

Class 11 Chemistry: Formulas and Named Reactions

All 9 NCERT chapters (Part I and Part II) · CBSE syllabus 2026-27 · NCERT textbook (Reprint 2026-27)

How to use this sheet. Every box covers the CBSE 2026-27 board-exam syllabus. Chapters 1, 2, 5 and 6 hold most of the numerical formulas. Chapters 3, 4, 7 and 8 list the rules, trends and tests. Chapter 9 lists every named reaction in the book. s- and p-Block Elements and The Gaseous State are assessed only formatively in 2026-27, and they are not in this textbook, so they are not covered here. Constants: $N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$, $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1} = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}$, $h = 6.626 \times 10^{-34} \text{ J s}$, $c = 3.0 \times 10^8 \text{ m s}^{-1}$, $m_e = 9.1 \times 10^{-31} \text{ kg}$, $e = 1.602 \times 10^{-19} \text{ C}$.

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1 Some Basic Concepts of Chemistry

LAWS OF CHEMICAL COMBINATION

- Conservation of mass; definite proportions; multiple proportions
- Gay Lussac: gases react in simple whole-number volume ratios (same T , p)
- Avogadro: equal volumes of gases at the same T , p contain equal numbers of molecules

ATOMIC AND MOLECULAR MASS

- $1 \text{ u (amu)} = \frac{1}{12}$ the mass of one ^{12}C atom
- Average atomic mass = $\sum (\text{fractional abundance} \times \text{isotopic mass})$
- Molecular mass = sum of atomic masses; formula mass for ionic solids

MOLE CONCEPT

- $1 \text{ mol} = 6.02214076 \times 10^{23}$ entities (N_A)
- $$n = \frac{\text{mass}}{\text{molar mass}} = \frac{N}{N_A}$$
- Molar mass in g mol^{-1} = molecular mass in u
- The book uses 22.7 L as the molar volume of a gas at STP (1 bar, 273.15 K).

COMPOSITION AND FORMULAE

- Mass % of an element = $\frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100$
- Empirical formula: simplest whole-number mole ratio
- Molecular formula = $n \times \text{empirical formula}$, $n = \frac{\text{molar mass}}{\text{empirical formula mass}}$

CONCENTRATION OF SOLUTIONS

- Mass % = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
- Mole fraction $x_A = \frac{n_A}{n_A + n_B}$; $x_A + x_B = 1$
- Molarity
- $$M = \frac{\text{moles of solute}}{\text{volume of solution in L}}$$
- Molality $m = \frac{\text{moles of solute}}{\text{mass of solvent in kg}}$
- Dilution: $M_1 V_1 = M_2 V_2$
- Molarity changes with temperature (volume changes); molality does not.

STOICHIOMETRY

- Balanced equation gives mole (and gas volume) ratios
- Limiting reagent: the reactant used up first; it fixes the amount of product
- e.g. $\text{CH}_4 + 2 \text{O}_2 \longrightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$: 1 mol CH_4 needs 2 mol O_2



2 Structure of Atom

SUB-ATOMIC PARTICLES

- $e/m_e = 1.758820 \times 10^{11} \text{ C kg}^{-1}$ (Thomson)
- $e = 1.602 \times 10^{-19} \text{ C}$ (Millikan); $m_e = 9.1094 \times 10^{-31} \text{ kg}$
- Mass number $A = Z + n$ (protons + neutrons)
- Isotopes: same Z ; isobars: same A

WAVES AND QUANTA

- $c = \nu\lambda$; wavenumber $\bar{\nu} = 1/\lambda$
- Planck:

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

- Photoelectric effect: $h\nu = h\nu_0 + \frac{1}{2}m_e v^2$ ($h\nu_0$ = work function)
- No electrons are emitted below the threshold frequency ν_0 , however intense the light.

HYDROGEN SPECTRUM

- $\bar{\nu} = 109677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{ cm}^{-1}$
- 109677 cm^{-1} = Rydberg constant for hydrogen; $n_2 = n_1 + 1, n_1 + 2, \dots$

Series	n_1	Region
Lyman	1	ultraviolet
Balmer	2	visible
Paschen	3	infrared
Brackett	4	infrared
Pfund	5	infrared

BOHR MODEL

- Angular momentum: $m_e v r = \frac{nh}{2\pi}$
- Energy: $E_n = -R_H \left(\frac{1}{n^2} \right)$, $R_H = 2.18 \times 10^{-18} \text{ J}$
- Hydrogen-like ion:

$$E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2} \right) \text{ J}$$

- Radius: $r_n = \frac{52.9 n^2}{Z} \text{ pm}$ ($a_0 = 52.9 \text{ pm}$)
- $\Delta E = 2.18 \times 10^{-18} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \text{ J}$; $\nu = \Delta E/h$

DUAL NATURE AND UNCERTAINTY

- de Broglie:

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

- Heisenberg:

$$\Delta x \times \Delta p_x \geq \frac{h}{4\pi}$$

$$\text{equivalently } \Delta x \times \Delta v_x \geq \frac{h}{4\pi m}$$

QUANTUM NUMBERS

- $n = 1, 2, 3, \dots$ (shell); max electrons in a shell = $2n^2$
- $l = 0, 1, \dots, n-1$ (s, p, d, f); orbitals in a subshell = $2l + 1$
- $m_l = -l, \dots, 0, \dots, +l$; $m_s = \pm \frac{1}{2}$
- Nodes: radial = $n - l - 1$, angular = l , total = $n - 1$

FILLING ORBITALS

- Aufbau: lower $(n + l)$ fills first; equal $(n + l)$: lower n first
- Order: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s
- Pauli: no two electrons have all four quantum numbers the same
- Hund: pairing starts only after each orbital of the subshell has one electron
- Half-filled and fully filled subshells are extra stable: Cr is $[\text{Ar}]3d^5 4s^1$, Cu is $[\text{Ar}]3d^{10} 4s^1$.



3 Classification of Elements and Periodicity

MODERN PERIODIC LAW

- Properties are periodic functions of atomic number
- 7 periods, 18 groups; period number = highest n
- s-block: ns^{1-2} ; p-block: ns^2np^{1-6} ; d-block: $(n-1)d^{1-10}ns^{0-2}$; f-block: $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$

NAMING ELEMENTS WITH $Z > 100$

0	1	2	3	4
nil (n)	un (u)	bi (b)	tri (t)	quad (q)
5	6	7	8	9
pent (p)	hex (h)	sept (s)	oct (o)	enn (e)

- Add "ium": $Z = 120$ is unbinilium (Ubn).

PERIODIC TRENDS

Property	Across a period (left to right)	Down a group
Atomic radius	decreases	increases
Ionization enthalpy	generally increases	decreases
Electron gain enthalpy	becomes more negative	becomes less negative
Electronegativity	increases	decreases
Metallic character	decreases	increases

- Cation smaller than its atom; anion larger. Isoelectronic species: higher Z means smaller radius ($O^{2-} > F^- > Na^+ > Mg^{2+}$)
- Exceptions: IE of B < Be and O < N (stable half-filled/filled subshells). Oxides on the left are basic, on the right acidic, in the centre amphoteric or neutral.

4 Chemical Bonding and Molecular Structure

FORMAL CHARGE

- $$F.C. = V - L - \frac{1}{2}S$$
- V = valence electrons in the free atom, L = non-bonding (lone pair) electrons, S = bonding (shared) electrons
- The lowest-energy Lewis structure has the smallest formal charges.

BOND PARAMETERS

- Bond order (Lewis) = number of bonds between two atoms
- Higher bond order: higher bond enthalpy, shorter bond length
- Dipole moment: $\mu = Q \times r$; 1 D = 3.33564×10^{-30} C m
- $\mu = 0$ for symmetric molecules (CO_2 , BF_3 , CH_4); $\mu(NH_3) > \mu(NF_3)$.

VSEPR SHAPES

Bond pairs + lone pairs	Shape	Examples
2 + 0	linear	$BeCl_2$
3 + 0 / 2 + 1	trigonal planar / bent	BF_3 / SO_2
4 + 0 / 3 + 1 / 2 + 2	tetrahedral / trigonal pyramidal / bent	CH_4 / NH_3 / H_2O
5 + 0 / 4 + 1 / 3 + 2	trigonal bipyramidal / see-saw / T-shape	PCl_5 / SF_4 / ClF_3
6 + 0 / 5 + 1 / 4 + 2	octahedral / square pyramidal / square planar	SF_6 / BrF_5 / XeF_4

- Repulsion order: lp-lp > lp-bp > bp-bp. This is why H_2O (104.5°) < NH_3 (107°) < CH_4 (109.5°).

HYBRIDISATION

Type	Geometry	Angle
sp	linear	180°
sp ²	trigonal planar	120°
sp ³	tetrahedral	109.5°
sp ³ d	trigonal bipyramidal	90° , 120°
sp ³ d ²	octahedral	90°

MOLECULAR ORBITAL THEORY

- Bond order = $\frac{1}{2}(N_b - N_a)$
- $N_b > N_a$: stable; $N_b \leq N_a$: unstable
- $H_2 = 1$, $He_2 = 0$, $Li_2 = 1$, $C_2 = 2$, $N_2 = 3$, $O_2 = 2$
- Unpaired electrons: paramagnetic (O_2); all paired: diamagnetic
- Up to N_2 , $\pi 2p_x = \pi 2p_y$ lies below $\sigma 2p_z$; for O_2 and F_2 , $\sigma 2p_z$ lies below.



HYDROGEN BONDING

- H bonded to F, O or N attracts a nearby electronegative atom
- Intermolecular: HF, water, alcohols
- Intramolecular: o-nitrophenol
- Strength order of H-bonding: solid > liquid > gas.

5 Thermodynamics**FIRST LAW**

- $\Delta U = q + w$
(q, w positive when added to the system)
- Expansion work: $w = -p_{ex}\Delta V$
- Isothermal reversible:

$$w_{rev} = -2.303 nRT \log \frac{V_f}{V_i}$$

- Free expansion (vacuum): $w = 0$
- U, H, S, G are state functions; q and w are not.

REACTION ENTHALPIES

- $\Delta_r H^\circ = \sum a_i \Delta_f H^\circ(\text{products}) - \sum b_i \Delta_f H^\circ(\text{reactants})$
- Gases: $\Delta_r H^\circ = \sum \text{bond enthalpies (reactants)} - \sum \text{bond enthalpies (products)}$
- Hess's law: $\Delta_r H = \Delta_r H_1 + \Delta_r H_2 + \dots$ (path-independent)
- $\Delta_f H^\circ$ of an element in its standard state = 0

ENTROPY AND GIBBS ENERGY

- $\Delta S = \frac{q_{rev}}{T}$; spontaneous change: $\Delta S_{total} > 0$
- $\Delta_r G = \Delta_r H - T\Delta_r S$
- $\Delta G < 0$: spontaneous; $\Delta G = 0$: equilibrium
- $\Delta_r G^\circ = -RT \ln K = -2.303 RT \log K$
- Third law: entropy of a perfect crystal $\rightarrow 0$ as $T \rightarrow 0$ K

ENTHALPY AND HEAT CAPACITY

- $\Delta U = q_V$ (constant volume); $\Delta H = q_p$ (constant pressure)
- $\Delta H = \Delta U + \Delta n_g RT$
- $q = C\Delta T = mc\Delta T$
- Ideal gas: $C_p - C_V = R$

TYPES OF ENTHALPY CHANGE

- Phase changes: fusion, vaporisation, sublimation (always positive, at constant T)
- Combustion, atomization, bond dissociation, solution, lattice enthalpy
- Lattice enthalpy from the Born-Haber cycle; $\Delta_{sol} H = \Delta_{lattice} H + \Delta_{hyd} H$

EFFECT OF TEMPERATURE ON SPONTANEITY

$\Delta_r H$	$\Delta_r S$	$\Delta_r G$
–	+	– at all T
–	–	– at low T
+	+	– at high T
+	–	+ at all T

6 Equilibrium**EQUILIBRIUM CONSTANTS**

- $aA + bB \rightleftharpoons cC + dD$:
$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$
- K_p : same form with partial pressures
- $K_p = K_c (RT)^{\Delta n}$
(Δn = moles of gaseous products – reactants)
- Reverse reaction: $K'_c = 1/K_c$; equation multiplied by n :
 $K''_c = K_c^n$
- Pure solids and liquids are left out of K (heterogeneous equilibria).

REACTION QUOTIENT AND GIBBS ENERGY

- $Q_c < K_c$: net forward; $Q_c > K_c$: net reverse; $Q_c = K_c$: at equilibrium
- $\Delta G = \Delta G^\circ + RT \ln Q$
- $\Delta G^\circ = -RT \ln K$, so $K = e^{-\Delta G^\circ / RT}$
- $\Delta G^\circ < 0$ gives $K > 1$ (products favoured).



LE CHATELIER'S PRINCIPLE

- Adding reactant / removing product: shifts forward
- Higher pressure: shifts to fewer gas moles
- Higher temperature: favours the endothermic direction (and changes K)
- Catalyst: no shift, equilibrium reached sooner
- Inert gas at constant volume: no effect

WATER AND PH

- $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ at 298 K
- $\text{pH} = -\log[\text{H}_3\text{O}^+]$
- $\text{pH} + \text{pOH} = \text{p}K_w = 14$ (298 K)
- Neutral pH 7; acidic < 7; basic > 7 (298 K)

SALT HYDROLYSIS AND BUFFERS

- Salt of weak acid + weak base:

$$\text{pH} = 7 + \frac{1}{2}(\text{p}K_a - \text{p}K_b)$$
- Strong acid + weak base salt: acidic; weak acid + strong base salt: basic
- Henderson-Hasselbalch (acidic buffer):

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$
- Basic buffer: $\text{pOH} = \text{p}K_b + \log \frac{[\text{BH}^+]}{[\text{B}]}$

ACIDS AND BASES

- Arrhenius: gives H^+ / OH^- in water
- Brønsted-Lowry: proton donor / proton acceptor; conjugate pairs differ by one H^+
- Lewis: electron-pair acceptor / electron-pair donor

WEAK ACIDS AND BASES

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

(α = degree of ionization); same form for K_b
- $\text{p}K_a = -\log K_a$; $\text{p}K_b = -\log K_b$
- Conjugate pair: $K_a \times K_b = K_w$; $\text{p}K_a + \text{p}K_b = 14$
- Adding a common ion suppresses ionization of a weak electrolyte.

SOLUBILITY PRODUCT

- $\text{M}_x\text{X}_y \rightleftharpoons x\text{M}^{p+} + y\text{X}^{q-}$: $K_{sp} = [\text{M}^{p+}]^x[\text{X}^{q-}]^y$
- With molar solubility S :

$$K_{sp} = x^x y^y S^{x+y}$$
- Precipitation when ionic product $Q_{sp} > K_{sp}$
- A common ion lowers solubility.

7 Redox Reactions**DEFINITIONS**

- Oxidation: loss of electrons, increase in oxidation number
- Reduction: gain of electrons, decrease in oxidation number
- Oxidant is itself reduced; reductant is itself oxidised
- e.g. $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \longrightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$

TYPES OF REDOX REACTIONS

- Combination: $\text{C(s)} + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g})$
- Decomposition: $2\text{H}_2\text{O(l)} \longrightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
- Displacement (metal or non-metal):

$$\text{Zn} + \text{CuSO}_4 \longrightarrow \text{ZnSO}_4 + \text{Cu}$$
- Disproportionation: $2\text{H}_2\text{O}_2(\text{aq}) \longrightarrow 2\text{H}_2\text{O(l)} + \text{O}_2(\text{g})$ (O: $-1 \rightarrow -2$ and 0)

OXIDATION NUMBER RULES

- Free element: 0; monatomic ion: its charge
- O: -2 (peroxides -1 , superoxides $-\frac{1}{2}$, OF_2 $+2$)
- H: $+1$ (metal hydrides -1); F: always -1
- Sum = 0 for a molecule, = charge for a polyatomic ion

BALANCING REDOX EQUATIONS

- Oxidation number method: equalise total increase and decrease in O.N.
- Half-reaction (ion-electron) method: balance each half for atoms, then O with H_2O , H with H^+ (acid) or OH^- (base), then charge with e^- ; add halves

REDOX REACTIONS AND ELECTRODE PROCESSES

- Daniell cell: Zn anode (oxidation), Cu cathode (reduction), salt bridge completes the circuit
- Standard electrode potential E° measured against the standard hydrogen electrode ($E^\circ = 0.00 \text{ V}$)
- More positive E° : stronger oxidising agent (F_2 strongest); more negative E° : stronger reducing agent (Li strongest)



8 Organic Chemistry: Basic Principles and Techniques

BOND FISSION AND INTERMEDIATES

- Homolytic fission: free radicals
- Heterolytic fission: carbocation (sp^2 , planar) and carbanion (sp^3 , pyramidal)
- Carbocation stability:

$$CH_3^+ < CH_3CH_2^+ < (CH_3)_2CH^+ < (CH_3)_3C^+$$
- Electrophile: electron-pair acceptor (H^+ , NO_2^+);
 nucleophile: electron-pair donor (OH^- , CN^- , NH_3)

ELECTRON DISPLACEMENT EFFECTS

- Inductive effect ($-I$: NO_2 , CN , halogens; $+I$: alkyl groups); permanent, along σ bonds
- Resonance effect ($+R$: halogens, OH , NH_2 ; $-R$: $COOH$, CHO , CN , NO_2)
- Electromeric effect: temporary, only while the reagent is present
- Hyperconjugation: $\sigma(C-H)$ delocalisation; more α -H gives more stability

ISOMERISM

- Structural: chain, position, functional group, metamerism
- Stereoisomerism: geometrical and optical

PURIFICATION

- Sublimation, crystallisation, distillation (fractional, under reduced pressure, steam), differential extraction, chromatography
- Retardation factor:

$$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$$

LASSAIGNE'S TEST (SODIUM FUSION EXTRACT)

Element	Test	Result
Nitrogen	boil with $FeSO_4$, acidify with conc. H_2SO_4	Prussian blue $Fe_4[Fe(CN)_6]_3 \cdot x H_2O$
Sulphur	acetic acid + lead acetate; or sodium nitroprusside	black PbS ; violet colour
N and S together	$FeCl_3$ with $NaSCN$ formed	blood red
Halogens	HNO_3 then $AgNO_3$	Cl: white, soluble in NH_4OH ; Br: pale yellow, sparingly soluble; I: yellow, insoluble
Phosphorus	Na_2O_2 , HNO_3 , ammonium molybdate	yellow ammonium phosphomolybdate

QUANTITATIVE ANALYSIS (MASS OF COMPOUND = m)

- Carbon: $\%C = \frac{12 \times m_1 \times 100}{44 \times m}$ (m_1 = mass of CO_2); Hydrogen: $\%H = \frac{2 \times m_2 \times 100}{18 \times m}$ (m_2 = mass of H_2O)
- Dumas: $\%N = \frac{28 \times V \times 100}{22400 \times m}$ (V = mL of N_2 at STP)
- Kjeldahl:

$$\%N = \frac{1.4 \times M \times 2 \left(V - \frac{V_1}{2}\right)}{m}$$

(V mL of M H_2SO_4 taken, V_1 mL of M $NaOH$ for the excess acid; not for nitro, azo or ring N)

- Carius (halogen): $\%X = \frac{\text{atomic mass of } X \times m_1 \times 100}{\text{molar mass of } AgX \times m}$
- Sulphur: $\%S = \frac{32 \times m_1 \times 100}{233 \times m}$ (m_1 = mass of $BaSO_4$)
- Phosphorus: $\%P = \frac{31 \times m_1 \times 100}{1877 \times m}$ (ammonium phosphomolybdate) or $\frac{62 \times m_1 \times 100}{222 \times m}$ ($Mg_2P_2O_7$)
- Oxygen: $\%O = \frac{32 \times m_1 \times 100}{88 \times m}$ (m_1 = mass of CO_2); usually found by difference



9 Hydrocarbons

PREPARATION OF ALKANES

- Hydrogenation: $\text{CH}_2=\text{CH}_2 + \text{H}_2 \xrightarrow{\text{Pt/Pd/Ni}} \text{CH}_3-\text{CH}_3$
- Alkyl halide reduction: $\text{CH}_3\text{Cl} + \text{H}_2 \xrightarrow{\text{Zn, H}^+} \text{CH}_4 + \text{HCl}$
- Wurtz reaction: $2\text{CH}_3\text{Br} + 2\text{Na} \xrightarrow{\text{dry ether}} \text{CH}_3-\text{CH}_3 + 2\text{NaBr}$ (even-carbon alkanes)
- Decarboxylation: $\text{CH}_3\text{COONa} + \text{NaOH} \xrightarrow[\Delta]{\text{CaO}} \text{CH}_4 + \text{Na}_2\text{CO}_3$
- Kolbe's electrolysis:
 $2\text{CH}_3\text{COONa} + 2\text{H}_2\text{O} \longrightarrow \text{CH}_3-\text{CH}_3 + 2\text{CO}_2 + \text{H}_2 + 2\text{NaOH}$
 (ethane at anode; methane cannot be made)

REACTIONS OF ALKANES

- Halogenation (free radical, $h\nu$):
 $\text{CH}_4 + \text{Cl}_2 \longrightarrow \text{CH}_3\text{Cl} + \text{HCl}$ and further to CCl_4 ; rate
 $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$; H replaced $3^\circ > 2^\circ > 1^\circ$
- Combustion: $\text{C}_n\text{H}_{2n+2} + \frac{3n+1}{2}\text{O}_2 \longrightarrow n\text{CO}_2 + (n+1)\text{H}_2\text{O}$
- Controlled oxidation:
 $2\text{CH}_4 + \text{O}_2 \xrightarrow{\text{Cu/523 K/100 atm}} 2\text{CH}_3\text{OH};$
 $\text{CH}_4 + \text{O}_2 \xrightarrow{\text{Mo}_2\text{O}_3} \text{HCHO} + \text{H}_2\text{O};$
 $2\text{CH}_3\text{CH}_3 + 3\text{O}_2 \xrightarrow{(\text{CH}_3\text{COO})_2\text{Mn}} 2\text{CH}_3\text{COOH} + 2\text{H}_2\text{O}$
- KMnO_4 : $(\text{CH}_3)_3\text{CH} \longrightarrow (\text{CH}_3)_3\text{COH}$
- Isomerisation (anhyd. AlCl_3/HCl): n-hexane $\longrightarrow$ 2- and 3-methylpentane
- Aromatization (V, Mo or Cr oxides on alumina, 773 K):
 n-hexane $\longrightarrow$ benzene
- Steam: $\text{CH}_4 + \text{H}_2\text{O} \xrightarrow[1273\text{ K}]{\text{Ni}} \text{CO} + 3\text{H}_2$; pyrolysis
 (cracking) of higher alkanes

CONFORMATIONS OF ETHANE

- Staggered (most stable) and eclipsed (least stable); skew in between
- Shown by Sawhorse and Newman projections
- Torsional strain: rotation barrier about 12.5 kJ mol^{-1}

PREPARATION OF ALKENES

- Alkyne + H_2 , Lindlar's catalyst (Pd/C poisoned): cis-alkene; Na in liquid NH_3 : trans-alkene
- Dehydrohalogenation (alc. KOH , β -elimination): rate $\text{I} > \text{Br} > \text{Cl}$ and $3^\circ > 2^\circ > 1^\circ$
- Dehalogenation:
 $\text{CH}_2\text{Br}-\text{CH}_2\text{Br} + \text{Zn} \longrightarrow \text{CH}_2=\text{CH}_2 + \text{ZnBr}_2$
- Acidic dehydration:
 $\text{CH}_3\text{CH}_2\text{OH} \xrightarrow[\Delta]{\text{conc. H}_2\text{SO}_4} \text{CH}_2=\text{CH}_2 + \text{H}_2\text{O}$

REACTIONS OF ALKENES

- Br_2/CCl_4 : vicinal dibromide; orange colour discharged (test for unsaturation)
- HX ($\text{HI} > \text{HBr} > \text{HCl}$): Markovnikov's rule,
 $\text{CH}_3\text{CH}=\text{CH}_2 + \text{HBr} \longrightarrow \text{CH}_3\text{CHBrCH}_3$
- Peroxide (Kharash) effect, HBr only:
 $\text{CH}_3\text{CH}=\text{CH}_2 + \text{HBr} \xrightarrow{(\text{C}_6\text{H}_5\text{CO})_2\text{O}_2} \text{CH}_3\text{CH}_2\text{CH}_2\text{Br}$
- Cold conc. H_2SO_4 : alkyl hydrogen sulphate; $\text{H}_2\text{O}/\text{H}^+$: alcohol (both Markovnikov)
- Baeyer's reagent (cold dil. KMnO_4): vicinal glycol; KMnO_4/H^+ :
 $\text{CH}_3\text{CH}=\text{CHCH}_3 \longrightarrow 2\text{CH}_3\text{COOH}$
- Ozonolysis: O_3 , then $\text{Zn}/\text{H}_2\text{O}$, gives aldehydes/ketones (locates the double bond)
- Polymerisation: $n\text{CH}_2=\text{CH}_2 \longrightarrow (-\text{CH}_2-\text{CH}_2-)_n$ (polythene)

ALKYNES

- From CaC_2 : $\text{CaC}_2 + 2\text{H}_2\text{O} \longrightarrow \text{Ca}(\text{OH})_2 + \text{C}_2\text{H}_2$
- From vicinal dihalide: alc. KOH , then NaNH_2
- Acidic H: $\text{HC}\equiv\text{CH} + \text{Na} \longrightarrow \text{HC}\equiv\text{C}^-\text{Na}^+ + \frac{1}{2}\text{H}_2$;
 acidity $\text{HC}\equiv\text{CH} > \text{H}_2\text{C}=\text{CH}_2 > \text{CH}_3-\text{CH}_3$
- 2 HBr : gem-dihalide,
 $\text{HC}\equiv\text{CH} \longrightarrow \text{CH}_2=\text{CHBr} \longrightarrow \text{CH}_3\text{CHBr}_2$
- Water ($\text{HgSO}_4/\text{dil. H}_2\text{SO}_4$, 333 K): carbonyl compound
- Cyclic polymerisation: $3\text{HC}\equiv\text{CH} \xrightarrow[873\text{ K}]{\text{red-hot Fe tube}} \text{C}_6\text{H}_6$



BENZENE AND AROMATICITY

- Hückel rule: planar, fully delocalised, $(4n + 2) \pi$ electrons
- All C-C bonds 139 pm (between 154 and 133 pm)
- Preparation: cyclic polymerisation of ethyne; sodium benzoate + soda lime; phenol over heated Zn dust

ADDITION AND DIRECTIVE INFLUENCE

- $\text{C}_6\text{H}_6 + 3 \text{H}_2 \xrightarrow{\text{Ni}} \text{C}_6\text{H}_{12}$ (cyclohexane);
 $\text{C}_6\text{H}_6 + 3 \text{Cl}_2 \xrightarrow{\text{UV}} \text{C}_6\text{H}_6\text{Cl}_6$ (BHC, gammaxane)
- Combustion: $\text{C}_x\text{H}_y + \left(x + \frac{y}{4}\right) \text{O}_2 \longrightarrow x \text{CO}_2 + \frac{y}{2} \text{H}_2\text{O}$
- o/p-directing, activating: $-\text{OH}$, $-\text{NH}_2$, $-\text{OCH}_3$, $-\text{CH}_3$
- o/p-directing, deactivating: halogens
- m-directing, deactivating: $-\text{NO}_2$, $-\text{CN}$, $-\text{CHO}$, $-\text{COOH}$, $-\text{SO}_3\text{H}$

ELECTROPHILIC SUBSTITUTION OF BENZENE

- Nitration (conc. HNO_3 + conc. H_2SO_4 , electrophile NO_2^+): nitrobenzene
- Halogenation (anhyd. $\text{FeCl}_3/\text{AlCl}_3$): chlorobenzene
- Sulphonation (oleum): benzenesulphonic acid
- Friedel-Crafts alkylation:
 $\text{C}_6\text{H}_6 + \text{CH}_3\text{Cl} \xrightarrow{\text{anhyd. AlCl}_3} \text{C}_6\text{H}_5\text{CH}_3 + \text{HCl}$
- Friedel-Crafts acylation:
 $\text{C}_6\text{H}_6 + \text{CH}_3\text{COCl} \xrightarrow{\text{anhyd. AlCl}_3} \text{C}_6\text{H}_5\text{COCH}_3 + \text{HCl}$

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