JEE MAIN · PAPER 1 · CHEMISTRY

02 Atomic Structure

Formula & concept sheet

How the atom is built, from light and the photoelectric effect through de Broglie waves and the uncertainty principle, the Bohr model and the hydrogen spectrum, to orbitals, quantum numbers and the rules for filling electrons. Each idea is explained first, then the formulas, shortcuts and traps.

INSIDE THIS SHEET

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READ ALONGSIDE NCERT XI · Structure of Atom

HOW TO USE THIS SHEET

1 **Skim**
Read the Key ideas, then the formula cards.

2 **Memorise**
Learn the formulas and check every Trap box.

3 **Test**
Do the worked examples, then this chapter's PYQs.

4 **Revise**
Use the last-minute checklist before the exam.

CHAPTER 02

Atomic Structure

NCERT XI · Structure of Atom

Covers: electromagnetic radiation and the photoelectric effect · hydrogen spectrum · Bohr model · dual nature and de Broglie relation · Heisenberg uncertainty principle · quantum mechanical model, ψ and ψ² · quantum numbers · shapes of s, p and d orbitals · spin · Aufbau, Pauli and Hund · half-filled and filled stability.

1 Light as wave and as photons

- Light is an electromagnetic wave: electric and magnetic fields oscillating at right angles, with wavelength λ, frequency ν and speed $c = \lambda\nu = 3 \times 10^{8}$ m s^{-1} in vacuum. The **wavenumber** $\bar{\nu} = 1/\lambda$ counts waves per metre (or per cm).
- It also behaves as a stream of quanta, **photons**: energy comes in whole packets $E = h\nu$ (Planck), and a photon carries momentum $p = h/\lambda$ but has no rest mass. Shorter wavelength means a more energetic photon.

WAVE RELATION

$$c = \lambda\nu, \quad \bar{\nu} = \frac{1}{\lambda}$$

ν in Hz (s^{-1}); $\bar{\nu}$ in m^{-1} or cm^{-1}; 1 nm $= 10^{-9}$ m, 1 Å $= 10^{-10}$ m

PHOTON ENERGY

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

$h = 6.626 \times 10^{-34}$ J s; $E(\text{eV}) = \dfrac{1240}{\lambda(\text{nm})}$; N photons carry $Nh\nu$

PER MOLE OF PHOTONS

$$E_m = N_A\frac{hc}{\lambda}$$

$= \dfrac{1.196 \times 10^{5}}{\lambda(\text{nm})}$ kJ mol^{-1}; 1 eV per atom $= 96.485$ kJ mol^{-1}

ELECTROMAGNETIC SPECTRUM (INCREASING WAVELENGTH, DECREASING ENERGY)

REGION	WAVELENGTH (ROUGH)	NOTE
γ rays	< 0.01 nm	nuclear transitions
X rays	0.01 to 10 nm	inner-shell electrons
ultraviolet	10 to 400 nm	Lyman series of H
visible	400 to 750 nm	violet (400) to red (750); Balmer lines
infrared	750 nm to 1 mm	Paschen, Brackett, Pfund; molecular vibrations
microwave and radio	> 1 mm	rotations; radio waves

PLANCK Radiation is emitted or absorbed only in whole quanta, $E = nh\nu$ ($n = 1, 2, \ldots$), not continuously; this ends the failure of classical physics on black-body radiation. $h = 6.626 \times 10^{-34}$ J s.

NUMBER OF PHOTONS

$$N = \frac{P\lambda}{hc} \text{ per second}$$

a source of power P emitting wavelength λ; energy divided by the energy of one photon

PHOTON MOMENTUM

$$p = \frac{h}{\lambda} = \frac{E}{c}$$

a photon has no rest mass; it can still push on a surface

KEY IDEA

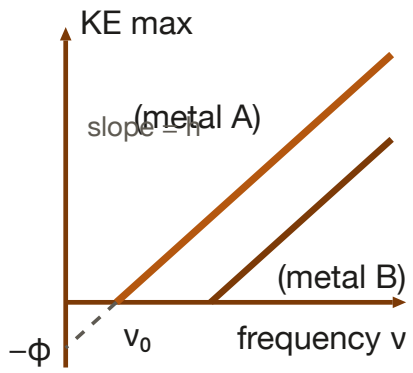
Frequency does not change when light enters a medium; wavelength and speed do. Colour is fixed by the frequency, and photon energy $h\nu$ is too.

QUICK CHECK

Energy of a photon of 400 nm in eV and in kJ per mole?

Answer: $\dfrac{1240}{400} = 3.1$ eV; $3.1 \times 96.485 \approx 299$ kJ mol^{-1}.

2 The photoelectric effect



slope is h for every metal; intercept on the energy axis is $-\phi$

PHOTOELECTRIC EFFECT Light of high enough frequency knocks electrons out of a metal surface. It cannot be explained by the wave picture (more intensity should eventually eject electrons at any frequency); Einstein explained it with photons: one photon gives its whole energy to one electron.

WHAT THE EXPERIMENT SHOWS

- No emission below the **threshold frequency** ν_0, however intense the light.
- Above ν_0 emission is instantaneous; KE_{max} grows with frequency, not with intensity.
- The **number** of photoelectrons (the current) is proportional to intensity.

EINSTEIN EQUATION

$$h\nu = \phi_0 + KE_{max}$$

work function $\phi_0 = h\nu_0 = \dfrac{hc}{\lambda_0}$ (eV: $\dfrac{1240}{\lambda_0 \text{ nm}}$)

STOPPING POTENTIAL

$$eV_0 = KE_{max}$$

$V_0 = \dfrac{h}{e}(\nu - \nu_0)$; the V_0 against ν line has slope h/e for every metal

TWO CASES OF ONE METAL

$$\frac{KE_1}{KE_2} = \frac{E_1 - \phi_0}{E_2 - \phi_0}$$

write $E - \phi_0$ for each, divide, and solve for ϕ_0

TRAP

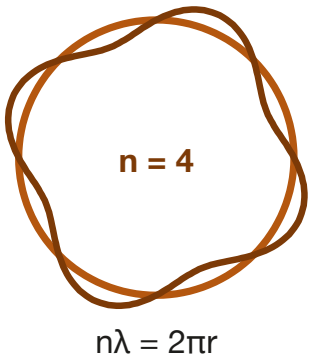
Doubling the intensity doubles the photocurrent, not KE_{max} or V_0. Doubling the frequency does **not** double KE_{max}: it is $h\nu - \phi_0$. Below threshold nothing comes out at any intensity.

QUICK CHECK

Light of 200 nm falls on a metal of threshold 300 nm. Stopping potential?

Answer: $\dfrac{1240}{200} - \dfrac{1240}{300} = 6.20 - 4.13 = 2.07$ eV, so $V_0 = 2.07$ V.

3 Dual nature of matter and the uncertainty principle



a Bohr orbit holds a whole number of wavelengths: $mvr = \dfrac{nh}{2\pi}$

DE BROGLIE RELATION Every moving particle has a wavelength $\lambda = h/p = h/mv$. It is large enough to matter only for tiny masses (electrons, atoms): a 0.1 kg ball at 30 m s^{-1} has $\lambda \approx 2 \times 10^{-34}$ m. This is what makes an electron in a Bohr orbit a standing wave: $2\pi r = n\lambda$.

DE BROGLIE WAVELENGTH

$$\lambda = \frac{h}{p} = \frac{h}{\sqrt{2m\,KE}}$$

SI units (kg, m s^{-1}, J); use the mass of the particle given, not the electron's

ELECTRON THROUGH V VOLTS

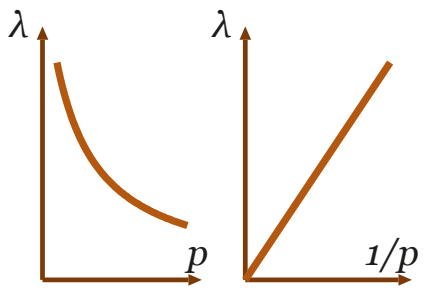
$$\lambda = \frac{h}{\sqrt{2meV}} \approx \frac{12.27}{\sqrt{V}}\ \text{Å}$$

$V = 100$ gives 1.23 Å

IN THE nTH BOHR ORBIT

$$\lambda_n = \frac{2\pi r_n}{n} = \frac{2\pi n\,a_0}{Z}$$

from $2\pi r_n = n\lambda$; H, $n = 2$: $4\pi a_0 = 6.65$ Å



λ against p λ against 1/p

$\lambda = h/p$

HEISENBERG UNCERTAINTY PRINCIPLE The position and momentum of a small particle cannot both be known exactly: $\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$. It is negligible for a cricket ball and decisive for an electron in an atom, which is why a fixed orbit (known r and v) cannot exist.

UNCERTAINTY

$$\Delta x\,\Delta p \geq \frac{h}{4\pi}, \quad \Delta v_{min} = \frac{h}{4\pi m\,\Delta x}$$

$\Delta x = 1$ Å for an electron: $\Delta v \approx 5.8 \times 10^5$ m s^{-1}

TRAP

The uncertainty bound uses $h/4\pi$, not $h/2\pi$. Convert Δx to metres and keep SI throughout. For protons and neutrons put their own mass in $\lambda = h/mv$.

QUICK CHECK

Which has the longer de Broglie wavelength at the same speed, an electron or a proton?

Answer: $\lambda \propto 1/m$, so the electron (about 1836 times longer).

4 The Bohr model and the hydrogen spectrum

BOHR'S POSTULATES The electron of a hydrogen-like species (nucleus Ze, one electron) moves in certain circular stationary orbits without radiating, those with angular momentum $mvr = \frac{nh}{2\pi}$ ($n = 1, 2, 3, \ldots$). It radiates or absorbs a photon only in a jump between orbits, with $h\nu = \Delta E$. It works only for one-electron species: H, He^+, Li^{2+}, Be^{3+}.

BOHR ORBITS: RADIUS AND ENERGY

$$\frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r^2}$$

Coulomb force supplies the centripetal force

$$mvr = \frac{nh}{2\pi}$$

quantisation of angular momentum

$$r_n = \frac{n^2}{Z}\,a_0, \quad a_0 = 0.529\ \text{Å} = 52.9\ \text{pm}$$

eliminate v

$$E_n = K + U = -K = -13.6\,\frac{Z^2}{n^2}\ \text{eV}$$

$K = \frac{Ze^2}{8\pi\varepsilon_0 r}$ and $U = -2K$

RADIUS

$$r_n = 52.9\,\frac{n^2}{Z}\ \text{pm}$$

grows as n^2; ground orbit of H: 52.9 pm

SPEED

$$v_n = 2.18 \times 10^6\,\frac{Z}{n}\ \text{m s}^{-1}$$

$\approx \frac{c}{137}\frac{Z}{n}$; falls as $1/n$

ENERGY

$$E_n = -13.6\,\frac{Z^2}{n^2}\ \text{eV}$$

$= -2.18 \times 10^{-18}\frac{Z^2}{n^2}$ J per atom $= -1312\frac{Z^2}{n^2}$ kJ mol^{-1}

ENERGY PARTS

$$KE = -E,\quad PE = 2E$$

so $KE : PE : E = 1 : -2 : -1$

SCALINGS

$$r \propto \frac{n^2}{Z},\ v \propto \frac{Z}{n},\ E \propto \frac{Z^2}{n^2}$$

take ratios: every constant cancels; period $\propto \frac{n^3}{Z^2}$, frequency $\propto \frac{Z^2}{n^3}$

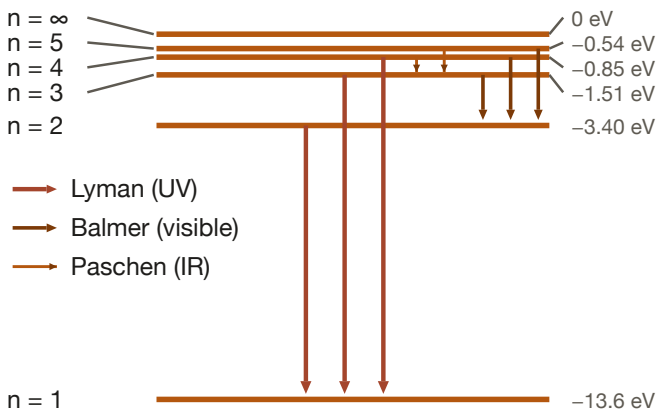
IONISATION ENERGY

$$IE_n = 13.6\,\frac{Z^2}{n^2}\ \text{eV}$$

H ground: 13.6 eV (1312 kJ mol^{-1}); He$^+$: 54.4 eV; H from $n = 2$: 3.4 eV

SOLVING A BOHR-MODEL QUESTION

1. Identify the species and its Z (H 1, He$^+$ 2, Li^{2+} 3, Be^{3+} 4); the model fails for anything with two or more electrons.
2. If a radius, speed or energy is given, divide by the hydrogen ground-state value to get $\frac{n^2}{Z}$, $\frac{Z}{n}$ or $\frac{Z^2}{n^2}$, then solve for whole-number n.
3. For a jump use $\Delta E = 13.6Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ eV, then $\lambda = \frac{1240}{\Delta E}$ nm; for a ratio of two lines the constants cancel.
4. Check units: eV per atom, J per atom, kJ per mole ($\times 96.485$ from eV); an energy **required** is positive.



lines of each series are the transitions to $n_1 = 1,\ 2,\ 3$: $\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n^2}\right)$

- A jump from a higher level n_2 to a lower n_1 emits a photon of $\Delta E = 13.6Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ eV.
- Lines ending on the same n_1 form a **series**; the first line (longest λ) has $n_2 = n_1 + 1$, the **series limit** (shortest λ) has $n_2 = \infty$.
- Absorption is the reverse: only the exact energy difference is taken up. A jump between degenerate orbitals ($2p_x \to 2p_y$) has $\Delta E = 0$ and gives no line.

RYDBERG FORMULA

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$R = 1.097 \times 10^7$ m^{-1} ($= 109677$ cm^{-1}), $n_2 > n_1$; $\lambda = \frac{1240}{\Delta E\ (\text{eV})}$ nm

LINES FROM LEVEL n

$$\frac{n(n-1)}{2}$$

all paths down to the ground state; between n_2 and n_1: $\frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$

SERIES LIMIT

$$\lambda_\infty = \frac{n_1^2}{RZ^2}$$

for an ion divide the hydrogen value by Z^2 (He$^+$ Lyman limit $91.2/4 = 22.8$ nm)

SERIES OF HYDROGEN

SERIES	n_1	REGION	FIRST LINE (NM)	LIMIT (NM)
Lyman	1	ultraviolet	121.5	91.2
Balmer	2	visible	656.3	364.6
Paschen	3	infrared	1875	820.4
Brackett	4	infrared	4051	1459
Pfund	5	infrared	7458	2279

The first Balmer line (H_α, red) is the $3 \to 2$ jump; the $1 \to 2$ excitation of H needs 10.2 eV (984 kJ mol^{-1}).

WHY BOHR SUCCEEDED

- Explains the sharp lines of H, He^+, Li^{2+} and the Rydberg formula.
- Gives the radius and ionisation energy of one-electron species.

WHY IT FAILS

- Multi-electron atoms (He, H^-): electron repulsion is ignored.
- Fine structure, Zeeman and Stark splitting; violates the uncertainty principle (exact r and v); no orbital shapes or bonding.

TRAP

$r \propto n^2/Z$ but $v \propto Z/n$ and $E \propto Z^2/n^2$. $n = 2$ is the **first** excited state. Levels are negative: "energy required" is the magnitude. Forgetting Z^2 in $1/\lambda$ for He^+ and Li^{2+} is the commonest slip.

KEY IDEA

Same-series lines of H: $\dfrac{\lambda_{4\to2}}{\lambda_{3\to2}} = \dfrac{1/4 - 1/9}{1/4 - 1/16} = \dfrac{20}{27}$, so 656 nm becomes 486 nm, without R. He^+ $4 \to 2$ equals H $2 \to 1$ in wavelength, since $4 \times \dfrac{3}{16} = \dfrac{3}{4}$.

QUICK CHECK

Which Be^{3+} orbit has the same radius as the first H orbit?

Answer: $r \propto \dfrac{n^2}{Z}$: $\dfrac{n^2}{4} = 1$ gives $n = 2$.

5 Quantum mechanical model, quantum numbers and orbitals

- The electron is a wave. Solving the Schrödinger equation gives a wave function ψ for each allowed energy; ψ itself has no meaning, but ψ^2 at a point is the **probability density** of finding the electron there. An **orbital** is a region of high probability, not a path; each holds at most two electrons.
- Every orbital is named by three quantum numbers that come out of the mathematics (n, l, m_l); the fourth, spin, is added from experiment.

THE FOUR QUANTUM NUMBERS

NUMBER	NAME	ALLOWED VALUES	TELLS
n	principal	$1, 2, 3, \ldots$	shell, size and energy
l	azimuthal	$0, 1, \ldots, n-1$ (s, p, d, f)	subshell, shape; $L = \sqrt{l(l+1)}\,\dfrac{h}{2\pi}$
m_l	magnetic	$-l, \ldots, 0, \ldots, +l$ ($2l+1$ values)	orientation: number of orbitals in the subshell
m_s	spin	$+\frac{1}{2}$, $-\frac{1}{2}$	spin direction of the electron

SHELLS AND SUBSHELLS

SHELL	n	SUBSHELLS	ORBITALS (n^2)	ELECTRONS ($2n^2$)
K	1	1s	1	2
L	2	2s 2p	4	8
M	3	3s 3p 3d	9	18
N	4	4s 4p 4d 4f	16	32

Orbitals in a subshell $2l + 1$ (s 1, p 3, d 5, f 7); electrons $2(2l + 1)$ $(2, 6, 10, 14)$.

NODES

$\text{radial} = n - l - 1, \quad \text{angular} = l$

total $n - 1$; 2s has 1 radial node, 3p has 1 radial and 1 angular

PROBABILITY

ψ^2 and $4\pi r^2\psi^2$

ψ^2 is per unit volume; $4\pi r^2\psi^2$ is the radial distribution (per unit r)

SPIN

$m_s = \pm\frac{1}{2}$

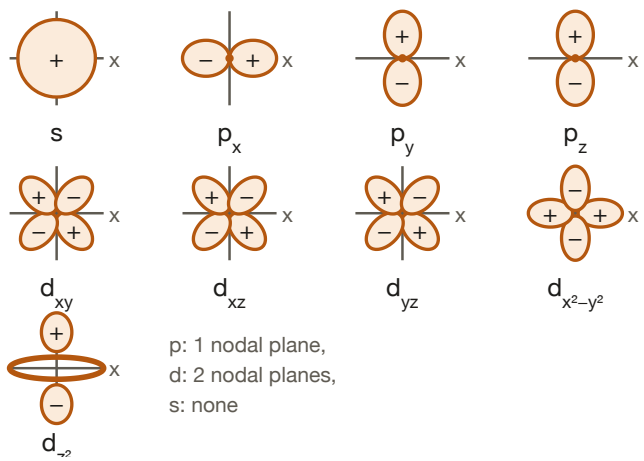
two electrons in an orbital must have opposite spins; spin is not a solution of the Schrödinger equation

ORBIT (BOHR)

- A definite circular path around the nucleus at fixed r and v.
- Ruled out by the uncertainty principle; n alone fixes it; holds $2n^2$ electrons in a shell.

ORBITAL (QUANTUM MECHANICS)

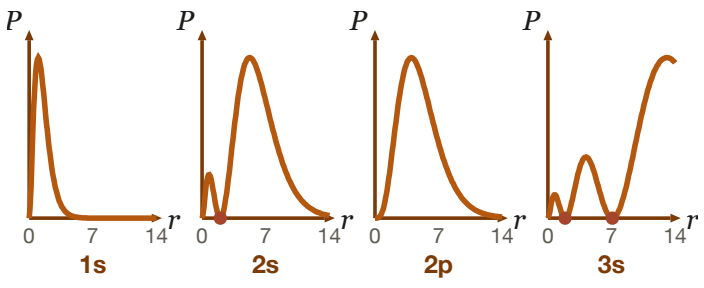
- A region where ψ^2 is large, with a shape fixed by l and m_l and no definite path.
- Described by n, l, m_l; holds at most 2 electrons of opposite spin.



angular shapes; lobes carry the sign of the wave function, nodes separate them

READING THE SHAPES

- **s**: a sphere, no angular node, same in every direction. **p**: two lobes along an axis, one nodal plane through the nucleus (3 orientations).
- **d**: four lobes (two nodal planes) for d_{xy}, d_{xz}, d_{yz} (between the axes) and $d_{x^2-y^2}$ (on the x and y axes); d_{z^2} has two lobes on the z axis and a ring in the xy plane (5 orientations).
- The **sign** of ψ changes across a node; ψ^2, the probability, is always positive.
- Same shape, bigger size and more nodes as n grows ($2s$ is a larger sphere than $1s$).



number of radial nodes n – l – 1; angular nodes l; total n – 1

RADIAL DISTRIBUTION

- The curve of $4\pi r^2\psi^2$ against r gives the chance of finding the electron at distance r in any direction; each orbital has $n - l$ peaks.
- $1s$ peaks at $a_0 = 52.9$ pm; $2s$ has a small inner peak, a node and a big outer peak; $3s$ has two nodes. An outer s electron therefore has some probability close to the nucleus (**penetration**), more than p or d of the same shell: the reason for $s < p < d < f$ energy order in a shell.
- $2p$ and $3d$ have no radial node ($n - l - 1 = 0$).

TRAP

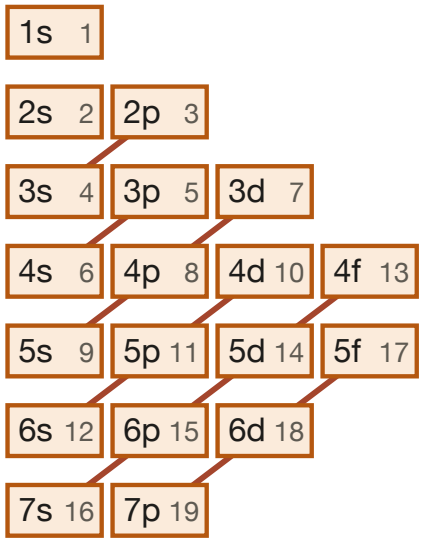
Check validity: $l < n$, $|m_l| \le l$, $m_s = \pm\frac{1}{2}$. $(n = 2, l = 2)$ is impossible (no 2d). In hydrogen the energy depends on n only: $3s = 3p = 3d$; in many-electron atoms it depends on $n + l$.

QUICK CHECK

How many electrons in an atom can have $n = 3, l = 1$?

Answer: The $3p$ subshell: 3 orbitals, so 6 electrons.

6 Filling electrons: configurations and magnetism



numbers give the order of filling; each diagonal has one value of n + l

THE THREE RULES

- **Aufbau**: fill the lowest energy first, in order of rising $(n + l)$; if two have the same $(n + l)$, the lower n first.
- **Pauli**: no two electrons share all four quantum numbers, so an orbital holds at most two, of opposite spin.
- **Hund**: degenerate orbitals are filled singly with parallel spins before any pairing.

ORDER OF FILLING

- 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p.
- 4s fills before 3d but is also the **first to be lost** on ionisation.

FIRST-ROW D-BLOCK: CONFIGURATIONS AND UNPAIRED ELECTRONS

ELEMENT	SC	TI	V	CR	MN	FE	CO	NI	CU	ZN
3d / 4s	1 / 2	2 / 2	3 / 2	5 / 1	5 / 2	6 / 2	7 / 2	8 / 2	10 / 1	10 / 2
unpaired	1	2	3	6	5	4	3	2	1	0

The two exceptions in the first row: Cr ($3d^5\ 4s^1$) and Cu ($3d^{10}\ 4s^1$), both gaining the extra stability of a half-filled or filled d subshell (more exchange energy, more symmetry).

OTHER EXCEPTIONS TO REMEMBER

ELEMENT	EXPECTED	ACTUAL	UNPAIRED
Nb (41)	$4d^3\,5s^2$	$4d^4\,5s^1$	5
Mo (42)	$4d^4\,5s^2$	$4d^5\,5s^1$	6
Ru (44)	$4d^6\,5s^2$	$4d^7\,5s^1$	4
Rh (45)	$4d^7\,5s^2$	$4d^8\,5s^1$	3
Pd (46)	$4d^8\,5s^2$	$4d^{10}\,5s^0$	0
Ag (47)	$4d^9\,5s^2$	$4d^{10}\,5s^1$	1
Pt (78)	$5d^8\,6s^2$	$5d^9\,6s^1$	2
Au (79)	$5d^9\,6s^2$	$5d^{10}\,6s^1$	1

IONS: TAKE ns FIRST

ION	CONFIGURATION	UNPAIRED
Fe^{2+}	$[Ar]\,3d^6$	4
Fe^{3+}, Mn^{2+}	$[Ar]\,3d^5$	5
Co^{2+}	$3d^7$	3
Ni^{2+}	$3d^8$	2
Cu^{2+}	$3d^9$	1
Cu^{+}, Zn^{2+}	$3d^{10}$	0
Cr^{3+}	$3d^3$	3
Ti^{3+}	$3d^1$	1

SPIN-ONLY MAGNETIC MOMENT

UNPAIRED n	M (BM)
1	1.73
2	2.83
3	3.87
4	4.90
5	5.92
6	6.93

$\mu = \sqrt{n(n+2)}$ BM; a species with no unpaired electron is **diamagnetic**, otherwise paramagnetic.

UNPAIRED ELECTRONS IN GROUND-STATE ATOMS

ATOM	OUTER CONFIGURATION	UNPAIRED
B, Al	ns^2np^1	1
C, Si	ns^2np^2	2
N, P	ns^2np^3	3
O, S	ns^2np^4	2
F, Cl	ns^2np^5	1
Ne, Ar	ns^2np^6	0

Hund: the three p orbitals fill singly before pairing, so N has 3 and O only 2.

ENERGY ORDER FROM $n + l$

SUBSHELL	$n + l$	RANK
5s	5	1
4d	6	2 (lower n than 5p, 6s)
5p	6	3
6s	6	4
4f	7	5

Same $n + l$: the lower n is lower in energy (4d $<$ 5p $<$ 6s).

KEY IDEA

To count unpaired electrons: write the configuration (for an ion remove the highest-n electrons first), draw one box per orbital, place electrons singly first (Hund), then count the single arrows.

TRAP

Fe^{2+} is $3d^6$, not $3d^4\,4s^2$. Pd has **no** $5s$ electron and no unpaired electron. The $(n + l)$ order applies to many-electron atoms only.

7 Worked examples

PHOTON IN HE^+

Wavelength of the photon emitted when He^+ falls from $n = 2$ to $n = 1$?

$\Delta E = 13.6 \times 4 \times \left(1 - \frac{1}{4}\right) = 40.8$ eV; $\lambda = \frac{1240}{40.8} = 30.4$ nm

IDENTIFY THE ORBIT

The radius of a Bohr orbit of He^+ is 105.8 pm ($a_0 = 52.9$ pm). Which orbit is it?

$r = \frac{n^2}{2} \times 52.9 = 105.8 \Rightarrow n^2 = 4, n = 2$

IONISATION OF BE^{3+}

Ionisation energy of Be^{3+} from its ground state in kJ mol^{-1}?

$Z = 4$: $13.6 \times 16 = 217.6$ eV; $217.6 \times 96.485 = 2.10 \times 10^4$ kJ mol^{-1}

EXCITATION OF H

Energy to excite one mole of H atoms from $n = 1$ to $n = 2$?

10.2 eV per atom; $10.2 \times 96.485 \approx 984$ kJ mol^{-1}

SERIES LIMIT

Balmer series limit of H ($R = 1.097 \times 10^7$ m^{-1})?

$\lambda = \dfrac{4}{R} = \dfrac{4}{1.097 \times 10^7} = 364.6$ nm

COUNTING LINES

Lines in the spectrum from atoms excited to $n = 5$ (all paths), and how many are in the Balmer series?

$\dfrac{5 \times 4}{2} = 10$ lines; Balmer lines end on $n = 2$ from $n = 3, 4, 5$: 3

DE BROGLIE WAVELENGTH

Wavelength of an electron accelerated through 100 V?

$\dfrac{12.27}{\sqrt{100}} = 1.23$ Å

UNCERTAINTY

Minimum uncertainty in the speed of an electron confined within 1×10^{-10} m?

$\Delta v = \dfrac{6.626 \times 10^{-34}}{4\pi \times 9.109 \times 10^{-31} \times 10^{-10}} = 5.8 \times 10^5$ m s^{-1}

ORBITAL ELECTRONS

How many electrons of a Cr atom have $m_l = 0$?

Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$. The four s subshells give $2 + 2 + 2 + 1 = 7$; $2p$ and $3p$ each put 2 in the $m_l = 0$ orbital: 4; $3d$ puts 1. Total 12

MAGNETIC MOMENT

Which has the larger spin-only moment, Fe^{2+} or Mn^{2+}?

Fe^{2+}: $3d^6$, 4 unpaired, $\sqrt{24} = 4.90$ BM; Mn^{2+}: $3d^5$, 5 unpaired, $\sqrt{35} = 5.92$ BM: Mn^{2+}

ANGULAR MOMENTUM

Orbital angular momentum of a $3d$ electron, in units of $\dfrac{h}{2\pi}$?

$l = 2$: $L = \sqrt{l(l+1)} = \sqrt{6}\,\dfrac{h}{2\pi} \approx 2.45\,\dfrac{h}{2\pi}$

ENERGY ORDER

Arrange $4d$, $5s$, $6s$, $4f$ in increasing energy.

$n + l$: $5s = 5$, $4d = 6$, $6s = 6$, $4f = 7$; tie $4d$ and $6s$ goes to the lower n: $5s < 4d < 6s < 4f$

BOHR ORBIT WAVELENGTH

de Broglie wavelength of the electron in the third orbit of H ($a_0 = 52.9$ pm)?

$r_3 = 9a_0 = 476$ pm; $\lambda = \dfrac{2\pi r_3}{3} = 997$ pm $= 9.97$ Å

THRESHOLD FREQUENCY

Threshold frequency of a metal with work function 4.13 eV?

$\nu_0 = \dfrac{\phi_0}{h} = \dfrac{4.13 \times 1.602 \times 10^{-19}}{6.626 \times 10^{-34}} = 1.0 \times 10^{15}$ Hz

8 Last-minute checklist

BEFORE YOU ANSWER

Which species (Z) and which levels (n_1, n_2)? Per atom or per mole? eV, J or erg? Is it a one-electron species at all?

- $E = h\nu = hc/\lambda$; $E(\text{eV}) = 1240/\lambda(\text{nm})$; $KE_{max} = h\nu - \phi_0$; $eV_0 = KE_{max}$; 1 eV per atom $= 96.485$ kJ mol^{-1}.
- $\lambda = h/mv$; $\Delta x\,\Delta p \geq h/4\pi$; Bohr: $2\pi r = n\lambda$, $mvr = nh/2\pi$.
- $r \propto n^2/Z$, $v \propto Z/n$, $E = -13.6Z^2/n^2$ eV; $KE = -E$, $PE = 2E$.
- $1/\lambda = RZ^2(1/n_1^2 - 1/n_2^2)$; lines $n(n-1)/2$; first line $n_2 = n_1 + 1$, limit $n_2 = \infty$; limit for an ion = H value$/Z^2$.
- $l = 0..n-1$, $m_l = -l..+l$, orbitals n^2, electrons $2n^2$; radial nodes $n - l - 1$, angular l.
- Aufbau $(n + l)$, Pauli, Hund; Cr, Cu (and Nb, Mo, Ru, Rh, Pd, Ag, Pt, Au) exceptions; ions lose ns first; $\mu = \sqrt{n(n+2)}$ BM.

Now test it. Practise this chapter with past-paper and bank questions on Prep Vault. **Practice this chapter →**