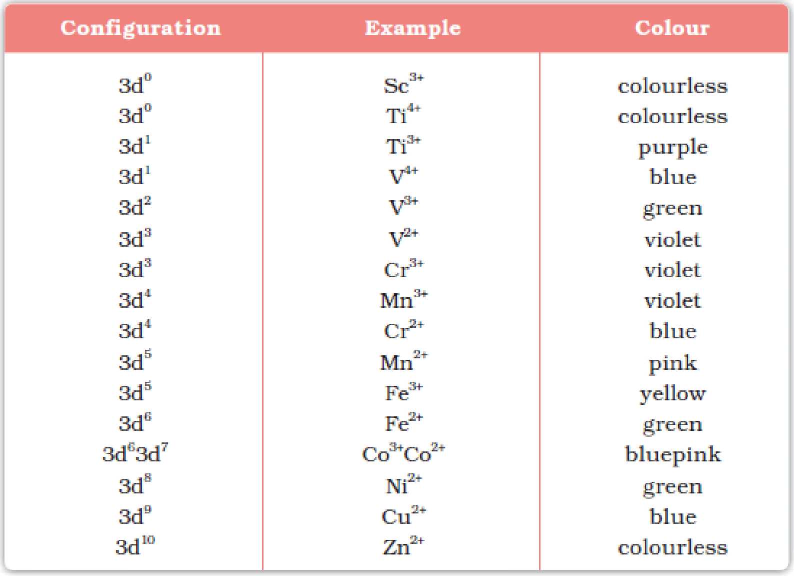
Transition elements exhibit variable oxidation states due to the involvement of both (n−1)d(n-1)d(n−1)d and nsnsns electrons in bonding.

Magnetic properties and magnetic character

Transition elements exhibit paramagnetic behaviour due to unpaired d-electrons, and their magnetic moment increases with the number of unpaired electrons.

$\mu = \sqrt{n(n+2)}$
where n is the number of unpaired electrons and μ is the magnetic moment in units of Bohr magneton (BM).

Formation of coloured ions



Configuration	Example	Colour
$3d^0$	Sc^{3+}	colourless
$3d^0$	Ti^{4+}	colourless
$3d^1$	Ti^{3+}	purple
$3d^1$	V^{4+}	blue
$3d^2$	V^{3+}	green
$3d^3$	V^{2+}	violet
$3d^3$	Cr^{3+}	violet
$3d^4$	Mn^{3+}	violet
$3d^4$	Cr^{2+}	blue
$3d^5$	Mn^{2+}	pink
$3d^5$	Fe^{3+}	yellow
$3d^6$	Fe^{2+}	green
$3d^6 3d^7$	$Co^{3+}Co^{2+}$	bluepink
$3d^8$	Ni^{2+}	green
$3d^9$	Cu^{2+}	blue
$3d^{10}$	Zn^{2+}	colourless

Formation of interstetial compounds

Interstitial compounds are those which are formed when small atoms like H, C or N are trapped inside the crystal lattices of metals

For example, TiC, Mn_4N, Fe_3H, $VH_{0.56}$ and $TiH_{1.7}$, etc.

Compounds of transitions elements

Oxides

They form oxides with oxidation states +1 to +7

Potassium permanganate

$2MnO_2 + 4KOH + O_2 \rightarrow 2\,K_2MnO_4 + 2H_2O$

$3MnO_4^{2-} + 4H^+ + O_2 \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$

$2Mn^{2+} + 5\,S_2O_8^{2-} + 8H_2O \rightarrow 2MnO_4 + 10SO_4^{2-} + 16H^+$

$2KMnO_4 \rightarrow K_2MnO_4 + MnO_2 + O_2$

A few important oxidising reactions of $KMnO_4$ are given below:
(1) In acidic solutions:

- $10I^- + 2MnO_4^- + 16H^+ \rightarrow 2Mn^{2+} + 8H_2O + 5I_2$
- $5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$

(2) In neutral or faintly alkaline solutions:

- $2MnO_4^- + H_2O + I^- \longrightarrow 2MnO_2 + 2OH^- + IO_3^-$
- $8MnO_4^- + 3\,S_2O_3^{2-} + H_2O \longrightarrow 8MnO_2 + 6SO_4^{2-} + 2OH^-$

Potassium dichromate

$4FeCr_2O_4 + 8Na_2CO_3 + 7O_2 \rightarrow 8Na_2CrO_4 + 2Fe_2O_3 + 8CO_2$

$2Na_2CrO_4 + 2H^+ \rightarrow Na_2Cr_2O_7 + 2Na^+ + H_2O$

$Na_2Cr_2O_7 + 2KCl \rightarrow K_2Cr_2O_7 + 2NaCl$

$2CrO_4^{2-} + 2H^+ \rightarrow Cr_2O_7^{2-} + H_2O$
$Cr_2O_7^{2-} + 2OH^- \rightarrow 2CrO_4^{2-} + H_2O$

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

$6I \rightarrow 3I_2 + 6e^-$
$3H_2\,S \rightarrow 6H^+ + 3\,S + 6e^-$
$3Sn^{2+} \rightarrow 3Sn^{4+} + 6e^-$
$6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^-$

$Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

Physical properties of f-block elements

Atomic and Ionic Sizes
The overall decrease in atomic and ionic radii from lanthanum to lutetium.

Oxidation States

f-block elements exhibit multiple oxidation states, but the +3 oxidation state is most stable and common for both lanthanides and actinides.

General Characteristics
All the lanthanides are metals are soft, malleable and ductile in nature, not good conductors of heat and electricity are highly dense

Colour
The colour is due to partially filled *f*-orbitals which allow *f-f* transitions. M^{3+} ions having $4f^0$, $4f^7$ or $4f^{14}$ configurations are colourless.

Magnetic Properties
Ions having unpaired electrons are paramagnetic while those having all the orbitals paired are diamagnetic.

Properties of Actinoids

Electronic Configurations
All the actinoids are believed to have the electronic configuration of $7s^2$ and variable occupancy of the 5f and 6d subshells.

Ionic Sizes

The general trend in lanthanoids is observable in the actinoids as well. There is a gradual decrease in the size of atoms or M^{3+} ions across the series.

CHAPTER-5 COORDINATION COMPOUNDS

Werner's theory of coordination compounds
Primary valency: Ionisable, corresponds to oxidation state of the central metal atom/ion, satisfied by negative ions, non-directional.
Secondary valency: Non-ionisable, corresponds to coordination number of the central metal atom/ion, satisfied by neutral molecules or ligands, fixed for a metal, directional, giving definite geometry to the complex.

Terminology of Coordination Compounds :
Central metal atom/ion: Central metal atom/ ion to which one or more neutral molecules or ions are linked by coordinate bonds in a definite geometrical arrangement around it. It is also referred as Lewis acid.
Ligands: Ions or molecules which are linked to central metal atom/ion through coordinate bond by donating lone pair of electrons present on its donor atom and their types are:

- Unidentate ligand: Ligand bound to a metal ion through a single donor atom e.g., Cl^-, H_2O, NH_3.
- Didentate ligand : Ligand bound to a metal ion through two donor atoms e.g., $H_2NCH_2CH_2NH_2, C_2O_4^{2-}$.
- Chelate ligand : A didentate or a polydentate ligand which forms two or more coordinate bonds with central metal ion in such a way that a five or six membered ring is formed.
- Ambidentate ligand : A unidentate ligand which can coordinate through two different atoms e.g. NO_2^-, SCN^-.
- **Coordination number:** It is the total number of coordinate bonds through which the central metal atom/ion is attached with ligands.
- **Coordination entity/Coordination sphere:** The central metal atom/ion and the ligands which are directly attached to it, are enclosed in square bracket and is collectively known as coordination sphere. It is non-ionisable.
- **Counter ions:** These are the ionisable groups written outside the bracket.
- **Coordination polyhedron:** It is the spatial arrangement of the ligands around the central metal/atoms or ions.
- **Oxidation number:** It is the number that represents the charge on the central metal atom, if all the ligand are removed along with the electron pairs that are shared with central atom.
- **Homoleptic complex** in which metal atom/ion is bound to only one kind of ligands.
- **Heteroleptic complex** in which metal atom/ion is bound to more than one kind of ligands.

Nomenclature of coordination compounds:
Rules for writing the formula of coordination compounds:

- Formula of the cation whether simple or complex must be written first followed by anion.
- The coordination sphere is written in square brackets.
- Within the coordination sphere the sequence of symbols is, first the metal atom followed by anionic ligand then neutral ligand finally cationic ligand. Ligands of same type are arranged alphabetically. Polyatomic ligands are enclosed in parentheses.
- The number of cations or anions to be written in the formula is calculated on the basis that total positive charge must be equal to the total negative charge, as the complex as a whole is electrically neutral.

Rules for naming coordination compounds:

- The cation is named first then the anion.
- In naming coordination sphere, ligands are named first in alphabetical order followed by metal atom and then oxidation state of metal by a roman numeral in parentheses.
- Name of coordination compounds is started with a small letter and the complex part is written as one word.

Naming of ligands:

- Name of anionic ligands end in -o. e.g., Cl^{-}: Chlorido
- Neutral ligands (with a few exceptions) retain their names e.g., NH_3 : Ammine
- Name of cationic ligands end in - ium. e.g., NO_2^{+}: Nitronium
- Certain ligands are represented by abbreviations in parentheses instead of their complex structural formulae. e.g., ethylenediamine(en).
- Ambidentate ligands are named by using different names of ligands or by placing the symbol of donor atom. e.g., $-\mathrm{SCN}^{-}$ (Thiocyanato-S or Thiocyanato), $-\mathrm{NCS}^{-}$(Thiocyanato-N or Isothiocyanato),$-\mathrm{ONO}^{-}$(Nitrito-O or Nitrito), $-\mathrm{NO}_2^{-}$(Nitrito-N or Nitro)

Isomerism in coordination compounds
Isomers are two or more compounds that have the same chemical formula but a different arrangement of atoms.
The two principal types of isomerism are:
1). Stereoisomerism
2). Structural isomerism
1). Stereoisomerism

Coordination compounds having different positions and arragements of ligands in space.

Structural isomerism

Coordination compounds having different ligands within their coordination spheres.

- **Ionisation isomerism:** Complexes that give different ions in solution. e.g., $[\mathrm{Co}(\mathrm{NH}_3)_5\mathrm{Br}]\mathrm{SO}_4$ $[\mathrm{Co}(\mathrm{NH}_3)_5\mathrm{SO}_4]\mathrm{Br}$

- **Solvate isomerism:** Complexes that differ in number of water molecules present as ligands (inside the coordination sphere) and as free molecules (outside the coordination sphere). e.g., $[\mathrm{Cr}(\mathrm{H}_2\mathrm{O})_6]\mathrm{Cl}_3$ $[\mathrm{Cr}(\mathrm{H}_2\mathrm{O})_5\mathrm{Cl}]\mathrm{Cl}_2 \cdot \mathrm{H}_2\mathrm{O}$

- **Linkage isomerism:** Complexes that differ in the atom of a ligand directly bonded to the metal atom, shown by complexes containing ambidentate ligands. e.g., $[\mathrm{Co}(\mathrm{NH}_3)_5\,(\mathrm{NO}_2)]\mathrm{Cl}_2$ $[\mathrm{Co}(\mathrm{NH}_3)_5(\mathrm{ONO})]\mathrm{Cl}_2$

- **Coordination isomerism:** Occurs when both positive and negative ions are complex ions and two isomers differ in ligand's distribution in two complex ions. e.g. $[\mathrm{Co}(\mathrm{NH}_3)_6]\,[\mathrm{Cr}(\mathrm{CN})_6]$ $[\mathrm{Cr}(\mathrm{NH}_3)_6]\,[\mathrm{Co}(\mathrm{CN})_6]$

Stereoisomerism

Coordination compounds having different positions and arragements of ligands in space.

- **Geometrical isomerism:** Arises in heteroleptic complexes due to different possible geometric arrangements of the ligands.
- **Optical isomerism:** Occurs when a complex is represented by two different structures which are non-superimposable mirror images of each other called enantiomers.

Valence Bond Theory

Coordination Number	Type of Hybridisation	Shape
4	sp^3	Tetrahedral
4	dsp^2	Square Planar
5	sp^3d	Trigonal Bipyramidal
6	sp^3d^2	Octahedral
6	d^2sp^3	Octahedral

Crystal field theory

Crystal Field Theory explains the splitting of degenerate d-orbitals of transition metal ions in a ligand field, affecting their colour, magnetism, and stability.

For octahedral complexes

$$\mathrm{CFSE} = \left(-0.4 \times n_{t_{2g}} + 0.6 \times n_{e_g}\right)\Delta_0$$

$n_{t_{2g}}$ = Number of electrons in t_{2g} orbitals
n_{e_g} = Number of electrons in e_g orbitals
Δ_0 = Crystal field splitting energy for octahedral complexes

Tetrahedral complexes

$$\text{CFSE} = (-0.6 \times n_{t_2} + 0.4 \times n_e)\Delta_t$$

n_{t_2} = Number of electrons in t_2 orbitals
n_e = Number of electrons in e orbitals
Δ_t = Crystal field splitting energy for tetrahedral complexes

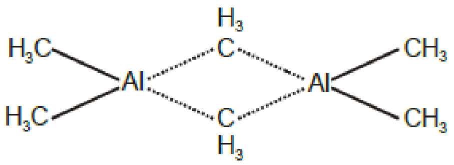
Colour of coordination compounds:

The magnitude of CFSE (Δ_o) for most of the complexes is of the same order as the energy of a photon of visible light. Hence, whenever $\boldsymbol{d}$ - $\boldsymbol{d}$ transition takes place, it imparts colour to the complex. The colour of the complex is the colour, complementary to the wavelength absorbed.

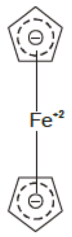
Bonding in Metal carbonyls:

Sigma(σ) bonded complexes:

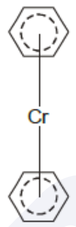
Bonding in Trimethyl aluminium is shown below



Pi(π) bonded complexes:

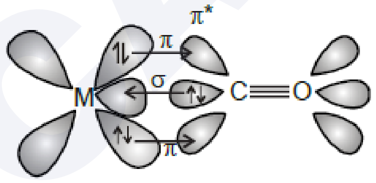


Ferrocene
$\mathrm{Fe}(\eta^5 - \mathrm{C_5H_5})_2$



Dibenzene chromium
$\mathrm{Cr}(\eta^6 - \mathrm{C_6H_6})_2$

Complexes containing both σ and π bonding characteristics:

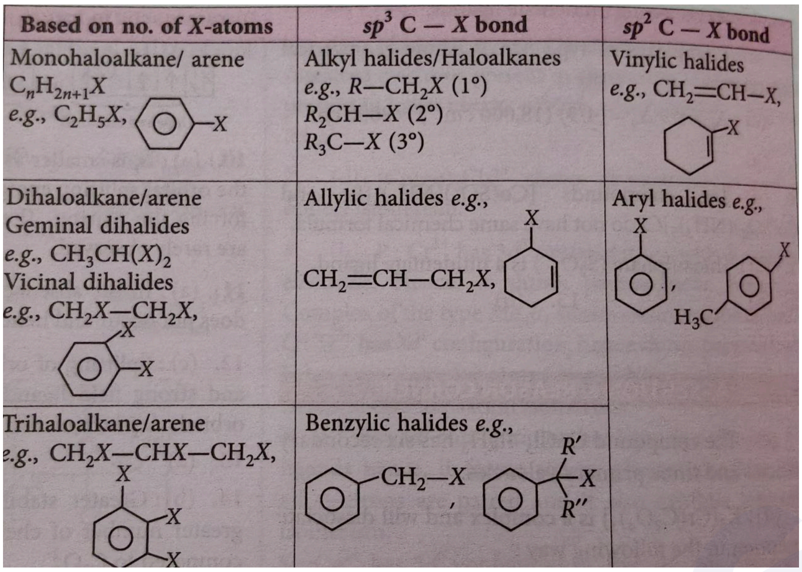


CHAPTER 6 - HALOALKANES AND HALOARENES

Haloalkanes are the compounds obtained from alkanes by the replacement of one or more hydrogens by corresponding number of halogen atoms ($-\mathrm{F}, -\mathrm{Cl}, -\mathrm{Br}_2, -\mathrm{I}$).

Haloarenes: The Hydrogen atom of the benzene nucleus is substituted by a halogen, and has the general formula C_6H_5X.
Aralkyl halide: Hydrogen atom of the side chain attached to benzene is substituted by a halogen e.g., benzyl halide $(C_6H_5CH_2X)$.

Classification



Based on no. of X-atoms	sp^3 C — X bond	sp^2 C — X bond
Monohaloalkane/ arene $C_nH_{2n+1}X$ e.g., C_2H_5X,	Alkyl halides/Haloalkanes e.g., $R-CH_2X$ (1°) R_2CH-X (2°) R_3C-X (3°)	Vinylic halides e.g., $CH_2=CH-X$,
Dihaloalkane/arene Geminal dihalides e.g., $CH_3CH(X)_2$ Vicinal dihalides e.g., CH_2X-CH_2X,	Allylic halides e.g., $CH_2=CH-CH_2X$,	Aryl halides e.g.,
Trihaloalkane/arene e.g., $CH_2X-CHX-CH_2X$,	Benzylic halides e.g.,	

Methods of Preparation of Haloalkanes

From Alcohols:

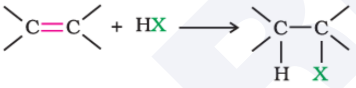
$$R-OH + HCl \xrightarrow{ZnCl_2} R-Cl + H_2O$$

From Hydrocarbons

(i) From alkanes by free radical halogenation

$$CH_3CH_2CH_2CH_3 \xrightarrow[\text{or heat}]{Cl_2/\text{UV light}} CH_3CH_2CH_2CH_2Cl + CH_3CH_2CHClCH_3$$

(II) From alkenes



$$CH_2=CH_2 + Br_2 \xrightarrow{CCl_4} \underset{\text{vic-Dibromide}}{BrCH_2-CH_2Br}$$

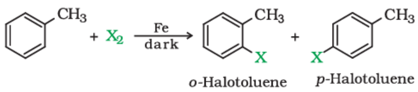
(III). From Halogen Exchange

$$R-X + NaI \longrightarrow R-I + NaX$$

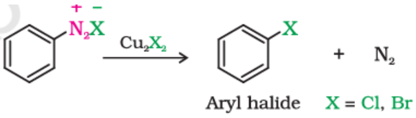
X=Cl, Br

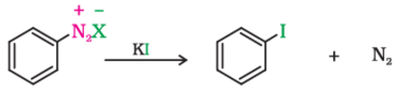
Preparation of Haloarenes

(1). From hydrocarbons by electrophilic substitution



(2). From amines by Sandmeyer's reaction





Physical properties of Haloalkanes and Haloarenes:

(1) These are less soluble in H_2O but more soluble in Organic solvents

(2) Their density follows the order:

Iodide > Bromide > Fluoride > Chloride

(3) Their boiling point follows the order:

Iodide > Bromide > Fluoride > Chloride

(4) Boiling point of Isomeric haloalkanes decreases with the increase in branching

(5) Boiling point of isomeric dihalobenzene is nearly the same and follows the order

para > ortho > meta

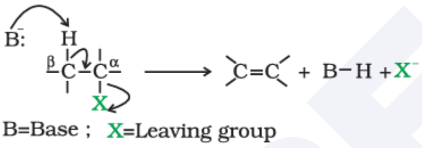
Chemical Reactions

(1). Nucleophilic substitution reactions

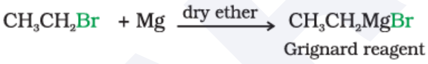
The general reaction occurs as follows:

$$R-CH_2-Cl+OH^- \rightarrow R-CH_2-OH$$

Elimination reactions



Reaction with metals



$$CH_3CH_2Br + Mg \xrightarrow{\text{dry ether}} \underset{\text{Grignard reagent}}{CH_3CH_2MgBr}$$

Wurtz reaction

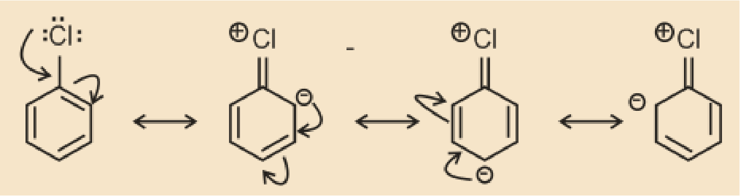
$$2RX + 2Na \longrightarrow RR + 2NaX$$

Reactions of Haloarenes

1. Nucleophilic substitution

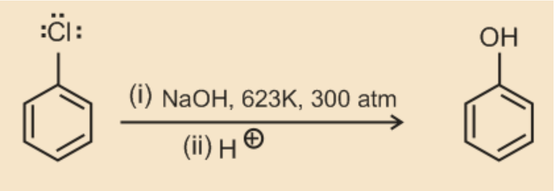
Aryl halides are extremely less reactive towards nucleophilic substitution reactions due to the following reasons:

- **Resonance effect:**

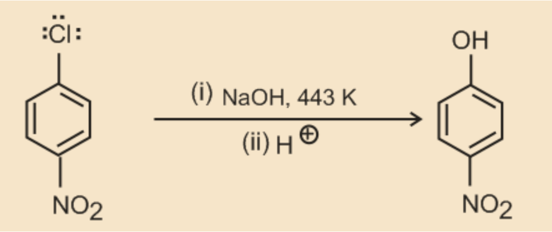


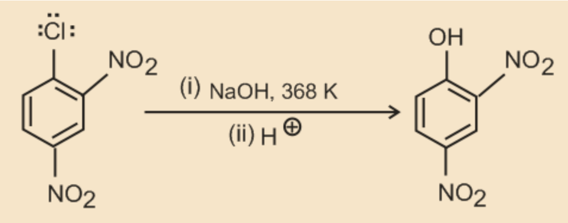
- **Difference in hybridisation of carbon atom in C—X bond**
- **Instability of phenyl cation**

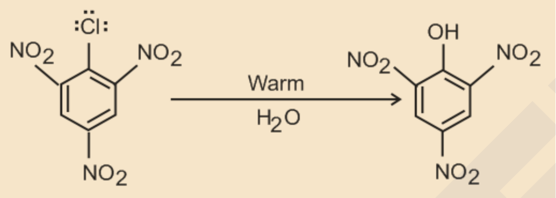
Chlorobenzene can be converted into phenol by heating in aqueous sodium hydroxide solution at a temperature of 623K and a pressure of 300 atmospheres.



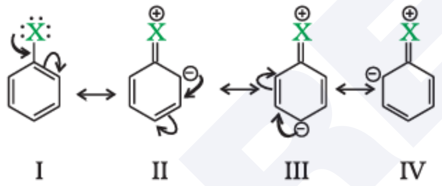
The presence of an electron withdrawing group ($-NO_2$) at ortho- and para-positions increases the reactivity of haloarenes.



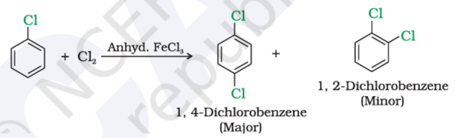




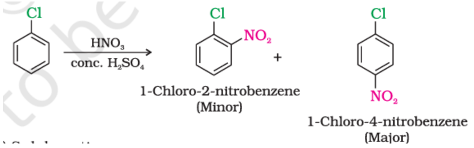
2. Electrophilic substitution reactions



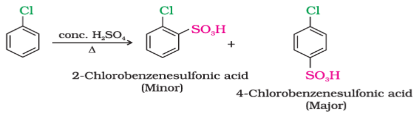
Halogination:



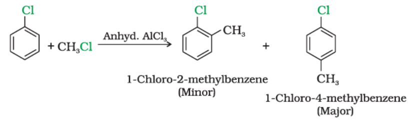
Nitration

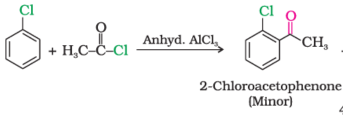


Sulphonation



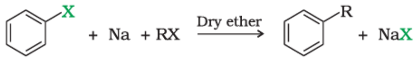
Friedel-Crafts reaction



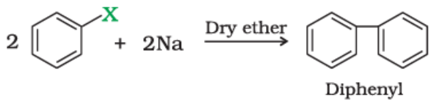


3. Reaction with metals

Wurtz-Fittig reaction



Fittig reaction



Polyhalogen compounds are the carbon compounds containing more than one halogen atom e.g.; $CHCl_3$, CH_2Cl_2.
Chloroform $(CHCl_3)$ **:**
Effects: It is oxidised to poisonous gas, carbonyl chloride, known as phosgene gas which causes liver and kidneys damage.

$$2CHCl_3 + O_2 \xrightarrow{\text{Light}} \underset{\text{Phosgene}}{2COCl_2} + 2HCl$$

Inhaling chloroform vapours depresses the central nervous system, causes dizziness, fatigue and headache.
Uses

- Its major use is in the production of freon refrigerant, R-22.
- It is used as a solvent for resins, rubbers, oils and fats, alkaloids, iodine and many other substances.
- It is used in preparation of chloretone (drug) and chloropicrin (insecticide) and used to preserve anatomical species.

Iodoform (CHI_3) is used as an antiseptic in dressing of wounds due to liberation of iodine and methylating agent in organic synthesis.
Freons: Chlorofluorocarbons are known as freons.
Effects: Freons cause disruption of ozone layer by initiating radical chain reactions in stratosphere and this anthropogenic compound is a greenhouse gas and effect is more than CO_2.
Uses : It is used as refrigerants, blowing agents, propellants in medical applications degreasing and solvent.

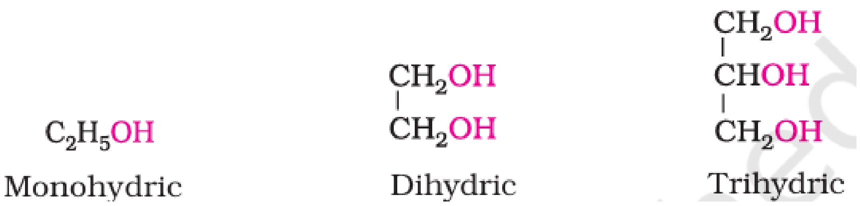
DDT (p, p'-Dichlorodiphenyltrichloroethane) :
Uses and Effects:

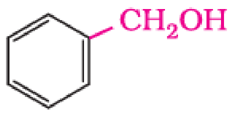
- In 1940, it was used as pesticides and is a persistent organic pollutant, strongly absorbed by soil.
- It is lipophilic so has a high potential to bioaccumulate.
- It may be directly genotoxic but may also induce enzymes to produce other genotoxic intermediates and DNA adducts.

Chapter - 7 Alcohols, phenols, and ethers

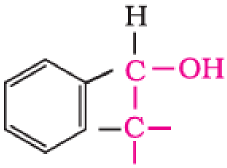
Alcohols are the hydroxy derivatives of alkanes having general formula $C_nH_{2n+1}OH$.
Alcohols are classified as monohydric and polyhydric alcohols depending upon the number of hydroxyl $(-OH)$ groups present in the molecule.

Alcohols Mono, Di, Tri or Polyhydric alcohols

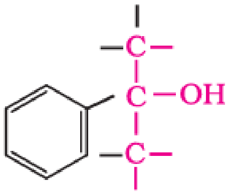




Primary

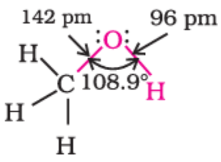


Secondary

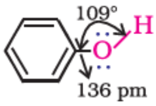


Tertiary

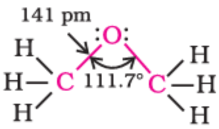
Structures of functional Groups



Methanol



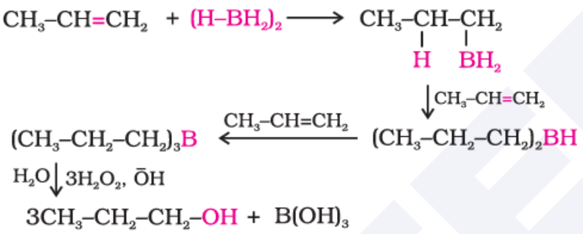
Phenol



Methoxymethane

Preparation of Alcohols

1. From alkenes



Propan-1-ol

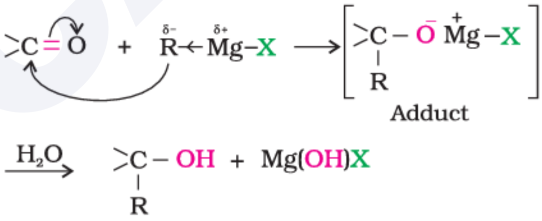
2. From carbonyl compounds

$$RCHO + H_2 \xrightarrow{Pd} RCH_2OH$$

$$RCOR' \xrightarrow{NaBH_4} R - \overset{CH}{\underset{OH}{OH}} - R'$$

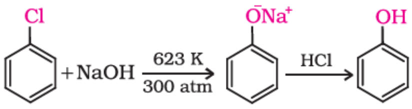
$$RCOOH \xrightarrow[(ii)\ H_2O]{(i)\ LiAlH_4} RCH_2OH$$

3. From Grignard reagents

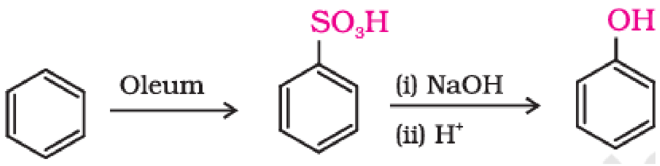


Preparation of Phenols

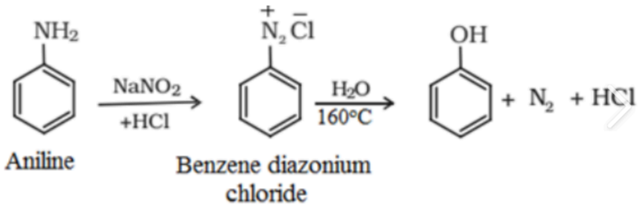
1. From haloarenes



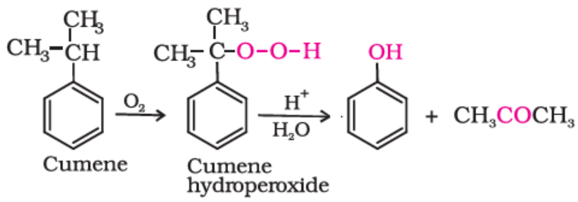
2. From benzenesulphonic acid



3. From diazonium salts



4. From cumene

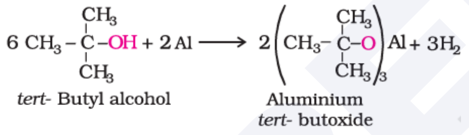


Chemical Reactions

(a) Reactions involving cleavage of $O-H$ bond

(i) Reaction with metals:

$$2R-O-H + 2Na \longrightarrow \underset{\text{Sodium alkoxide}}{2R-O-Na} + H_2$$

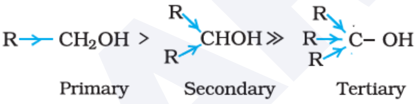


$$6\ CH_3-\overset{CH_3}{\underset{CH_3}{C}}-OH + 2\,Al \longrightarrow 2\left(CH_3-\overset{CH_3}{\underset{CH_3}{C}}-O\right)_3 Al + 3H_2$$

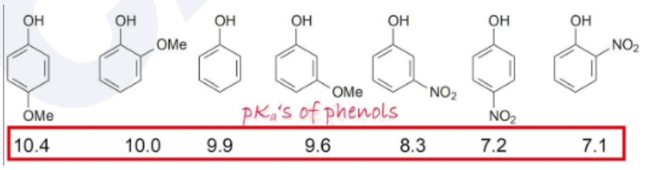
tert- Butyl alcohol

Aluminium *tert*- butoxide

(ii) Acidity of alcohols:



(iii) Acidity of phenols:



2. Esterification

$$Ar/ROH + R'-COOH \overset{H^+}{\rightleftharpoons} Ar/ROCOR' + H_2O$$

$$Ar/R-OH + (R'CO)_2O \overset{H^+}{\rightleftharpoons} Ar/ROCOR' + R'COOH$$

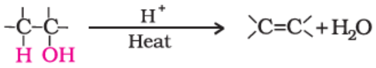
$$R/ArOH + R'COCl \xrightarrow{\text{Pyridine}} R/ArOCOR' + HCl$$

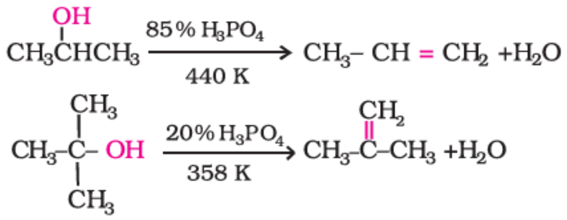
(b) Reactions involving cleavage of carbon - oxygen ($C-O$) bond in alcohols

1. Reaction with hydrogen halides:

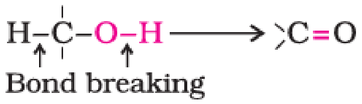
$ROH + HX \rightarrow R - X + H_2O$

2. Dehydration:





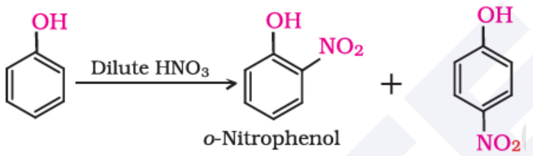
4. Oxidation:



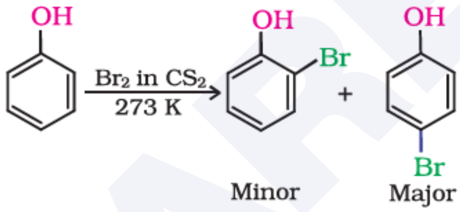
(c) Reactions of phenols

1. Electrophilic aromatic substitution

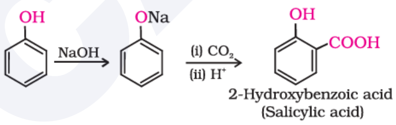
(i) Nitration:



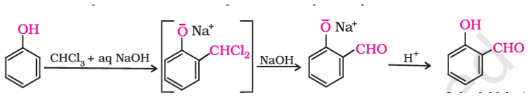
(ii) Halogenation:



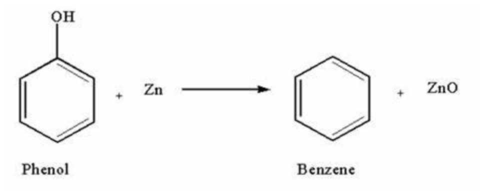
2. Kolbe's reaction



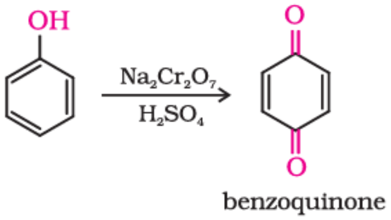
3. Reimer-Tiemann reaction



4. Reaction of phenol with zinc dust

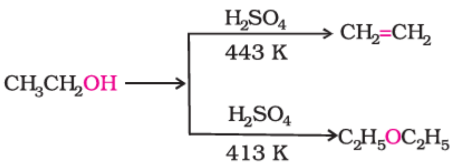


5. Oxidation



Preparation of Ethers

1. By dehydration of alcohols



2. Williamson synthesis

$R{-}X + R'{-}\overset{-}{\ddot{O}}\,\overset{+}{Na} \longrightarrow R{-}\ddot{O}{-}R' + Na\,X$

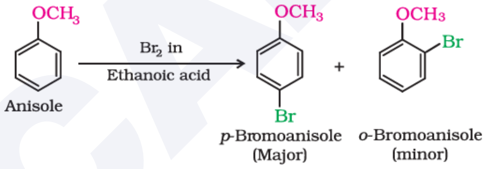
Chemical Reactions

1. Cleavage of $C-O$ bond in ethers

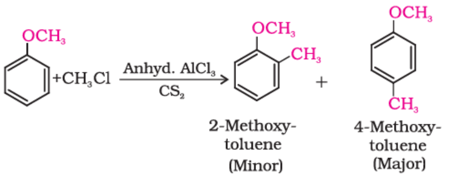
$$R-O-R+HX \longrightarrow RX+R-OH$$
$$R-OH+HX \longrightarrow R-X+H_2O$$

2. Electrophilic substitution

(i) Halogenation:



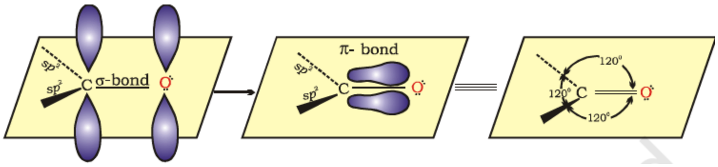
(ii) Friedel-Crafts reaction:



Chapter-8 Aldehydes, Ketones and Carboxylic acid

General formula: These are the compounds with carbon-oxygen double bond $(C=O)$ called carbonyl group having general formula $C_nH_{2n}O$.

Structure of the Carbonyl Group



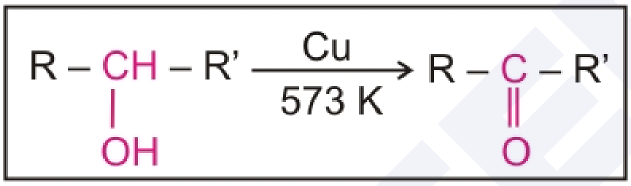
Preperation of Aldehydes and Ketones

Oxidation

$$CH_3-CH=CH-CH_2OH \xrightarrow{PCC} CH_3-CH=CH-CHO$$

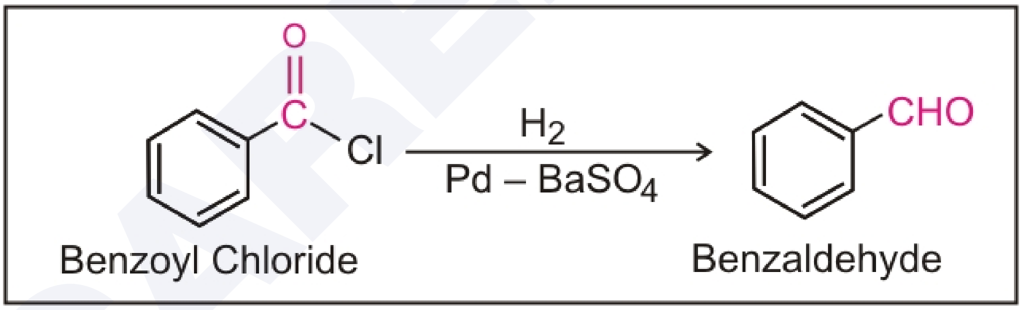
Dehydrogenation

$$RCH_2OH \xrightarrow[573\ K]{Cu} RCHO$$

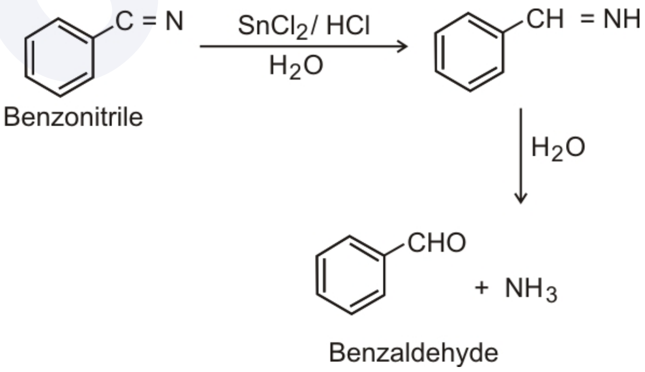


$$R-\underset{\displaystyle OH}{\underset{|}{CH}}-R' \xrightarrow[573\ K]{Cu} R-\underset{\displaystyle O}{\underset{\|}{C}}-R'$$

Rosenmund Reduction



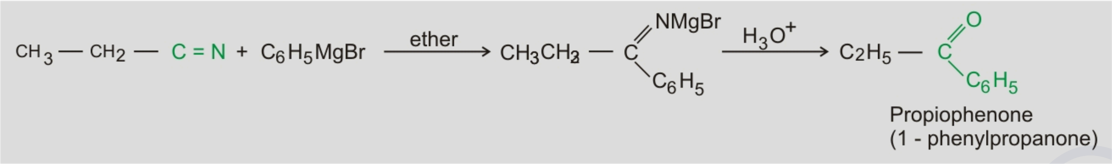
Stephen Reduction



From Acyl chlorides

$$2\,R'-\underset{\underset{O}{\|}}{C}-Cl + R_2Cd \longrightarrow 2\,R'-\underset{\underset{O}{\|}}{C}-R + CdCl_2$$

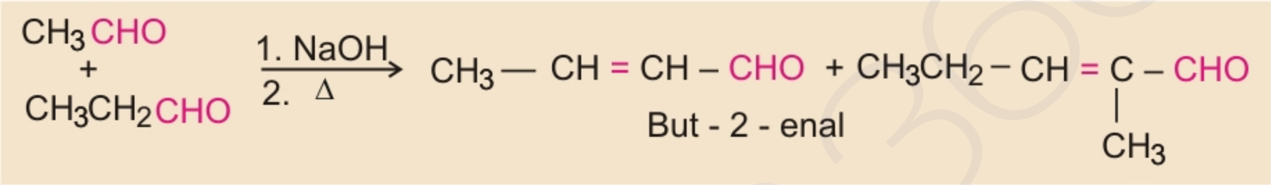
From Nitriles

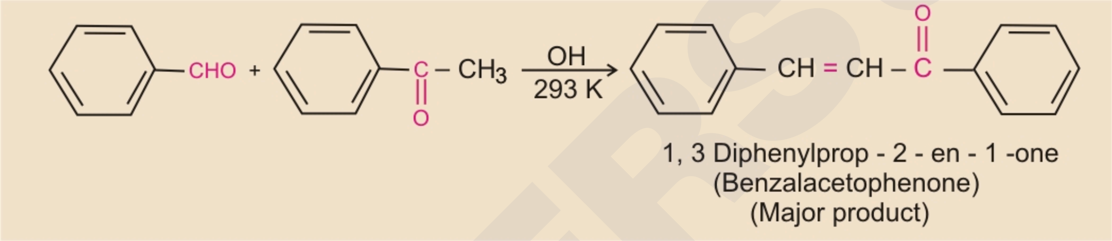


Propiophenone
(1 - phenylpropanone)

Chemical Reactions

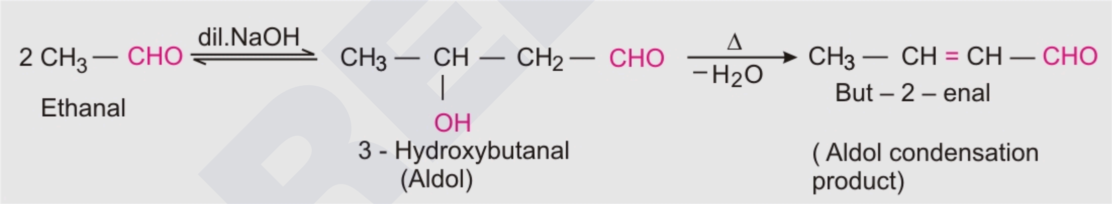
Intermolecular Aldol Condensation

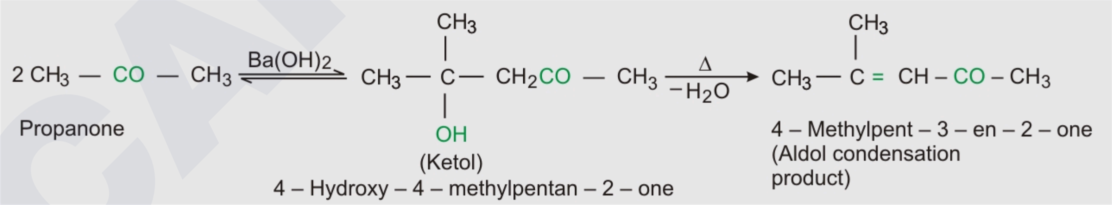




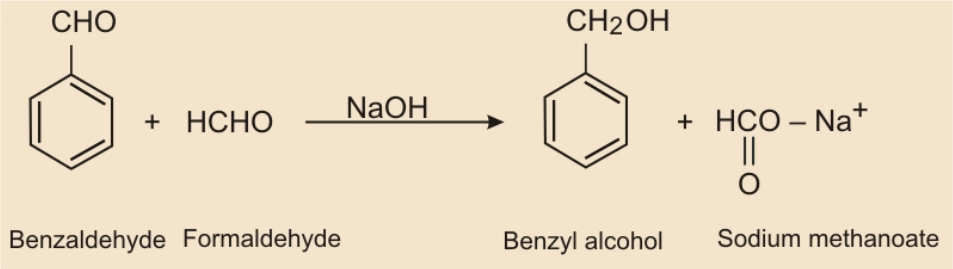
1, 3 Diphenylprop - 2 - en - 1 -one
(Benzalacetophenone)
(Major product)

Intramolecular Aldol Condensation

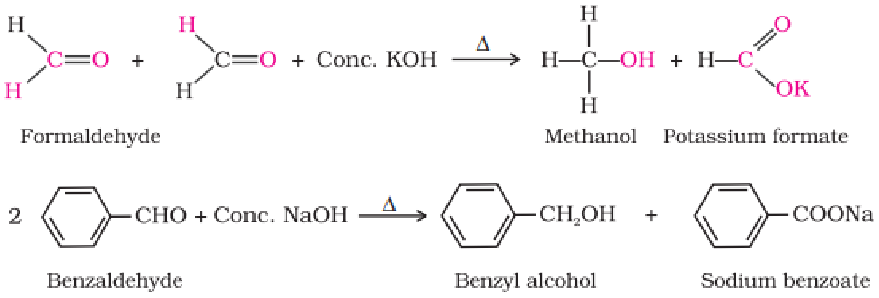




Intermolecular Cannizzaro Reaction



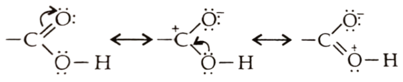
Intramolecular Cannizzaro Reaction



Carboxylic acid

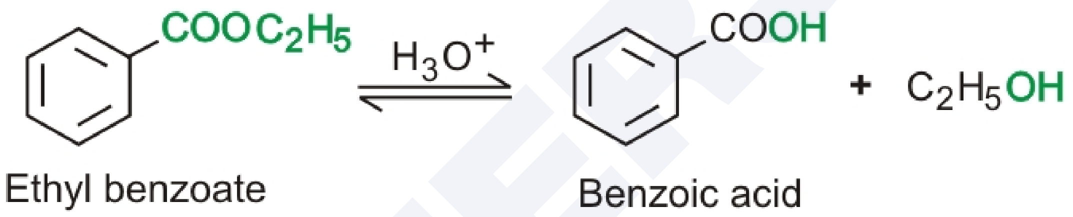
Compounds having - COOH group and general formula, $C_nH_{2n}O_2$

Structure:

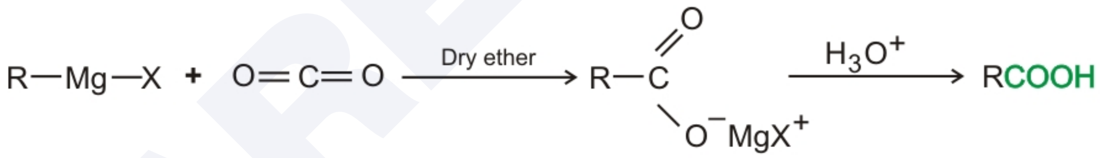


Methods of preparation of Carboxylic acids

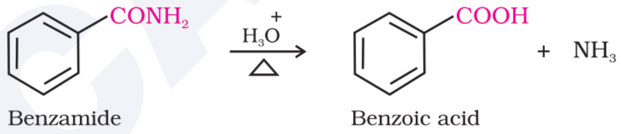
From Esters



From Grignard reagents



From Nitriles and Amides



Physical properties

Physical state : The lower fatty acids upto C_9 are colourless liquids. The higher ones are colourless waxy solids.
Odour: The first three members have a sharp pungent odour. The middle ones, C_4 to C_9, have an odour of rancid butter. The higher members do not possess any smell.
Boiling points: Carboxylic acids are higher boiling liquids than aldehydes, ketones and even alcohols of comparable molecular masses due to more extensive association of their molecules through intermolecular hydrogen bonding. The H-bonds are not broken completely even in the vapour phase.

Chemical properties

Formation of Anhydride
Carboxylic acid on heating with H_2SO_4 or P_2O_5 gives corresponding anhydride. The reaction occurs as follows:

Reaction with Ammonia

$$CH_3COOH + NH_3 \rightleftharpoons \underset{\text{Ammonium acetate}}{CH_3CO\overset{-}{O}\overset{+}{N}H_4} \xrightarrow[-H_2O]{\Delta} CH_3CONH_2$$

Reduction

$$R-COOH \xrightarrow[\text{(ii) } H_3O^+]{\text{(i) } LiAlH_4/\text{ ether or } B_2H_6} R-CH_2OH$$

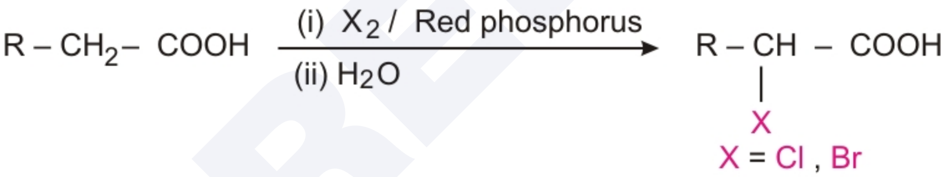
Decarboxylation

$$R-COONa \xrightarrow[\text{Heat}]{\text{NaOH \& CaO}} R-H + Na_2CO_3$$

Kolbe's electrolysis

$$\underset{\text{Sodium acetate}}{2\,CH_3COONa\,(aq.)} \xrightarrow{\text{Electrolysis}} \underbrace{CH_3-CH_3 + 2CO_2}_{\text{At anode}} + \underbrace{2\,NaOH + H_2}_{\text{At cathode}}$$

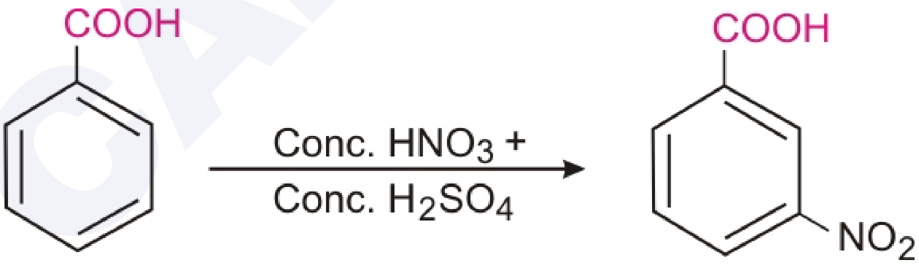
Halogenation:



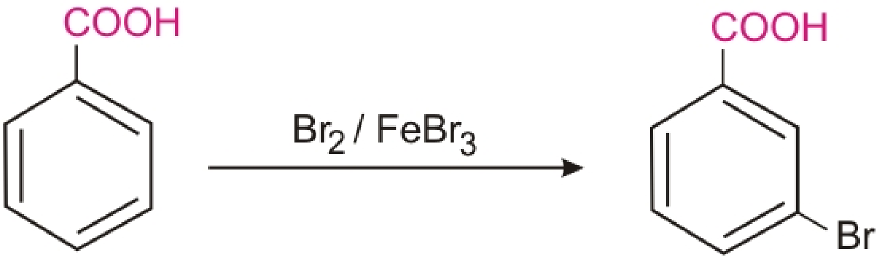
$$R-CH_2-COOH \xrightarrow[\text{(ii) } H_2O]{\text{(i) } X_2/\text{ Red phosphorus}} R-\underset{\underset{X}{|}}{CH}-COOH$$

X = Cl , Br

Ring substitution:

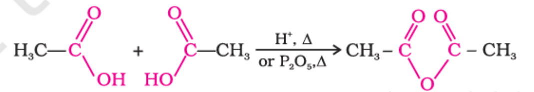


m - Nitrobenzoic acid



m - Bromobenzoic acid

Reactions Involving Cleavage of $C-OH$ Bond



Ethanoic acid

Ethanoic anhydride

Esterification

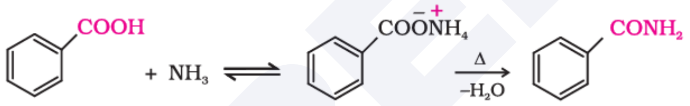
$$RCOOH + R'OH \overset{H^+}{\rightleftharpoons} RCOOR' + H_2O$$

Reactions with PCl_5, PCl_3 and $SOCl_2$

$RCOOH + PCl_5 \longrightarrow RCOCl + POCl_3 + HCl$
$3RCOOH + PCl_3 \longrightarrow 3RCOCl + H_3PO_3$
$RCOOH + SOCl_2 \longrightarrow RCOCl + SO_2 + HCl$

Reaction with ammonia

$$\underset{}{CH_3COOH} + NH_3 \rightleftharpoons \underset{\text{Ammonium acetate}}{CH_3\overset{-}{CO}O\overset{+}{N}H_4} \xrightarrow[-H_2O]{\Delta} \underset{\text{Acetamide}}{CH_3CONH_2}$$



Ammonium benzoate Benzamide

Reactions Involving $-COOH$ Group

Reduction

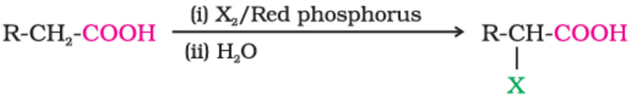
$$\text{R-COOH} \xrightarrow[\text{(ii) } H_3O^+]{\text{(i) } LiAlH_4/\text{ ether or } B_2H_6} R-CH_2OH$$

Decarboxylation

$$\text{R-COONa} \xrightarrow[\text{Heat}]{NaOH\&CaO} R-H + Na_2CO_3$$

Substitution Reactions in the Hydrocarbon Part

1. Halogenation

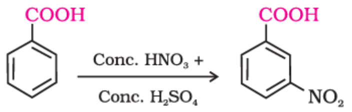


$$\text{R-CH}_2\text{-COOH} \xrightarrow[\text{(ii) } H_2O]{\text{(i) } X_2/\text{Red phosphorus}} \underset{\underset{X}{|}}{\text{R-CH-COOH}}$$

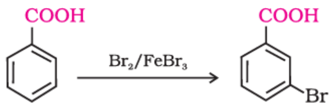
X = Cl, Br

α – Halocarboxylic acid

2. Ring substitution



m-Nitrobenzoic acid

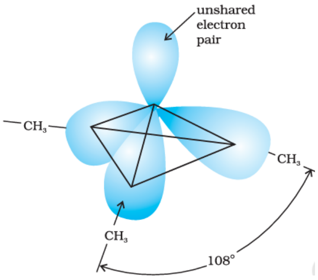


m-Bromobenzoic acid

Chapter - 9 Amines

Amines are the ammonia derivatives obtained by replacing one, two or three hydogen atoms by alkyl/aryl groups.

Structure of Amines.

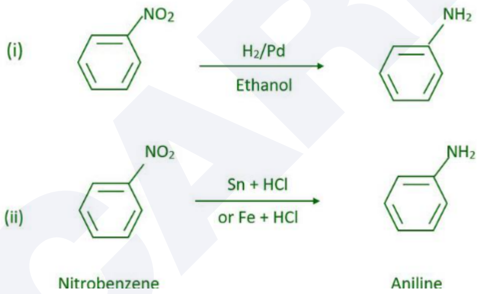


Classification:

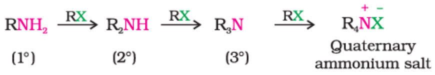
Amines can be classified as $1^{\circ}, 2^{\circ}$ or 3° amines depending on the number of alkyl/aryl groups present in ammonia molecule, RNH_2 (1° amine), $\mathrm{R}_2\mathrm{NH}$ (2° amine) and R_3 N (3° amine).

Preparation of Amines

1. Reduction of nitro compounds



2. Ammonolysis of alkyl halides



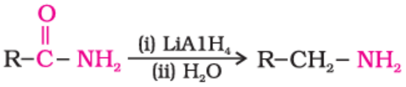
$$\underset{(1^\circ)}{RNH_2} \xrightarrow{RX} \underset{(2^\circ)}{R_2NH} \xrightarrow{RX} \underset{(3^\circ)}{R_3N} \xrightarrow{RX} \underset{\text{Quaternary ammonium salt}}{R_4\overset{+}{N}\overset{-}{X}}$$

$$R\text{–}\overset{+}{N}H_3\overset{-}{X} + NaOH \longrightarrow R\text{–}NH_2 + H_2O + \overset{+}{Na}\overset{-}{X}$$

3. Reduction of nitriles

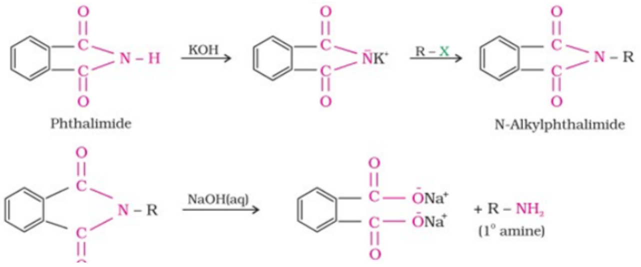
$$\mathrm{R}-\mathrm{C} \equiv \mathrm{N} \xrightarrow[\mathrm{Na(Hg)}/\mathrm{C_2H_5OH}]{\mathrm{H_2/Ni}} \mathrm{R}-\mathrm{CH}_2-\mathrm{NH}_2$$

4. Reduction of amides

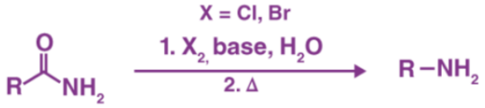


$$R-\overset{O}{\overset{\|}{C}}-NH_2 \xrightarrow[\text{(ii) } H_2O]{\text{(i) } LiAlH_4} R-CH_2-NH_2$$

5. Gabriel phthalimide synthesis



6. Hoffmann bromamide degradation reaction

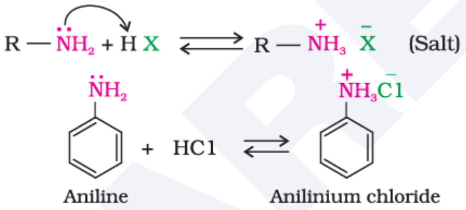


Physical Properties:

- Lower aliphatic amines are gases with ammoniacal smell, primary amines with C_3 or more are liquids with fishy odour.
- Lower aromatic amines are liquids with characteristic unpleasant odour but higher ones are solids which are odourless.
- **Colour:** Pure amines are colourless but get coloured due to atmospheric oxidation.
- **Solubility:** Amines are soluble in water as they can form hydrogen bonds with water and solubility decreases as the size of alkyl group increases, higher amines and aromatic amines are insoluble in water but soluble in organic solvents.
- **Boiling points:** Amines being polar, form intermolecular hydrogen bonds and exist as associated molecules, therefore, shows higher boiling points than hydrocarbons of comparable molecular masses. The order of boiling points of isomeric amines is $1^\circ > 2^\circ > 3^\circ$ due to the presence of two hydrogen atoms in 1° amine, one hydrogen atom in 2° amine and no hydrogen atom in 3° amine.

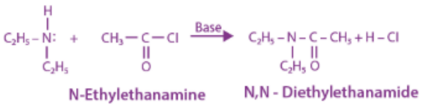
Chemical Reactions

1. Basic character of amines



2. Alkylation

3. Acylation



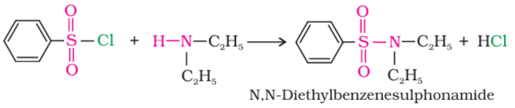
4. Carbylamine reaction

$$R-NH_2 + CHCl_3 + 3KOH \xrightarrow{\text{Heat}} R-NC + 3KCl + 3H_2O$$

5. Reaction with nitrous acid

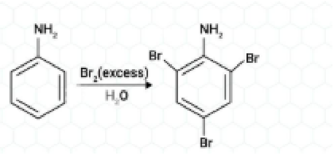
$$R-NH_2+HNO_2\xrightarrow{NaNO_2+HCl}\left[R-\overset{+}{N_2}\,Cl\right]\xrightarrow{H_2O}ROH+N_2+HCl$$

6. Reaction with arylsulphonyl chloride

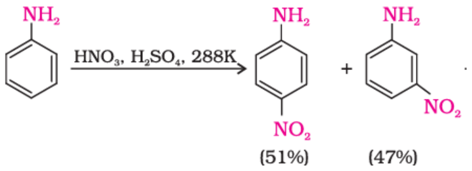


7. Electrophilic substitution

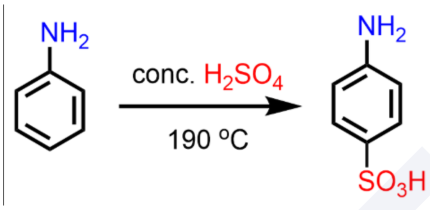
(a) Bromination



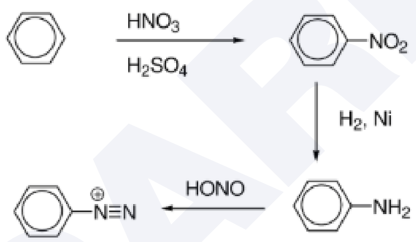
(b) Nitration:



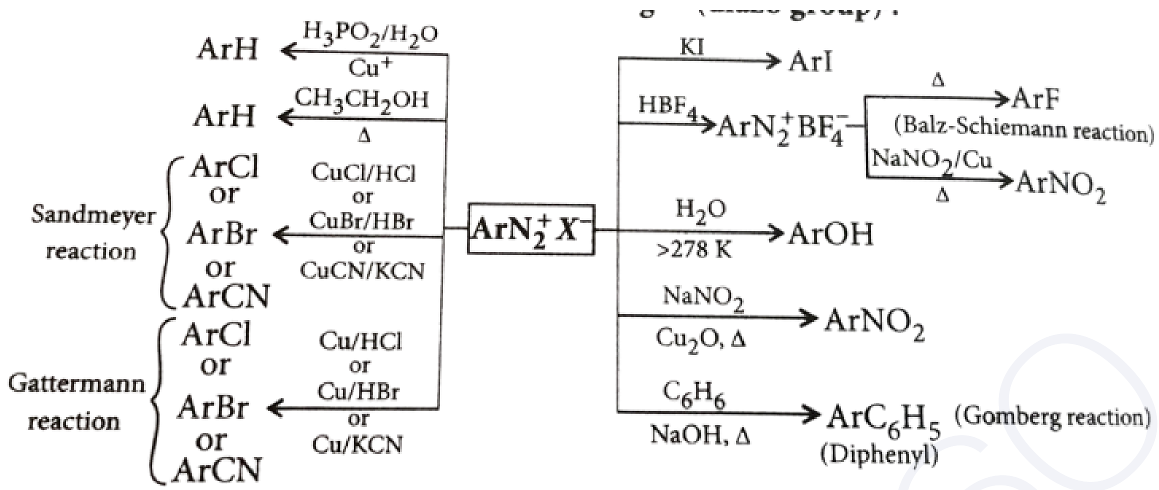
(c) Sulphonation:



II. DIAZONIUM SALTS



Chemical Reactions



Chapter - 10 Biomolecules

Carbohydrates:

They are optically active polyhydroxy aldehydes or ketones or the compounds that produce such units on hydrolysis. Most of them have a general formula, $C_x(H_2O)_y$ and are called as hydrates of carbon.

Classification of Carbohydrates

(i) Monosaccharides

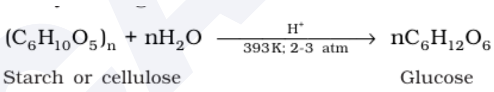
Cannot be hydrolysed further. $(CH_2O)_n$, where $n = 3 - 7$ e.g. glucose, fructose, ribose, etc.

Preperation of Glucose

1. From sucrose (Cane sugar):

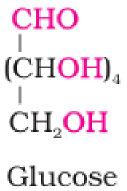
$$C_{12}H_{22}O_{11} + H_2O \xrightarrow{H^+} C_6H_{12}O_6 + C_6H_{12}O_6$$

2. From starch:

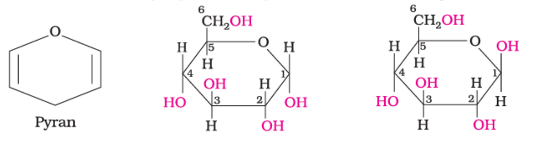


$$\underset{\text{Starch or cellulose}}{(C_6H_{10}O_5)_n} + nH_2O \xrightarrow[393\,K;\ 2\text{-}3\ atm]{H^+} \underset{\text{Glucose}}{nC_6H_{12}O_6}$$

Structure of Glucose



Cyclic Structure of Glucose



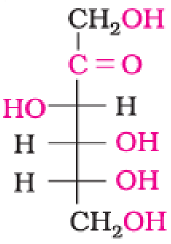
Pyran

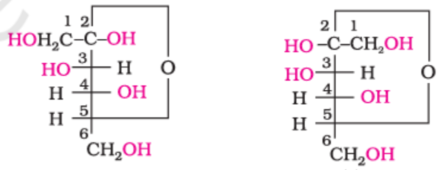
α – D – (+) – Glucopyranose

β – D – (+) – Glucopyranose

Fructose

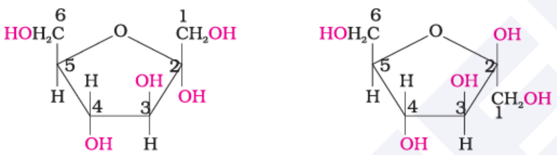
It is a keto hexoses as it contains six carbon atoms and ketonic group.





– D – (–) – Fructofuranose

β – D – (–) – Fructofuranose



α – D – (–) – Fructofuranose

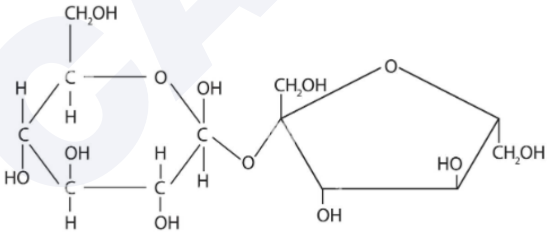
β – D – (–) – Fructofuranose

(ii) Oligosaccharides:

Yield 2-10 monosaccharides on hydrolysis. $(C_6H_{10}O_5)_n$, where $n = 2 - 10$ e.g. sucrose, maltose, etc.

1. Sucrose:

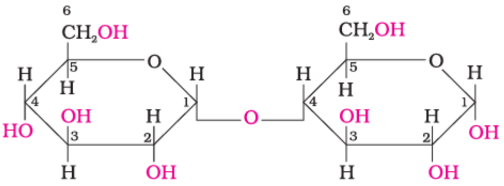
$$\underset{\text{Sucrose}}{C_{12}H_{22}O_{11}} + H_2O \longrightarrow \underset{\text{D-(+)-Glucose}}{C_6H_{12}O_6} + \underset{\text{D-(–)-Fructose}}{C_6H_{12}O_6}$$



Glucose

Fructose

2. Maltose:

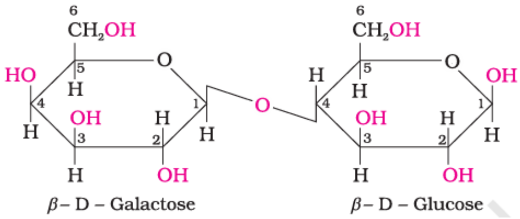


(I)
α – D – Glucose

(II)
α – D – Glucose

Maltose

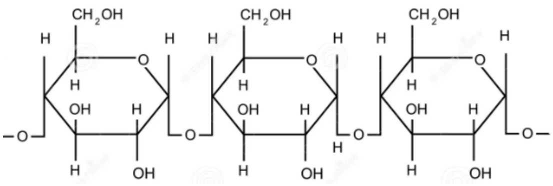
3. Lactose:



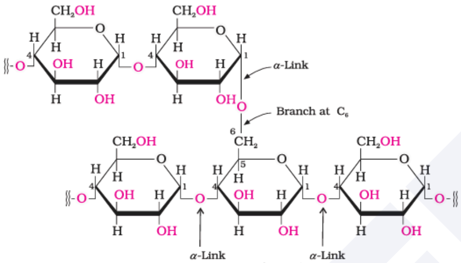
(iii)Polysaccharides:

Yield large number of monosaccharides on hydrolysis. $(C_6H_{10}O_5)_n$ where $n = 100 - 3000$ e.g. starch, cellulose, etc.

1. Starch:

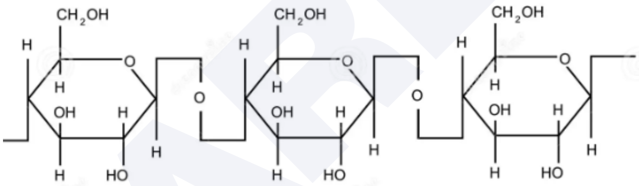


Amylose



Amylopectin

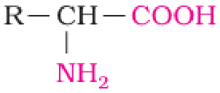
2. Cellulose:



Protein

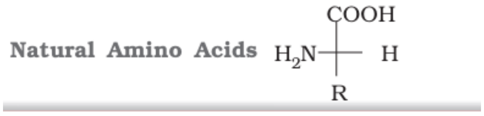
They are biomoolecules of living system made up of nitrogeneous organic compounds by condensation polymerisation of a- amino acids.

1. Amino acids



$R-\underset{\underset{NH_2}{|}}{CH}-COOH$

α-amino acid
(R = side chain)



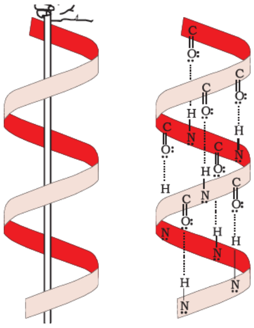
2. Proteins

They are classified into two types

(a) Fibrous proteins

(b) Globular proteins

Structure of a- helical protein



3. Enzymes

They are the biocatalysts products by livig cell which catalyse biochemical reaction in living organisms.

4. Vitamins

These are the complex organic molecules which cannot be produced by the body and must be supplied in small amount in diet.

Classification of vitamins

i). Fat soluble

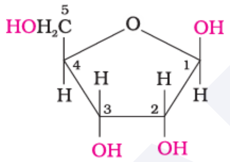
ii). Water soluble

5. Nucleic acids

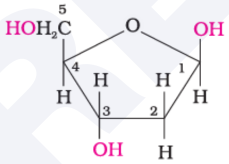
Polymers of nucleotides present in nucleous of all living cells.

Chemical Composition of Nucleic Acids

Complete hydrolysis of DNA (or RNA) yields a pentose sugar, phosphoric acid and nitrogen containing heterocyclic compounds (called bases).

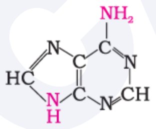


β–D–ribose

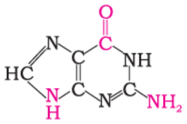


β–D–2–deoxyribose

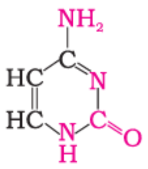
4 bases of DNA



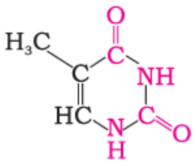
Adenine (A)



Guanine (G)

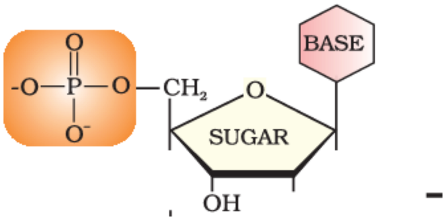


Cytosine (C)



Thymine (T)

Structure of Nucleic Acids



Hormones

They are the molecules that acts as intercellular messanger and are poured directly in the blood stream by endocrine gland.