

CHSE Odisha · +2 Science

CHEMISTRY

+2 1st Year (Class XI) · As per CHSE Syllabus 2023-24

Theory in simple language · Formula boxes
Solved problems · Exam Q&A · Practice & revision

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CHEMISTRY

+2 1st Year (Class XI) · Theory 70 + Practical 30

Exam Pattern & How to Study

Chemistry - Unit Weightage

Class XI Unit	Marks	Class XII Unit	Marks
Some Basic Concepts of Chemistry	7	Solutions	7
Structure of Atom	9	Electrochemistry	9
Classification of Elements	6	Chemical Kinetics	7
Chemical Bonding	7	d- and f-Block Elements	7
Chemical Thermodynamics	9	Coordination Compounds	7
Equilibrium	7	Haloalkanes and Haloarenes	6
Redox Reactions	4	Alcohols, Phenols and Ethers	6
Organic Chemistry - Basic Principles	11	Aldehydes, Ketones, Carboxylic Acids	8
Hydrocarbons	10	Amines	6
		Biomolecules	7

GOLDEN RULES FOR WRITING ANSWERS

- Always write the **formula first**, then substitute values, then the answer **with units**.
- Underline or box the final answer. Examiners mark step-by-step, so even a wrong final answer earns step marks.
- For theory answers, use short points and draw a diagram wherever possible.
- Attempt questions you know best first, but keep their numbering exactly as in the paper.

XI-1. Some Basic Concepts of Chemistry

Importance and scope: Chemistry is the study of matter — what things are made of, and how they change. It gives us medicines (paracetamol, antibiotics), fertilisers (urea) for food, fuels, plastics, soaps, cement and clean water. It also helps solve problems such as pollution and ozone depletion.

Nature of matter

State	Shape	Volume	Particles
Solid	Fixed	Fixed	Closely packed, only vibrate
Liquid	Takes shape of container	Fixed	Close but can move
Gas	No fixed shape	No fixed volume	Far apart, move freely

Matter is classified as **pure substances** (elements like Cu, O₂; compounds like H₂O) and **mixtures** (homogeneous like salt solution; heterogeneous like sand + water).

Laws of chemical combination

Law	Given by	Statement (simple)	Example
Conservation of mass	Lavoisier	Mass is neither created nor destroyed in a reaction.	Mass of reactants = mass of products
Definite proportions	Proust	A compound always contains the same elements in the same ratio by mass.	Water always has H : O = 1 : 8
Multiple proportions	Dalton	When two elements form more than one compound, the masses of one that combine with a fixed mass of the other are in a simple whole-number ratio.	H ₂ O and H ₂ O ₂ : O ratio 16 : 32 = 1 : 2
Gay-Lussac's gaseous volumes	Gay-Lussac	Gases combine in simple volume ratios (same T, P).	2 vol H ₂ + 1 vol O ₂ → 2 vol steam
Avogadro's law	Avogadro	Equal volumes of gases at the same T and P contain equal numbers of molecules.	—

Dalton's atomic theory

1. Matter is made of tiny indivisible particles called atoms.
2. Atoms of the same element are identical in mass and properties; atoms of different elements differ.
3. Compounds form when atoms of different elements combine in fixed whole-number ratios.
4. Chemical reactions only rearrange atoms; atoms are neither created nor destroyed.

Limitations: Atoms are actually divisible (electrons, protons, neutrons), and isotopes of the same element have different masses.

Atomic mass, molecular mass, mole

DEFINITIONS

- **Atomic mass unit (u):** exactly $1/12$ of the mass of one C-12 atom = 1.66×10^{-24} g.
- **Average atomic mass:** weighted average of isotopes, e.g. Cl = 35.5 u.
- **Molecular mass:** sum of atomic masses in a molecule. $\text{H}_2\text{O} = 2(1) + 16 = 18$ u.
- **Mole:** amount of substance containing 6.022×10^{23} particles (Avogadro number N_A).
- **Molar mass:** mass of 1 mole in grams = molecular mass in g. (H_2O : 18 g mol^{-1})
- **Equivalent mass:** mass of a substance that combines with or displaces 1.008 g of hydrogen or 8 g of oxygen or 35.5 g of chlorine.

MOLE FORMULAS

Number of moles $n = \text{given mass} / \text{molar mass} = \text{number of particles} / N_A = \text{volume of gas at STP (L)} / 22.4$

Equivalent mass of element = Atomic mass / Valency

Eq. mass of acid = Molar mass / Basicity Eq. mass of base = Molar mass / Acidity

Eq. mass of salt = Molar mass / total positive charge Eq. mass of oxidant/reductant = Molar mass / change in oxidation number per formula unit

Mass percent of element = (mass of element in 1 mol / molar mass) $\times 100$

Concentration of solutions

CONCENTRATION TERMS

Mass % = (mass of solute / mass of solution) $\times 100$

Mole fraction $x_A = n_A / (n_A + n_B)$ ($x_A + x_B = 1$)

Molarity $M = \text{moles of solute} / \text{volume of solution in litres}$

Molality $m = \text{moles of solute} / \text{mass of solvent in kg}$

Normality $N = \text{gram equivalents of solute} / \text{volume of solution in litres}$; $N = M \times \text{n-factor}$

Dilution: $M_1V_1 = M_2V_2$ Titration: $N_1V_1 = N_2V_2$

MOLARITY VS MOLALITY

Molarity changes with temperature (volume expands). Molality does not (mass doesn't change). So molality is preferred for precise work.

Empirical and molecular formula

Empirical formula = simplest whole-number ratio of atoms (CH_2O for glucose).

Molecular formula = actual number of atoms ($\text{C}_6\text{H}_{12}\text{O}_6$). Molecular formula = $n \times$ empirical formula, where $n = \text{molar mass} / \text{empirical formula mass}$.

STEPS TO FIND EMPIRICAL FORMULA

(1) Write % of each element (2) Divide by atomic mass $\rightarrow$ moles (3) Divide all by the smallest $\rightarrow$ ratio (4) If needed, multiply to get whole numbers.

Stoichiometry

A balanced equation tells us the mole ratio. In $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$: 2 mol H_2 react with 1 mol O_2 to give 2 mol H_2O . The reactant that is used up first is the **limiting reagent**; it decides how much product forms.

Example 1. Calculate the molecular mass of glucose $\text{C}_6\text{H}_{12}\text{O}_6$.

$$6(12) + 12(1) + 6(16) = 72 + 12 + 96 = \mathbf{180 \text{ u.}}$$

Example 2. How many moles and molecules are present in 36 g of water?

$$n = 36/18 = 2 \text{ mol. Molecules} = 2 \times 6.022 \times 10^{23} = \mathbf{1.204 \times 10^{24}}.$$

Example 3. Calculate the mass percent of each element in sodium sulphate Na_2SO_4 .

$$\text{Molar mass} = 46 + 32 + 64 = 142 \text{ g. Na} = 46/142 \times 100 = 32.39\%; \text{S} = 32/142 \times 100 = 22.54\%; \text{O} = 64/142 \times 100 = 45.07\%.$$

Example 4. A compound contains 4.07% H, 24.27% C and 71.65% Cl. Its molar mass is 98.96 g. Find its empirical and molecular formula.

$$\text{Moles: H} = 4.07/1 = 4.07; \text{C} = 24.27/12 = 2.02; \text{Cl} = 71.65/35.5 = 2.02. \text{ Divide by } 2.02: \text{H} = 2, \text{C} = 1, \text{Cl} = 1. \text{ Empirical formula } \text{CH}_2\text{Cl} \text{ (mass} = 49.5\text{). } n = 98.96/49.5 \approx 2. \text{ Molecular formula } \mathbf{\text{C}_2\text{H}_4\text{Cl}_2}.$$

Example 5. Calculate the molarity of NaOH in a solution made by dissolving 4 g of it in 250 mL of solution.

$$\text{Moles} = 4/40 = 0.1. \text{ V} = 0.25 \text{ L. M} = 0.1/0.25 = \mathbf{0.4 \text{ M.}}$$

Example 6. Calculate the amount of water produced by burning 16 g of methane.



Example 7. 50 kg of N_2 and 10 kg of H_2 are mixed to make NH_3 . Find the limiting reagent and the NH_3 formed.

$$\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3. \text{ Moles } \text{N}_2 = 50000/28 = 1786; \text{H}_2 = 10000/2 = 5000. 1786 \text{ mol } \text{N}_2 \text{ needs } 5358 \text{ mol } \text{H}_2 \text{ but only } 5000 \text{ is present} \Rightarrow \mathbf{\text{H}_2 \text{ is limiting.}} \text{NH}_3 = (2/3) \times 5000 = 3333 \text{ mol} = 3333 \times 17 \text{ g} \approx \mathbf{56.7 \text{ kg.}}$$

Example 8. Find the equivalent mass of H_2SO_4 and of Ca(OH)_2 .

$$\text{H}_2\text{SO}_4: 98/2 = \mathbf{49}. \text{Ca(OH)}_2: 74/2 = \mathbf{37}.$$

Example 9. What volume does 4.4 g of CO_2 occupy at STP?

$$n = 4.4/44 = 0.1 \text{ mol. V} = 0.1 \times 22.4 = \mathbf{2.24 \text{ L.}}$$

XI-2. Structure of Atom

Discovery of sub-atomic particles

Particle	Discovered by	Experiment	Charge	Mass
Electron (e)	J.J. Thomson (1897)	Cathode ray tube	$-1.6 \times 10^{-19} \text{ C}$	$9.1 \times 10^{-31} \text{ kg}$
Proton (p)	Goldstein (anode rays), Rutherford	Canal rays	$+1.6 \times 10^{-19} \text{ C}$	$1.672 \times 10^{-27} \text{ kg}$
Neutron (n)	Chadwick (1932)	Beryllium bombarded by α -particles	0	$1.675 \times 10^{-27} \text{ kg}$

TERMS

- **Atomic number Z** = number of protons. **Mass number A** = protons + neutrons. Neutrons = $A - Z$.
- **Isotopes**: same Z, different A (^1H , ^2H , ^3H). **Isobars**: same A, different Z (^{40}Ar , ^{40}Ca).
- **Isotones**: same number of neutrons (^{14}C and ^{16}O both have 8 n).

Atomic models

Model	Main idea	Limitation
Thomson (plum pudding)	Atom is a sphere of positive charge with electrons embedded like seeds in a watermelon.	Could not explain Rutherford's scattering result.
Rutherford (nuclear)	α -particles fired at gold foil: most passed through, a few bounced back. So the atom is mostly empty, with a tiny dense positive nucleus ; electrons revolve around it.	A revolving electron should lose energy and spiral into the nucleus — atom would be unstable. Also could not explain line spectra.
Bohr	Electrons move in fixed circular orbits (energy levels). They gain or lose energy only when jumping between orbits. Angular momentum $mvr = nh/2\pi$.	Works only for one-electron species (H , He^+); ignores the wave nature of electrons and the uncertainty principle; cannot explain splitting of lines in a magnetic field (Zeeman effect).

BOHR MODEL AND LIGHT FORMULAS

Energy of n-th orbit (H): $E_n = -2.18 \times 10^{-18} (Z^2/n^2) \text{ J} = -13.6 Z^2/n^2 \text{ eV}$

Radius: $r_n = 0.529 n^2/Z \text{ \AA}$

Energy of photon: $E = h\nu = hc/\lambda$ $c = \nu\lambda$ $h = 6.626 \times 10^{-34} \text{ J s}$

Rydberg: $1/\lambda = R_H(1/n_1^2 - 1/n_2^2)$, $R_H = 109677 \text{ cm}^{-1}$

de Broglie: $\lambda = h/mv = h/p$

Heisenberg: $\Delta x \cdot \Delta p \geq h/4\pi$

Spectral series of hydrogen: Lyman ($n_1 = 1$, UV), Balmer ($n_1 = 2$, visible), Paschen (3), Brackett (4), Pfund (5) — the last three in infrared.

Dual nature and uncertainty

Dual nature: Light behaves like a wave (interference) and like particles called photons (photoelectric effect). de Broglie said *matter* also has a wave nature, $\lambda = h/mv$. This matters only for tiny particles such as electrons.

Heisenberg's uncertainty principle: It is impossible to know both the exact position and the exact momentum of an electron at the same time. Therefore, fixed orbits (Bohr) are not possible. Instead we talk of **orbitals** — regions where the probability of finding the electron is highest.

Quantum numbers

Quantum number	Symbol	Tells us	Values
Principal	n	Shell, size and energy	1, 2, 3, ... (K, L, M, ...)
Azimuthal	l	Subshell and shape	0 to ($n - 1$): 0 = s, 1 = p, 2 = d, 3 = f
Magnetic	m_l	Orientation of orbital	-1 ... 0 ... +1 ($2l + 1$ values)
Spin	m_s	Spin of electron	$+\frac{1}{2}$ or $-\frac{1}{2}$

Shapes: s - spherical; p - dumb-bell (p_x , p_y , p_z); d - double dumb-bell / clover-leaf (five d orbitals; d_{z^2} has a ring around the middle).

Maximum electrons: subshell s = 2, p = 6, d = 10, f = 14. Shell n holds $2n^2$ electrons.

Rules for filling electrons

- **Aufbau principle:** fill orbitals in order of increasing energy, decided by the ($n + l$) rule. If ($n + l$) is the same, lower n fills first. Order: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p...
- **Pauli's exclusion principle:** no two electrons in an atom can have all four quantum numbers the same — an orbital holds at most 2 electrons with opposite spins.
- **Hund's rule:** in orbitals of equal energy (e.g. three p orbitals), electrons first occupy them singly with parallel spins; pairing starts only after each has one.

EXCEPTIONS - EXTRA STABILITY

Half-filled and completely filled subshells are extra stable (symmetry and exchange energy). So Cr = [Ar] $3d^5 4s^1$ (not $3d^4 4s^2$) and Cu = [Ar] $3d^{10} 4s^1$ (not $3d^9 4s^2$).

Element	Z	Configuration
N	7	$1s^2 2s^2 2p^3$
Na	11	$1s^2 2s^2 2p^6 3s^1$
Cl	17	$[\text{Ne}] 3s^2 3p^5$
Fe	26	$[\text{Ar}] 3d^6 4s^2$
Cr	24	$[\text{Ar}] 3d^5 4s^1$
Cu	29	$[\text{Ar}] 3d^{10} 4s^1$

Example 1. Calculate the energy of a photon of wavelength 500 nm.

$$E = hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^8)/(500 \times 10^{-9}) = \mathbf{3.98 \times 10^{-19} \text{ J}}.$$

Example 2. Find the wavelength of the electron moving with velocity 2.05×10^7 m/s.

$$\lambda = h/mv = 6.626 \times 10^{-34}/(9.1 \times 10^{-31} \times 2.05 \times 10^7) = \mathbf{3.55 \times 10^{-11} \text{ m}}.$$

Example 3. What is the energy required to remove the electron from the first orbit of hydrogen (ionisation energy)?

$$\Delta E = 0 - (-2.18 \times 10^{-18}) = \mathbf{2.18 \times 10^{-18} \text{ J per atom}} \text{ (13.6 eV)}.$$

Example 4. Calculate the wavelength of the line in the Balmer series when the electron falls from $n = 4$.

$$1/\lambda = 109677(1/4 - 1/16) = 109677 \times 3/16 = 20564 \text{ cm}^{-1} \Rightarrow \lambda = 4.86 \times 10^{-5} \text{ cm} = \mathbf{486 \text{ nm}} \text{ (blue-green)}.$$

Example 5. Write the four quantum numbers of the outermost electron of sodium.

$$3s^1: n = 3, l = 0, m_l = 0, m_s = +\frac{1}{2}.$$

Example 6. How many neutrons, protons and electrons are in $^{80}_{35}\text{Br}$?

$$p = 35, e = 35, n = 80 - 35 = \mathbf{45}.$$

Exam Q&A - Basic Concepts & Atomic Structure

Q1. What is the SI unit of amount of substance? [1]

The mole (mol).

Q2. Which has more atoms: 1 g of Na or 1 g of Mg? [1]

1 g Na ($1/23$ mol) has more atoms than 1 g Mg ($1/24$ mol).

Q3. State the $(n + l)$ rule. [2]

Orbitals are filled in increasing order of $(n + l)$. If two orbitals have the same $(n + l)$, the one with the lower n is filled first. E.g. $4s$ ($4 + 0 = 4$) fills before $3d$ ($3 + 2 = 5$).

Q4. Why is the electronic configuration of Cr $3d^5 4s^1$? [2]

A half-filled d subshell is more stable due to symmetrical distribution and greater exchange energy, so one 4s electron moves to 3d.

Q5. Which quantum numbers are not possible: (a) $n = 2, l = 2$ (b) $n = 3, l = 1, m = -2$ [2]

(a) Not possible: l must be less than n . (b) Not possible: for $l = 1$, m can only be $-1, 0, +1$.

Q6. Describe Rutherford's α -scattering experiment and its conclusions. [5]

A thin gold foil was bombarded with fast α -particles and a ZnS screen detected them. Observations: (i) most α -particles passed straight through; (ii) a few were deflected by small angles; (iii) very few (about 1 in 20,000) bounced back. Conclusions: (i) most of the atom is empty space; (ii) positive charge and nearly all mass are concentrated in a very small nucleus; (iii) the nucleus is about 10^{-15} m across while the atom is about 10^{-10} m. Electrons revolve round the nucleus like planets round the sun.

XI-3. Classification of Elements and Periodicity in Properties

Why classify? More than 118 elements are known. Studying each separately is hard. Arranging them in a table where similar elements fall together makes study easy and helps predict properties of new elements.

Brief history

Scientist	Contribution
Döbereiner (1829)	Triads: groups of three similar elements; middle one's atomic mass $\approx$ average of the other two (Li, Na, K).
Newlands (1865)	Law of octaves: every eighth element resembles the first, like musical notes. Worked only up to Ca.
Mendeleev (1869)	Periodic law: properties are periodic functions of atomic mass . Left gaps and predicted eka-aluminium (Ga) and eka-silicon (Ge).
Moseley (1913)	Showed atomic number is the more basic property (X-ray study).

MODERN PERIODIC LAW

The physical and chemical properties of elements are periodic functions of their **atomic numbers**.

Present form - the long form of the periodic table

- 18 vertical columns called **groups**, 7 horizontal rows called **periods**.
- Period number = number of shells; elements in a group have the same outer electronic configuration.
- **s-block** (groups 1, 2): reactive metals; ns^{1-2} .
- **p-block** (groups 13-18): metals, non-metals, metalloids; ns^2np^{1-6} . Group 18 = noble gases.
- **d-block** (groups 3-12): transition metals; $(n-1)d^{1-10}ns^{0-2}$.
- **f-block**: lanthanoids and actinoids (inner transition), placed separately at the bottom.

Naming elements with $Z > 100$ (IUPAC)

Digit	0	1	2	3	4	5	6	7	8	9
Root	nil	un	bi	tri	quad	pent	hex	sept	oct	enn

Join the roots and add "-ium". Example: $Z = 104 \rightarrow$ un-nil-quad-ium = **Unnilquadium (Unq)**. $Z = 120 \rightarrow$ Unbinilium (Ubn).

Periodic trends

Property	Meaning	Across a period (→)	Down a group (↓)
Atomic radius	Size of atom	Decreases (more nuclear charge pulls electrons in)	Increases (new shells added)
Ionisation enthalpy	Energy to remove an electron from a gaseous atom	Generally increases	Decreases
Electron gain enthalpy	Energy change when a gaseous atom gains an electron	Becomes more negative	Becomes less negative
Electronegativity	Tendency of a bonded atom to attract shared electrons	Increases	Decreases
Metallic character	Tendency to lose electrons	Decreases	Increases

IMPORTANT POINTS

- **Ionic radius:** cation is smaller than its atom ($\text{Na}^+ < \text{Na}$); anion is bigger than its atom ($\text{Cl}^- > \text{Cl}$).
- **Isoelectronic species** (same number of electrons): size decreases as nuclear charge increases. $\text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+}$.
- **Inert gas radius** is a van der Waals radius, so it is larger than the covalent radius of the neighbouring halogen.
- **Exceptions in ionisation enthalpy:** $\text{Be} > \text{B}$ and $\text{N} > \text{O}$, because Be ($2s^2$) is fully filled and N ($2p^3$) is half-filled — extra stable.
- Highest electronegativity: F (4.0, Pauling scale). Most negative electron gain enthalpy: **Cl** (not F, because F is so small that the incoming electron is repelled).
- **Valency** of s- and p-block = number of valence electrons, or 8 minus valence electrons.

Example 1. Arrange in increasing order of size: Mg, Mg^{2+} , Al, Al^{3+} .

$\text{Al}^{3+} < \text{Mg}^{2+} < \text{Al} < \text{Mg}$.

Example 2. Predict the position of the element with $Z = 17$ in the periodic table.

$1s^2 2s^2 2p^6 3s^2 3p^5$: period 3 ($n = 3$), p-block, group $10 + 7 = 17$ (halogens). It is chlorine.

Example 3. Why is the first ionisation enthalpy of Na lower than that of Mg, but its second ionisation enthalpy is higher?

Na ($3s^1$) easily loses one electron to get the stable Ne configuration. After that, removing an electron from Na^+ means breaking a stable noble gas core, so its second ionisation enthalpy is very high. For Mg, the second electron still comes from 3s.

XI-4. Chemical Bonding and Molecular Structure

Idea: Atoms join together (bond) to become more stable — usually by getting 8 electrons in their outer shell (octet rule, Kössel-Lewis). They do this by giving/taking electrons (ionic bond) or sharing them (covalent bond).

Ionic bond

Formed by complete **transfer** of electrons from a metal to a non-metal. Example: $\text{Na} \rightarrow \text{Na}^+ + e^-$; $\text{Cl} + e^- \rightarrow \text{Cl}^-$; Na^+Cl^- . Favoured by low ionisation enthalpy of the metal, high (negative) electron gain enthalpy of the non-metal, and high lattice enthalpy.

Properties: hard crystalline solids, high melting point, conduct electricity in molten state or solution, soluble in water.

Covalent bond

Formed by **sharing** of electron pairs. Single (H-H), double (O=O), triple ($\text{N}\equiv\text{N}$). A **coordinate bond** is a covalent bond in which both shared electrons come from one atom ($\text{NH}_3 \rightarrow \text{H}^+$ gives NH_4^+).

Lewis structure: shows valence electrons as dots. Formal charge = (valence electrons) – (non-bonding electrons) – $\frac{1}{2}$ (bonding electrons).

LIMITATIONS OF THE OCTET RULE

Incomplete octet (BeCl_2 , BF_3), odd-electron molecules (NO , NO_2), expanded octet (PCl_5 , SF_6), and it does not explain shape or energy of molecules.

Bond parameters

Parameter	Meaning
Bond length	Equilibrium distance between the nuclei of two bonded atoms. Triple < double < single.
Bond angle	Angle between two bonds around the central atom (H_2O : 104.5°).
Bond enthalpy	Energy needed to break one mole of bonds. Stronger bond $\rightarrow$ larger value.
Bond order	Number of bonds between two atoms. Higher bond order $\rightarrow$ shorter, stronger bond.

Polar covalent bond and covalent character of ionic bond

If two atoms with different electronegativity share electrons, the shared pair shifts towards the more electronegative atom: $\text{H}^{\delta+}-\text{Cl}^{\delta-}$. This creates a **dipole moment** $\mu = \text{charge} \times \text{distance}$ (unit Debye). CO_2 and BF_3 have zero dipole moment because their symmetrical shapes cancel bond dipoles; H_2O and NH_3 have net dipole moments.

Fajans' rules – an ionic bond gets covalent character when: the cation is small, the cation has high charge, and the anion is large. So LiI is more covalent than NaCl ; AlCl_3 is more covalent than NaCl .

VSEPR theory (shape of molecules)

Electron pairs around the central atom repel each other and arrange themselves as far apart as possible. Repulsion order: lone pair-lone pair > lone pair-bond pair > bond pair-bond pair. Lone pairs squeeze bond angles.

Bond pairs + lone pairs	Shape	Example	Angle
2 + 0	Linear	BeCl_2 , CO_2	180°
3 + 0	Trigonal planar	BF_3	120°
4 + 0	Tetrahedral	CH_4	109.5°
3 + 1	Trigonal pyramidal	NH_3	107°
2 + 2	Bent (V-shape)	H_2O	104.5°
5 + 0	Trigonal bipyramidal	PCl_5	120° , 90°
6 + 0	Octahedral	SF_6	90°
4 + 1	See-saw	SF_4	—
3 + 2	T-shape	ClF_3	—
4 + 2	Square planar	XeF_4	90°

Valence bond theory and hybridisation

VBT: a covalent bond forms when half-filled atomic orbitals of two atoms overlap. Head-on (end-to-end) overlap gives a strong **sigma (σ) bond**; sideways overlap of p orbitals gives a weaker **pi (π) bond**. A double bond = $1\sigma + 1\pi$; a triple bond = $1\sigma + 2\pi$.

Hybridisation: mixing of atomic orbitals of slightly different energies on the same atom to form an equal number of new identical (hybrid) orbitals.

Hybridisation	Orbitals mixed	Shape	Example
sp	1s + 1p	Linear, 180°	BeCl ₂ , C ₂ H ₂
sp ²	1s + 2p	Trigonal planar, 120°	BF ₃ , C ₂ H ₄
sp ³	1s + 3p	Tetrahedral, 109.5°	CH ₄ , NH ₃ , H ₂ O
sp ³ d	1s + 3p + 1d	Trigonal bipyramidal	PCl ₅
sp ³ d ²	1s + 3p + 2d	Octahedral	SF ₆

QUICK WAY TO FIND HYBRIDISATION

$H = \frac{1}{2}[V + M - C + A]$, where V = valence electrons of central atom, M = monovalent atoms attached, C = cation charge, A = anion charge. $H = 2 \rightarrow sp$, $3 \rightarrow sp^2$, $4 \rightarrow sp^3$, $5 \rightarrow sp^3d$, $6 \rightarrow sp^3d^2$. Example NH₄⁺: $\frac{1}{2}[5 + 4 - 1] = 4 \rightarrow sp^3$.

Resonance

When one Lewis structure cannot describe a molecule, it is shown as a blend (resonance hybrid) of two or more structures. Example: O₃ has two structures where the double bond is on either side; actually both O-O bonds are equal (128 pm). Resonance makes molecules more stable. Benzene and CO₃²⁻ also show resonance.

Molecular orbital theory (qualitative)

Atomic orbitals combine to form molecular orbitals spread over the whole molecule. Adding orbitals gives a lower-energy **bonding MO**; subtracting gives a higher-energy **antibonding MO** (*).

MOT FORMULAS

Bond order = $\frac{1}{2}(N_b - N_a)$ (bonding minus antibonding electrons)

Bond order > 0 → molecule exists; = 0 → does not exist (He₂)

Unpaired electrons present → paramagnetic; all paired → diamagnetic

Order (up to N₂, 14 electrons): $\sigma 1s$ $\sigma^* 1s$ $\sigma 2s$ $\sigma^* 2s$ ($\pi 2p_x = \pi 2p_y$) $\sigma 2p_z$ ($\pi^* 2p_x = \pi^* 2p_y$) $\sigma^* 2p_z$

For O₂, F₂: $\sigma 2p_z$ comes before $\pi 2p$

Molecule	Electrons	Bond order	Magnetic nature
H ₂	2	1	Diamagnetic
He ₂	4	0 (does not exist)	—
N ₂	14	3	Diamagnetic
O ₂	16	2	Paramagnetic (2 unpaired electrons in π^*)
O ₂ ⁺	15	2.5	Paramagnetic
O ₂ ⁻	17	1.5	Paramagnetic

MOT explains why O₂ is paramagnetic — something Lewis theory cannot.

Hydrogen bond

A weak attraction between a hydrogen atom bonded to a highly electronegative atom (F, O, N) and another electronegative atom nearby. **Intermolecular** (between molecules, e.g. H_2O , HF) raises boiling point; **intramolecular** (inside one molecule, e.g. o-nitrophenol) lowers it. Water is a liquid at room temperature and ice floats because of hydrogen bonding.

Example 1. Calculate the bond order of N_2 and say whether it is para- or diamagnetic.

N_2 : $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p^4 \sigma 2p^2$. $N_b = 10$, $N_a = 4$. B.O. = $\frac{1}{2}(10 - 4) = 3$. All paired $\Rightarrow$ diamagnetic.

Example 2. Why is NH_3 pyramidal while BF_3 is planar?

N in NH_3 has 3 bond pairs + 1 lone pair (sp^3), so the shape is pyramidal. B in BF_3 has 3 bond pairs and no lone pair (sp^2), so it is trigonal planar.

Example 3. Find the number of σ and π bonds in $\text{CH}_2=\text{CH}-\text{C}\equiv\text{CH}$.

C-H bonds: 4 σ . C=C: 1 σ + 1 π . C-C: 1 σ . C $\equiv$ C: 1 σ + 2 π . Total **7 σ and 3 π** .

Example 4. Why does H_2O have a higher boiling point than H_2S ?

O is highly electronegative, so water molecules form strong intermolecular hydrogen bonds; S is not electronegative enough. Extra energy is needed to break these H-bonds.

Exam Q&A - Periodicity & Bonding

Q1. Which has the highest electronegativity? [1]

Fluorine.

Q2. What is the shape of SF_6 ? [1]

Octahedral (sp^3d^2).

Q3. Why is the dipole moment of BeF_2 zero? [2]

BeF_2 is linear; the two equal Be-F bond dipoles point in opposite directions and cancel.

Q4. Why does nitrogen have a higher first ionisation enthalpy than oxygen? [2]

N ($2p^3$) has a half-filled, extra-stable p subshell. O ($2p^4$) loses its fourth p electron easily to become half-filled.

Q5. Distinguish between σ and π bonds. [3]

σ : head-on overlap, strong, free rotation possible, can exist alone, involves s or p orbitals.
 π : sideways overlap of p orbitals, weaker, restricts rotation, forms only in addition to a σ bond.

Q6. Explain sp^3 hybridisation with the example of methane. [5]

Ground state C: $2s^2 2p^2$. In the excited state one 2s electron is promoted: $2s^1 2p_x^1 2p_y^1 2p_z^1$. One s and three p orbitals mix to give four identical sp^3 orbitals pointing to the corners of a tetrahedron (109.5°). Each overlaps with the 1s orbital of an H atom to form four C-H σ bonds. Hence CH_4 is tetrahedral with all bonds equal.

XI-5. Chemical Thermodynamics

Idea: Thermodynamics studies energy changes (heat and work) in physical and chemical processes. It answers: "Will this reaction happen on its own?" and "How much heat will it give?"

BASIC TERMS

- **System:** the part we study (e.g., reaction in a beaker). **Surroundings:** everything else. System + surroundings = universe.
- **Open system:** exchanges both energy and matter (open beaker). **Closed:** energy only (closed flask). **Isolated:** neither (ideal thermos flask).
- **State functions:** depend only on initial and final state, not on path — U, H, S, G, P, V, T. **Path functions:** heat (q) and work (w).
- **Extensive properties** depend on amount (mass, volume, U, H). **Intensive** do not (temperature, pressure, density, molar heat capacity).
- **Processes:** isothermal (T constant), adiabatic (no heat exchange, $q = 0$), isobaric (P constant), isochoric (V constant), reversible, irreversible.

First law of thermodynamics

Energy can neither be created nor destroyed; it can only change form. Mathematically, $\Delta U = q + w$.

SIGN CONVENTION (IUPAC) AND FORMULAS

q positive: heat absorbed by system; q negative: heat released

w positive: work done *on* the system; w negative: work done *by* the system

Work of expansion against constant pressure: $w = -P_{\text{ext}}\Delta V$

Isothermal reversible expansion of ideal gas: $w = -2.303 nRT \log(V_2/V_1)$

Enthalpy: $H = U + PV$; $\Delta H = \Delta U + P\Delta V = \Delta U + \Delta n_g RT$

Heat capacity: $q = C \Delta T$; specific heat: $q = m c \Delta T$; $C_p - C_v = R$ (per mole, ideal gas)

At constant volume $q_v = \Delta U$; at constant pressure $q_p = \Delta H$

$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$

Measurement: ΔU is measured with a **bomb calorimeter** (constant volume, sealed steel vessel). ΔH is measured with a **coffee-cup calorimeter** (constant pressure).

Exothermic: heat released, ΔH negative (burning). **Endothermic:** heat absorbed, ΔH positive (melting ice).

Types of enthalpy changes

Enthalpy of...	Meaning (per mole)
Formation ($\Delta_f H^\circ$)	Forming 1 mol of compound from its elements in standard states. $\Delta_f H^\circ$ of elements = 0.
Combustion ($\Delta_c H^\circ$)	Complete burning of 1 mol of substance in oxygen (always negative).
Bond dissociation	Breaking 1 mol of a particular bond in the gas phase.
Atomisation	Breaking 1 mol of substance completely into gaseous atoms.
Sublimation	Solid directly to vapour.
Phase transition	Fusion (solid→liquid) and vaporisation (liquid→gas).
Ionisation	Removing electrons from gaseous atoms.
Solution	Dissolving 1 mol of substance in a large amount of solvent.
Dilution	Diluting a solution from one concentration to another.

CALCULATING ΔH

$$\Delta_r H^\circ = \sum \Delta_f H^\circ(\text{products}) - \sum \Delta_f H^\circ(\text{reactants})$$

$$\Delta_r H^\circ = \sum \text{bond enthalpies (reactants)} - \sum \text{bond enthalpies (products)}$$

Hess's law: total enthalpy change is the same whether a reaction occurs in one step or many steps. So equations can be added like algebra.

Second and third laws, entropy and Gibbs energy

Entropy (S) measures disorder or randomness. Gas > liquid > solid. Entropy is a state function. $\Delta S = q_{\text{rev}}/T$.

Second law: The entropy of the universe always increases in a spontaneous process ($\Delta S_{\text{total}} > 0$).

Third law: The entropy of a perfectly crystalline substance is zero at absolute zero (0 K).

GIBBS ENERGY — THE DECIDER OF SPONTANEITY

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

$$\Delta G < 0 \rightarrow \text{spontaneous}; \Delta G > 0 \rightarrow \text{non-spontaneous}; \Delta G = 0 \rightarrow \text{at equilibrium}$$

$$\Delta G^\circ = -2.303 RT \log K \quad \Delta G^\circ = -nFE^\circ_{\text{cell}}$$

ΔH	ΔS	Spontaneous?
–	+	Always
+	–	Never
–	–	At low temperature
+	+	At high temperature

Example 1. A gas expands from 2 L to 6 L against a constant pressure of 1 atm. Find the work done (1 L atm = 101.3 J).

$$w = -P\Delta V = -1 \times 4 = -4 \text{ L atm} = -4 \times 101.3 = \mathbf{-405.2 \text{ J}}$$
 (work done by the gas).

Example 2. For $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, $\Delta U = -87.4 \text{ kJ}$ at 298 K. Find ΔH .

$$\Delta n_g = 2 - 4 = -2. \Delta H = \Delta U + \Delta n_g RT = -87.4 + (-2)(8.314 \times 10^{-3})(298) = -87.4 - 4.96 = \mathbf{-92.36 \text{ kJ}}.$$

Example 3. Find the heat needed to raise the temperature of 60 g of aluminium from 35 °C to 55 °C (molar heat capacity $24 \text{ J mol}^{-1} \text{ K}^{-1}$).

$$n = 60/27 = 2.22 \text{ mol. } q = nC\Delta T = 2.22 \times 24 \times 20 = \mathbf{1066.7 \text{ J} \approx 1.07 \text{ kJ}}.$$

Example 4. Use Hess's law: Find $\Delta_f H$ of CO given $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$, $\Delta H = -393.5 \text{ kJ}$ and $\text{CO} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}_2$, $\Delta H = -283.0 \text{ kJ}$.

$$\text{Target: } \text{C} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO. Subtract equation 2 from equation 1: } \Delta H = -393.5 - (-283.0) = \mathbf{-110.5 \text{ kJ mol}^{-1}}.$$

Example 5. For a reaction, $\Delta H = 400 \text{ kJ}$ and $\Delta S = 0.2 \text{ kJ K}^{-1}$. Above what temperature will it be spontaneous?

$$\text{At equilibrium } \Delta G = 0 \Rightarrow T = \Delta H / \Delta S = 400 / 0.2 = 2000 \text{ K. Spontaneous } \mathbf{\text{above } 2000 \text{ K}}.$$

Example 6. Find ΔG° for a reaction with $K = 10$ at 300 K.

$$\Delta G^\circ = -2.303 \times 8.314 \times 300 \times \log 10 = \mathbf{-5744 \text{ J} \approx -5.74 \text{ kJ}}.$$

XI-6. Equilibrium

Idea: In a closed bottle of soda, CO_2 keeps dissolving and escaping at the same rate, so nothing seems to change. That is equilibrium — not a stop, but **two opposite processes at equal speed**. That is why it is called **dynamic** equilibrium.

Physical equilibrium examples: ice $\rightleftharpoons$ water at 0 °C; water $\rightleftharpoons$ vapour in a closed vessel; solid sugar $\rightleftharpoons$ dissolved sugar in a saturated solution. **Chemical equilibrium:** $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$.

Law of mass action and equilibrium constant

Rate of a reaction is proportional to the product of the active masses (molar concentrations) of reactants, each raised to its coefficient.

EQUILIBRIUM CONSTANT

For $aA + bB \rightleftharpoons cC + dD$: $K_c = [C]^c[D]^d / [A]^a[B]^b$

For gases: $K_p = K_c(RT)^{\Delta n}$, where $\Delta n = (c + d) - (a + b)$ of gaseous species

Reverse reaction: $K' = 1/K$; equation multiplied by n : $K' = K^n$

Pure solids and liquids are not written in K

Reaction quotient Q : if $Q < K$ reaction goes forward; $Q > K$ goes backward; $Q = K$ at equilibrium

Large $K (>10^3)$ $\rightarrow$ mostly products; small $K (<10^{-3})$ $\rightarrow$ mostly reactants.

Le Chatelier's principle

If a system at equilibrium is disturbed (by changing concentration, pressure or temperature), it shifts in the direction that **cancels the disturbance**.

Change	Effect (example: $N_2 + 3H_2 \rightleftharpoons 2NH_3$, $\Delta H = -92 \text{ kJ}$)
Add N_2 or H_2	Shifts forward, more NH_3
Remove NH_3	Shifts forward
Increase pressure	Shifts to fewer gas moles $\rightarrow$ forward
Increase temperature	Favours endothermic direction $\rightarrow$ backward (less NH_3)
Catalyst	No shift; equilibrium is only reached faster
Inert gas at constant volume	No effect

Ionic equilibrium**ACIDS AND BASES**

- **Arrhenius**: acid gives H^+ in water; base gives OH^- .
- **Brønsted-Lowry**: acid = proton donor; base = proton acceptor. $HCl + H_2O \rightarrow H_3O^+ + Cl^-$; Cl^- is the conjugate base of HCl .
- **Lewis**: acid = electron-pair acceptor (BF_3 , $AlCl_3$); base = electron-pair donor (NH_3).
- **Strong electrolytes** ionise almost completely (HCl , $NaOH$, $NaCl$). **Weak electrolytes** ionise partially (CH_3COOH , NH_4OH).
- **Degree of ionisation (α)**: fraction of molecules ionised.
- **Polybasic acids** ionise in steps: $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ (K_{a1}); $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$ (K_{a2}); always $K_{a1} > K_{a2}$.
- **Acid strength** increases with larger K_a (smaller pK_a).

IONIC EQUILIBRIUM FORMULAS

$$K_w = [H^+][OH^-] = 1 \times 10^{-14} \text{ at } 298 \text{ K}$$

$$pH = -\log[H^+] \quad pOH = -\log[OH^-] \quad pH + pOH = 14$$

$$\text{Ostwald dilution law (weak acid): } K_a = C\alpha^2 \quad (\alpha \text{ small}) \Rightarrow \alpha = \sqrt{(K_a/C)}, [H^+] = \sqrt{(K_a C)}$$

$$K_a \times K_b = K_w \text{ (conjugate pair)}$$

$$\text{Henderson equation (acidic buffer): } pH = pK_a + \log([salt]/[acid])$$

$$\text{Basic buffer: } pOH = pK_b + \log([salt]/[base])$$

$$\text{Solubility product of } A_xB_y: K_{sp} = (xs)^x(ys)^y; \text{ for AB: } K_{sp} = s^2; \text{ AB}_2: K_{sp} = 4s^3$$

$$\text{Precipitation occurs when ionic product} > K_{sp}$$

pH scale: $pH < 7$ acidic, 7 neutral, > 7 basic (at 25°C). Blood ≈ 7.4 , lemon juice ≈ 2.2 .

Hydrolysis of salts: Salt of strong acid + strong base (NaCl) $\rightarrow$ neutral. Strong acid + weak base (NH_4Cl) $\rightarrow$ acidic. Weak acid + strong base (CH_3COONa) $\rightarrow$ basic. Weak acid + weak base ($\text{CH}_3\text{COONH}_4$) $\rightarrow$ depends on K_a and K_b .

Buffer solution: resists change in pH on adding a little acid or base. Acidic buffer: $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$. Basic buffer: $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$. Our blood is a buffer ($\text{H}_2\text{CO}_3/\text{HCO}_3^-$).

Common ion effect: the ionisation of a weak electrolyte is suppressed by adding a strong electrolyte having a common ion. Example: adding CH_3COONa to CH_3COOH lowers its ionisation. Used in qualitative analysis: NH_4Cl suppresses NH_4OH ionisation so that only Group III hydroxides precipitate; HCl suppresses H_2S ionisation in Group II.

Example 1. Write K_c for $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ and the relation between K_p and K_c .

$$K_c = [\text{SO}_3]^2/([\text{SO}_2]^2[\text{O}_2]). \Delta n = 2 - 3 = -1, \text{ so } K_p = K_c(RT)^{-1}.$$

Example 2. At equilibrium $[\text{N}_2] = 0.03$, $[\text{H}_2] = 0.02$, $[\text{NH}_3] = 0.003 \text{ mol/L}$. Find K_c .

$$K_c = (0.003)^2/[(0.03)(0.02)^3] = 9 \times 10^{-6}/(0.03 \times 8 \times 10^{-6}) = 9/0.24 = \mathbf{37.5}.$$

Example 3. Find the pH of 0.001 M HCl .

$$[H^+] = 10^{-3} \Rightarrow pH = \mathbf{3}.$$

Example 4. Find the pH of 0.01 M NaOH .

$$[OH^-] = 10^{-2}, pOH = 2, pH = \mathbf{12}.$$

Example 5. Find the pH of 0.1 M acetic acid ($K_a = 1.8 \times 10^{-5}$).

$$[H^+] = \sqrt{(1.8 \times 10^{-5} \times 0.1)} = \sqrt{(1.8 \times 10^{-6})} = 1.34 \times 10^{-3}. pH = 3 - \log 1.34 = 3 - 0.127 = \mathbf{2.87}.$$

Example 6. Find the pH of a buffer containing 0.1 M CH_3COOH and 0.2 M CH_3COONa ($\text{pK}_a = 4.74$).

$$\text{pH} = 4.74 + \log(0.2/0.1) = 4.74 + 0.30 = \mathbf{5.04}.$$

Example 7. The solubility of AgCl is 1.06×10^{-5} mol/L. Find K_{sp} .

$$K_{\text{sp}} = s^2 = (1.06 \times 10^{-5})^2 = \mathbf{1.12 \times 10^{-10}}.$$

Example 8. K_{sp} of CaF_2 is 3.2×10^{-11} . Find its solubility.

$$K_{\text{sp}} = 4s^3 \Rightarrow s^3 = 8 \times 10^{-12} \Rightarrow s = \mathbf{2 \times 10^{-4} \text{ mol/L}}.$$

XI-7. Redox Reactions

CONCEPTS

	Oxidation	Reduction
Classical	Addition of O / removal of H	Removal of O / addition of H
Electronic	Loss of electrons	Gain of electrons
Oxidation number	Increase in O.N.	Decrease in O.N.

Memory aid: **OIL RIG** – Oxidation Is Loss, Reduction Is Gain. The substance that gets oxidised is the **reducing agent**; the one reduced is the **oxidising agent**.

Rules for oxidation number

- Free element: 0 (Na , O_2 , P_4).
- Monoatomic ion: equals its charge ($\text{Na}^+ = +1$).
- O is -2 (except peroxides -1 , OF_2 $+2$, superoxides $-\frac{1}{2}$).
- H is $+1$ (except metal hydrides like NaH , -1).
- F is always -1 . Alkali metals $+1$, alkaline earth $+2$.
- Sum of O.N. in a neutral molecule = 0; in an ion = its charge.

Example 1. Find the O.N. of Cr in $\text{K}_2\text{Cr}_2\text{O}_7$ and Mn in KMnO_4 .

$$2(+1) + 2x + 7(-2) = 0 \Rightarrow x = \mathbf{+6}.$$
 For KMnO_4 : $1 + x - 8 = 0 \Rightarrow x = \mathbf{+7}.$

Example 2. Find the O.N. of S in H_2SO_4 and of C in $\text{C}_6\text{H}_{12}\text{O}_6$.

$$\text{H}_2\text{SO}_4: 2 + x - 8 = 0 \Rightarrow \mathbf{+6}.$$
 Glucose: $6x + 12 - 12 = 0 \Rightarrow x = 0.$

Balancing redox reactions

Oxidation number method: (1) Write skeleton equation (2) find atoms whose O.N. changes (3) equalise increase and decrease by multiplying (4) balance other atoms, then O with H_2O and H with H^+ (acid) or OH^- (base).

Ion-electron (half-reaction) method: split into oxidation and reduction halves, balance each (atoms, then O with H_2O , H with H^+ , then charge with electrons), multiply to equalise electrons, add.

Example 3. Balance $\text{MnO}_4^- + \text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+}$ (acidic medium).

Reduction: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$.

Oxidation: $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^- (\times 5)$.

Add: **$\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$.**

Example 4. Balance $\text{Cr}_2\text{O}_7^{2-} + \text{I}^- \rightarrow \text{Cr}^{3+} + \text{I}_2$ (acidic).

$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$; $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^- (\times 3)$.

$\text{Cr}_2\text{O}_7^{2-} + 6\text{I}^- + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 3\text{I}_2 + 7\text{H}_2\text{O}$.

Types: combination, decomposition, displacement ($\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$), disproportionation (same element is both oxidised and reduced, e.g. $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$).

Applications: batteries and cells, extraction of metals, photosynthesis, respiration, bleaching, titrations (KMnO_4 , $\text{K}_2\text{Cr}_2\text{O}_7$), rusting and its prevention.

Exam Q&A - Thermodynamics, Equilibrium, Redox

Q1. Is heat a state function? [1]

No, it is a path function.

Q2. What is ΔG for a reaction at equilibrium? [1]

Zero.

Q3. State Hess's law. [2]

The total enthalpy change of a reaction is the same whether it takes place in one step or several steps.

Q4. What is the conjugate base of H_2O and the conjugate acid of NH_3 ? [2]

OH^- and NH_4^+ .

Q5. Why is an aqueous solution of NH_4Cl acidic? [2]

NH_4^+ (from weak base) hydrolyses: $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4\text{OH} + \text{H}^+$, releasing H^+ .

Q6. What is the oxidation number of O in H_2O_2 ? [1]

-1.

Q7. Explain the effect of temperature and pressure on the Haber process using Le Chatelier's principle. [4]

The reaction is exothermic and reduces gas moles ($4 \rightarrow 2$). High pressure favours NH_3 . Low temperature favours NH_3 but makes the reaction slow, so an optimum of about 700 K, 200 atm and an iron catalyst is used.

XI-8. Organic Chemistry - Some Basic Principles and Techniques

Idea: Organic chemistry is the chemistry of carbon compounds (except CO, CO₂, carbonates, cyanides). Carbon is special because it can form four strong bonds and long chains or rings with itself (**catenation**). Food, fuels, medicines, plastics and our own body are made of organic compounds.

Methods of purification

Method	Used for	Example
Sublimation	Solids that go directly to vapour	Camphor, naphthalene, iodine
Crystallisation	Solids; based on different solubility	Sugar, benzoic acid
Simple distillation	Liquids with large difference in boiling points	Chloroform (334 K) + aniline (457 K)
Fractional distillation	Liquids with close boiling points	Crude oil refining
Distillation under reduced pressure	Liquids that decompose at their boiling point	Glycerol
Steam distillation	Steam-volatile, water-immiscible liquids	Aniline from water, essential oils
Differential extraction	Separating a compound from water using a solvent	Using ether in a separating funnel
Chromatography	Separating mixtures based on adsorption/partition	Plant pigments (paper / column / TLC)

Qualitative analysis (detecting elements)

- **C and H:** heat with CuO. C → CO₂ turns lime water milky; H → H₂O turns anhydrous CuSO₄ blue.
- **Lassaigne's test** (N, S, halogens): fuse compound with sodium metal, extract with water (sodium fusion extract).
- **N:** boil extract with FeSO₄, acidify with H₂SO₄ → Prussian blue colour.
- **S:** add sodium nitroprusside → violet colour; or lead acetate → black PbS.
- **Halogens:** acidify with HNO₃, add AgNO₃: white ppt (Cl, soluble in NH₃), pale yellow (Br, sparingly soluble), yellow (I, insoluble).
- **P:** oxidise to phosphate; with ammonium molybdate → yellow ppt.

Quantitative analysis

FORMULAS (ESTIMATION OF ELEMENTS)

$\%C = (12/44) \times (\text{mass of CO}_2/\text{mass of compound}) \times 100$ (Liebig's method)

$\%H = (2/18) \times (\text{mass of H}_2\text{O}/\text{mass of compound}) \times 100$

$\%N \text{ (Dumas)} = (28/22400) \times (\text{vol. of N}_2 \text{ at STP in mL/mass}) \times 100$

$\%N \text{ (Kjeldahl)} = 1.4 \times N \times V / \text{mass of compound}$ (N = normality, V = volume of acid used for NH_3)

$\%Cl \text{ (Carius)} = (35.5/143.5) \times (\text{mass of AgCl}/\text{mass}) \times 100$

$\%S = (32/233) \times (\text{mass of BaSO}_4/\text{mass}) \times 100$

Example 1. 0.246 g of an organic compound gave 0.198 g CO_2 and 0.1014 g H_2O . Find %C and %H.

$\%C = (12/44)(0.198/0.246) \times 100 = 21.95\%$. $\%H = (2/18)(0.1014/0.246) \times 100 = 4.58\%$.

Example 2. In Carius method, 0.1890 g of compound gave 0.2870 g AgCl. Find %Cl.

$\%Cl = (35.5/143.5)(0.2870/0.1890) \times 100 = 37.57\%$.

Classification and IUPAC nomenclature

Organic compounds are classified as **acyclic (open chain)** and **cyclic** (alicyclic or aromatic). A **functional group** is the atom/group that gives a compound its characteristic properties. A **homologous series** is a family with the same functional group in which successive members differ by $-\text{CH}_2-$.

Class	Functional group	Suffix/Prefix	Example (IUPAC)
Alkane	C-C	-ane	CH_3CH_3 ethane
Alkene	C=C	-ene	$\text{CH}_2=\text{CH}_2$ ethene
Alkyne	$\text{C}\equiv\text{C}$	-yne	$\text{HC}\equiv\text{CH}$ ethyne
Haloalkane	-X	halo-	CH_3Cl chloromethane
Alcohol	-OH	-ol	$\text{CH}_3\text{CH}_2\text{OH}$ ethanol
Aldehyde	-CHO	-al	CH_3CHO ethanal
Ketone	$>\text{C}=\text{O}$	-one	CH_3COCH_3 propanone
Carboxylic acid	-COOH	-oic acid	CH_3COOH ethanoic acid
Ester	-COOR	alkyl -oate	$\text{CH}_3\text{COOC}_2\text{H}_5$ ethyl ethanoate
Amine	$-\text{NH}_2$	-amine	CH_3NH_2 methanamine
Ether	-O-	alkoxy-	CH_3OCH_3 methoxymethane

IUPAC NAMING STEPS

1. Find the longest carbon chain containing the functional group (word root: meth 1, eth 2, prop 3, but 4, pent 5, hex 6, hept 7, oct 8, non 9, dec 10).
2. Number the chain so that the functional group (then multiple bond, then substituents) gets the lowest number.
3. Name substituents alphabetically with their position numbers (di-, tri- for repeats).
4. Write: prefixes + word root + suffix.

Example: $\text{CH}_3\text{-CH}(\text{CH}_3)\text{-CH}_2\text{-OH} \rightarrow 2\text{-methylpropan-1-ol}$.

Electronic displacements in covalent bonds

Effect	What happens	Key points
Inductive (I)	Permanent shift of σ electrons along a chain towards a more electronegative atom	–I groups pull: $-\text{NO}_2 > -\text{CN} > -\text{COOH} > -\text{F} > -\text{Cl} > -\text{Br} > -\text{I}$. +I groups push: alkyl groups ($-\text{CH}_3$). Decreases with distance.
Electromeric (E)	Temporary, complete shift of π electrons to one atom when an attacking reagent approaches	Disappears when the reagent is removed. +E and –E.
Resonance (mesomeric, R/M)	Permanent delocalisation of π electrons or lone pairs in conjugated systems	+R: $-\text{OH}$, $-\text{NH}_2$, $-\text{Cl}$ (donate to ring). –R: $-\text{NO}_2$, $-\text{CHO}$, $-\text{COOH}$.
Hyperconjugation	Delocalisation of σ electrons of C–H bonds adjacent to a carbocation or double bond	"No-bond resonance". More α -hydrogens $\rightarrow$ more stability. Explains $3^\circ > 2^\circ > 1^\circ$ carbocation stability.

Bond fission and reactive intermediates

- **Homolytic fission:** each atom takes one electron $\rightarrow$ **free radicals** ($\text{CH}_3^\bullet$). Occurs in presence of light/heat/peroxide. Stability: $3^\circ > 2^\circ > 1^\circ > \text{CH}_3^\bullet$.
- **Heterolytic fission:** one atom takes both electrons $\rightarrow$ ions. **Carbocation** (C^+ , sp^2 , planar; stability $3^\circ > 2^\circ > 1^\circ$) and **carbanion** (C^- , sp^3 , pyramidal; stability $1^\circ > 2^\circ > 3^\circ$).
- **Electrophile** ("electron-loving"): electron-deficient species, attacks electron-rich sites: H^+ , NO_2^+ , Cl^+ , BF_3 , AlCl_3 , carbocations.
- **Nucleophile** ("nucleus-loving"): electron-rich species with lone pair/negative charge: OH^- , CN^- , Cl^- , NH_3 , H_2O .

Types of organic reactions

Type	Description	Example
Substitution	One atom/group replaced by another	$\text{CH}_4 + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \text{HCl}$
Addition	Atoms added across a multiple bond	$\text{CH}_2=\text{CH}_2 + \text{Br}_2 \rightarrow \text{CH}_2\text{Br}-\text{CH}_2\text{Br}$
Elimination	Atoms removed to form a multiple bond	$\text{C}_2\text{H}_5\text{OH} \rightarrow \text{CH}_2=\text{CH}_2 + \text{H}_2\text{O}$
Rearrangement	Atoms rearrange within the molecule	Isomerisation of butane

XI-9. Hydrocarbons

Compounds of only carbon and hydrogen. Main source: petroleum and natural gas. Classified as **alkanes** (single bonds, $\text{C}_n\text{H}_{2n+2}$), **alkenes** (one double bond, C_nH_{2n}), **alkynes** (one triple bond, $\text{C}_n\text{H}_{2n-2}$) and **aromatic** (benzene ring).

Alkanes

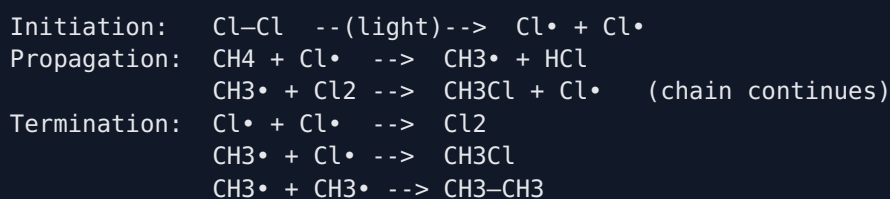
Isomerism: butane has two chain isomers — n-butane and 2-methylpropane (isobutane). Pentane has three.

Conformations of ethane: rotation about the C-C σ bond gives different arrangements. **Staggered** (H atoms as far apart as possible; most stable) and **eclipsed** (H atoms directly behind each other; least stable). The energy difference (torsional strain) is about 12.5 kJ mol^{-1} . Shown by Sawhorse and Newman projections.

Physical properties: non-polar; C1-C4 gases, C5-C17 liquids, higher ones solids. Boiling point increases with molecular mass; branching lowers boiling point (less surface area).

Preparation: (i) hydrogenation of alkenes/alkynes with Ni/Pt/Pd; (ii) Wurtz reaction: $2\text{CH}_3\text{Br} + 2\text{Na} \rightarrow \text{CH}_3-\text{CH}_3 + 2\text{NaBr}$ (dry ether); (iii) decarboxylation: $\text{CH}_3\text{COONa} + \text{NaOH (CaO, heat)} \rightarrow \text{CH}_4 + \text{Na}_2\text{CO}_3$; (iv) Kolbe's electrolysis.

Free radical halogenation of methane (mechanism)



Combustion: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + \text{heat}$. **Pyrolysis (cracking):** heating higher alkanes without air breaks them into smaller alkanes and alkenes — used to get petrol from heavy oil.

Alkenes

Structure of ethene: each C is sp^2 hybridised; C=C has one σ and one π bond; planar molecule, bond angle about 120° , C=C bond length 134 pm.

Geometrical (cis-trans) isomerism: because rotation about C=C is restricted. In but-2-ene, cis has both CH₃ groups on the same side; trans has them on opposite sides. Cis is more polar (higher boiling point); trans packs better (higher melting point).

Preparation: (i) partial hydrogenation of alkynes with Lindlar's catalyst (Pd/BaSO₄) → cis-alkene; (ii) dehydrohalogenation of alkyl halides with alcoholic KOH; (iii) dehydration of alcohols with conc. H₂SO₄ at 443 K; (iv) dehalogenation of vicinal dihalides with Zn.

Chemical reactions of alkenes (electrophilic addition)

- **H₂:** CH₂=CH₂ + H₂ (Ni) → C₂H₆.
- **Halogen:** + Br₂ → 1,2-dibromoethane. Red-brown bromine water is decolourised — **test for unsaturation**.
- **Water:** with dil. H₂SO₄ → alcohol.
- **HX — Markovnikov's rule:** In addition of HX to an unsymmetrical alkene, the negative part (X) goes to the carbon having **fewer hydrogens**. CH₃-CH=CH₂ + HBr → CH₃-CHBr-CH₃ (2-bromopropane). Reason: the more stable 2° carbocation forms.
- **Peroxide effect (Kharasch, anti-Markovnikov):** in presence of peroxide, HBr (only HBr) adds the opposite way: CH₃CH=CH₂ + HBr → CH₃CH₂CH₂Br (1-bromopropane), by a free-radical mechanism.
- **Ozonolysis:** O₃ then Zn/H₂O breaks the C=C, giving aldehydes/ketones. Used to locate the double bond. Propene → ethanal + methanal.
- **Oxidation:** cold dilute alkaline KMnO₄ (Baeyer's reagent) → glycol; the purple colour disappears — test for unsaturation. Hot acidic KMnO₄ → acids/ketones.
- **Polymerisation:** n CH₂=CH₂ → polythene.

Mechanism of electrophilic addition (HBr to propene)

Step 1: HBr → H⁺ + Br⁻
 CH₃-CH=CH₂ + H⁺ → CH₃-CH(+)-CH₃ (2° carbocation, more stable)
 Step 2: CH₃-CH(+)-CH₃ + Br⁻ → CH₃-CHBr-CH₃ (2-bromopropane)

Alkynes

Structure of ethyne: C is sp hybridised; C≡C = 1σ + 2π; linear (180°); bond length 120 pm.

Preparation: (i) CaC₂ + 2H₂O → C₂H₂ + Ca(OH)₂ (calcium carbide); (ii) dehydrohalogenation of vicinal dihalides with alcoholic KOH then NaNH₂.

Acidic character: the H on a triple-bonded carbon (sp, 50% s-character) is slightly acidic. HC≡CH + 2Na → NaC≡CNa + H₂. Terminal alkynes give a white ppt with ammoniacal AgNO₃ and red ppt with ammoniacal Cu₂Cl₂. Acidity: HC≡CH > CH₂=CH₂ > CH₃-CH₃.

Addition reactions: H_2 ($\rightarrow$ alkene $\rightarrow$ alkane); Br_2 ($\rightarrow$ 1,1,2,2-tetrabromoethane); HX (Markovnikov, two steps); water with $\text{HgSO}_4/\text{H}_2\text{SO}_4$ at 333 K $\rightarrow$ ethanal (CH_3CHO). Polymerisation: $3\text{HC}\equiv\text{CH} \rightarrow$ benzene (red-hot iron tube, 873 K).

Aromatic hydrocarbons - Benzene

Structure: C_6H_6 , a flat hexagon. Kekulé proposed alternate single and double bonds, but all C-C bonds are actually equal (139 pm) because of **resonance**: the six π electrons are delocalised over the ring, shown by a circle inside the hexagon.

AROMATICITY - HÜCKEL'S RULE

A compound is aromatic if it is cyclic, planar, fully conjugated and has **$(4n + 2) \pi$ electrons** ($n = 0, 1, 2, \dots$). Benzene has 6 π electrons ($n = 1$).

Chemical properties: despite having "double bonds", benzene prefers **electrophilic substitution** (to keep its stable ring) rather than addition.

Reaction	Reagents	Electrophile	Product
Nitration	conc. HNO_3 + conc. H_2SO_4	NO_2^+	Nitrobenzene
Halogenation	Cl_2 + anhydrous FeCl_3	Cl^+	Chlorobenzene
Sulphonation	fuming H_2SO_4	SO_3	Benzenesulphonic acid
Friedel-Crafts alkylation	CH_3Cl + anhyd. AlCl_3	CH_3^+	Toluene
Friedel-Crafts acylation	CH_3COCl + anhyd. AlCl_3	CH_3CO^+	Acetophenone

Mechanism of electrophilic substitution (3 steps)

- Generation of electrophile:** $\text{HNO}_3 + 2\text{H}_2\text{SO}_4 \rightarrow \text{NO}_2^+ + \text{H}_3\text{O}^+ + 2\text{HSO}_4^-$.
- Formation of carbocation (arenium ion / σ -complex):** NO_2^+ attacks the ring; aromaticity is temporarily lost.
- Removal of proton:** HSO_4^- removes H^+ , aromaticity is restored $\rightarrow$ nitrobenzene.

Directive influence of groups in monosubstituted benzene

Type	Groups	Effect on ring
Ortho- and para-directing (activating)	$-\text{OH}$, $-\text{NH}_2$, $-\text{OCH}_3$, $-\text{CH}_3$	Increase electron density at o and p positions
Ortho- and para-directing (deactivating)	Halogens $-\text{Cl}$, $-\text{Br}$	Strong $-\text{I}$ but $+\text{R}$ directs o/p
Meta-directing (deactivating)	$-\text{NO}_2$, $-\text{CN}$, $-\text{CHO}$, $-\text{COOH}$, $-\text{SO}_3\text{H}$	Withdraw electrons; m-position is least deactivated

Carcinogenicity and toxicity: benzene and many polynuclear hydrocarbons (with more than two fused rings, e.g. 1,2-benzanthracene, 3-methylcholanthrene) formed by incomplete burning of tobacco, coal and petroleum are toxic and can cause cancer. They damage DNA.

Example 1. Write the IUPAC name of $\text{CH}_3\text{-CH=CH-CH}_2\text{-CH}_3$ and show its geometrical isomers.

Pent-2-ene. It has cis-pent-2-ene (CH_3 and C_2H_5 on the same side) and trans-pent-2-ene (opposite sides).

Example 2. An alkene on ozonolysis gives only propanone (CH_3COCH_3). Identify it.

Join two propanone units at the C=O carbons: $(\text{CH}_3)_2\text{C}=\text{C}(\text{CH}_3)_2$, **2,3-dimethylbut-2-ene**.

Example 3. How will you convert ethyne to benzene, and benzene to toluene?

$3\text{HC}\equiv\text{CH} \rightarrow \text{C}_6\text{H}_6$ (red-hot iron tube, 873 K). $\text{C}_6\text{H}_6 + \text{CH}_3\text{Cl} \rightarrow \text{C}_6\text{H}_5\text{CH}_3 + \text{HCl}$ (anhyd. AlCl_3 , Friedel-Crafts).

Example 4. Is cyclopentadienyl anion aromatic?

It is cyclic, planar, conjugated and has 6 π electrons ($4n + 2$, $n = 1$). So **yes, it is aromatic**.

Exam Q&A - Organic Chemistry

Q1. What is the hybridisation of carbon in a carbocation? [1]

sp^2 .

Q2. Which is more stable: $(\text{CH}_3)_3\text{C}^+$ or CH_3CH_2^+ ? Why? [2]

$(\text{CH}_3)_3\text{C}^+$ (3°), due to +I effect of three methyl groups and greater hyperconjugation (9 α -hydrogens).

Q3. Why is a mixture of HNO_3 and H_2SO_4 used in nitration? [2]

H_2SO_4 protonates HNO_3 to produce the nitronium ion NO_2^+ , the actual electrophile.

Q4. State Markovnikov's rule with an example. [2]

See the alkene section: in HBr addition to propene, Br goes to the middle carbon giving 2-bromopropane.

Q5. Why does benzene undergo substitution rather than addition? [2]

Addition would destroy the very stable delocalised π -electron (aromatic) system; substitution keeps it intact.

Q6. Write the IUPAC names: (a) $\text{CH}_3\text{CH}(\text{Cl})\text{CH}_3$ (b) $(\text{CH}_3)_2\text{CHCH}_2\text{CH}_3$ (c) $\text{CH}\equiv\text{C-CH}_2\text{-CH}_3$.

[3]

(a) 2-chloropropane (b) 2-methylbutane (c) but-1-yne.

Chemistry Practical – Class XI (30 marks)

Item	Marks
Salt analysis – acid radical (Dry test 4 + Wet test 6)	10
Crystallisation / Single titration / Equivalent mass / Solubility	10
Viva-voce	6
Record	4

Identification of acid radicals (key tests)

Radical	Dry test (with dil. H_2SO_4)	Confirmatory wet test
CO_3^{2-}	Brisk effervescence; colourless gas turns lime water milky	$\text{BaCl}_2 \rightarrow$ white ppt soluble in dil. HCl
SO_3^{2-}	Gas with burning sulphur smell; turns acidified $\text{K}_2\text{Cr}_2\text{O}_7$ green	$\text{BaCl}_2 \rightarrow$ white ppt soluble in HCl
S^{2-}	Rotten-egg smell; lead acetate paper turns black	Sodium nitroprusside $\rightarrow$ violet colour
NO_2^-	Light brown fumes	$\text{KI} + \text{starch} \rightarrow$ blue colour
Cl^- (conc. H_2SO_4)	Colourless pungent gas, white fumes with NH_3 rod	$\text{AgNO}_3 \rightarrow$ white ppt soluble in NH_4OH
Br^-	Reddish-brown vapours	$\text{AgNO}_3 \rightarrow$ pale yellow ppt partly soluble in NH_4OH
I^-	Violet vapours	$\text{AgNO}_3 \rightarrow$ yellow ppt insoluble in NH_4OH
NO_3^-	Brown fumes on heating with Cu turnings + conc. H_2SO_4	Ring test: brown ring with FeSO_4
SO_4^{2-}	No gas with dil. or conc. acid	$\text{BaCl}_2 \rightarrow$ white ppt insoluble in HCl
PO_4^{3-}	—	Ammonium molybdate + $\text{HNO}_3 \rightarrow$ canary yellow ppt

Volumetric analysis – single titration

Use $N_1V_1 = N_2V_2$. Indicator: phenolphthalein (colourless in acid, pink in base) for strong base titrations; methyl orange (red in acid, yellow in base) for strong acid-weak base. Strength (g/L) = Normality $\times$ Equivalent mass.

Example. 20 mL of NaOH solution needs 25 mL of N/10 oxalic acid. Find the normality and strength of NaOH.

$N_1 \times 20 = 0.1 \times 25 \Rightarrow N_1 = \mathbf{0.125\ N}$. Strength = $0.125 \times 40 = \mathbf{5\ g/L}$.

Viva questions

Why is oxalic acid used as a primary standard?

It is pure, stable, non-hygroscopic and its solution of exact strength can be made by direct weighing.

Why are blue vitriol crystals blue?

Because of five molecules of water of crystallisation ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$); anhydrous CuSO_4 is white.

What is the equivalent mass of Mg found experimentally?

About 12 (atomic mass 24 / valency 2), found by measuring H_2 displaced.

CHAPTER-WISE QUESTION BANK

Chemistry · +2 1st Year · MCQs, short and long answers, numericals

HOW TO USE

Learn the key points first, then answer the questions without looking. For numericals always write: given data → formula → substitution → answer with unit.

QB-1. Some Basic Concepts of Chemistry

Key points

- 1 mol = 6.022×10^{23} particles = molar mass in grams = 22.4 L of gas at STP.
- SI base units: metre, kilogram, second, kelvin, mole, ampere, candela.
- Significant figures: all non-zero digits are significant; zeros between them are significant; leading zeros are not; trailing zeros after a decimal point are significant.
- Molarity depends on temperature; molality and mole fraction do not.
- Limiting reagent is the one that gives the least amount of product.

A. Multiple choice questions

1. The number of moles in 22 g of CO_2 is (a) 1 (b) 0.5 (c) 2 (d) 0.25

(b) 22/44.

2. Which has the maximum number of molecules? (a) 1 g H_2 (b) 1 g O_2 (c) 1 g N_2 (d) 1 g CO_2

(a) smallest molar mass gives most moles.

3. The law of multiple proportions was given by (a) Lavoisier (b) Proust (c) Dalton (d) Avogadro

(c).

4. Number of significant figures in 0.0025 is (a) 4 (b) 2 (c) 3 (d) 5

(b).

5. The empirical formula of C_6H_6 is (a) CH (b) C_2H_2 (c) C_6H_6 (d) CH_2

(a).

6. Molarity of a solution containing 5.85 g NaCl in 500 mL is (a) 0.1 M (b) 0.2 M (c) 1 M (d) 0.5 M

(b) 0.1 mol/0.5 L.

7. Which is temperature-independent? (a) molarity (b) normality (c) molality (d) formality

(c).

8. Volume occupied by 2 mol of a gas at STP is (a) 22.4 L (b) 44.8 L (c) 11.2 L (d) 2 L

(b).

9. Equivalent mass of Mg (atomic mass 24) is (a) 24 (b) 12 (c) 48 (d) 8

(b) $24/2$.

10. Percentage of oxygen in water is about (a) 11.1% (b) 88.9% (c) 50% (d) 66.7%

(b) $16/18 \times 100$.

11. The mass of one atom of carbon-12 is (a) 12 g (b) 1.99×10^{-23} g (c) 6.022×10^{23} g (d) 1 u

(b) $12/6.022 \times 10^{23}$.

12. In $2H_2 + O_2 \rightarrow 2H_2O$, if 4 g H_2 and 40 g O_2 react, the limiting reagent is (a) H_2 (b) O_2 (c) both (d) none

(a) 2 mol H_2 needs only 1 mol (32 g) O_2 ; O_2 is in excess.

B. Very short answer questions

1. Define one atomic mass unit.

Exactly one-twelfth of the mass of one carbon-12 atom.

2. What is Avogadro's number?

6.022×10^{23} , the number of particles in one mole.

3. Give an example of a homogeneous and a heterogeneous mixture.

Sugar solution; sand in water.

4. What is a limiting reagent?

The reactant that is completely used up first and limits the amount of product.

5. Write the relation between molarity and normality.

Normality = Molarity $\times$ n-factor.

6. Express 0.000045 in scientific notation.

4.5×10^{-5} .

C. Short answer questions

1. State and illustrate the law of definite proportions.

A given compound always contains exactly the same elements combined in the same fixed ratio by mass, whatever its source or method of preparation. Example: CO_2 from burning charcoal or from heating limestone always has C : O = 3 : 8 by mass.

2. State the postulates of Dalton's atomic theory and its two limitations.

See chapter notes (four postulates). Limitations: atoms are divisible into sub-atomic particles; isotopes of the same element have different masses (and isobars of different elements have the same mass).

3. Calculate the number of atoms in (i) 52 mol Ar (ii) 52 u He (iii) 52 g He.

(i) $52 \times 6.022 \times 10^{23} = 3.13 \times 10^{25}$ atoms. (ii) One He atom = 4 u $\Rightarrow 52/4 = 13$ atoms. (iii) $52/4 = 13$ mol $\Rightarrow 13 \times 6.022 \times 10^{23} = 7.83 \times 10^{24}$ atoms.

4. What is the concentration of sugar (C₁₂H₂₂O₁₁) in mol/L if 20 g are dissolved in enough water to make 2 L?

Molar mass 342 g. $n = 20/342 = 0.0585$ mol. $M = 0.0585/2 = \mathbf{0.0292\ M}$.

5. If the density of methanol is 0.793 kg/L, what volume is needed to make 2.5 L of its 0.25 M solution?

Molarity of pure methanol = $793/32 = 24.78$ M. $M_1V_1 = M_2V_2$: $24.78 V_1 = 0.25 \times 2.5 \Rightarrow V_1 = 0.0252$ L = **25.2 mL**.

6. Calculate the mass of sodium acetate (CH₃COONa) needed to make 500 mL of 0.375 M solution.

$n = 0.375 \times 0.5 = 0.1875$ mol; molar mass 82.0 g $\Rightarrow$ mass = **15.4 g**.

D. Numerical / long answer questions**1. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% oxygen by mass (Fe = 55.85).**

Fe: $69.9/55.85 = 1.25$; O: $30.1/16 = 1.88$. Ratio $1.25 : 1.88 = 1 : 1.5 = 2 : 3$. Empirical formula **Fe₂O₃**.

2. Calculate the amount of carbon dioxide produced when (i) 1 mol carbon burns in air (ii) 1 mol carbon burns in 16 g O₂ (iii) 2 mol carbon burn in 16 g O₂.

$C + O_2 \rightarrow CO_2$. (i) 1 mol $CO_2 = 44$ g. (ii) 16 g $O_2 = 0.5$ mol is limiting $\Rightarrow 0.5$ mol $CO_2 = 22$ g. (iii) O_2 still limiting $\Rightarrow 22$ g.

3. Calculate the mass of CaO obtained by heating 1 kg of 90% pure limestone (CaCO₃).

Pure $CaCO_3 = 900$ g. $CaCO_3 \rightarrow CaO + CO_2$: 100 g gives 56 g. Mass $CaO = 900 \times 56/100 = \mathbf{504\ g}$.

4. A compound contains 40% C, 6.67% H and 53.33% O. Its vapour density is 30. Find its molecular formula.

C: $40/12 = 3.33$; H: $6.67/1 = 6.67$; O: $53.33/16 = 3.33$. Ratio 1 : 2 : 1 $\Rightarrow$ CH_2O (30).

Molar mass = $2 \times \text{VD} = 60 \Rightarrow n = 2$. **$\text{C}_2\text{H}_4\text{O}_2$.**

5. Calculate the molality and mole fraction of a 10% (w/w) aqueous NaOH solution.

100 g solution: 10 g NaOH (0.25 mol), 90 g water (5 mol). Molality = $0.25/0.09 = 2.78$

m. $x(\text{NaOH}) = 0.25/5.25 = 0.048$.

6. How many grams of HCl react with 5 g of MnO_2 ? $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + 2\text{H}_2\text{O} + \text{Cl}_2$ (Mn = 55).

$\text{MnO}_2 = 87$ g. 87 g reacts with $4 \times 36.5 = 146$ g HCl. 5 g reacts with $146 \times 5/87 = 8.4$ g.

QB-2. Structure of Atom

Key points

- e/m of electron = 1.758×10^{11} C/kg (Thomson); charge measured by Millikan's oil-drop experiment.
- Number of orbitals in a shell = n^2 ; electrons = $2n^2$; orbitals in a subshell = $2l + 1$.
- Nodes: radial = $n - l - 1$; angular = l ; total = $n - 1$.
- Energy order is decided by $(n + l)$; same $(n + l) \rightarrow$ lower n first.
- Number of spectral lines when an electron falls from n to 1 = $n(n - 1)/2$.

A. Multiple choice questions

1. The number of neutrons in $^{35}_{17}\text{Cl}$ is (a) 17 (b) 18 (c) 35 (d) 52

(b).

2. Which pair are isobars? (a) ^{12}C , ^{14}C (b) ^{40}Ar , ^{40}Ca (c) ^1H , ^2H (d) O_2 , O_3

(b).

3. The maximum number of electrons in the M shell is (a) 8 (b) 18 (c) 32 (d) 2

(b) 2×3^2 .

4. For $l = 2$, the number of orbitals is (a) 3 (b) 5 (c) 7 (d) 1

(b).

5. Which orbital is dumb-bell shaped? (a) s (b) p (c) d_{xy} (d) f

(b).

6. Which set of quantum numbers is not possible? (a) $n = 3, l = 2, m = 0$ (b) $n = 2, l = 0, m = -1$ (c) $n = 4, l = 3, m = -3$ (d) $n = 1, l = 0, m = 0$

(b) for $l = 0$, m must be 0.

7. The configuration of Cu ($Z = 29$) is (a) $[\text{Ar}]3d^94s^2$ (b) $[\text{Ar}]3d^{10}4s^1$ (c) $[\text{Ar}]3d^84s^24p^1$ (d) $[\text{Ar}]4s^23d^9$

(b) fully filled d is stable.

8. Which series of hydrogen lies in the visible region? (a) Lyman (b) Balmer (c) Paschen (d) Pfund

(b).

9. de Broglie wavelength is (a) h/mv (b) mv/h (c) h/m (d) h/mv

(a).

10. Number of unpaired electrons in N ($Z = 7$) is (a) 1 (b) 2 (c) 3 (d) 5

(c) $2p^3$ with Hund's rule.

11. Which orbital is filled after 4s? (a) 4p (b) 3d (c) 4d (d) 5s

(b).

12. The radius of the first Bohr orbit of hydrogen is (a) $0.529 \text{ \AA}$ (b) $2.12 \text{ \AA}$ (c) $1 \text{ \AA}$ (d) $0.0529 \text{ \AA}$

(a).

B. Very short answer questions

1. Who discovered the neutron?

James Chadwick (1932).

2. What are isotones? Give an example.

Atoms with the same number of neutrons, e.g. ^{14}C and ^{16}O (8 neutrons each).

3. State Pauli's exclusion principle.

No two electrons in an atom can have the same set of all four quantum numbers.

4. How many electrons can have $n = 3$, $l = 1$?

3p subshell: 6 electrons.

5. Write the configuration of Fe^{3+} ($Z = 26$).

$[\text{Ar}]3d^5$.

6. What is the photoelectric effect?

Emission of electrons from a metal surface when light of suitable frequency falls on it.

C. Short answer questions

1. What are the main observations and conclusions of the cathode ray experiment?

Rays travel in straight lines from cathode to anode, cast shadows, rotate a light paddle wheel (they have mass and energy), are deflected towards the positive plate (negatively charged), and their nature does not depend on the gas or electrode material. Conclusion: cathode rays are streams of negatively charged particles (electrons), present in all atoms.

2. Explain Hund's rule with the example of nitrogen and oxygen.

In degenerate orbitals electrons occupy singly with parallel spins before pairing. N ($2p^3$): $\uparrow \uparrow \uparrow$ — three unpaired. O ($2p^4$): $\uparrow \downarrow \uparrow \uparrow$ — pairing starts only with the 4th electron, two unpaired.

3. Why are half-filled and fully filled orbitals more stable?

(i) Symmetrical distribution of electrons gives lower energy. (ii) Exchange energy: electrons with parallel spins in degenerate orbitals can exchange positions; the greater the number of exchanges (maximum for d^5 , d^{10}), the greater the stability.

4. Calculate the frequency and wave number of radiation of wavelength 5800 Å.

$\lambda = 5.8 \times 10^{-7} \text{ m}$. $\nu = c/\lambda = 3 \times 10^8 / 5.8 \times 10^{-7} = 5.17 \times 10^{14} \text{ Hz}$. Wave number = $1/\lambda = 1.72 \times 10^6 \text{ m}^{-1}$.

5. How many spectral lines are produced when an electron in hydrogen falls from $n = 5$ to $n = 1$?

$n(n-1)/2 = 5 \times 4/2 = 10$ lines.

6. Give the number of radial and angular nodes in 3p and 4d orbitals.

3p: radial $3 - 1 - 1 = 1$, angular 1. 4d: radial $4 - 2 - 1 = 1$, angular 2.

D. Numerical / long answer questions**1. Calculate the energy associated with the first orbit of He^+ , and the radius of this orbit.**

$E_1 = -2.18 \times 10^{-18} \times Z^2/n^2 = -2.18 \times 10^{-18} \times 4 = -8.72 \times 10^{-18} \text{ J}$. $r = 0.529 \times 1/2 = 0.2645 \text{ Å}$.

2. The threshold frequency of a metal is $7.0 \times 10^{14} \text{ Hz}$. Calculate the kinetic energy of an electron emitted when radiation of frequency $1.0 \times 10^{15} \text{ Hz}$ hits it.

$\text{KE} = h(\nu - \nu_0) = 6.626 \times 10^{-34} \times 3 \times 10^{14} = 1.99 \times 10^{-19} \text{ J}$.

3. Calculate the wavelength of the line emitted when the electron in hydrogen jumps from $n = 4$ to $n = 2$.

$1/\lambda = R(1/4 - 1/16) = 109677 \times 3/16 = 20564 \text{ cm}^{-1} \Rightarrow \lambda = 486 \text{ nm}$ (blue-green, Balmer series).

4. Calculate the mass of a photon with wavelength 3.6 Å.

$m = h/(\lambda c) = 6.626 \times 10^{-34} / (3.6 \times 10^{-10} \times 3 \times 10^8) = 6.13 \times 10^{-33} \text{ kg}$.

5. A golf ball has a mass of 40 g and speed 45 m/s. If the speed can be measured within an accuracy of 2%, calculate the uncertainty in position.

$\Delta v = 0.9 \text{ m/s}$. $\Delta x = h/(4\pi m \Delta v) = 6.626 \times 10^{-34} / (4 \times 3.14 \times 0.04 \times 0.9) = 1.46 \times 10^{-33} \text{ m}$ — negligible, so the uncertainty principle doesn't matter for large objects.

6. Explain the four quantum numbers in detail with their significance and allowed values.

n : size and energy of the shell, 1, 2, 3, ...; l : shape of subshell, 0 to $n - 1$ (s, p, d, f); m_l : orientation, $-l$ to $+l$; m_s : spin, $+\frac{1}{2}$ or $-\frac{1}{2}$. Together they give the "address" of an electron; no two electrons in an atom share all four (Pauli).

QB-3. Classification of Elements and Periodicity

Key points

- Modern periodic law is based on atomic number (Moseley).
- Period number = principal quantum number of the valence shell. Group: s-block = valence electrons; p-block = 10 + valence electrons; d-block = $(n - 1)d + ns$ electrons.
- Atomic radius: decreases across a period, increases down a group. Noble gases have the largest radius in their period (van der Waals radius).
- $IE_1 < IE_2 < IE_3$ always. Noble gases have the highest IE in each period.
- Electron gain enthalpy of noble gases is positive; Cl has the most negative value.
- Diagonal relationship: Li-Mg, Be-Al, B-Si show similar properties.

A. Multiple choice questions

1. Which element has the largest atomic radius? (a) Na (b) K (c) Mg (d) Li
(b) lowest in the group.
2. The most electronegative element is (a) O (b) Cl (c) F (d) N
(c).
3. Which has the highest first ionisation enthalpy? (a) Na (b) Mg (c) Al (d) Ar
(d) noble gas.
4. Element with configuration $[Ne]3s^2 3p^3$ belongs to group (a) 3 (b) 13 (c) 15 (d) 5
(c) phosphorus.
5. IUPAC name of element 109 is (a) Unnilennium (b) Unnilnonium (c) Ununnilium (d) Unnilnilium
(a) 1-0-9 = un-nil-enn + ium.
6. Which ion is the smallest? (a) Na^+ (b) Mg^{2+} (c) Al^{3+} (d) F^-
(c) isoelectronic, highest nuclear charge.
7. Which has the most negative electron gain enthalpy? (a) F (b) Cl (c) Br (d) O
(b).
8. Which is a metalloid? (a) Na (b) Si (c) S (d) Cl
(b).
9. Across a period, metallic character (a) increases (b) decreases (c) stays same (d) first increases then decreases
(b).
10. Mendeleev predicted eka-silicon, which is now (a) Ga (b) Ge (c) Sc (d) Si
(b).

B. Very short answer questions

1. State the modern periodic law.

Properties of elements are periodic functions of their atomic numbers.

2. Why do noble gases have positive electron gain enthalpy?

They have stable ns^2np^6 configuration; the added electron must enter a new, higher shell.

3. Name the elements of the 2nd period with half-filled and fully filled p orbitals.

N ($2p^3$) and Ne ($2p^6$).

4. What is the valency of an element with $Z = 16$?

2 (configuration 2, 8, 6).

C. Short answer questions

1. Why is the second ionisation enthalpy of sodium very high?

After losing one electron Na^+ has the stable neon configuration; removing another electron means breaking a complete octet from an inner shell close to the nucleus, which needs very high energy.

2. Arrange in increasing order of the property indicated: (i) F, Cl, Br, I (electron gain enthalpy - negative value) (ii) B, C, N, O (first IE).

(i) $I < Br < F < Cl$. (ii) $B < C < O < N$.

3. Explain why cations are smaller and anions larger than their parent atoms.

A cation has fewer electrons with the same nuclear charge, so the effective attraction per electron increases and the size shrinks (often a whole shell is lost). An anion has more electrons, increased repulsion and lower effective nuclear charge per electron, so the electron cloud expands.

4. What are the major differences between Mendeleev's and the modern periodic table?

Mendeleev: based on atomic mass, 8 groups, gaps for undiscovered elements, isotopes and anomalous pairs (Ar-K) not explained. Modern: based on atomic number, 18 groups and 7 periods, blocks s, p, d, f, anomalies removed, isotopes placed together.

5. Predict the formula of a stable compound formed by an element with $Z = 114$ and fluorine.

$Z = 114$ is in group 14 (like Pb). Inert pair effect makes +2 stable: formula MF_2 (MF_4 also possible).

D. Long answer question

What is ionisation enthalpy? Discuss the factors affecting it and its variation in the periodic table.

Ionisation enthalpy is the energy required to remove the most loosely bound electron from an isolated gaseous atom in its ground state: $X(g) \rightarrow X^+(g) + e^-$. **Factors:** (1) Atomic size — larger atom, lower IE. (2) Nuclear charge — higher charge, higher IE. (3) Shielding by inner electrons — more shielding, lower IE. (4) Penetration — $s > p > d > f$; an s electron is harder to remove. (5) Stable configurations (half/fully filled) raise IE. **Variation:** increases across a period (with exceptions $Be > B$, $N > O$, $Mg > Al$); decreases down a group because size and shielding increase.

QB-4. Chemical Bonding and Molecular Structure

Key points

- Lone pairs reduce bond angles: CH_4 $109.5^\circ > NH_3$ $107^\circ > H_2O$ 104.5° .
- Dipole moment zero: CO_2 , BF_3 , CH_4 , CCl_4 , $BeCl_2$. Non-zero: H_2O , NH_3 , $CHCl_3$, HCl .
- NH_3 (1.47 D) has a larger dipole moment than NF_3 (0.23 D): in NH_3 the lone-pair moment adds to the bond moments; in NF_3 it opposes them.
- Bond order $\propto$ bond strength $\propto$ 1/bond length.
- Paramagnetic species in MOT: B_2 , O_2 , O_2^+ , O_2^- ; diamagnetic: N_2 , C_2 , F_2 , O_2^{2-} .

A. Multiple choice questions

1. Shape of NH_3 is (a) planar (b) pyramidal (c) tetrahedral (d) linear

(b).

2. Hybridisation of carbon in C_2H_2 is (a) sp^3 (b) sp^2 (c) sp (d) dsp^2

(c).

3. Which molecule has zero dipole moment? (a) H_2O (b) NH_3 (c) CO_2 (d) HCl

(c).

4. Bond order of O_2 is (a) 1 (b) 2 (c) 2.5 (d) 3

(b).

5. Which has a coordinate bond? (a) NH_4^+ (b) CH_4 (c) H_2O (d) $NaCl$

(a).

6. Number of σ and π bonds in ethene: (a) 5σ , 1π (b) 4σ , 2π (c) 6σ , 0π (d) 3σ , 1π

(a).

7. Which is paramagnetic? (a) N_2 (b) O_2 (c) F_2 (d) H_2

(b).

8. Shape of SF_4 is (a) tetrahedral (b) square planar (c) see-saw (d) octahedral

(c) 4 bp + 1 lp.

9. Which shows hydrogen bonding? (a) CH_4 (b) HF (c) H_2S (d) HCl

(b).

10. Which is the most covalent? (a) LiCl (b) NaCl (c) KCl (d) CsCl

(a) smallest cation (Fajans).

11. Hybridisation of P in PCl_5 : (a) sp^3 (b) sp^3d (c) sp^3d^2 (d) sp^2

(b).

12. He_2 does not exist because its bond order is (a) 1 (b) 0 (c) 2 (d) 0.5

(b).

B. Very short answer questions

1. Define bond enthalpy.

Energy needed to break one mole of a particular bond in gaseous molecules.

2. Write the Lewis structure of CO_2 .

$\text{:}\ddot{\text{O}}=\text{C}=\ddot{\text{O}}\text{:}$ (each O with two lone pairs, double bonds to C).

3. What is the formal charge on each O in O_3 (structure $\text{O}=\text{O}-\text{O}$)?

Central O: +1; double-bonded O: 0; single-bonded O: -1.

4. Which hybrid orbitals are used by Xe in XeF_4 ?

sp^3d^2 (square planar).

5. Why is o-nitrophenol more volatile than p-nitrophenol?

o-Nitrophenol has intramolecular H-bonding (no association), while p-nitrophenol forms intermolecular H-bonds.

C. Short answer questions

1. Explain the formation of the ionic bond in MgO with electron-dot diagram.

Mg (2, 8, 2) loses 2 electrons $\rightarrow \text{Mg}^{2+}$; O (2, 6) gains 2 electrons $\rightarrow \text{O}^{2-}$ (both reach the neon configuration). Electrostatic attraction between Mg^{2+} and O^{2-} forms the ionic bond: $\text{Mg:} + \cdot\ddot{\text{O}}\cdot \rightarrow \text{Mg}^{2+}[\text{:}\ddot{\text{O}}\text{:}]^{2-}$.

2. Although both CO_2 and H_2O are triatomic, the dipole moment of H_2O is non-zero. Explain.

CO_2 is linear, so the two C=O dipoles are equal and opposite and cancel. H_2O is bent (104.5°), so the O-H dipoles add up to a net dipole moment (1.85 D).

3. Draw the resonance structures of the carbonate ion.

Three structures, each with one C=O double bond and two C-O⁻ single bonds, the double bond rotating among the three O atoms. The real ion is a hybrid; all C-O bonds are equal (bond order 4/3).

4. Use VSEPR to predict the shapes of BeCl_2 , BCl_3 , SiCl_4 , AsF_5 , H_2S , PH_3 .

BeCl_2 linear; BCl_3 trigonal planar; SiCl_4 tetrahedral; AsF_5 trigonal bipyramidal; H_2S bent; PH_3 pyramidal.

5. Compare the relative stability of O_2^+ , O_2 , O_2^- , O_2^{2-} on the basis of MOT.

Bond orders: 2.5, 2, 1.5, 1. Stability: $\text{O}_2^+ > \text{O}_2 > \text{O}_2^- > \text{O}_2^{2-}$.

6. Explain the hybridisation and shape of PCl_5 . Why are axial bonds longer than equatorial bonds?

P: $3s^2 3p^3 \rightarrow$ excited $3s^1 3p^3 3d^1 \rightarrow sp^3d$ hybridisation $\rightarrow$ trigonal bipyramidal: three equatorial P-Cl bonds at 120° and two axial bonds at 90° to the plane. Each axial bond pair is repelled by three equatorial bond pairs at 90° , so axial bonds are longer and weaker (PCl_5 is reactive).

D. Long answer questions**1. Describe the molecular orbital diagram of N_2 and O_2 . Calculate their bond orders and explain their magnetic behaviour.**

N_2 (14 e): $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 (\pi 2p_x^2 = \pi 2p_y^2) \sigma 2p_z^2$. $N_b = 10$, $N_a = 4$, B.O. = 3; all paired $\Rightarrow$ diamagnetic; very strong triple bond.

O_2 (16 e): $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 (\pi 2p_x^2 = \pi 2p_y^2) (\pi^* 2p_x^1 = \pi^* 2p_y^1)$. $N_b = 10$, $N_a = 6$, B.O. = 2; two unpaired electrons in $\pi^* \Rightarrow$ paramagnetic (explains why liquid O_2 is attracted by a magnet).

2. What is a hydrogen bond? Explain its types and its effect on properties with examples.

A weak electrostatic attraction between H bonded to a highly electronegative atom (F, O, N) and another electronegative atom. **Intermolecular** (between molecules: H_2O , HF, alcohols) — raises boiling point and solubility in water; explains why water is liquid and ice is less dense than water (open cage structure). **Intramolecular** (within a molecule: o-nitrophenol, salicylaldehyde) — lowers boiling point and water solubility. Strength $\sim 10\text{--}40$ kJ/mol, weaker than covalent bonds but stronger than van der Waals forces. It is important in DNA base pairing and protein structure.

QB-5. Chemical Thermodynamics

Key points

- State functions: U, H, S, G, P, V, T . Path functions: q, w .
- $\Delta U = q + w$; $w = -P_{\text{ext}}\Delta V$; free expansion (into vacuum) $w = 0$.
- $\Delta H = \Delta U + \Delta n_g RT$; if $\Delta n_g = 0$, $\Delta H = \Delta U$.
- Standard enthalpy of formation of an element in its standard state is zero.
- $\Delta G = \Delta H - T\Delta S$; spontaneous if $\Delta G < 0$.
- Entropy increases: solid $\rightarrow$ liquid $\rightarrow$ gas; dissolving; increase in number of gas molecules; rise in temperature.

A. Multiple choice questions

1. Which is an intensive property? (a) volume (b) mass (c) temperature (d) enthalpy
(c).
2. For an isolated system, (a) $q = 0, w = 0$ (b) only $q = 0$ (c) only $w = 0$ (d) $\Delta U > 0$
(a) no exchange of energy or matter.
3. In an adiabatic process, (a) $\Delta T = 0$ (b) $q = 0$ (c) $\Delta V = 0$ (d) $\Delta P = 0$
(b).
4. For $N_2 + 3H_2 \rightarrow 2NH_3$, $\Delta H - \Delta U =$ (a) RT (b) $-2RT$ (c) $2RT$ (d) 0
(b) $\Delta n_g = -2$.
5. Standard enthalpy of formation is zero for (a) CO_2 (b) O_3 (c) diamond (d) graphite
(d) graphite is the standard state of carbon.
6. A reaction with $\Delta H < 0$ and $\Delta S > 0$ is (a) always spontaneous (b) never spontaneous (c) spontaneous at low T only (d) spontaneous at high T only
(a).
7. Entropy of a perfect crystal at 0 K is (a) 1 (b) 0 (c) infinity (d) negative
(b) third law.
8. Hess's law is based on (a) first law (b) second law (c) third law (d) zeroth law
(a) energy conservation; H is a state function.
9. Which process has $\Delta S < 0$? (a) melting of ice (b) evaporation (c) freezing of water (d) dissolving sugar
(c).
10. Enthalpy of combustion is always (a) positive (b) negative (c) zero (d) either
(b).

B. Very short answer questions

1. State the first law of thermodynamics.

Energy can neither be created nor destroyed; $\Delta U = q + w$.

2. What is the sign of w when a gas expands?

Negative (work done by the system).

3. Define enthalpy of atomisation.

Enthalpy change on breaking one mole of bonds completely to get atoms in the gas phase.

4. Give one example each of a state function and a path function.

Enthalpy; work.

5. What happens to ΔG at equilibrium?

$\Delta G = 0$.

C. Short answer questions**1. Distinguish between extensive and intensive properties with examples.**

Extensive properties depend on the amount of matter (mass, volume, internal energy, enthalpy, heat capacity). Intensive properties do not (temperature, pressure, density, viscosity, molar heat capacity, refractive index). Ratio of two extensive properties is intensive (density = mass/volume).

2. Calculate the work done when 2 mol of an ideal gas expands isothermally and reversibly from 10 L to 20 L at 300 K.

$$w = -2.303 nRT \log(V_2/V_1) = -2.303 \times 2 \times 8.314 \times 300 \times \log 2 = -2.303 \times 4988.4 \times 0.301 = -3458 \text{ J} \approx -3.46 \text{ kJ}.$$

3. The enthalpy of combustion of methane, graphite and hydrogen are -890.3 , -393.5 and -285.8 kJ/mol. Find the enthalpy of formation of methane.

$$\text{C} + 2\text{H}_2 \rightarrow \text{CH}_4. \Delta_f H = [\Delta_c H(\text{C}) + 2\Delta_c H(\text{H}_2)] - \Delta_c H(\text{CH}_4) = (-393.5 - 571.6) - (-890.3) = -74.8 \text{ kJ/mol}.$$

4. For the reaction $2\text{A}(\text{g}) + \text{B}(\text{g}) \rightarrow 2\text{D}(\text{g})$, $\Delta U^\circ = -10.5$ kJ and $\Delta S^\circ = -44.1$ J/K at 298 K. Calculate ΔG° and predict spontaneity.

$$\Delta n_g = 2 - 3 = -1. \Delta H = \Delta U + \Delta n_g RT = -10.5 + (-1)(8.314 \times 10^{-3})(298) = -12.98 \text{ kJ}. \Delta G = \Delta H - T\Delta S = -12.98 - 298(-0.0441) = -12.98 + 13.14 = +0.16 \text{ kJ} \Rightarrow \text{non-spontaneous}.$$

5. Predict the sign of ΔS : (i) $2\text{Cl}(\text{g}) \rightarrow \text{Cl}_2(\text{g})$ (ii) $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ (iii) $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$.

(i) negative (fewer gas particles) (ii) positive (gas formed) (iii) positive.

D. Long answer questions

1. Explain how ΔU and ΔH are measured experimentally.

ΔU - bomb calorimeter: a strong steel vessel (bomb) containing the sample in oxygen is immersed in a known mass of water. The sample is ignited electrically; heat released raises the water temperature. Since volume is constant, $q_v = \Delta U = -C_{\text{cal}} \times \Delta T$. **ΔH - coffee-cup calorimeter:** reaction is done in an insulated vessel open to the atmosphere (constant pressure); $q_p = \Delta H = -(m c \Delta T)$ of the solution.

2. State Hess's law and calculate the enthalpy of formation of CO from: $\text{C(s)} + \text{O}_2 \rightarrow \text{CO}_2$, $\Delta H = -393.5 \text{ kJ}$; $\text{CO} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}_2$, $\Delta H = -283.0 \text{ kJ}$. Mention two applications of Hess's law.

Hess's law: total enthalpy change is independent of the path. Subtracting: $\text{C} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}$, $\Delta H = -110.5 \text{ kJ}$. Applications: (i) calculating enthalpies that cannot be measured directly (formation of CO, CH_4); (ii) finding enthalpies of transition (graphite $\rightarrow$ diamond), lattice enthalpy (Born-Haber cycle).

3. What is Gibbs energy? Explain the effect of temperature on spontaneity with examples.

$G = H - TS$; $\Delta G = \Delta H - T\Delta S$ is the useful work available and the criterion of spontaneity at constant T and P. (i) $\Delta H -$, $\Delta S +$: always spontaneous (combustion). (ii) $\Delta H +$, $\Delta S -$: never spontaneous. (iii) $\Delta H -$, $\Delta S -$: spontaneous at low T (freezing of water below 0°C). (iv) $\Delta H +$, $\Delta S +$: spontaneous at high T (melting of ice above 0°C , decomposition of CaCO_3).

QB-6. Equilibrium

Key points

- K is constant at a fixed temperature; it changes only with temperature.
- For exothermic reactions K decreases with rise in temperature; for endothermic reactions it increases.
- pH of pure water at $25^\circ\text{C} = 7$; at higher temperature it is less than 7 but water is still neutral.
- pH of a buffer = pK_a when $[\text{salt}] = [\text{acid}]$.
- Precipitation occurs when ionic product $> K_{\text{sp}}$.

A. Multiple choice questions

1. For $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$, K_p and K_c are related as (a) $K_p = K_c$ (b) $K_p = K_c RT$ (c) $K_p > K_c$ (d) $K_p = K_c/RT$

(a) $\Delta n = 0$.

2. Addition of a catalyst (a) shifts equilibrium forward (b) increases K (c) helps reach equilibrium faster (d) shifts it backward

(c).

3. pH of 10^{-3} M NaOH is (a) 3 (b) 11 (c) 10 (d) 7

(b).

4. The conjugate base of HSO_4^- is (a) H_2SO_4 (b) SO_4^{2-} (c) SO_3 (d) OH^-

(b).

5. Which is a Lewis acid? (a) NH_3 (b) H_2O (c) BF_3 (d) Cl^-

(c).

6. Aqueous solution of CH_3COONa is (a) acidic (b) basic (c) neutral (d) amphoteric

(b).

7. K_{sp} of AB_2 in terms of solubility s is (a) s^2 (b) $2s^2$ (c) $4s^3$ (d) $27s^4$

(c).

8. In $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$, increasing pressure will (a) increase NO_2 (b) decrease NO_2 (c) no effect (d) double K

(b) shifts to fewer gas moles.

9. Which mixture is a buffer? (a) $\text{HCl} + \text{NaCl}$ (b) $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$ (c) $\text{NaOH} + \text{NaCl}$ (d) $\text{H}_2\text{SO}_4 + \text{Na}_2\text{SO}_4$

(b).

10. The pH of a solution with $[\text{H}^+] = 10^{-9}$ M is (a) 9 (b) 5 (c) -9 (d) 14

(a).

B. Very short answer questions

1. State the law of mass action.

Rate of a reaction is proportional to the product of molar concentrations of reactants, each raised to its stoichiometric coefficient.

2. Write K_c for $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$.

$K_c = [\text{CO}_2]$ (solids omitted).

3. What is the ionic product of water at 298 K?

1.0×10^{-14} .

4. Name a salt whose aqueous solution is acidic.

NH_4Cl (or CuSO_4).

5. Why is NH_4Cl added before NH_4OH in Group III analysis?

Common ion NH_4^+ suppresses ionisation of NH_4OH , so $[\text{OH}^-]$ is low and only Group III hydroxides (low K_{sp}) precipitate.

C. Short answer questions

1. At 450 K, $K_p = 2.0 \times 10^{10}$ /bar for $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$. Find K_c ($R = 0.0831 \text{ L bar K}^{-1} \text{ mol}^{-1}$).

$$K_c = K_p(RT)^{-\Delta n}, \Delta n = -1 \Rightarrow K_c = K_p \times RT = 2 \times 10^{10} \times 0.0831 \times 450 = 7.48 \times 10^{11} \text{ L/mol.}$$

2. The value of K_c for $\text{A} + \text{B} \rightleftharpoons \text{C} + \text{D}$ is 4. If 1 mol each of A and B are taken in 1 L, find the equilibrium concentrations.

$$\text{Let } x \text{ react: } K = x^2/(1-x)^2 = 4 \Rightarrow x/(1-x) = 2 \Rightarrow x = 2/3. [\text{C}] = [\text{D}] = 0.667 \text{ M; } [\text{A}] = [\text{B}] = 0.333 \text{ M.}$$

3. Calculate the pH of 0.005 M $\text{Ba}(\text{OH})_2$.

$$[\text{OH}^-] = 0.01 \text{ M} \Rightarrow \text{pOH} = 2 \Rightarrow \text{pH} = 12.$$

4. The ionisation constant of HF is 3.2×10^{-4} . Calculate the degree of ionisation and pH in its 0.02 M solution.

$$\alpha = \sqrt{(K_a/C)} = \sqrt{(1.6 \times 10^{-2})} = 0.126 \text{ (12.6\%; approximate). } [\text{H}^+] = C\alpha = 2.53 \times 10^{-3} \Rightarrow \text{pH} \approx 2.6. \text{ (Exact quadratic solution gives } \alpha \approx 0.12, \text{ pH} \approx 2.62.)$$

5. Explain the common ion effect with an example.

Suppression of the ionisation of a weak electrolyte by adding a strong electrolyte having a common ion. $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$; adding CH_3COONa raises $[\text{CH}_3\text{COO}^-]$, equilibrium shifts left and $[\text{H}^+]$ falls (pH rises). Used in salt analysis and in purifying common salt (passing HCl gas into brine precipitates pure NaCl).

6. The solubility product of Ag_2CrO_4 is 1.1×10^{-12} . Find its molar solubility.

$$K_{sp} = (2s)^2s = 4s^3 \Rightarrow s^3 = 2.75 \times 10^{-13} \Rightarrow s = 6.5 \times 10^{-5} \text{ M.}$$

D. Long answer questions

1. State Le Chatelier's principle. Discuss the effects of concentration, pressure, temperature and catalyst on $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$, $\Delta H = +180 \text{ kJ}$.

If an equilibrium is disturbed, it shifts so as to undo the change. (i) Adding N_2 or $\text{O}_2 \rightarrow$ forward; removing $\text{NO} \rightarrow$ forward. (ii) Pressure: $\Delta n_g = 0$, so no effect. (iii)

Temperature: reaction is endothermic, so higher temperature shifts forward (more NO — happens in lightning and car engines). (iv) Catalyst: no shift; equilibrium is reached sooner.

2. What is a buffer solution? How does an acetate buffer resist change in pH? Calculate the pH of a buffer made of 0.2 M NH_4OH and 0.4 M NH_4Cl ($\text{pK}_b = 4.75$).

Acetate buffer: $\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$. Added H^+ is consumed by CH_3COO^- ($\rightarrow \text{CH}_3\text{COOH}$, weakly ionised); added OH^- is neutralised by CH_3COOH ($\rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$). So pH hardly changes.

$$\text{pOH} = \text{pK}_b + \log([\text{salt}]/[\text{base}]) = 4.75 + \log 2 = 5.05 \Rightarrow \text{pH} = \mathbf{8.95}.$$

QB-7. Redox Reactions

A. Multiple choice questions

1. Oxidation number of S in $\text{Na}_2\text{S}_2\text{O}_3$ (average) is (a) +2 (b) +4 (c) +6 (d) -2

(a) $2 + 2x - 6 = 0$.

2. Oxidation number of Cr in CrO_5 (butterfly structure, four peroxide O) is (a) +10 (b) +6 (c) +5 (d) +3

(b) four O are -1, one O is -2: $x - 4 - 2 = 0$.

3. In $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$, the oxidising agent is (a) Zn (b) Cu^{2+} (c) SO_4^{2-} (d) Cu

(b) it is reduced.

4. Which is a disproportionation reaction? (a) $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ (b) $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$ (c) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ (d) $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

(a) O goes from -1 to -2 and 0.

5. Oxidation number of N in NH_4^+ is (a) +3 (b) -3 (c) +5 (d) -4

(b).

6. Oxidation number of O in OF_2 is (a) -2 (b) +2 (c) -1 (d) +1

(b).

B. Short answer questions

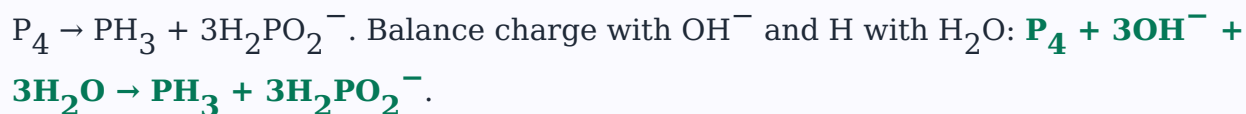
1. Identify the substance oxidised, reduced, oxidising agent and reducing agent:



I^- oxidised ($-1 \rightarrow 0$) — reducing agent. Fe^{3+} reduced ($+3 \rightarrow +2$) — oxidising agent.

2. Balance by oxidation number method: $\text{P}_4 + \text{OH}^- \rightarrow \text{PH}_3 + \text{H}_2\text{PO}_2^-$ (basic medium).

P goes $0 \rightarrow -3$ (PH_3) and $0 \rightarrow +1$ (H_2PO_2^-). For balance 1 P reduced per 3 P oxidised:



3. Balance by ion-electron method: $\text{MnO}_4^- + \text{SO}_2 \rightarrow \text{Mn}^{2+} + \text{HSO}_4^-$ (acidic).

Reduction: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ ($\times 2$). Oxidation: $\text{SO}_2 + 2\text{H}_2\text{O} \rightarrow \text{HSO}_4^- + 3\text{H}^+ + 2\text{e}^-$ ($\times 5$). Add: $2\text{MnO}_4^- + 16\text{H}^+ + 5\text{SO}_2 + 10\text{H}_2\text{O} \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{HSO}_4^- + 15\text{H}^+$. Simplify: **$2\text{MnO}_4^- + 5\text{SO}_2 + 2\text{H}_2\text{O} + \text{H}^+ \rightarrow 2\text{Mn}^{2+} + 5\text{HSO}_4^-$** .

4. Why does H_2O_2 act as both oxidising and reducing agent, whereas HNO_3 acts only as an oxidising agent?

In H_2O_2 , O has -1 , intermediate between -2 and 0 , so it can be reduced or oxidised. In HNO_3 , N has $+5$, its maximum, so it can only be reduced.

5. Give two applications of redox reactions in daily life.

Batteries/cells produce electricity; extraction of metals (Fe_2O_3 reduced by CO); bleaching; respiration and photosynthesis; corrosion and its prevention.

QB-8. Organic Chemistry – Some Basic Principles and Techniques

Key points

- Order of priority of functional groups (for IUPAC): $-\text{COOH} > -\text{SO}_3\text{H} > -\text{COOR} > -\text{COCl} > -\text{CONH}_2 > -\text{CN} > -\text{CHO} > >\text{C}=\text{O} > -\text{OH} > -\text{NH}_2 > \text{C}=\text{C} > \text{C}\equiv\text{C}$.
- Isomers: chain, position, functional, metamerism (structural); geometrical and optical (stereo).
- Carbocation stability: $3^\circ > 2^\circ > 1^\circ > \text{CH}_3^+$; benzyl and allyl cations are extra stable by resonance.
- Electrophiles are Lewis acids; nucleophiles are Lewis bases.
- Lassaigne's test is not given by hydrazine (no carbon) — carbon is needed to form CN^- .

A. Multiple choice questions

1. IUPAC name of $\text{CH}_3\text{CH}_2\text{CHO}$ is (a) propanol (b) propanal (c) propanone (d) ethanal (b).
2. Which method separates a mixture of two liquids with close boiling points? (a) simple distillation (b) fractional distillation (c) sublimation (d) crystallisation (b).
3. Prussian blue in Lassaigne's test indicates (a) S (b) N (c) Cl (d) P (b).
4. Hybridisation of carbon in a carbocation is (a) sp^3 (b) sp^2 (c) sp (d) sp^3d (b).
5. Which is an electrophile? (a) NH_3 (b) OH^- (c) NO_2^+ (d) CN^- (c).
6. Homolytic fission produces (a) ions (b) free radicals (c) carbocations (d) carbanions (b).
7. $-\text{NO}_2$ group shows (a) +I (b) -I (c) +R (d) no effect (b) (also -R).
8. Ethanol and dimethyl ether are (a) chain isomers (b) position isomers (c) functional isomers (d) metamers (c).
9. Most stable carbocation: (a) CH_3^+ (b) CH_3CH_2^+ (c) $(\text{CH}_3)_2\text{CH}^+$ (d) $(\text{CH}_3)_3\text{C}^+$ (d).
10. Steam distillation is used to purify (a) glycerol (b) aniline (c) sugar (d) naphthalene (b).
11. In Kjeldahl's method nitrogen is converted to (a) N_2 (b) NH_3 (c) NaCN (d) NO_2 (b) (as ammonium sulphate, then NH_3).

12. Which is not aromatic? (a) benzene (b) toluene (c) cyclohexane (d) naphthalene
(c) it has no π electrons.

B. Very short answer questions

1. What is catenation?

The ability of carbon atoms to link with one another to form long chains and rings.

2. Write the IUPAC name of $\text{CH}_3\text{-CH=CH-CH}_2\text{-OH}$.

But-2-en-1-ol.

3. Write the IUPAC name of $(\text{CH}_3)_2\text{CH-COOH}$.

2-Methylpropanoic acid.

4. Define a homologous series.

A series of compounds with the same functional group whose successive members differ by a $-\text{CH}_2-$ unit and show gradation in physical properties.

5. Why is sodium fusion done in Lassaigne's test?

To convert covalent N, S and halogens into ionic NaCN , Na_2S and NaX that can be detected in water.

6. Name the technique used to separate plant pigments.

Chromatography.

C. Short answer questions

1. Explain why $(\text{CH}_3)_3\text{C}^+$ is more stable than CH_3CH_2^+ .

Three methyl groups release electrons by +I effect and provide 9 α -hydrogens for hyperconjugation, dispersing the positive charge; ethyl cation has only one methyl group (3 α -H).

2. Differentiate between inductive and electromeric effects.

Inductive: permanent, involves σ electrons, transmitted along the chain, decreases with distance. Electromeric: temporary, involves complete transfer of π electrons, occurs only in presence of an attacking reagent, disappears when the reagent is removed.

3. In Dumas method, 0.3 g of an organic compound gave 50 mL of N_2 collected at 300 K and 715 mm pressure. Find %N (aqueous tension at 300 K = 15 mm).

Dry pressure = 700 mm. Volume at STP = $(700 \times 50 \times 273)/(760 \times 300) = 41.9$ mL.
Mass of $\text{N}_2 = 28 \times 41.9/22400 = 0.0524$ g. %N = $0.0524/0.3 \times 100 = 17.46\%$.

4. In Kjeldahl's estimation, the ammonia from 0.5 g of a compound neutralised 10 mL of 1 M H_2SO_4 . Find %N.

1 M $\text{H}_2\text{SO}_4 = 2$ N. %N = $1.4 \times N \times V/w = 1.4 \times 2 \times 10/0.5 = 56\%$.

5. 0.468 g of a compound gave 0.668 g BaSO₄ in Carius estimation. Find %S.

$$\%S = (32/233)(0.668/0.468) \times 100 = \mathbf{19.6\%}.$$

6. Classify as nucleophile or electrophile: HS⁻, BF₃, C₂H₅O⁻, (CH₃)₃N:, Cl⁺, CH₃C⁺=O, H₂N⁻, NO₂⁺.

Nucleophiles: HS⁻, C₂H₅O⁻, (CH₃)₃N:, H₂N⁻. Electrophiles: BF₃, Cl⁺, CH₃C⁺=O, NO₂⁺.

D. Long answer questions

1. Describe the principle and method of chromatography. Name its types.

Principle: components of a mixture distribute differently between a stationary phase and a mobile phase, so they move at different rates and get separated. Method (column chromatography): a column is packed with adsorbent (alumina/silica gel); the mixture is placed at the top and a solvent (eluent) is passed; the less strongly adsorbed component moves faster and comes out first. Types: adsorption chromatography (column, thin-layer), partition chromatography (paper). In TLC and paper chromatography, R_f = distance moved by substance/distance moved by solvent.

2. Explain resonance and hyperconjugation with suitable examples. How do they affect stability?

Resonance: delocalisation of π electrons/lone pairs over a conjugated system, shown by canonical structures (benzene, carboxylate ion, phenoxide ion); resonance energy stabilises the molecule. Hyperconjugation: delocalisation of σ electrons of an α C-H bond into an adjacent empty p orbital or π bond (in carbocations, alkenes, toluene); more α -H means more stability. Thus but-2-ene is more stable than but-1-ene, and 3° carbocation is more stable than 1°.

QB-9. Hydrocarbons

Key points

- Alkanes: substitution (free radical); alkenes and alkynes: addition (electrophilic); benzene: electrophilic substitution.
- Markovnikov's rule applies to unsymmetrical alkenes; peroxide effect only with HBr.
- Terminal alkynes are acidic (sp carbon); they react with Na, ammoniacal AgNO₃ and Cu₂Cl₂.
- Aromatic: cyclic, planar, conjugated, $(4n + 2)$ π electrons.
- Ortho/para directors: -OH, -NH₂, -CH₃, -X. Meta directors: -NO₂, -CHO, -COOH, -CN, -SO₃H.

A. Multiple choice questions

1. The general formula of alkynes is (a) C_nH_{2n+2} (b) C_nH_{2n} (c) C_nH_{2n-2} (d) C_nH_n
(c).
2. Which conformation of ethane is most stable? (a) eclipsed (b) staggered (c) skew (d) all equal
(b).
3. Which shows geometrical isomerism? (a) propene (b) but-1-ene (c) but-2-ene (d) ethene
(c).
4. The product of HBr addition to propene in presence of peroxide is (a) 1-bromopropane (b) 2-bromopropane (c) 1,2-dibromopropane (d) propane
(a).
5. Ozonolysis of ethene gives (a) ethanal (b) methanal (c) propanone (d) ethanol
(b) two molecules of HCHO.
6. Which is most acidic? (a) ethane (b) ethene (c) ethyne (d) benzene
(c).
7. Hydration of ethyne ($HgSO_4/H_2SO_4$) gives (a) ethanol (b) ethanal (c) acetone (d) acetic acid
(b).
8. Number of π electrons in benzene is (a) 2 (b) 4 (c) 6 (d) 12
(c).
9. The electrophile in nitration is (a) NO_2^- (b) NO_2^+ (c) NO^+ (d) HNO_3
(b).
10. Friedel-Crafts alkylation requires (a) anhydrous $AlCl_3$ (b) Ni (c) $Pd/BaSO_4$ (d) Na/ether
(a).
11. Which group is meta-directing? (a) $-OH$ (b) $-CH_3$ (c) $-NO_2$ (d) $-Cl$
(c).
12. Wurtz reaction of CH_3Br with Na gives (a) methane (b) ethane (c) propane (d) ethene
(b).
13. Lindlar's catalyst converts alkynes to (a) trans-alkenes (b) cis-alkenes (c) alkanes (d) aldehydes
(b).
14. Baeyer's reagent is (a) alkaline $KMnO_4$ (cold, dilute) (b) acidic $K_2Cr_2O_7$ (c) bromine water (d) ammoniacal $AgNO_3$
(a).

B. Very short answer questions

1. Write the IUPAC name of $CH_3-C\equiv C-CH(CH_3)_2$.
4-Methylpent-2-yne.

2. How many structural isomers does pentane have? Name them.

Three: n-pentane, 2-methylbutane, 2,2-dimethylpropane.

3. Why does branching lower the boiling point of alkanes?

Branched molecules are more compact, have less surface area and weaker van der Waals forces.

4. What is the Kharasch effect?

Anti-Markovnikov addition of HBr to unsymmetrical alkenes in presence of peroxide.

5. Name a test to distinguish ethene from ethyne.

Ethyne gives a white ppt with ammoniacal AgNO_3 ; ethene does not.

C. Short answer questions**1. How will you convert: (i) ethanoic acid to methane (ii) ethyne to ethene (iii) benzene to nitrobenzene?**

(i) $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa}$; then soda lime, heat $\rightarrow \text{CH}_4$. (ii) H_2 , Lindlar's catalyst. (iii) Conc. HNO_3 + conc. H_2SO_4 , 323–333 K.

2. Write the mechanism of chlorination of methane.

Initiation: $\text{Cl}_2 \rightarrow 2\text{Cl}\cdot$ (light). Propagation: $\text{CH}_4 + \text{Cl}\cdot \rightarrow \text{CH}_3\cdot + \text{HCl}$; $\text{CH}_3\cdot + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \text{Cl}\cdot$. Termination: combination of any two radicals.

3. An alkene "A" on ozonolysis gives a mixture of ethanal and pentan-3-one. Write the structure and IUPAC name of A.

Join $\text{CH}_3\text{CH}=\text{}$ with $=\text{C}(\text{C}_2\text{H}_5)_2$: $\text{CH}_3\text{CH}=\text{C}(\text{C}_2\text{H}_5)_2$, **3-ethylpent-2-ene**.

4. Why is benzene extraordinarily stable though it contains three double bonds?

The six π electrons are delocalised over the whole ring (resonance). The resonance energy (about 150 kJ/mol) makes benzene more stable than a hypothetical cyclohexatriene, so it resists addition reactions.

5. Explain why -OH group is o/p directing while $-\text{NO}_2$ is m-directing.

-OH donates its lone pair into the ring by resonance (+R), increasing electron density at ortho and para positions. $-\text{NO}_2$ withdraws electrons (-R, -I), decreasing density mainly at ortho and para; the meta position is relatively richer, so electrophiles attack there.

6. What are the carcinogenic hydrocarbons? Give examples.

Polynuclear aromatic hydrocarbons with more than two fused benzene rings, formed by incomplete combustion of coal, tobacco and petroleum, e.g. 1,2-benzanthracene, 3-methylcholanthrene, benzopyrene. They damage DNA and can cause cancer.

D. Long answer questions

1. Describe the preparation and four chemical properties of ethyne.

Preparation: $\text{CaC}_2 + 2\text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_2 + \text{Ca(OH)}_2$; or from 1,2-dibromoethane with alcoholic KOH followed by NaNH_2 . Properties: (1) acidic: with Na $\rightarrow$ sodium acetylide + H_2 ; with ammoniacal $\text{AgNO}_3 \rightarrow$ white silver acetylide. (2) $\text{H}_2/\text{Ni} \rightarrow$ ethane; $\text{H}_2/\text{Lindlar} \rightarrow$ ethene. (3) $\text{Br}_2 \rightarrow$ 1,1,2,2-tetrabromoethane (decolourises bromine water). (4) $\text{HBr} \rightarrow \text{CH}_3\text{CHBr}_2$ (Markovnikov, two steps). (5) $\text{H}_2\text{O}/\text{HgSO}_4, \text{H}_2\text{SO}_4 \rightarrow$ ethanal. (6) Cyclic polymerisation at 873 K $\rightarrow$ benzene.

2. Explain the mechanism of electrophilic substitution in benzene taking Friedel-Crafts alkylation as an example.

Step 1 - electrophile: $\text{CH}_3\text{Cl} + \text{AlCl}_3 \rightarrow \text{CH}_3^+ + \text{AlCl}_4^-$. Step 2 - the π electrons of benzene attack CH_3^+ forming a resonance-stabilised arenium ion (σ -complex), in which one carbon is sp^3 and aromaticity is lost. Step 3 - AlCl_4^- removes H^+ from this carbon: aromaticity is restored, giving toluene, HCl and regenerating AlCl_3 .

SAMPLE PAPERS WITH SOLUTIONS

Chemistry · +2 1st Year · Theory 70 marks · 3 hours

Sample Paper 1

Group A - Objective (1 × 10)

1. Number of molecules in 18 g of water.
2. The quantum number that determines the shape of an orbital.
3. The element with the highest electron gain enthalpy (most negative).
4. Shape of the BF_3 molecule.
5. Sign of ΔS when water freezes.
6. pH of 0.01 M HCl.
7. Oxidation number of Mn in MnO_2 .
8. IUPAC name of CH_3COCH_3 .
9. The product of ozonolysis of 2-butene.
10. Number of σ bonds in benzene.

Answers: 1. 6.022×10^{23} 2. Azimuthal (l) 3. Cl 4. Trigonal planar 5. Negative 6. 2 7. +4 8. Propanone 9. Ethanal (2 molecules) 10. 12 (6 C-C + 6 C-H).

Group B - Short (2 × 8)

11. State the law of multiple proportions with an example.

When two elements combine to form two or more compounds, the masses of one element that combine with a fixed mass of the other bear a simple whole-number ratio. CO and CO_2 : 12 g C combines with 16 g and 32 g O, ratio 1 : 2.

12. Write the electronic configurations of Cr (24) and Cu (29).

Cr: $[\text{Ar}]3d^5 4s^1$; Cu: $[\text{Ar}]3d^{10} 4s^1$.

13. Why does the atomic radius decrease across a period?

Nuclear charge increases by one at each step while electrons are added to the same shell; the increased attraction pulls the electron cloud inward.

14. Explain sp^2 hybridisation with an example.

One s and two p orbitals mix to give three equivalent sp^2 orbitals at 120° in a plane. BF_3 : B ($2s^2 2p^1 \rightarrow 2s^1 2p^2$) uses three sp^2 orbitals to bond with three F atoms — trigonal planar.

15. Define enthalpy of formation and give an example.

Enthalpy change when one mole of a compound is formed from its elements in their standard states. $C(\text{graphite}) + O_2(g) \rightarrow CO_2(g)$, $\Delta_f H^\circ = -393.5 \text{ kJ/mol}$.

16. What is a buffer solution? Give one example.

A solution that resists change in pH on adding small amounts of acid or base or on dilution; e.g. $CH_3COOH + CH_3COONa$.

17. What is hyperconjugation?

Delocalisation of σ electrons of a C-H bond on a carbon next to a carbocation or a double bond into the empty p orbital/ π system ("no-bond resonance").

18. State Markovnikov's rule.

In the addition of HX to an unsymmetrical alkene, the negative part of the reagent attaches to the carbon of the double bond that has fewer hydrogen atoms.

Group C - Short (3 × 8)**19. Calculate the molarity of a solution of 0.25 mol NaOH in 250 mL. Also find the mass of NaOH needed.**

$M = 0.25/0.25 = 1 \text{ M}$. Mass = $0.25 \times 40 = 10 \text{ g}$.

20. Calculate the wavelength of an electron moving with velocity $3 \times 10^6 \text{ m/s}$.

$\lambda = 6.626 \times 10^{-34} / (9.1 \times 10^{-31} \times 3 \times 10^6) = 2.43 \times 10^{-10} \text{ m} = 2.43 \text{ \AA}$.

21. Explain the variation of electronegativity in the periodic table.

Increases across a period (smaller size, higher nuclear charge), decreases down a group (larger size, more shielding). F is highest (4.0), Cs/Fr are lowest.

22. Draw the MO diagram of O_2 and explain its paramagnetism.

See QB-4 Long answer 1: two unpaired electrons in $\pi^* 2p_x$ and $\pi^* 2p_y$; B.O. = 2; paramagnetic.

23. For a reaction $\Delta H = -10 \text{ kJ}$ and $\Delta S = -20 \text{ J/K}$. At what temperature will it stop being spontaneous?

$\Delta G = 0$ at $T = \Delta H/\Delta S = 10000/20 = 500 \text{ K}$. Spontaneous below 500 K, non-spontaneous above.

24. Write the expression of K_p and K_c for $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$ and the relation between them.

$K_c = [\text{PCl}_3][\text{Cl}_2]/[\text{PCl}_5]$; $K_p = p_{\text{PCl}_3}p_{\text{Cl}_2}/p_{\text{PCl}_5}$; $\Delta n = 1 \Rightarrow K_p = K_c RT$.

25. Balance: $\text{Cu} + \text{HNO}_3 (\text{dilute}) \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO} + \text{H}_2\text{O}$.

Cu: $0 \rightarrow +2$ (loses 2e); N: $+5 \rightarrow +2$ (gains 3e). Multiply Cu by 3 and NO by 2: $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$.

26. How will you convert benzene into (i) toluene (ii) acetophenone (iii) chlorobenzene?

(i) CH_3Cl , anhyd. AlCl_3 (ii) CH_3COCl , anhyd. AlCl_3 (iii) Cl_2 , anhyd. FeCl_3 .

Group D - Long (5 × 4)

27. Discuss Bohr's model of the hydrogen atom. What are its limitations? OR Discuss the quantum numbers and the rules for filling electrons in orbitals.

Postulates (stationary orbits, $mvr = nh/2\pi$, energy emitted/absorbed in jumps, $E_n = -13.6/n^2 \text{ eV}$) and its explanation of the hydrogen line spectrum. Limitations: fails for multi-electron atoms, cannot explain fine structure, Zeeman and Stark effects, ignores dual nature and uncertainty principle. OR: four quantum numbers + Aufbau, Pauli and Hund's rules (see notes).

28. Explain VSEPR theory and predict the shapes of CH_4 , NH_3 and H_2O .

Electron pairs around the central atom repel and stay as far apart as possible; $\text{lp-lp} > \text{lp-bp} > \text{bp-bp}$. CH_4 : 4 bp $\rightarrow$ tetrahedral 109.5° . NH_3 : 3 bp + 1 lp $\rightarrow$ pyramidal 107° . H_2O : 2 bp + 2 lp $\rightarrow$ bent 104.5° .

29. Explain the first law of thermodynamics and derive $\Delta H = \Delta U + \Delta n_g RT$.

$H = U + PV \Rightarrow \Delta H = \Delta U + P\Delta V$ at constant P. For ideal gases $PV = nRT \Rightarrow P\Delta V = \Delta n_g RT$. Hence $\Delta H = \Delta U + \Delta n_g RT$.

30. Describe the mechanism of nitration of benzene. OR Write the methods of preparation and three chemical properties of ethene.

Nitration: $\text{HNO}_3 + 2\text{H}_2\text{SO}_4 \rightarrow \text{NO}_2^+ + \text{H}_3\text{O}^+ + 2\text{HSO}_4^-$; attack gives arenium ion; loss of H^+ gives nitrobenzene. OR: from ethanol (conc. H_2SO_4 , 443 K), ethyl bromide (alc. KOH), ethyne (Lindlar). Properties: addition of H_2 , Br_2 , HBr ; oxidation with Baeyer's reagent $\rightarrow$ ethylene glycol; ozonolysis $\rightarrow$ HCHO ; polymerisation $\rightarrow$ polythene.

Sample Paper 2

Group A - Objective (1 × 10)

1. Mass of 0.1 mol of CaCO_3 .
2. Maximum number of electrons in a d subshell.
3. Name the scientist who proposed the modern periodic law.
4. Bond order of N_2 .
5. The unit of entropy.
6. The conjugate acid of NH_3 .
7. Oxidation number of Fe in Fe_3O_4 .
8. The reagent used in Lassaigne's test.
9. The type of isomerism shown by but-2-ene.
10. The hybridisation of carbon in benzene.

Answers: 1. 10 g 2. 10 3. Moseley 4. 3 5. $\text{J K}^{-1} \text{mol}^{-1}$ 6. NH_4^+ 7. +8/3 (average) 8. Sodium metal 9. Geometrical (cis-trans) 10. sp^2 .

Group B - Short (2 × 8)

11. Distinguish between molarity and molality.

Molarity = moles of solute per litre of solution (temperature dependent). Molality = moles of solute per kg of solvent (temperature independent).

12. State Heisenberg's uncertainty principle.

It is impossible to determine simultaneously the exact position and exact momentum of a microscopic particle like an electron: $\Delta x \cdot \Delta p \geq h/4\pi$.

13. Give the IUPAC name and symbol of the element with Z = 107.

Unnilseptium, Uns (now named bohrium, Bh).

14. Why is the bond angle of H_2O less than that of NH_3 ?

H_2O has two lone pairs and NH_3 has one; greater lp-lp and lp-bp repulsion in water compresses the bond angle more (104.5° vs 107°).

15. State the third law of thermodynamics.

The entropy of a perfectly crystalline substance at absolute zero is zero.

16. What is the effect of temperature on the equilibrium constant of an exothermic reaction?

K decreases with increase in temperature (equilibrium shifts backward).

17. What is a free radical? Give an example.

An atom or group with an unpaired electron, formed by homolytic fission, e.g. $\text{CH}_3^\bullet$, $\text{Cl}^\bullet$.

18. Why are alkynes acidic?

The sp -hybridised carbon (50% s character) is more electronegative and holds the C-H bonding pair closer, so H^+ is released easily.

Group C - Short (3 × 8)**19. A compound has 4.07% H, 24.27% C, 71.65% Cl; molar mass 98.96 g. Find its molecular formula.**

$\text{C}_2\text{H}_4\text{Cl}_2$ (worked in chapter notes).

20. Write the four quantum numbers of the 19th electron of potassium.

$4s^1$: $n = 4$, $l = 0$, $m = 0$, $s = +\frac{1}{2}$.

21. Explain the anomalous behaviour of IE_1 of N and O.

N ($2p^3$) is half-filled and stable; O ($2p^4$) has one paired electron whose removal reduces repulsion and gives the stable half-filled $2p^3$. So $\text{IE}_1(\text{N}) > \text{IE}_1(\text{O})$.

22. Compare σ and π bonds (any three points).

σ : axial overlap, stronger, free rotation possible. π : sideways overlap, weaker, rotation restricted; π forms only after σ .

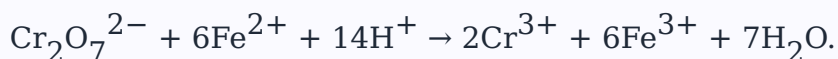
23. The enthalpies of formation of CO_2 , H_2O and $\text{C}_2\text{H}_5\text{OH}$ are -393.5 , -285.8 and -277.7 kJ/mol. Find the enthalpy of combustion of ethanol.

$\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$. $\Delta H = [2(-393.5) + 3(-285.8)] - (-277.7) = (-787 - 857.4) + 277.7 = -1366.7$ kJ/mol.

24. Calculate the pH of 0.1 M NH_4OH ($K_b = 1.8 \times 10^{-5}$).

$[\text{OH}^-] = \sqrt{(1.8 \times 10^{-6})} = 1.34 \times 10^{-3} \Rightarrow \text{pOH} = 2.87 \Rightarrow \text{pH} = 11.13$.

25. Balance in acidic medium: $\text{Cr}_2\text{O}_7^{2-} + \text{Fe}^{2+} \rightarrow \text{Cr}^{3+} + \text{Fe}^{3+}$.



26. Draw the Newman projections of the staggered and eclipsed conformations of ethane.

Look along the C-C bond: front carbon shown as a point with three H at 120° , back carbon as a circle with three H. Staggered: back H's lie exactly between front H's (60° apart) — most stable. Eclipsed: back H's directly behind front H's — least stable (torsional strain ≈ 12.5 kJ/mol).

Group D - Long (5 × 4)

27. Describe Rutherford's α -scattering experiment. What are the drawbacks of his model?

See chapter Q&A 6. Drawbacks: revolving electron should radiate energy and fall into the nucleus (instability); does not explain line spectra or distribution of electrons.

28. What is resonance? Explain with O_3 and CO_3^{2-} .

When a molecule can be represented by two or more structures differing only in electron positions, the real structure is a resonance hybrid. O_3 : two structures $\text{O}=\text{O}-\text{O} \leftrightarrow \text{O}-\text{O}=\text{O}$; both O-O bonds are equal (128 pm). CO_3^{2-} : three equivalent structures; all C-O bonds equal (129 pm). Resonance lowers energy and increases stability.

29. Explain the ionisation of weak acids and derive Ostwald's dilution law.

$\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$. Initial C, 0, 0; at equilibrium $C(1 - \alpha)$, $C\alpha$, $C\alpha$. $K_a = C\alpha^2/(1 - \alpha)$. For weak acids $\alpha \ll 1$: $K_a = C\alpha^2 \Rightarrow \alpha = \sqrt{K_a/C}$ — degree of ionisation increases on dilution.

30. Write a note on the inductive effect and the electromeric effect with examples.

See QB-8 Short answer 2 and chapter notes; include examples $\text{CH}_3\text{-CH}_2\text{-Cl}$ ($-I$) and addition to $\text{C}=\text{O}$ ($+E/-E$).

OBJECTIVE BANK & PRACTICAL GUIDE

Chemistry · +2 1st Year

100 One-Mark Questions with Answers

#	Question	Answer
1	SI unit of amount of substance	mole
2	Molar volume of a gas at STP	22.4 L
3	Mass of 1 mol of electrons	about 0.55 mg ($9.1 \times 10^{-31} \times 6.022 \times 10^{23}$ kg)
4	Law stating mass is conserved	Law of conservation of mass (Lavoisier)
5	Empirical formula of glucose	CH ₂ O
6	Number of significant figures in 2.050	4
7	Equivalent mass of H ₃ PO ₄ (basicity 3)	98/3 = 32.67
8	Molality is expressed in	mol kg ⁻¹
9	e/m ratio was measured by	J.J. Thomson
10	Charge on an electron was measured by	R.A. Millikan
11	Particle with no charge	Neutron
12	³ H is called	Tritium
13	Lyman series lies in	Ultraviolet region
14	Value of Planck's constant	6.626×10^{-34} J s
15	Shape of s-orbital	Spherical
16	Number of orbitals in n = 3	9
17	Electronic configuration of Na ⁺	1s ² 2s ² 2p ⁶
18	Which is larger, 4s or 3d energy (in neutral K)?	3d is higher (4s fills first)
19	Number of groups in the modern periodic table	18
20	Block of element Z = 26	d-block
21	Lanthanoids belong to period	6
22	The most metallic element (stable) of group 1	Cs
23	Element with lowest ionisation enthalpy (stable)	Cs
24	Halogen that is a liquid at room temperature	Bromine
25	Diagonal partner of Li	Mg
26	Bond in NaCl	Ionic
27	Bond angle in BeCl ₂	180°
28	Hybridisation of S in SF ₆	sp ³ d ²
29	Unit of dipole moment	Debye (D)
30	Shape of XeF ₂	Linear
31	Bond order of He ₂ ⁺	0.5
32	Strongest hydrogen bond is in	HF
33	Number of lone pairs on Xe in XeF ₄	2

34	Molecule with a coordinate bond	NH_4^+ , H_3O^+ , CO
35	Example of an open system	Water boiling in an open beaker
36	Heat exchanged at constant pressure equals	ΔH
37	Heat exchanged at constant volume equals	ΔU
38	Device to measure ΔU	Bomb calorimeter
39	For an ideal gas $C_p - C_v =$	R
40	Spontaneity criterion	$\Delta G < 0$
41	Entropy is a measure of	Disorder/randomness
42	ΔG° and K are related by	$\Delta G^\circ = -2.303RT \log K$
43	Equilibrium is dynamic because	Forward and backward reactions continue at equal rates
44	K for reverse reaction	1/K
45	Haber process catalyst	Iron (with Mo promoter)
46	pH of blood	About 7.4
47	pH + pOH at 25 °C	14
48	Lewis base example	NH_3
49	Salt of strong acid and strong base is	Neutral
50	Henderson equation for acid buffer	$\text{pH} = \text{pK}_a + \log[\text{salt}]/[\text{acid}]$
51	K_{sp} of AgCl in terms of s	s^2
52	The ion common in NH_4OH and NH_4Cl	NH_4^+
53	Oxidation means (electronic concept)	Loss of electrons
54	O.N. of H in NaH	-1
55	O.N. of Cl in HClO_4	+7
56	O.N. of C in CO_2	+4
57	Strongest oxidising agent among halogens	F_2
58	In a redox reaction, the reducing agent	gets oxidised
59	Reaction where the same element is oxidised and reduced	Disproportionation
60	Father of organic chemistry (synthesised urea)	F. Wöhler
61	Purification of camphor	Sublimation
62	Separation of glycerol from spent lye	Distillation under reduced pressure
63	Test for halogens in organic compounds (flame)	Beilstein's test (green flame with Cu wire)
64	Method to estimate halogens	Carius method
65	IUPAC name of $\text{CH}_3\text{CH}(\text{OH})\text{CH}_3$	Propan-2-ol
66	IUPAC name of HCOOH	Methanoic acid

67	Functional group of esters	-COOR
68	Hybridisation of carbanion carbon	sp^3
69	Type of reaction: $CH_2=CH_2 + Br_2$	Addition
70	$-CH_3$ group shows which inductive effect	+I
71	Isomers differing in position of functional group	Position isomers
72	Main constituent of LPG	Butane
73	Main constituent of CNG/natural gas	Methane
74	Product of decarboxylation of sodium propanoate	Ethane
75	C-C bond length in ethane	154 pm
76	C=C bond length	134 pm
77	C≡C bond length	120 pm
78	C-C bond length in benzene	139 pm
79	Cis or trans: which has higher dipole moment?	Cis
80	Reagent to convert alkyl halide to alkene	Alcoholic KOH
81	Product of HBr + propene (no peroxide)	2-Bromopropane
82	Gas from calcium carbide and water	Ethyne (acetylene)
83	Red-hot iron tube converts ethyne into	Benzene
84	Hückel's rule	$(4n + 2) \pi$ electrons
85	Electrophile in sulphonation	SO_3
86	Electrophile in chlorination of benzene	Cl^+
87	Product of Friedel-Crafts acylation of benzene with CH_3COCl	Acetophenone
88	-Cl on benzene is	Deactivating, o/p directing
89	Benzene reacts with Cl_2 in UV light to give	Benzene hexachloride (addition)
90	Polymer of ethene	Polythene
91	Colour of copper sulphate crystals	Blue
92	Indicator for strong acid vs strong base	Phenolphthalein or methyl orange
93	Gas turning lime water milky	CO_2
94	Gas with rotten-egg smell	H_2S
95	Brown ring test detects	Nitrate ion
96	Chromyl chloride test detects	Chloride ion
97	Colour of AgI precipitate	Yellow
98	Reagent for sulphate ion	$BaCl_2$ (white ppt insoluble in HCl)
99	Ammonium molybdate test detects	Phosphate
100		0.2 N

Normality of 0.1 M H ₂ SO ₄

Practical Guide – Class XI

1. Basic laboratory techniques (non-evaluative)

- **Bunsen burner:** parts — base, tube (barrel), air holes with collar, gas nipple. With air holes open, a non-luminous blue flame forms; it has an inner cone (unburnt gas), middle zone (reducing zone, partially burnt) and outer zone (oxidising, hottest). With air holes closed, a luminous yellow, sooty flame forms.
- **Chemical balance:** weighing by the method of equal oscillations — the pointer swings equally on both sides of the zero when weights balance. Never weigh hot objects or chemicals directly on the pan.
- **Glass work:** cut a glass tube by making a scratch with a file and snapping; bend by heating evenly in a flat (fish-tail) flame; draw a jet by heating the middle and pulling; bore a cork with a cork borer of slightly smaller diameter.

2. Crystallisation - CuSO₄·5H₂O from CuCO₃

Principle: $\text{CuCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O} + \text{CO}_2 \uparrow$. The solution is concentrated and cooled to give blue crystals.

Procedure: Add dilute H₂SO₄ to a little warm water; add CuCO₃ in small portions with stirring until effervescence stops (slight excess of carbonate). Filter. Heat the filtrate in a china dish until the crystallisation point (a drop on a glass rod forms crystals on cooling). Cover and cool slowly. Decant the mother liquor, dry crystals between filter papers, record yield.

Precautions: add carbonate slowly; do not over-heat (anhydrous white CuSO₄ may form); cool slowly for large crystals.

3. Volumetric analysis - single titration

Step	What to do
Rinse	Burette with the solution it will hold; pipette with the solution to be pipetted; conical flask with distilled water only.
Fill	Fill burette, remove air bubble from the nozzle, note the initial reading (lower meniscus for colourless solutions).
Pipette	Take 10/20 mL of the other solution into the flask, add 1-2 drops of indicator.
Titrate	Add from burette with swirling till the end point (phenolphthalein: colourless → light pink with base in burette).
Repeat	Take at least three concordant readings (differing by ≤ 0.1 mL).
Calculate	$N_1 V_1 = N_2 V_2$; strength (g/L) = $N \times \text{equivalent mass}$.

Model calculation: 10 mL of oxalic acid (N/10) needs 12.5 mL of NaOH. Find the normality and strength of NaOH.

$$N(\text{NaOH}) = 0.1 \times 10/12.5 = 0.08 \text{ N. Strength} = 0.08 \times 40 = 3.2 \text{ g/L.}$$

4. Equivalent mass of Mg by hydrogen displacement

$\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. A known mass of Mg ribbon reacts with dilute HCl in an eudiometer; the volume of H_2 is measured and reduced to STP (correcting for aqueous tension). Equivalent mass = (mass of Mg $\times$ 11200)/(volume of H_2 at STP in mL).

Example: 0.03 g Mg gave 28 mL of H_2 at STP. Find the equivalent mass.

$$E = 0.03 \times 11200/28 = 12.$$

5. Solubility of K_2SO_4 at room temperature

Prepare a saturated solution, weigh an empty china dish (W_1), add some clear saturated solution and weigh (W_2), evaporate to dryness and weigh (W_3). Mass of salt = $W_3 - W_1$; mass of water = $W_2 - W_3$. Solubility = $[(W_3 - W_1)/(W_2 - W_3)] \times 100$ g per 100 g water.

6. Salt analysis - dry tests for basic radicals

Test	Observation	Inference
Colour of salt	Blue / green / pink / light green	Cu^{2+} / Ni^{2+} (or Fe^{2+}) / Co^{2+} or Mn^{2+} / Fe^{2+}
Heating in dry tube	Sublimate white	NH_4^+ salts, HgCl_2
Flame test	Golden yellow / lilac / brick red / crimson / apple green	Na^+ / K^+ / Ca^{2+} / Sr^{2+} / Ba^{2+}
Borax bead test (oxidising flame)	Blue / brown (hot) / green	Co^{2+} / Ni^{2+} / Cr^{3+}
Charcoal cavity + Na_2CO_3	Residue yellow when hot, white when cold	Zn^{2+}
Cobalt nitrate test (on residue)	Blue / green / pink	Al^{3+} / Zn^{2+} / Mg^{2+}

Viva-voce questions with answers

1. Why is a burette not rinsed with water just before use?

Water would dilute the solution and change its concentration.

2. Why is the conical flask not rinsed with the solution?

It would add extra solute beyond the measured volume.

3. Why does phenolphthalein turn pink in alkali?

It is a weak acid indicator; in basic solution it forms an ion with a quinonoid structure which is pink.

4. What is a standard solution?

A solution of exactly known concentration.

5. What are concordant readings?

Successive titre values that agree within 0.1 mL.

6. Why is dilute H_2SO_4 preferred over HCl in preparing CuSO_4 ?

To get the sulphate salt; HCl would give CuCl_2 .

7. What is the mother liquor?

The solution left after crystals separate.

8. Why is the flame test done with conc. HCl ?

It converts salts into volatile chlorides which give colour to the flame easily.

9. Why is a platinum wire used in the flame test?

Platinum does not react and does not impart its own colour to the flame.

10. What is the meniscus? How is it read?

The curved surface of liquid in a tube; read the lower meniscus for colourless solutions and the upper for coloured ones (like KMnO_4), at eye level.

11. Why is sodium carbonate extract prepared for acid radicals?

To remove basic radicals as insoluble carbonates and bring acid radicals into solution as sodium salts.

12. What is the role of conc. H_2SO_4 in the brown ring test?

It forms a dense lower layer and reacts with nitrate to produce NO , which forms $[\text{Fe}(\text{H}_2\text{O})_5\text{NO}]^{2+}$ (brown ring) at the junction.

REASONING QUESTIONS & NUMERICAL PRACTICE

Chemistry · +2 1st Year

"Give Reasons" Questions (Very Important)

1. Molality is preferred over molarity for expressing concentration.

Molality uses mass of solvent, which does not change with temperature; molarity uses volume, which expands or contracts.

2. The atomic mass of chlorine is 35.5 and not a whole number.

It is the weighted average of isotopes ^{35}Cl (75%) and ^{37}Cl (25%): $0.75 \times 35 + 0.25 \times 37 = 35.5$.

3. Bohr's model could not explain the spectra of multi-electron atoms.

It neglects electron-electron repulsion and the existence of subshells; it works only for one-electron species.

4. 4s orbital is filled before 3d.

4s has lower $(n + l) = 4$ than 3d (5), so lower energy.

5. Electrons are not found in the nucleus.

By the uncertainty principle, an electron confined to nuclear dimensions (10^{-15} m) would have an enormous uncertainty in velocity, far greater than the speed of light.

6. Noble gases have large atomic radii.

Their radii are van der Waals radii (non-bonded), which are larger than covalent radii.

7. Electron gain enthalpy of F is less negative than that of Cl.

F is very small; the incoming electron faces strong repulsion from the compact 2p electrons.

8. Be and Mg do not impart colour to the flame.

Their electrons are strongly bound; the energy of the flame is not enough to excite them to higher levels (or the emitted light is outside the visible region).

9. NaCl is ionic but AlCl_3 is covalent.

Al^{3+} is small and highly charged; it polarises Cl^- strongly (Fajans' rules), giving covalent character.

10. CO_2 is a gas but SiO_2 is a solid.

CO_2 consists of discrete small molecules (C forms p π -p π double bonds); Si cannot form such π bonds, so SiO_2 forms a giant 3-D network of Si-O single bonds.

11. H_2O is a liquid while H_2S is a gas.

Strong intermolecular hydrogen bonding in water.

12. Ice floats on water.

In ice, H-bonds hold molecules in an open cage-like structure with empty spaces, so its density is lower than liquid water.

13. NF_3 has a smaller dipole moment than NH_3 .

In NF_3 the lone-pair moment opposes the N-F bond moments; in NH_3 they act in the same direction.

14. σ bond is stronger than π bond.

Head-on overlap is greater than sideways overlap.

15. N_2 is chemically inert.

It has a very strong triple bond (bond order 3, bond enthalpy 946 kJ/mol).

16. Heat and work are not state functions.

Their values depend on the path followed between initial and final states.

17. The enthalpy of neutralisation of strong acids with strong bases is constant (-57.1 kJ/mol).

All such reactions are essentially $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$, as the acids and bases are completely ionised.

18. Enthalpy of neutralisation of acetic acid with NaOH is less than 57.1 kJ/mol.

Part of the energy released is used to ionise the weak acid.

19. Melting of ice is spontaneous above 0°C even though it is endothermic.

ΔS is positive and $T\Delta S > \Delta H$ above 273 K, making ΔG negative.

20. A catalyst does not change the equilibrium constant.

It speeds up the forward and backward reactions equally.

21. Aqueous solution of Na_2CO_3 is alkaline.

CO_3^{2-} (from weak H_2CO_3) hydrolyses: $\text{CO}_3^{2-} + \text{H}_2\text{O} \rightleftharpoons \text{HCO}_3^- + \text{OH}^-$.

22. Pure NaCl is precipitated when HCl gas is passed into a saturated NaCl solution.

Common ion Cl^- increases the ionic product $[\text{Na}^+][\text{Cl}^-]$ above K_{sp} .

23. HCl is added before passing H_2S in Group II.

H^+ suppresses ionisation of H_2S , keeping $[\text{S}^{2-}]$ low so that only Group II sulphides (very low K_{sp}) precipitate.

24. F₂ is the strongest oxidising agent.

It has the highest electrode potential due to low bond dissociation energy and high hydration enthalpy of F⁻.

25. Carbon forms a large number of compounds.

Catenation, tetravalency, ability to form multiple bonds and strong bonds with many elements.

26. Lassaigne's test is not shown by diazonium salts.

They lose N₂ on heating before reacting with sodium, so NaCN is not formed.

27. Tertiary carbocations are more stable than primary ones.

+I effect and hyperconjugation of more alkyl groups.

28. Alkanes undergo substitution while alkenes undergo addition reactions.

Alkanes are saturated (only σ bonds); alkenes have a reactive π bond which electrophiles attack.

29. Ethyne is more acidic than ethene.

sp carbon (50% s) is more electronegative than sp² carbon (33% s).

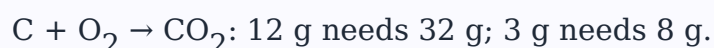
30. Benzene undergoes electrophilic substitution rather than addition.

To retain the resonance-stabilised aromatic sextet.

Numerical Practice with Solutions

1. What is the mass of 3.011×10^{23} molecules of NH₃?

$$n = 0.5 \text{ mol} \times 17 \text{ g} = \mathbf{8.5 \text{ g}}.$$

2. How many grams of oxygen are needed to burn 3 g of carbon completely?**3. Calculate the number of electrons in 1.6 g of CH₄.**

$$0.1 \text{ mol CH}_4; \text{ each has } 10 \text{ electrons} \Rightarrow 1 \text{ mol electrons} = 6.022 \times 10^{23}.$$

4. 20 mL of 0.5 M HCl is diluted to 500 mL. Find the new molarity.

$$M_2 = 0.5 \times 20/500 = 0.02 \text{ M}.$$

5. Find the energy of one mole of photons of wavelength 600 nm.

$$E \text{ per photon} = hc/\lambda = 3.31 \times 10^{-19} \text{ J}; \text{ per mole} \times 6.022 \times 10^{23} = 1.99 \times 10^5 \text{ J} = \mathbf{199 \text{ kJ/mol}}.$$

6. The ionisation energy of hydrogen is 13.6 eV. What is the energy of the electron in the 3rd orbit?

$$E_3 = -13.6/9 = \mathbf{-1.51 \text{ eV}}.$$

7. A 25 W bulb emits monochromatic light of 0.57 μm . Find the number of photons emitted per second.

$$E \text{ per photon} = 6.626 \times 10^{-34} \times 3 \times 10^8 / 0.57 \times 10^{-6} = 3.49 \times 10^{-19} \text{ J. Number} = 25 / 3.49 \times 10^{-19} = \mathbf{7.17 \times 10^{19}}.$$

8. Calculate the enthalpy change for C(graphite) $\rightarrow$ C(diamond) given combustion enthalpies -393.5 and -395.4 kJ/mol.

$$\Delta H = -393.5 - (-395.4) = \mathbf{+1.9 \text{ kJ/mol}}.$$

9. Calculate the entropy change for melting 1 mol ice at 0 $^{\circ}\text{C}$ ($\Delta H_{\text{fus}} = 6.0$ kJ/mol).

$$\Delta S = 6000 / 273 = \mathbf{21.98 \text{ J K}^{-1} \text{ mol}^{-1}}.$$

10. For the reaction at 298 K: $2A + B \rightarrow C$, $\Delta H = 400$ kJ, $\Delta S = 0.2$ kJ/K. Is it spontaneous at 298 K?

$$\Delta G = 400 - 298 \times 0.2 = 340.4 \text{ kJ} > 0 \Rightarrow \text{not spontaneous; it becomes spontaneous above } 2000 \text{ K}.$$

11. 0.5 mol H_2 and 0.5 mol I_2 are heated in a 1 L vessel. At equilibrium 0.8 mol HI is present. Find K_c .

$$\text{HI formed } 0.8 \Rightarrow \text{H}_2 \text{ and I}_2 \text{ used } 0.4 \text{ each, left } 0.1 \text{ each. } K_c = 0.64 / 0.01 = \mathbf{64}.$$

12. Find the pH of a solution obtained by mixing 50 mL of 0.1 M HCl and 50 mL of 0.2 M NaOH.

$$\text{H}^+ = 5 \text{ mmol; OH}^- = 10 \text{ mmol; excess OH}^- = 5 \text{ mmol in } 100 \text{ mL} = 0.05 \text{ M} = \text{pOH} = 1.30 \Rightarrow \text{pH} = \mathbf{12.70}.$$

13. The K_{sp} of PbCl_2 is 1.6×10^{-5} . Will a precipitate form when 0.1 M $\text{Pb}(\text{NO}_3)_2$ and 0.01 M NaCl are mixed in equal volumes?

$$\text{After mixing: } [\text{Pb}^{2+}] = 0.05, [\text{Cl}^-] = 0.005. \text{ Ionic product} = 0.05 \times (0.005)^2 = 1.25 \times 10^{-6} < K_{\text{sp}} \Rightarrow \mathbf{\text{no precipitate}}.$$

14. Calculate the oxidation number of: (i) P in $\text{H}_4\text{P}_2\text{O}_7$ (ii) Mn in K_2MnO_4 (iii) Fe in $[\text{Fe}(\text{CN})_6]^{4-}$ (iv) S in $\text{H}_2\text{S}_2\text{O}_8$ (peroxo, two O as -1).

$$\text{(i) } 4 + 2x - 14 = 0 \Rightarrow +5 \text{ (ii) } 2 + x - 8 = 0 \Rightarrow +6 \text{ (iii) } x - 6 = -4 \Rightarrow +2 \text{ (iv) } 2 + 2x - 12 - 2 = 0 \Rightarrow +6.$$

15. 0.2475 g of an organic compound gave 0.4950 g CO_2 and 0.2025 g H_2O . Find %C and %H.

$$\%C = (12/44)(0.4950/0.2475) \times 100 = 54.55\%. \%H = (2/18)(0.2025/0.2475) \times 100 = 9.09\%.$$

16. Calculate the volume of CO₂ at STP produced by complete combustion of 5.6 L of propane.

$\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$. Volume ratio 1 : 3 $\Rightarrow$ **16.8 L**.

Exam-Style Short Notes (write 4-6 lines each)

1. Aufbau principle

Electrons fill orbitals in increasing order of energy: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d ... according to the (n + l) rule; when (n + l) is equal, the orbital with lower n is filled first. It explains the configurations of elements, with exceptions such as Cr and Cu due to stability of half/fully filled subshells.

2. Lattice enthalpy

Energy released when one mole of an ionic solid is formed from its gaseous ions (or needed to separate it). Higher lattice enthalpy means a more stable ionic compound; it increases with higher ionic charges and smaller ionic sizes (MgO > NaCl).

3. Electronegativity

Tendency of an atom in a molecule to attract the shared electron pair. Pauling scale: F = 4.0, O = 3.5, N = 3.0, Cl = 3.0, C = 2.5, H = 2.1. Difference in electronegativity determines bond polarity; difference above 1.7 gives mostly ionic bonds.

4. Hess's law

Enthalpy change of a reaction is the same whether it occurs in one step or several. Thermochemical equations can be added or subtracted like algebraic equations. Used to calculate enthalpies of formation, transition and lattice enthalpies.

5. Le Chatelier's principle

If a system in equilibrium is subjected to a change of concentration, pressure or temperature, it shifts in the direction that tends to counteract the change. Applied in the Haber process, contact process and in everyday systems like soda bottles.

6. Ozonolysis

Treatment of an alkene with ozone forms an ozonide, which on reduction with Zn/H₂O gives aldehydes and/or ketones. By identifying the products, the position of the double bond can be located. E.g. propene $\rightarrow$ ethanal + methanal.

7. Aromaticity

Special stability of cyclic, planar, fully conjugated systems with (4n + 2) π electrons (Hückel's rule). Benzene, naphthalene, pyridine and the cyclopentadienyl anion are aromatic; such compounds undergo substitution rather than addition.

8. Conformations of ethane

Different spatial arrangements due to rotation about the C-C σ bond. Staggered (most stable, H atoms farthest apart) and eclipsed (least stable). Energy difference 12.5 kJ/mol is small, so rotation is almost free at room temperature; shown by Sawhorse and Newman projections.

Sample Paper 3 (with answers)

Group A (1 × 10)

1. How many moles of O atoms are in 1 mol of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$?
2. Which quantum number distinguishes the two electrons of the 1s orbital?
3. Write the IUPAC symbol for element 116.
4. Name the hybridisation of O in H_2O .
5. Is enthalpy an extensive or intensive property?
6. Write the unit of K_c for $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$.
7. What is the O.N. of S in SO_4^{2-} ?
8. Name the technique to separate o- and p-nitrophenol.
9. Write the structure of 2,2-dimethylpropane.
10. Which gas is formed when sodium acetate is heated with soda lime?

Answers: 1. 9 2. Spin quantum number 3. Uuh (now Lv) 4. sp^3 5. Extensive 6. $\text{L}^2 \text{mol}^{-2}$
7. +6 8. Steam distillation 9. $\text{C}(\text{CH}_3)_4$ 10. Methane.

Group B (2 × 8)

11. Define mole fraction. What is the sum of mole fractions in a solution?

Ratio of moles of a component to total moles. The sum of all mole fractions is 1.

12. Write two differences between an orbit and an orbital.

Orbit: definite circular path (Bohr), holds up to $2n^2$ electrons, planar, against the uncertainty principle. Orbital: 3-D region of high probability of finding an electron, holds at most 2 electrons, has definite shape (s, p, d).

13. Why do elements in the same group have similar properties?

They have the same number of valence electrons and similar outer electronic configurations.

14. What is formal charge? Calculate it for N in NH_4^+ .

$$\text{FC} = V - N - B/2 = 5 - 0 - 8/2 = +1.$$

15. Explain why ΔH and ΔU are nearly equal for reactions involving only solids and liquids.

Volume change is negligible, so $P\Delta V \approx 0$ and $\Delta H \approx \Delta U$.

16. What is the effect of adding an inert gas at constant volume on an equilibrium?

No effect, because the partial pressures (concentrations) of reacting gases do not change.

17. Write the IUPAC name of $\text{CH}_3\text{-CH(Br)-CH}_2\text{-CH(CH}_3\text{)-CH}_3$.

2-Bromo-4-methylpentane.

18. What happens when propene is treated with HBr in presence of benzoyl peroxide?

1-Bromopropane is formed (anti-Markovnikov, peroxide effect).

Group C (3 × 8)

19. Calculate the percentage composition of urea, NH_2CONH_2 .

Molar mass 60. N: $28/60 = 46.67\%$; C: $12/60 = 20\%$; O: $16/60 = 26.67\%$; H: $4/60 = 6.67\%$.

20. An electron is in the 4f orbital. What are the possible values of n, l and m_l ?

$n = 4$, $l = 3$, $m_l = -3, -2, -1, 0, +1, +2, +3$.

21. Arrange in order: (i) increasing size: Na^+ , F^- , O^{2-} , Mg^{2+} (ii) increasing IE: Li, Na, K.

(i) $\text{Mg}^{2+} < \text{Na}^+ < \text{F}^- < \text{O}^{2-}$ (ii) $\text{K} < \text{Na} < \text{Li}$.

22. Give the hybridisation and shape of: C_2H_6 , C_2H_4 , C_2H_2 .

sp^3 , tetrahedral around each C; sp^2 , planar; sp , linear.

23. Calculate the bond enthalpy of C-H in CH_4 if $\Delta_f H(\text{CH}_4) = -74.8 \text{ kJ/mol}$, $\Delta_{\text{sub}} H(\text{C}) = 716.7 \text{ kJ/mol}$ and bond enthalpy of H-H = 436 kJ/mol .

$\text{CH}_4(\text{g}) \rightarrow \text{C}(\text{g}) + 4\text{H}(\text{g})$: $\Delta H = 716.7 + 2(436) + 74.8 = 1663.5 \text{ kJ}$. C-H = $1663.5/4 = 415.9 \text{ kJ/mol}$.

24. 0.2 M solution of a weak monobasic acid has pH 3. Find K_a .

$[\text{H}^+] = 10^{-3} = \sqrt{(K_a \times 0.2)} \Rightarrow K_a = 10^{-6}/0.2 = 5 \times 10^{-6}$.

25. Balance by ion-electron method (basic medium): $\text{MnO}_4^- + \text{I}^- \rightarrow \text{MnO}_2 + \text{I}_2$.

$\text{MnO}_4^- + 2\text{H}_2\text{O} + 3\text{e}^- \rightarrow \text{MnO}_2 + 4\text{OH}^-$ ($\times 2$); $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$ ($\times 3$). Sum: $2\text{MnO}_4^- + 6\text{I}^- + 4\text{H}_2\text{O} \rightarrow 2\text{MnO}_2 + 3\text{I}_2 + 8\text{OH}^-$.

26. Complete: (i) $\text{CH}_3\text{CH}_2\text{Br} + \text{alc. KOH} \rightarrow ?$ (ii) $\text{HC}\equiv\text{CH} + \text{H}_2\text{O} (\text{HgSO}_4, \text{H}_2\text{SO}_4) \rightarrow ?$ (iii) $\text{C}_6\text{H}_6 + \text{conc. H}_2\text{SO}_4 \rightarrow ?$

(i) $\text{CH}_2=\text{CH}_2 + \text{KBr} + \text{H}_2\text{O}$ (ii) CH_3CHO (iii) $\text{C}_6\text{H}_5\text{SO}_3\text{H} + \text{H}_2\text{O}$.

Group D (5 × 4)**27. Explain the dual behaviour of matter and derive de Broglie's relation.**

Light shows wave (interference) and particle (photoelectric effect) nature; de Broglie suggested matter also has wave nature. From $E = mc^2$ and $E = h\nu = hc/\lambda$: $mc^2 = hc/\lambda \Rightarrow \lambda = h/mc$; for a particle of velocity v , $\lambda = h/mv = h/p$. Confirmed by electron diffraction (Davisson-Germer); used in electron microscopes.

28. Explain the formation of the H_2 molecule by valence bond theory, with the potential-energy curve.

As two H atoms approach, attractive forces (nucleus-electron of other atom) and repulsive forces (nucleus-nucleus, electron-electron) act. Energy decreases until a minimum at 74 pm, where attraction balances repulsion and 1s orbitals overlap to form a σ bond; bond enthalpy 435.8 kJ/mol. Closer approach raises energy sharply due to repulsion.

29. Define pH. Calculate the pH of (i) 10^{-8} M HCl (ii) 0.001 M H_2SO_4 .

$pH = -\log[H^+]$. (i) Must include H^+ from water: $[H^+] \approx 10^{-8} + 10^{-7} = 1.1 \times 10^{-7} \Rightarrow pH \approx 6.96$ (slightly acidic, not 8). (ii) $[H^+] = 0.002 \Rightarrow pH = 2.70$.

30. Write three methods of preparation of alkanes and describe their halogenation and combustion.

Preparation: hydrogenation of alkenes/alkynes (Ni, 523 K); Wurtz reaction; decarboxylation of sodium salts; reduction of alkyl halides with Zn/HCl. Halogenation: free-radical substitution in light (mechanism: initiation, propagation, termination). Combustion: $C_nH_{2n+2} + (3n + 1)/2 O_2 \rightarrow nCO_2 + (n + 1)H_2O + \text{heat}$; incomplete combustion gives CO and soot.

Chapter-wise Revision Charts

Chapter	Most important points to revise
Basic concepts	Laws of combination; mole concept; empirical/molecular formula; molarity, molality, mole fraction; limiting reagent; stoichiometry numericals.
Structure of atom	Discovery of particles; Rutherford and Bohr models; spectral series; photoelectric effect; de Broglie, Heisenberg; quantum numbers; Aufbau, Pauli, Hund; configurations and exceptions.
Periodicity	Modern periodic law; blocks; IUPAC names for $Z > 100$; trends in radius, IE, EGE, EN; exceptions with reasons.
Bonding	Octet rule; ionic and covalent bonds; Fajans' rules; dipole moment; VSEPR shapes; hybridisation; resonance; MOT for H_2 to Ne_2 ; hydrogen bonding.
Thermodynamics	System types; state functions; first law; $\Delta H = \Delta U + \Delta n_g RT$; Hess's law; enthalpy types; entropy; Gibbs energy and spontaneity table.
Equilibrium	K_c , K_p ; Le Chatelier; acid-base concepts; pH; hydrolysis; buffers and Henderson equation; K_{sp} ; common ion effect.
Redox	O.N. rules; oxidising and reducing agents; balancing by both methods; types of redox reactions.
Organic basics	Purification methods; detection and estimation of elements; IUPAC naming; isomerism; inductive, electromeric, resonance, hyperconjugation; reaction intermediates; reagents.
Hydrocarbons	Preparation and reactions of alkanes, alkenes, alkynes; Markovnikov and peroxide effect; ozonolysis; benzene structure; electrophilic substitution mechanism; directive influence.

Common Mistakes in Chemistry Answers

Mistake	Correct practice
Writing "mole" values without units	Always write mol, g, L, mol/L, J etc.
Using 22.4 L for gases not at STP	Use $PV = nRT$ for other conditions.
Filling 3d before 4s in K, Ca	4s fills first; but 4s electrons are removed first when forming ions ($\text{Fe}^{2+} = 3d^6$).
Writing the shape of H_2O as tetrahedral	Electron-pair geometry is tetrahedral; molecular shape is bent.
Forgetting sign convention of work	Expansion work done by system is negative.
Including solids/liquids in K	Omit pure solids and liquids.
Taking pH of 10^{-8} M HCl as 8	Include water's H^+ ; $\text{pH} \approx 6.96$.
Ignoring charge balance in redox equations	Check atoms and charges on both sides.
Numbering the chain from the wrong end	Lowest locant to the principal functional group first.
Applying the peroxide effect to HCl or HI	Only HBr shows it.

PRACTICE WORKSHEETS

Chemistry · +2 1st Year · Chapter-wise with answers

Worksheet 1 – Some Basic Concepts

1. Calculate the molar mass of (i) H_2O (ii) CO_2 (iii) CH_4 (iv) $\text{C}_{12}\text{H}_{22}\text{O}_{11}$.
2. How many significant figures: (i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000 (v) 500.0 (vi) 2.0034?
3. Round off to 3 significant figures: (i) 34.216 (ii) 10.4107 (iii) 0.04597 (iv) 2808.
4. Calculate the number of moles in 1 g of each: H_2 , O_2 , CO_2 .
5. What is the molality of a solution of 0.4 mol glucose in 500 g water?
6. If 4 g NaOH dissolves in 36 g H_2O , find the mole fraction of each and the molarity (density of solution = 1 g/mL).
7. In $3\text{A} + 2\text{B} \rightarrow 2\text{C}$, what is the limiting reagent if (i) 300 atoms A + 200 molecules B (ii) 2 mol A + 3 mol B?

Answers: 1. 18, 44, 16, 342 g/mol 2. 2, 3, 4, 3, 4, 5 3. 34.2, 10.4, 0.0460, 2810 4. 0.5, 0.03125, 0.0227 mol 5. 0.8 m 6. $n(\text{NaOH}) = 0.1$, $n(\text{H}_2\text{O}) = 2$; $x(\text{NaOH}) = 0.048$, $x(\text{H}_2\text{O}) = 0.952$; volume 40 mL $\Rightarrow M = 2.5 \text{ M}$ 7. (i) exact ratio — no limiting reagent (ii) A is limiting: 2 mol A needs only $\frac{4}{3}$ mol B, so B is in excess.

Worksheet 2 – Structure of Atom

1. Calculate the number of protons, neutrons and electrons in $^{80}_{35}\text{Br}$ and $^{56}_{26}\text{Fe}^{3+}$.
2. Write the electronic configurations of: Na^+ , O^{2-} , F^- , Cr^{3+} , Mn^{2+} .
3. What is the total number of orbitals associated with $n = 3$? How many electrons can be accommodated?
4. Using s, p, d notation describe orbitals with: (i) $n = 1, l = 0$ (ii) $n = 3, l = 1$ (iii) $n = 4, l = 2$ (iv) $n = 4, l = 3$.
5. Explain why the electronic configuration of Cr is $3d^5 4s^1$.

- Find the energy required to excite an electron in hydrogen from $n = 1$ to $n = 2$.
- Calculate the de Broglie wavelength of (i) a ball of mass 0.1 kg moving at 35 m/s (ii) an electron moving at 1×10^6 m/s.

Answers: 1. Br: 35, 45, 35; Fe^{3+} : 26, 30, 23 2. $1s^2 2s^2 2p^6$ (Na^+ , O^{2-} , F^-); Cr^{3+} $[\text{Ar}]3d^3$; Mn^{2+} $[\text{Ar}]3d^5$ 3. 9 orbitals, 18 electrons 4. 1s, 3p, 4d, 4f 5. Extra stability of half-filled d^5 6. $\Delta E = 2.18 \times 10^{-18}(1 - 1/4) = 1.635 \times 10^{-18}$ J 7. Ball: $\lambda = 6.626 \times 10^{-34}/35 = 1.89 \times 10^{-34}$ m; electron: $6.626 \times 10^{-34}/(9.1 \times 10^{-31} \times 10^6) = 7.28 \times 10^{-10}$ m.

Worksheet 3 – Periodicity

- Write the IUPAC names of elements with $Z = 101, 105, 110, 118$.
- Predict the period, group and block of elements with $Z = 13, 21, 35, 56$.
- Which of the following pairs has the larger size? (i) K or K^+ (ii) Br or Br^- (iii) O^{2-} or F^- (iv) Li^+ or Na^+ .
- Why is the first IE of Mg higher than that of Al?
- Arrange in increasing order of metallic character: Si, Be, Mg, Na, P.

Answers: 1. Unnilunium, Unnilpentium, Ununnilium, Ununoctium 2. $Z = 13$: period 3, group 13, p; 21: period 4, group 3, d; 35: period 4, group 17, p; 56: period 6, group 2, s 3. K, Br^- , O^{2-} , Na^+ 4. Mg has $3s^2$ (fully filled, penetrating s electron); Al's $3p^1$ electron is shielded and easier to remove 5. $\text{P} < \text{Si} < \text{Be} < \text{Mg} < \text{Na}$.

Worksheet 4 – Chemical Bonding

- Write Lewis dot structures of: H_2S , SiCl_4 , BeF_2 , CO_3^{2-} , HCOOH .
- Predict the shapes: SO_2 , XeF_4 , ClF_3 , NH_4^+ , PCl_5 , SF_6 .
- Arrange in increasing order of ionic character: N-H, F-H, C-H, O-H.
- Write the hybridisation of each carbon in $\text{CH}_2=\text{C}=\text{O}$, CH_3CN , HCONH_2 .
- Compare bond orders of N_2 , N_2^+ and N_2^- .
- Which has higher boiling point: ethanol or dimethyl ether? Why?

Answers: 2. Bent, square planar, T-shaped, tetrahedral, trigonal bipyramidal, octahedral 3. $\text{C-H} < \text{N-H} < \text{O-H} < \text{F-H}$ 4. $\text{CH}_2 = sp^2$, $=\text{C} = sp$; $\text{CH}_3 sp^3$, $\text{C}\equiv\text{N} sp$; HCONH_2 carbon sp^2 5. 3, 2.5, 2.5 6. Ethanol, due to intermolecular hydrogen bonding.

Worksheet 5 – Thermodynamics

- A system absorbs 300 J of heat and does 100 J of work. Find ΔU .
- For the combustion of 1 mol benzene at 298 K, $\Delta U = -3267$ kJ. Find ΔH ($\text{C}_6\text{H}_6(\text{l}) + 15/2 \text{O}_2 \rightarrow 6\text{CO}_2 + 3\text{H}_2\text{O}(\text{l})$).

- Calculate the heat needed to raise the temperature of 100 g water from 25 °C to 75 °C ($c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$).
- $\Delta_f H$ of Fe_2O_3 and Al_2O_3 are -824.2 and -1675.7 kJ/mol . Find ΔH of $2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}$.
- Predict whether ΔS is +ve or -ve: (i) $\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$ (ii) $2\text{HgO}(\text{s}) \rightarrow 2\text{Hg}(\text{l}) + \text{O}_2(\text{g})$.

Answers: 1. $\Delta U = 300 - 100 = 200 \text{ J}$ 2. $\Delta n_g = 6 - 7.5 = -1.5 \Rightarrow \Delta H = -3267 - 1.5 \times 8.314 \times 10^{-3} \times 298 = -3270.7 \text{ kJ}$ 3. $100 \times 4.18 \times 50 = 20900 \text{ J} = 20.9 \text{ kJ}$ 4. $-1675.7 + 824.2 = -851.5 \text{ kJ}$ 5. (i) negative (ii) positive.

Worksheet 6 - Equilibrium

- Write K_c for: (i) $2\text{NOCl}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Cl}_2(\text{g})$ (ii) $\text{Fe}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightleftharpoons \text{Fe}(\text{OH})_3(\text{s})$.
- Which way will the equilibrium $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$ (ΔH -ve) shift on (i) adding O_2 (ii) raising temperature (iii) raising pressure?
- Calculate the pH of 0.0001 M $\text{Ca}(\text{OH})_2$.
- Write the conjugate acids of H_2O , HCO_3^{-} , NH_2^{-} and conjugate bases of HF , H_2SO_4 , HCO_3^{-} .
- Predict whether solutions are acidic, basic or neutral: NaCl , KBr , NaCN , NH_4NO_3 , NaNO_2 , KF .
- The solubility of BaSO_4 is $2.42 \times 10^{-3} \text{ g/L}$ at 298 K. Calculate K_{sp} (molar mass 233).

Answers: 1. (i) $[\text{NO}]^2[\text{Cl}_2]/[\text{NOCl}]^2$ (ii) $1/([\text{Fe}^{3+}][\text{OH}^{-}]^3)$ 2. forward, backward, forward 3. $[\text{OH}^{-}] = 2 \times 10^{-4} \Rightarrow \text{pOH} = 3.70 \Rightarrow \text{pH} = 10.30$ 4. H_3O^{+} , H_2CO_3 , NH_3 ; F^{-} , HSO_4^{-} , CO_3^{2-} 5. neutral, neutral, basic, acidic, basic, basic 6. $s = 2.42 \times 10^{-3}/233 = 1.04 \times 10^{-5} \text{ M} = K_{sp} = s^2 = 1.08 \times 10^{-10}$.

Worksheet 7 - Redox Reactions

- Find the O.N. of P in NaH_2PO_4 , S in NaHSO_4 , P in $\text{H}_4\text{P}_2\text{O}_7$, Mn in K_2MnO_4 , B in NaBH_4 .
- Identify the species oxidised and reduced: $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$.
- Balance: $\text{Fe}^{2+} + \text{H}_2\text{O}_2 \rightarrow \text{Fe}^{3+} + \text{H}_2\text{O}$ (acidic).
- Balance: $\text{Zn} + \text{NO}_3^{-} \rightarrow \text{Zn}^{2+} + \text{NH}_4^{+}$ (acidic).

Answers: 1. +5, +6, +5, +6, +3 2. H_2 oxidised, CuO reduced 3. $2\text{Fe}^{2+} + \text{H}_2\text{O}_2 + 2\text{H}^{+} \rightarrow 2\text{Fe}^{3+} + 2\text{H}_2\text{O}$ 4. $4\text{Zn} + \text{NO}_3^{-} + 10\text{H}^{+} \rightarrow 4\text{Zn}^{2+} + \text{NH}_4^{+} + 3\text{H}_2\text{O}$.

Worksheet 8 – Organic Chemistry Basics

1. Write IUPAC names: (i) $\text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{CH}_2\text{OH}$ (ii) $\text{CH}_3\text{COCH}_2\text{CH}_3$ (iii) $\text{CH}_3\text{CH}_2\text{COOCH}_3$ (iv) $\text{CH}_2=\text{CHCHO}$ (v) ClCH_2COOH .
2. Draw structures: (i) 2-chlorohexane (ii) pent-4-en-2-ol (iii) 3-nitrocyclohexene (iv) 2,5,6-trimethyloctane.
3. Which bond is more polar: C-Cl or C-Mg? Show direction of polarity.
4. Give an example each of chain, position and functional isomerism.
5. 0.15 g of an organic compound gave 0.12 g AgBr. Find %Br.

Answers: 1. 3-methylbutan-1-ol; butan-2-one; methyl propanoate; prop-2-enal; 2-chloroethanoic acid 3. C-Mg ($\text{C}^{\delta-}-\text{Mg}^{\delta+}$; difference 1.3) vs $\text{C}^{\delta+}-\text{Cl}^{\delta-}$ (0.5) 4. butane/isobutane; propan-1-ol/propan-2-ol; ethanol/dimethyl ether 5. $\% \text{Br} = (80/188)(0.12/0.15) \times 100 = 34.04\%$.

Worksheet 9 – Hydrocarbons

1. Write the structures and names of all isomers of C_4H_8 (alkenes only).
2. What happens when: (i) ethyne is passed through ammoniacal Cu_2Cl_2 (ii) propyne reacts with water ($\text{HgSO}_4/\text{H}_2\text{SO}_4$) (iii) benzene reacts with Cl_2 in UV light?
3. Write the ozonolysis products of (i) pent-2-ene (ii) 2-methylpropene.
4. Arrange in decreasing order of reactivity towards electrophiles: benzene, toluene, nitrobenzene, phenol.
5. Convert: (i) propene to 1-bromopropane (ii) benzene to m-dinitrobenzene.

Answers: 1. But-1-ene, cis-but-2-ene, trans-but-2-ene, 2-methylpropene 2. (i) red ppt of copper acetylide (ii) propanone (iii) benzene hexachloride ($\text{C}_6\text{H}_6\text{Cl}_6$) 3. (i) ethanal + propanal (ii) propanone + methanal 4. phenol > toluene > benzene > nitrobenzene 5. (i) HBr with peroxide (ii) nitrate twice: second $-\text{NO}_2$ enters meta.

HOTS, FORMULA SHEET & FINAL MOCK

Chemistry · +2 1st Year

Higher-Order Thinking Questions

1. Two elements A and B form compounds A_2B_3 and AB_2 . If 0.15 mol A_2B_3 weighs 15.9 g and 0.15 mol AB_2 weighs 9.3 g, find the atomic masses of A and B.

Molar masses: $A_2B_3 = 106$, $AB_2 = 62$. $2a + 3b = 106$; $a + 2b = 62 \Rightarrow 2a + 4b = 124 \Rightarrow b = 18$, $a = 26$.

2. Why is the third ionisation enthalpy of Mn very high while that of Fe is lower?

Mn^{2+} is $3d^5$ (half-filled, very stable), so removing a third electron is hard. Fe^{2+} is $3d^6$; losing one electron gives stable $3d^5$, so IE_3 is lower.

3. The dipole moment of $CHCl_3$ is less than that of CH_2Cl_2 . Explain.

In $CHCl_3$ the three C-Cl dipoles partly cancel each other (closer to symmetric), while in CH_2Cl_2 two C-Cl dipoles add up at the tetrahedral angle, giving a larger resultant.

4. Why is the bond angle in PH_3 (93.5°) much smaller than in NH_3 (107°)?

P is larger and less electronegative; bond pairs lie farther from P and repel less, and P uses nearly pure p orbitals for bonding, so the angle approaches 90° .

5. A reaction is spontaneous at all temperatures. What are the signs of ΔH and ΔS ?

ΔH negative and ΔS positive.

6. The pH of an acetic acid-sodium acetate buffer is 4.74. What happens to its pH on dilution with water? Why?

Practically unchanged: dilution changes [salt] and [acid] equally, so their ratio (and $pH = pK_a + \log \text{ratio}$) remains the same.

7. Why does $AgCl$ dissolve in aqueous ammonia but AgI does not?

NH_3 forms $[\text{Ag}(\text{NH}_3)_2]^+$, lowering $[\text{Ag}^+]$. For AgCl ($K_{\text{sp}} 10^{-10}$) the ionic product falls below K_{sp} and it dissolves; AgI has a much lower K_{sp} (10^{-17}), so $[\text{Ag}^+]$ cannot be reduced enough.

8. Why is 2-methylpropene more reactive than ethene towards HBr?

Protonation gives a stable 3° carbocation from 2-methylpropene but only a 1° carbocation from ethene.

9. How would you distinguish but-1-yne from but-2-yne?

But-1-yne (terminal) gives a white ppt with ammoniacal AgNO_3 ; but-2-yne does not.

10. Benzene is more stable than expected from Kekulé's structure by about 150 kJ/mol. How is this known?

Enthalpy of hydrogenation of cyclohexene is -120 kJ/mol; for a hypothetical cyclohexatriene it would be $3 \times -120 = -360$ kJ/mol, but benzene's is about -208 kJ/mol. The difference (≈ 150 kJ/mol) is resonance energy.

Chemistry Formula Sheet - Class XI

Topic	Formula
Mole	$n = m/M = N/N_A = V(\text{STP})/22.4$
Concentration	$M = n/V(L)$; $m = n/\text{kg solvent}$; $x_A = n_A/\Sigma n$; $N = M \times \text{n-factor}$; $M_1V_1 = M_2V_2$
Formula	Molecular formula = $n \times \text{empirical formula}$; $n = M/(\text{EF mass})$; $M = 2 \times \text{vapour density}$
Light	$c = \nu\lambda$; $E = h\nu = hc/\lambda$; wave number = $1/\lambda$
Bohr	$E_n = -13.6Z^2/n^2 \text{ eV} = -2.18 \times 10^{-18}Z^2/n^2 \text{ J}$; $r_n = 0.529n^2/Z \text{ \AA}$; $mvr = nh/2\pi$
Spectrum	$1/\lambda = R_H Z^2(1/n_1^2 - 1/n_2^2)$; lines = $n(n-1)/2$
Photoelectric	$h\nu = h\nu_0 + \frac{1}{2}mv^2$
de Broglie, Heisenberg	$\lambda = h/mv$; $\Delta x \cdot \Delta p \geq h/4\pi$
Nodes	radial = $n - l - 1$; angular = l
Bonding	Bond order = $\frac{1}{2}(N_b - N_a)$; formal charge = $V - L - B/2$; $\mu = q \times d$
Hybridisation shortcut	$H = \frac{1}{2}(V + M - C + A)$
Thermodynamics	$\Delta U = q + w$; $w = -P_{\text{ext}}\Delta V$; $w_{\text{rev}} = -2.303nRT \log(V_2/V_1)$; $\Delta H = \Delta U + \Delta n_g RT$; $q = mc\Delta T$; $\Delta S = q_{\text{rev}}/T$; $\Delta G = \Delta H - T\Delta S$; $\Delta G^\circ = -2.303RT \log K$
Hess / bond enthalpy	$\Delta_r H = \Sigma \Delta_f H(P) - \Sigma \Delta_f H(R) = \Sigma BE(R) - \Sigma BE(P)$
Equilibrium	$K_p = K_c(RT)^{\Delta n}$; $K(\text{reverse}) = 1/K$
Ionic	$\text{pH} = -\log[H^+]$; $K_w = 10^{-14}$; $\alpha = \sqrt{(K_a/C)}$; $[H^+] = \sqrt{(K_a C)}$; $K_a K_b = K_w$
Buffer	$\text{pH} = \text{p}K_a + \log([\text{salt}]/[\text{acid}])$; $\text{pOH} = \text{p}K_b + \log([\text{salt}]/[\text{base}])$
Solubility	$AB: K_{sp} = s^2$; $AB_2/A_2B: 4s^3$; $AB_3: 27s^4$; $A_2B_3: 108s^5$
Organic estimation	$\%C = 12/44 \times m(\text{CO}_2)/w \times 100$; $\%H = 2/18 \times m(\text{H}_2\text{O})/w \times 100$; $\%N$ (Kjeldahl) = $1.4NV/w$; $\%X = (\text{at. mass } X/\text{mol. mass } AgX) \times m(AgX)/w \times 100$; $\%S = 32/233 \times m(\text{BaSO}_4)/w \times 100$

Final Mock Test - Chemistry Class XI (Full Syllabus)

Group A - fill in the blanks (1 × 10)

- The number of atoms in 4.25 g of NH_3 is _____.
- The subshell with $n = 5$, $l = 1$ is _____.
- The element with configuration $[\text{Kr}]4d^{10}5s^25p^5$ is in group _____.
- The shape of ClF_3 is _____.
- For a cyclic process $\Delta U =$ _____.

6. The pH of a neutral solution at 25 °C is _____.
7. In $2\text{KMnO}_4 + 16\text{HCl} \rightarrow \dots$, KMnO_4 acts as _____ agent.
8. The IUPAC name of $\text{CH}_2=\text{CH}-\text{CH}=\text{CH}_2$ is _____.
9. The reagent that converts alkynes into cis-alkenes is _____.
10. The $-\text{CHO}$ group on benzene is _____ directing.

Answers: 1. $0.25 \text{ mol} \times 4 \text{ atoms} \times 6.022 \times 10^{23} = 6.022 \times 10^{23}$ 2. 5p 3. 17 4. T-shaped
5. zero 6. 7 7. oxidising 8. Buta-1,3-diene 9. Lindlar's catalyst (H_2 , Pd/BaSO_4) 10. meta.

Group B - very short answers (2 × 8)

1. Define limiting reagent with an example.
2. Why does an electron revolving in an orbit not radiate energy (Bohr)?
3. Why is K more reactive than Na?
4. What is the hybridisation of N in NO_3^- ?
5. State the second law of thermodynamics.
6. Define K_{sp} .
7. What is the difference between homolytic and heterolytic fission?
8. What is the peroxide effect?

Answers: 11. The reactant consumed first; e.g. in burning 1 mol C in 16 g O_2 , O_2 is limiting. 12. It is a postulate: electrons in stationary orbits do not lose energy. 13. K is larger with lower IE, so it loses its electron more easily. 14. sp^2 . 15. In a spontaneous process the entropy of the universe increases. 16. Product of the molar concentrations of ions in a saturated solution, each raised to its stoichiometric power. 17. Homolytic: each atom gets one electron $\rightarrow$ radicals; heterolytic: one atom takes both $\rightarrow$ ions. 18. Anti-Markovnikov addition of HBr in presence of peroxide.

Group C - short answers (3 × 8)

1. Calculate the mass of 2.5 mol of Na_2CO_3 and the number of Na^+ ions in it.
2. Calculate the wavelength for the transition $n = 3 \rightarrow n = 2$ in hydrogen.
3. Explain why Cu^+ is diamagnetic but Cu^{2+} is paramagnetic.
4. Draw the structure and give the hybridisation of SF_4 .
5. Calculate ΔG° at 298 K for a reaction with $K = 1.8 \times 10^{-5}$.
6. Calculate pH of 0.05 M acetic acid ($K_a = 1.8 \times 10^{-5}$).
7. Write the structures of all isomers of C_5H_{12} with IUPAC names.
8. Write the mechanism of addition of HBr to propene.

Answers: 19. 265 g; $5 \text{ mol } \text{Na}^+ = 3.011 \times 10^{24}$ ions. 20. $1/\lambda = 109677(1/4 - 1/9) = 15233 \text{ cm}^{-1} \Rightarrow \lambda = 656 \text{ nm}$ (red, H_α). 21. $\text{Cu}^+ 3d^{10}$ (all paired); $\text{Cu}^{2+} 3d^9$ (one unpaired). 22. sp^3d , see-saw. 23. $\Delta G^\circ = -2.303 \times 8.314 \times 298 \times \log(1.8 \times 10^{-5}) =$

$-5706 \times (-4.745) = +27.07 \text{ kJ/mol}$. 24. $[\text{H}^+] = \sqrt{9 \times 10^{-7}} = 9.49 \times 10^{-4} \Rightarrow \text{pH} = 3.02$.
25. Pentane, 2-methylbutane, 2,2-dimethylpropane. 26. See chapter notes.

Group D - long answers (5 × 4)

1. Explain the shapes of s, p and d orbitals with diagrams and the significance of nodes.
2. Explain molecular orbital theory. Draw the MO energy diagrams of N_2 and O_2 .
3. State Le Chatelier's principle and apply it to the Haber process and to the contact process ($2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$, ΔH negative).
4. Explain the structure of benzene. Describe the mechanism of its halogenation and Friedel-Crafts acylation.

Guidance: Use the chapter notes; draw labelled diagrams; for 29, both reactions are exothermic with fewer gas moles on the product side — high pressure and moderate temperature with a catalyst (Fe for Haber, V_2O_5 for contact) give good yield at a reasonable rate.

REACTIONS & CONVERSIONS

Chemistry · +2 1st Year · Organic and inorganic practice

Organic Conversions (2–3 marks each)

#	Conversion	Steps (reagent/condition)
1	Methane → ethane	$\text{CH}_4 + \text{Cl}_2 \xrightarrow{h\nu} \text{CH}_3\text{Cl}$; Wurtz: $2\text{CH}_3\text{Cl} + 2\text{Na}$ (dry ether) $\rightarrow \text{C}_2\text{H}_6$
2	Ethane → methane	Not by a single standard reaction in syllabus; via: $\text{C}_2\text{H}_6 \rightarrow \text{C}_2\text{H}_5\text{Cl} \rightarrow \text{C}_2\text{H}_5\text{OH} \rightarrow \text{CH}_3\text{COOH}$ (oxidation) $\rightarrow \text{CH}_3\text{COONa} \rightarrow$ (soda lime) CH_4
3	Ethanol → ethene	Conc. H_2SO_4 , 443 K
4	Ethene → ethyne	$\text{Br}_2 \rightarrow \text{CH}_2\text{BrCH}_2\text{Br}$; alc. KOH then $\text{NaNH}_2 \rightarrow \text{HC}\equiv\text{CH}$
5	Ethyne → ethanal	H_2O , HgSO_4 /dil. H_2SO_4 , 333 K
6	Ethyne → benzene	Red-hot iron tube, 873 K
7	Benzene → toluene	CH_3Cl , anhydrous AlCl_3
8	Benzene → nitrobenzene	Conc. HNO_3 + conc. H_2SO_4 , 323–333 K
9	Benzene → benzenesulphonic acid	Fuming H_2SO_4 (oleum)
10	Benzene → acetophenone	CH_3COCl , anhyd. AlCl_3
11	Benzene → p-nitrobromobenzene	First $\text{Br}_2/\text{FeBr}_3$ (o/p director), then nitration $\rightarrow$ p-isomer (major)
12	Benzene → m-nitrobromobenzene	First nitration (m-director), then $\text{Br}_2/\text{FeBr}_3$
13	Propene → propan-2-ol	H_2O , dil. H_2SO_4 (Markovnikov)
14	Propene → 1-bromopropane	HBr , benzoyl peroxide
15	Propene → 2-bromopropane	HBr (no peroxide)
16	Propyne → propanone	H_2O , $\text{HgSO}_4/\text{H}_2\text{SO}_4$
17	But-2-yne → cis-but-2-ene	H_2 , Lindlar's catalyst
18	Ethene → ethane-1,2-diol	Cold dilute alkaline KMnO_4
19	Sodium benzoate → benzene	Soda lime, heat (decarboxylation)
20	Calcium carbide → ethane	$\text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_2$; then H_2/Ni (2 mol)

Name Reactions – Equation Practice

Wurtz reaction

$2\text{CH}_3\text{CH}_2\text{Br} + 2\text{Na} \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3 + 2\text{NaBr}$ (dry ether). Used to make higher alkanes with an even number of carbons.

Kolbe's electrolytic method

$2\text{CH}_3\text{COONa} + 2\text{H}_2\text{O} \rightarrow \text{CH}_3\text{--CH}_3 + 2\text{CO}_2 + 2\text{NaOH} + \text{H}_2$. Ethane at anode; H_2 at cathode.

Decarboxylation

$\text{CH}_3\text{COONa} + \text{NaOH} (\text{CaO}, \Delta) \rightarrow \text{CH}_4 + \text{Na}_2\text{CO}_3$. The alkane has one carbon less.

Friedel-Crafts alkylation and acylation

$\text{C}_6\text{H}_6 + \text{C}_2\text{H}_5\text{Cl} \rightarrow \text{C}_6\text{H}_5\text{C}_2\text{H}_5 + \text{HCl}$; $\text{C}_6\text{H}_6 + (\text{CH}_3\text{CO})_2\text{O} \rightarrow \text{C}_6\text{H}_5\text{COCH}_3 + \text{CH}_3\text{COOH}$ (anhyd. AlCl_3).

Ozonolysis

$\text{CH}_3\text{CH}=\text{CH}_2 + \text{O}_3 \rightarrow \text{ozonide}$; $\text{Zn}/\text{H}_2\text{O} \rightarrow \text{CH}_3\text{CHO} + \text{HCHO}$.

Dehydrohalogenation (Saytzeff)

$\text{CH}_3\text{CHBrCH}_2\text{CH}_3 + \text{alc. KOH} \rightarrow \text{CH}_3\text{CH}=\text{CHCH}_3$ (major) + $\text{CH}_2=\text{CHCH}_2\text{CH}_3$ (minor).

Inorganic and Physical Chemistry – Equation Practice

1. Write balanced equations for the action of heat on: (i) CaCO_3 (ii) KClO_3 (with MnO_2) (iii) NH_4NO_2 .

(i) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ (ii) $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$ (iii) $\text{NH}_4\text{NO}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}$.

2. Write the thermochemical equation for the formation of water.

$\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta_f\text{H}^\circ = -285.8 \text{ kJ/mol}$.

3. Write the hydrolysis equations of NH_4Cl and CH_3COONa .

$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4\text{OH} + \text{H}^+$ (acidic); $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$ (basic).

4. Write the reactions of the dry tests for CO_3^{2-} and S^{2-} .

$\text{Na}_2\text{CO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O} + \text{CO}_2\uparrow$; $\text{CO}_2 + \text{Ca}(\text{OH})_2 \rightarrow \text{CaCO}_3\downarrow$ (milky) + H_2O . $\text{Na}_2\text{S} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{S}\uparrow$; $\text{H}_2\text{S} + (\text{CH}_3\text{COO})_2\text{Pb} \rightarrow \text{PbS}\downarrow$ (black) + $2\text{CH}_3\text{COOH}$.

5. Write the chemistry of the brown ring test.

$\text{NaNO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{NaHSO}_4 + \text{HNO}_3$; $6\text{FeSO}_4 + 2\text{HNO}_3 + 3\text{H}_2\text{SO}_4 \rightarrow 3\text{Fe}_2(\text{SO}_4)_3 + 2\text{NO} + 4\text{H}_2\text{O}$; $\text{FeSO}_4 + \text{NO} + 5\text{H}_2\text{O} \rightarrow [\text{Fe}(\text{H}_2\text{O})_5(\text{NO})]\text{SO}_4$ (brown ring).

6. Write the confirmatory test reaction of chloride.

$\text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl}\downarrow$ (white) + NaNO_3 ; $\text{AgCl} + 2\text{NH}_4\text{OH} \rightarrow [\text{Ag}(\text{NH}_3)_2]\text{Cl} + 2\text{H}_2\text{O}$ (dissolves).

7. Write the reaction for the confirmatory test of sulphate.

$\text{Na}_2\text{SO}_4 + \text{BaCl}_2 \rightarrow \text{BaSO}_4\downarrow$ (white, insoluble in HCl) + 2NaCl .

8. Write the titration reaction of oxalic acid with NaOH .

$(\text{COOH})_2 + 2\text{NaOH} \rightarrow (\text{COONa})_2 + 2\text{H}_2\text{O}$. Oxalic acid dihydrate: molar mass 126, equivalent mass 63.

Diagram-Based Questions

1. Sketch the shapes of the five d orbitals.

d_{xy} , d_{yz} , d_{xz} : four lobes between the axes in their planes. $d_{x^2-y^2}$: four lobes along the x and y axes. d_{z^2} : two lobes along the z-axis with a doughnut-shaped ring in the xy-plane.

2. Draw the energy level diagram showing the hydrogen spectral series.

Horizontal lines for $n = 1, 2, 3, 4, 5, 6, \infty$ getting closer together upward. Arrows ending at $n = 1$ (Lyman, UV), $n = 2$ (Balmer, visible), $n = 3$ (Paschen), $n = 4$ (Brackett), $n = 5$ (Pfund) — the last three in IR.

3. Draw the orbital overlap diagram of ethene.

Two sp^2 carbons in one plane; each forms two σ bonds with H ($s-sp^2$) and one C-C σ bond (sp^2-sp^2); unhybridised p orbitals above and below the plane overlap sideways to form the π bond (two lobes above and below the C-C axis).

4. Draw the apparatus for steam distillation and label it.

Steam generator (flask with water and safety tube) $\rightarrow$ steam passes through a delivery tube into the distillation flask containing the mixture $\rightarrow$ vapours pass through a water condenser (water in at bottom, out at top) $\rightarrow$ distillate collected in a receiver.

5. Draw a sketch showing how the free energy changes with reaction progress at equilibrium.

Plot G (y-axis) against extent of reaction (x-axis): G falls from pure reactants to a minimum and then rises towards pure products; the minimum point is equilibrium, where $\Delta G = 0$.

Glossary of Important Definitions

Term	Definition
Atomic number	Number of protons in the nucleus of an atom.
Mass number	Total number of protons and neutrons.
Orbital	Region of space around the nucleus where the probability of finding an electron is maximum.
Degenerate orbitals	Orbitals having the same energy (e.g. three 2p orbitals).
Ionisation enthalpy	Energy needed to remove the outermost electron from an isolated gaseous atom.
Electron gain enthalpy	Enthalpy change when an electron is added to an isolated gaseous atom.
Electronegativity	Ability of a bonded atom to attract the shared pair towards itself.
Bond length	Equilibrium distance between the nuclei of two bonded atoms.
Hybridisation	Mixing of atomic orbitals of comparable energy to form equivalent new orbitals.
Resonance	Representation of a molecule by two or more structures differing only in electron positions.
System / surroundings	Part of universe under study / everything else.
Internal energy	Total energy stored in a system (kinetic + potential of all particles).
Enthalpy	$H = U + PV$; heat content at constant pressure.
Entropy	Measure of randomness or disorder of a system.
Gibbs energy	$G = H - TS$; its decrease measures useful work and decides spontaneity.
Chemical equilibrium	State where forward and backward reaction rates are equal and concentrations remain constant.
Buffer	Solution resisting change in pH.
Hydrolysis of salt	Reaction of salt ions with water to produce an acidic or basic solution.
Solubility product	Product of ionic concentrations in a saturated solution of a sparingly soluble salt.
Oxidation number	Charge an atom appears to have when all bonds are treated as ionic.
Functional group	Atom or group that determines the characteristic properties of an organic compound.
Electrophile	Electron-deficient species that attacks electron-rich centres.
Nucleophile	Electron-rich species that attacks electron-deficient centres.
Aromaticity	Extra stability of cyclic planar conjugated systems with $(4n + 2) \pi$ electrons.
Isomers	Compounds with the same molecular formula but different structures or arrangements.

PRACTICE & REVISION

Chemistry · +2 1st Year (Class XI)

Chemistry – Numerical Problem Bank

1. How many atoms are present in 0.5 mol of O_2 ?

Molecules = $0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$; atoms = $2 \times$ that = **6.022×10^{23}** .

2. What volume of 0.5 M HCl is needed to make 250 mL of 0.1 M HCl?

$M_1V_1 = M_2V_2 \Rightarrow 0.5 V_1 = 0.1 \times 250 \Rightarrow V_1 =$ **50 mL**.

3. Calculate the molality of a solution containing 18 g glucose in 500 g water.

$n = 18/180 = 0.1$ mol. $m = 0.1/0.5 =$ **0.2 mol/kg**.

4. Find the mole fraction of ethanol in a solution of 46 g ethanol and 90 g water.

$n(\text{ethanol}) = 1$, $n(\text{water}) = 5$. $x = 1/6 =$ **0.167**.

5. The frequency of a radiation is 6×10^{14} Hz. Find its wavelength and energy per photon.

$\lambda = c/\nu = 3 \times 10^8 / 6 \times 10^{14} = 5 \times 10^{-7}$ m (500 nm). $E = h\nu = 6.626 \times 10^{-34} \times 6 \times 10^{14} =$ **3.98×10^{-19} J**.

6. Calculate the radius of the 2nd Bohr orbit of hydrogen.

$r = 0.529 \times 4 =$ **2.116 Å**.

7. A ball of mass 0.1 kg moves at 10 m/s. Find its de Broglie wavelength. Why is it not noticed?

$\lambda = 6.626 \times 10^{-34} / (0.1 \times 10) = 6.6 \times 10^{-34}$ m — far too small to detect, so wave nature of big objects is not observed.

8. Calculate the uncertainty in velocity of an electron if $\Delta x = 1 \times 10^{-10}$ m.

$$\Delta v \geq h/(4\pi m \Delta x) = 6.626 \times 10^{-34} / (4 \times 3.14 \times 9.1 \times 10^{-31} \times 10^{-10}) \approx \mathbf{5.8 \times 10^5 \text{ m/s.}}$$

9. Calculate $\Delta_r H$ for $C_2H_4 + H_2 \rightarrow C_2H_6$ using bond enthalpies: C=C 614, C-C 347, C-H 414, H-H 435 kJ/mol.

Bonds broken: C=C + 4 C-H + H-H = 614 + 1656 + 435 = 2705. Bonds formed: C-C + 6 C-H = 347 + 2484 = 2831. $\Delta H = 2705 - 2831 = \mathbf{-126 \text{ kJ/mol.}}$

10. Find the pH of a solution with $[H^+] = 3.8 \times 10^{-3}$ M.

$$\text{pH} = 3 - \log 3.8 = 3 - 0.58 = \mathbf{2.42.}$$

11. The pH of a solution is 4.5. Find $[H^+]$ and $[OH^-]$.

$$[H^+] = 10^{-4.5} = 3.16 \times 10^{-5} \text{ M. } [OH^-] = 10^{-14} / 3.16 \times 10^{-5} = \mathbf{3.16 \times 10^{-10} \text{ M.}}$$

12. For $PCl_5 \rightleftharpoons PCl_3 + Cl_2$, $K_c = 8.3 \times 10^{-3}$. If $[PCl_3] = [Cl_2] = 0.1$ M, find $[PCl_5]$.

$$[PCl_5] = (0.1 \times 0.1) / 8.3 \times 10^{-3} = \mathbf{1.2 \text{ M.}}$$

Chemistry – Important Name Reactions & Tests (Class XI)

Name	Reaction
Wurtz reaction	$2RX + 2Na \rightarrow R-R + 2NaX$ (dry ether)
Decarboxylation	$RCOONa + NaOH$ (CaO) $\rightarrow$ RH + Na_2CO_3
Kolbe's electrolysis	$2CH_3COONa + 2H_2O \rightarrow C_2H_6 + 2CO_2 + 2NaOH + H_2$
Markovnikov addition	$CH_3CH=CH_2 + HBr \rightarrow CH_3CHBrCH_3$
Kharasch (peroxide) effect	$CH_3CH=CH_2 + HBr$ (peroxide) $\rightarrow$ $CH_3CH_2CH_2Br$
Ozonolysis	Alkene + O_3 , then Zn/ H_2O $\rightarrow$ carbonyl compounds
Friedel-Crafts	$C_6H_6 + RCl$ / $RCOCl$ (anhyd. $AlCl_3$) $\rightarrow$ alkyl / acyl benzene
Baeyer's test	Cold dil. alkaline $KMnO_4$ decolourised by unsaturated compounds
Bromine water test	Red-brown colour discharged by alkenes/alkynes
Lassaigne's test	Sodium fusion extract $\rightarrow$ Prussian blue (N), violet (S), AgX ppt (halogens)

Chemistry – Important Long-Answer Questions (Model Answers)

Q1. State Bohr's postulates. Derive the expression for the radius of the n-th orbit (outline). [5]

Postulates: (1) Electrons revolve round the nucleus in fixed circular paths called orbits or stationary states, each with definite energy. (2) While in an orbit, an electron does not lose energy. (3) Only those orbits are allowed for which angular momentum $mvr = nh/2\pi$ ($n = 1, 2, 3, \dots$). (4) Energy is absorbed or emitted only when the electron jumps from one orbit to another; $\Delta E = E_2 - E_1 = h\nu$.

Radius: Centripetal force = electrostatic attraction: $mv^2/r = kZe^2/r^2$. Using $mvr = nh/2\pi$ to eliminate v gives $r_n = n^2h^2/(4\pi^2mkZe^2) = 0.529 n^2/Z \text{ \AA}$.

Q2. What is hybridisation? Explain the formation of ethene and ethyne on its basis. [5]

Hybridisation is the intermixing of atomic orbitals of slightly different energies to form an equal number of new identical orbitals. **Ethene:** each C is sp^2 ; two sp^2 orbitals form σ bonds with H atoms and one forms a C-C σ bond; the unhybridised 2p orbitals on the two carbons overlap sideways to form a π bond. The molecule is planar with angles of about 120° . **Ethyne:** each C is sp ; one sp orbital bonds with H, the other forms the C-C σ bond; two unhybridised p orbitals on each carbon form two π bonds. The molecule is linear (180°).

Q3. State and explain the first law of thermodynamics. Show that $q_p = \Delta H$. [5]

Energy of an isolated system is constant; it can be converted from one form to another. $\Delta U = q + w$. At constant pressure, $w = -P\Delta V$, so $\Delta U = q_p - P\Delta V \Rightarrow q_p = \Delta U + P\Delta V$. Since $H = U + PV$, at constant P: $\Delta H = \Delta U + P\Delta V$. Therefore $q_p = \Delta H$.

Q4. What is a buffer solution? Derive the Henderson equation for an acidic buffer. [5]

A buffer resists change in pH on adding small amounts of acid or base. For $HA \rightleftharpoons H^+ + A^-$: $K_a = [H^+][A^-]/[HA] \Rightarrow [H^+] = K_a[HA]/[A^-]$. Taking $-\log$: $pH = pK_a + \log([A^-]/[HA])$. In the buffer, the weak acid is hardly ionised, so $[HA] \approx [\text{acid}]$, and the salt is fully ionised, so $[A^-] \approx [\text{salt}]$. Hence $pH = pK_a + \log([\text{salt}]/[\text{acid}])$.

Q5. Explain inductive effect and hyperconjugation with examples. [5]

Inductive effect: permanent displacement of σ electrons along a carbon chain due to an electronegative atom/group. In CH_3-CH_2-Cl , Cl pulls electrons ($-I$), making C1 $\delta+$, C2 $\delta\delta+$. Alkyl groups push electrons ($+I$). It explains why $ClCH_2COOH$ is a stronger acid than CH_3COOH . **Hyperconjugation:** delocalisation of σ electrons of C-H bonds on the carbon next to a carbocation or double bond into the empty p orbital or π system. The more α -hydrogens, the greater the stability: $(CH_3)_3C^+$ (9 α -H) $>$ $(CH_3)_2CH^+$ (6) $>$ $CH_3CH_2^+$ (3).

Q6. Explain the mechanism of addition of HBr to propene in the presence and absence of peroxide. [5]

Absence of peroxide (ionic, Markovnikov): H^+ adds to the terminal C, forming the more stable 2° carbocation $CH_3-C^+H-CH_3$; Br^- attacks it $\rightarrow$ 2-bromopropane. **Presence of peroxide (free radical, anti-Markovnikov):** peroxide gives $RO\cdot$, which takes H from HBr to give $Br\cdot$. $Br\cdot$ adds to the terminal C forming the more stable 2° radical $CH_3-C\cdot H-CH_3$.

CH_2Br , which takes H from $\text{HBr} \rightarrow 1\text{-bromopropane}$, regenerating $\text{Br}\cdot$. HCl and HI do not show this effect.

Q7. Explain the variation of atomic radius and ionisation enthalpy in the periodic table.

[5]

Across a period, nuclear charge increases while electrons enter the same shell, so attraction increases: radius decreases and ionisation enthalpy increases (with exceptions at Be/B and N/O due to stable configurations). Down a group, new shells are added and shielding increases, so radius increases and ionisation enthalpy decreases.

Last-Minute Exam Tips

Chemistry

- Always write units: g mol^{-1} , mol L^{-1} , J, kJ mol^{-1} , m s^{-1} .
- Remember exceptions: $\text{Cr } 3\text{d}^5 4\text{s}^1$, $\text{Cu } 3\text{d}^{10} 4\text{s}^1$; ionisation enthalpy $\text{Be} > \text{B}$, $\text{N} > \text{O}$.
- In thermodynamics, follow the IUPAC sign convention (work done on the system is positive).
- In equilibrium, do not write pure solids/liquids in K; use Δn of gases only for $K_p = K_c(RT)^{\Delta n}$.
- Draw mechanisms step by step with curved arrows (free radical halogenation, electrophilic addition and substitution).
- In the practical, record observations in a three-column table: Experiment - Observation - Inference.

THE NIGHT BEFORE THE EXAM

Revise only formula boxes, name reactions, grammar rules and your own mistake list. Sleep well — a fresh mind solves problems faster than a tired one. Carry admit card, two pens, pencil, scale, eraser and a geometry box. Good luck!