CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 2 : Structure of Atom (NCERT)

2.1 Subatomic Particles & Atomic Terms

Particle / Term	Key Values / Definition
Electron (Thomson)	Charge = -1.602×10^{-19} C ; mass = 9.1×10^{-31} kg ; e/m = 1.758×10^{11} C/kg
Proton	Charge = $+1.602 \times 10^{-19}$ C ; mass = 1.672×10^{-27} kg
Neutron (Chadwick)	Charge = 0 ; mass = 1.675×10^{-27} kg
Atomic number (Z)	Z = number of protons = number of electrons (neutral atom)
Mass number (A)	A = protons + neutrons → Neutrons = A − Z
Isotopes	Same Z, different A (e.g. ^{1}H, ^{2}H, ^{3}H)
Isobars	Same A, different Z (e.g. ^{14}C and ^{14}N)
Isotones	Same number of neutrons (e.g. ^{14}C and ^{16}O — 8 neutrons each)

2.2 Electromagnetic Radiation & Planck's Theory

Quantity	Formula / Expression
Speed–frequency–wavelength	$c = v \times \lambda$ ($c = 3 \times 10^{8}$ m/s)
Wave number ($\bar{v}$)	$\bar{v} = 1/\lambda$ (unit: m^{-1} or cm^{-1})
Energy of one photon	$E = hv = hc/\lambda$ ($h = 6.626 \times 10^{-34}$ J·s)
Energy of n photons	$E = nhv$
Handy shortcut	E (eV) = 12400 ÷ λ (Å) or E (J) = 1.986×10^{-25} ÷ λ (m)
Photoelectric effect (Einstein)	$hv = hv_0 + \frac{1}{2}mv^2 \rightarrow KE_{max} = hv - W_0$ (W_0 = work function = hv_0)
Threshold condition	Electron ejected only if $v \geq v_0$ (higher intensity → more electrons, NOT higher energy)

2.3 Bohr's Model (H and H-like species: He^{+}, Li^{2+})

Quantity	Formula	Quick Values
Radius of nth orbit	$r_n = 52.9 \times n^2/Z$ pm	$r_1(H)$ = 52.9 pm = 0.529 Å
Energy of nth orbit	$E_n = -13.6 \times Z^2/n^2$ eV $= -2.18 \times 10^{-18} \times Z^2/n^2$ J	$E_1(H) = -13.6$ eV
Velocity of electron	$v_n = 2.18 \times 10^{6} \times Z/n$ m/s	$v_1(H) \approx c/137$
Angular momentum	$mvr = nh/2\pi$ (quantised)	n = 1, 2, 3 …
Ionisation energy	$IE = E_\infty - E_1 = +13.6\ Z^2$ eV	H: 13.6 eV; He^{+}: 54.4 eV
Energy gap	$\Delta E = E_2 - E_1 = 13.6\ Z^2 (1/n_1^2 - 1/n_2^2)$ eV	= hv of emitted photon

2.4 Hydrogen Spectrum — Rydberg Formula

Concept	Formula / Detail
Rydberg formula	$1/\lambda = R_H \times Z^2 (1/n_1^2 - 1/n_2^2)$; $R_H = 109677\ cm^{-1}$

Concept	Formula / Detail
Lyman series	$n_1 = 1$, $n_2 = 2,3,4\ldots$ → Ultraviolet region
Balmer series	$n_1 = 2$, $n_2 = 3,4,5\ldots$ → VISIBLE region (only visible series)
Paschen series	$n_1 = 3$ → Infrared
Brackett series	$n_1 = 4$ → Infrared
Pfund series	$n_1 = 5$ → Infrared
Max. spectral lines	From level n: total lines = $n(n-1)/2$; between n_2 and n_1: $(n_2-n_1)(n_2-n_1+1)/2$

2.5 Dual Nature & Heisenberg Uncertainty Principle

Concept	Formula / Expression
de Broglie wavelength	$\lambda = h/mv = h/p$
With kinetic energy	$\lambda = h \div \sqrt{(2m \cdot KE)}$
Charged particle (charge q, accelerated by V volts)	$\lambda = h \div \sqrt{(2mqV)}$; for electron: $\lambda = 12.27/\sqrt{V}$ Å
Heisenberg uncertainty principle	$\Delta x \cdot \Delta p \geq h/4\pi$ or $\Delta x \cdot \Delta v \geq h/(4\pi m)$
Significance	Meaningful only for microscopic particles; rules out fixed Bohr orbits → 'orbitals' (probability regions)

2.6 Quantum Numbers

Quantum Number	Symbol & Values	What It Tells
Principal	n = 1, 2, 3, … (K, L, M, N…)	Shell, size & energy of orbital
Azimuthal (angular)	l = 0 to (n−1) ; s=0, p=1, d=2, f=3	Subshell & shape of orbital
Magnetic	$m_l = -l \ldots 0 \ldots +l$ (total 2l+1 values)	Orientation of orbital
Spin	$m_s = +½$ or $-½$	Spin of electron (↑ or ↓)

Counting Formula	Expression
Orbitals in a shell	n^2 ; Max electrons in shell = $2n^2$
Orbitals in a subshell	2l + 1 ; Max electrons in subshell = 2(2l + 1) → s:2, p:6, d:10, f:14
Radial nodes	n − l − 1
Angular nodes	l ; Total nodes = n − 1
Orbital angular momentum	$= \sqrt{(l(l+1))} \times h/2\pi$
Spin angular momentum	$= \sqrt{(s(s+1))} \times h/2\pi$, s = ½
Spin magnetic moment	$\mu = \sqrt{(n(n+2))}$ BM (n = number of unpaired electrons)

2.7 Rules for Filling Electrons

Rule	Statement
Aufbau principle	Electrons fill orbitals in order of increasing energy: lower (n + l) first; if (n + l) equal → lower n first
Energy order	1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s
Pauli exclusion principle	No two electrons in an atom can have all four quantum numbers identical (max 2 e^- per orbital, opposite spins)
Hund's rule	Pairing starts only after each orbital of a subshell gets one electron (maximum multiplicity)
Exceptional configs	Cr (Z=24): $[Ar]3d^5 4s^1$; Cu (Z=29): $[Ar]3d^{10} 4s^1$ — extra stability of half/fully-filled subshells
Cation formation	Remove electrons from HIGHEST n first (4s before 3d): $Fe^{2+} = [Ar]3d^6$

★ QUICK MEMORY AIDS & TRICKS

Spectrum series order	'Lyman Bhai Pass Bhi Phir-se' → Lyman, Balmer, Paschen, Brackett, Pfund (n_1 = 1,2,3,4,5)
−13.6 eV table	E_n(H): n=1 → −13.6 ; n=2 → −3.4 ; n=3 → −1.51 ; n=4 → −0.85 eV (divide by n^2)
Lines shortcut	Drop from n=4 to 1 → 4(3)/2 = 6 lines.
Nodes	Radial = n − l − 1, Angular = l, Total = n − 1 → '3p: 1 radial + 1 angular = 2 total'.
de Broglie quick	Heavier or faster particle → smaller λ (why cricket balls don't show waves!)
12400 rule	E(eV) × λ(Å) = 12400 — fastest way for photon numericals.
Orbital capacity	s-2, p-6, d-10, f-14 → 'Spdf: 2, 6, 10, 14 — add 4 each time'.