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PHYSICAL

CH 08

# Atomic Structure — Bohr Model, Quantum Numbers & Orbitals

Subatomic particles, electromagnetic radiation and the photoelectric effect, the Bohr model and hydrogen spectrum, de Broglie and Heisenberg, quantum numbers, orbital shapes and nodes, and electronic configuration — with every numerical type solved, an exam trap gallery, graded MCQs and a one-page revision sheet.

## WHAT YOU WILL BE ABLE TO DO

- Relate frequency, wavelength, wavenumber and energy, and solve photoelectric-effect problems
- Use Bohr's formulas for radius, energy and velocity of hydrogen-like species
- Calculate wavelengths of spectral lines and the number of lines from a given level
- Apply de Broglie's relation and Heisenberg's uncertainty principle
- Assign quantum numbers, count nodes, and write configurations including the exceptions

## INSIDE

A The big picture
B Subatomic particles and early models
C Electromagnetic radiation and quanta
D The Bohr model

HOW TO USE THESE NOTES

| | | |
|---|---|---|
| **IN ONE LINE** The core idea of a section | **KEY IDEA** The concept to remember | **EXAM TRAP** Where marks are usually lost |
| **SOLVED EXAMPLE** Worked step by step | **CHECK YOURSELF** Quick self-test | **MEMORY HOOK** A line that sticks |

SECTION A

# The big picture

A

*From particles to waves.*

**IN ONE LINE**

Electrons are both particles and waves: Bohr's fixed orbits explain the hydrogen spectrum, but only the quantum-mechanical model of orbitals and quantum numbers explains configurations and the periodic table.

| Question type | Formula |
|---|---|
| Light | $c = \nu\lambda$; $\bar{\nu} = 1/\lambda$; $E = h\nu = hc/\lambda$ |
| Photoelectric effect | $h\nu = h\nu_0 + \frac{1}{2}mv^2$ |
| Bohr (H-like) | $r_n = 52.9\ n^2/Z$ pm; $E_n = -13.6\ Z^2/n^2$ eV; $v_n = 2.18 \times 10^6\ Z/n$ m s$^{-1}$ |
| Spectral line | $1/\lambda = R_h Z^2 (1/n_1^2 - 1/n_2^2)$; $R_h = 109\ 677$ cm$^{-1}$ |
| Matter waves | $\lambda = h/mv = h/p$ |
| Uncertainty | $\Delta x \cdot \Delta p \geq h/4\pi$ |

**THINK OF IT LIKE THIS**

Think of an electron in an atom as a guitar string fixed at both ends. Only certain wavelengths fit — those are the allowed energy levels. Pluck it harder and it jumps to a higher "note"; when it falls back, it gives out light of exactly the energy difference. That is why each element has its own line spectrum, like a fingerprint.

SECTION B

# Subatomic particles and early models

B

*What the atom is made of.*

| Particle | Charge / C | Mass / kg | Discovered by |
|---|---|---|---|
| Electron | $-1.602 \times 10^{-19}$ | $9.109 \times 10^{-31}$ | J. J. Thomson (cathode rays, e/m = $1.758820 \times 10^{11}$ C kg$^{-1}$) |
| Proton | $+1.602 \times 10^{-19}$ | $1.673 \times 10^{-27}$ | Goldstein (canal rays) |

| Particle | Charge / C | Mass / kg | Discovered by |
|---|---|---|---|
| Neutron | 0 | $1.675 \times 10^{-27}$ | Chadwick (1932) |

- **Millikan's oil-drop experiment** measured the charge of the electron ($1.602 \times 10^{-19}$ C); with e/m this gave its mass.
- **Thomson's model** (plum pudding): positive sphere with embedded electrons — could not explain α-particle scattering.
- **Rutherford's α-scattering:** most α particles passed straight through gold foil; a few were deflected; about 1 in 20 000 bounced back. So the atom is mostly empty, with a tiny, dense, positive **nucleus** (radius about $10^{-15}$ m, atom about $10^{-10}$ m).
- **Failure of Rutherford's model:** an orbiting charge should radiate energy and spiral into the nucleus; it also could not explain line spectra.
- **Isotopes** (same Z, different A: $^{1}H$, $^{2}H$, $^{3}H$); **isobars** (same A, different Z: $^{14}C$, $^{14}N$); **isotones** (same number of neutrons: $^{14}C$, $^{16}O$).

SECTION C

# Electromagnetic radiation and quanta

*Light as a wave and as a stream of photons.*

$$c = \nu \lambda \quad \bar{\nu} = 1/\lambda \quad E = h\nu = hc/\lambda$$

$c = 3.0 \times 10^{8}$ m $s^{-1}$; $h = 6.626 \times 10^{-34}$ J s

- Order of wavelength: radio > microwave > infrared > visible (400–750 nm) > ultraviolet > X-rays > γ-rays. Energy increases in the opposite order.
- **Planck's quantum theory:** energy is emitted or absorbed in discrete packets (quanta) of $E = h\nu$. For light these packets are photons.
- **Black-body radiation** — the shift of peak emission to shorter wavelength at higher temperature — could not be explained by wave theory; Planck's quanta explained it.

## Photoelectric effect

$$h\nu = h\nu_0 + \tfrac{1}{2} m v^2$$

$h\nu_0$ = work function (threshold energy); no electrons are emitted below $\nu_0$, however intense the light

- Electrons are emitted **instantly** when $\nu > \nu_0$.
- The **number** of electrons depends on intensity; their **kinetic energy** depends only on frequency.
- Alkali metals (K, Cs) have low work functions and are used in photocells.

**SOLVED EXAMPLE 01** — PHOTONS IN A BEAM

How many photons of wavelength 500 nm carry 1.0 J of energy?

SOLUTION

1. $E$ per photon $= hc/\lambda = (6.626 \times 10^{-34} \times 3.0 \times 10^{8}) / (5.00 \times 10^{-7}) = 3.98 \times 10^{-19}$ J.
2. $N = 1.0 / 3.98 \times 10^{-19}$.

ANSWER About $2.5 \times 10^{18}$ photons.

**SOLVED EXAMPLE 02** — PHOTOELECTRON ENERGY

The threshold frequency of a metal is $7.0 \times 10^{14}\ s^{-1}$. Find the kinetic energy of electrons emitted by light of frequency $1.0 \times 10^{15}\ s^{-1}$.

SOLUTION

1. $KE = h(\nu - \nu_0) = 6.626 \times 10^{-34} \times (10.0 - 7.0) \times 10^{14}$.

ANSWER $KE \approx 1.99 \times 10^{-19}$ J.

SECTION D

D

# The Bohr model

*Fixed orbits with quantised angular momentum.*

- Electrons move in circular orbits (stationary states) of fixed energy without radiating.
- **Angular momentum is quantised:** $mvr = nh/2\pi$.
- Energy is absorbed or emitted only when an electron jumps between orbits: $\Delta E = h\nu$.
- Valid only for **one-electron species**: H, $He^{+}$, $Li^{2+}$, $Be^{3+}$.

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

$$E_n = -2.18 \times 10^{-18}\ Z^2/n^2\ \text{J} = -13.6\ Z^2/n^2\ \text{eV}$$

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

Radius $\propto n^2/Z$; energy $\propto -Z^2/n^2$; velocity $\propto Z/n$

KEY IDEA

**Energy is negative** because the electron is bound: zero energy means a free electron at infinity. Ionisation energy of H = +13.6 eV = $2.18 \times 10^{-18}$ J per atom = 1312 kJ $mol^{-1}$. Kinetic energy = −E; potential energy = 2E.

**SOLVED EXAMPLE 03** — BOHR RADIUS AND ENERGY OF $He^+$

Find the radius and energy of the first orbit of $He^+$.

SOLUTION

1. Z = 2, n = 1.
2. $r = 52.9 \times 1^2/2 = 26.45$ pm.
3. $E = -13.6 \times 2^2/1^2 = -54.4$ eV.

ANSWER r = 26.45 pm; E = −54.4 eV.

**SOLVED EXAMPLE 04** — IONISATION ENERGY OF $Li^{2+}$

Find the energy needed to remove the electron from $Li^{2+}$ in its ground state.

SOLUTION

1. Z = 3, n = 1: $E_1 = -13.6 \times 9 = -122.4$ eV.
2. Ionisation energy = $0 - E_1$.

ANSWER 122.4 eV per ion.

### Limitations

- Cannot explain spectra of multi-electron atoms.
- Cannot explain splitting of lines in magnetic (Zeeman) or electric (Stark) fields.
- Ignores the dual nature of the electron and contradicts the uncertainty principle (exact orbits).
- Cannot explain chemical bonding or molecular shape.

SECTION E

E

## The hydrogen spectrum

*Each line is one jump between two levels.*

$$\bar{\nu} = 1/\lambda = R_h Z^2 (1/n_1^2 - 1/n_2^2)$$

$R_h$ = 109 677 $cm^{-1}$ (Rydberg constant); $n_2 > n_1$

| Series | $n_1$ (lower level) | $n_2$ | Region |
|---|---|---|---|
| Lyman | 1 | 2, 3, 4… | Ultraviolet |
| Balmer | 2 | 3, 4, 5… | Visible |
| Paschen | 3 | 4, 5, 6… | Infrared |
| Brackett | 4 | 5, 6… | Infrared |

| Series | $n_1$ (lower level) | $n_2$ | Region |
|---|---|---|---|
| Pfund | 5 | 6, 7… | Infrared |

- **Number of lines** when an electron falls from level n to the ground state (many atoms): $n(n-1)/2$. Between levels $n_2$ and $n_1$: $(n_2 - n_1)(n_2 - n_1 + 1)/2$.
- A single atom with one electron can give at most $n - 1$ lines in one cascade.
- The **series limit** ($n_2 = \infty$) gives the shortest wavelength of a series.

**SOLVED EXAMPLE 05** — THE RED BALMER LINE

Find the wavelength of the line for $n = 3 \rightarrow n = 2$ in hydrogen.

SOLUTION

1. $\bar{\nu} = 109\,677 \times (1/4 - 1/9) = 109\,677 \times 5/36 = 15\,233\ \text{cm}^{-1}$.
2. $\lambda = 1 / 15\,233\ \text{cm} = 6.565 \times 10^{-5}\ \text{cm}$.

ANSWER $\lambda \approx 656.5$ nm (red, the Hα line).

**SOLVED EXAMPLE 06** — COUNTING LINES

Electrons in a sample of hydrogen atoms are excited to $n = 5$. How many lines can appear in the emission spectrum, and how many are in the Balmer series?

SOLUTION

1. Total $= 5 \times 4 / 2 = 10$.
2. Balmer lines end on $n = 2$: from 5, 4 and 3 $\rightarrow$ 3 lines.

ANSWER 10 lines; 3 in the Balmer series.

**SOLVED EXAMPLE 07** — SAME LINE FROM $He^+$

Which transition in $He^+$ has the same wavelength as the $n = 2 \rightarrow 1$ transition in hydrogen?

SOLUTION

1. For H: $Z^2 (1/1 - 1/4) = 3/4$.
2. For $He^+$ ($Z^2 = 4$): $4 (1/n_1^2 - 1/n_2^2) = 3/4 \rightarrow 1/n_1^2 - 1/n_2^2 = 3/16$.
3. $n_1 = 2$, $n_2 = 4$ gives $1/4 - 1/16 = 3/16$.

ANSWER $n = 4 \rightarrow n = 2$ in $He^+$.

CHECK YOURSELF

C1 Which series of hydrogen lies in the visible region?

C2 How many spectral lines appear when electrons fall from $n = 4$ to $n = 1$?

*Answers at the end of the chapter.*

SECTION F

# Dual nature and uncertainty

F

*Electrons behave as waves too.*

$$\lambda = h / mv = h / p$$

de Broglie relation; significant only for very small masses

$$\Delta x \cdot \Delta p \geq h / 4\pi \quad (\Delta x \cdot m\,\Delta v \geq h / 4\pi)$$

Heisenberg's uncertainty principle: position and momentum cannot both be known exactly

- Electron diffraction (Davisson–Germer) confirmed the wave nature; the electron microscope uses it.
- For everyday objects λ is far too small to detect, and the uncertainty is negligible.
- The uncertainty principle rules out exact Bohr orbits — only probabilities (orbitals) make sense.
- In Bohr's orbit, $2\pi r = n\lambda$: the circumference holds a whole number of de Broglie waves.

**SOLVED EXAMPLE 08** — DE BROGLIE WAVELENGTH

Compare the de Broglie wavelengths of (a) an electron moving at $1.0 \times 10^{6}$ m s$^{-1}$ and (b) a 0.10 kg ball at 10 m s$^{-1}$.

SOLUTION

1. (a) $\lambda = 6.626 \times 10^{-34} / (9.109 \times 10^{-31} \times 1.0 \times 10^{6}) = 7.27 \times 10^{-10}$ m.
2. (b) $\lambda = 6.626 \times 10^{-34} / (0.10 \times 10) = 6.6 \times 10^{-34}$ m.

ANSWER (a) 727 pm — comparable to atomic sizes, so it matters. (b) $6.6 \times 10^{-34}$ m — undetectable.

**SOLVED EXAMPLE 09** — UNCERTAINTY IN VELOCITY

The position of an electron is known to within 0.1 nm. Find the minimum uncertainty in its velocity.

SOLUTION

1. $\Delta v \geq h / (4\pi\, m\, \Delta x) = 6.626 \times 10^{-34} / (4 \times 3.142 \times 9.109 \times 10^{-31} \times 1.0 \times 10^{-10})$.
2. $= 6.626 \times 10^{-34} / 1.145 \times 10^{-39}$.

ANSWER $\Delta v \geq 5.8 \times 10^{5}$ m s$^{-1}$ — huge, so an exact orbit is meaningless.

SECTION G

# Quantum numbers and orbitals

G

*Four numbers describe each electron.*

| Quantum number | Symbol | Values | Describes |
|---|---|---|---|
| Principal | n | 1, 2, 3… | Shell: size and energy |
| Azimuthal | l | 0 to n − 1 (s, p, d, f = 0, 1, 2, 3) | Subshell: shape; orbital angular momentum $\sqrt{[l(l+1)]}\ h/2\pi$ |
| Magnetic | $m_l$ | −l to +l (2l + 1 values) | Orientation of the orbital |
| Spin | $m_s$ | +½, −½ | Spin of the electron |

- Orbitals in shell n: $n^2$; electrons in shell n: $2n^2$; orbitals in a subshell: 2l + 1; electrons in a subshell: 2(2l + 1).
- **Shapes:** s spherical; p dumbbell (px, py, pz along the three axes); d cloverleaf (dxy, dyz, dxz between axes; $dx^2-y^2$ on the axes; $dz^2$ a dumbbell with a ring).
- **Nodes:** radial nodes = n − l − 1; angular nodes (nodal planes) = l; total nodes = n − 1.
- 1s has no node; 2s has one radial node; 2p has one nodal plane; 3d has two nodal planes and no radial node.

| Orbital | n | l | Radial nodes | Angular nodes |
|---|---|---|---|---|
| 1s | 1 | 0 | 0 | 0 |
| 2s | 2 | 0 | 1 | 0 |
| 2p | 2 | 1 | 0 | 1 |
| 3s | 3 | 0 | 2 | 0 |
| 3p | 3 | 1 | 1 | 1 |
| 3d | 3 | 2 | 0 | 2 |
| 4f | 4 | 3 | 0 | 3 |

**SOLVED EXAMPLE 10** — VALID QUANTUM NUMBERS

Which set is not possible: (a) n = 3, l = 2, $m_l$ = −2; (b) n = 2, l = 2, $m_l$ = 0; (c) n = 4, l = 0, $m_l$ = 0?

SOLUTION

1. l must be less than n.
2. (b) has l = n = 2.

ANSWER (b) is impossible; (a) is a 3d orbital and (c) a 4s orbital.

CHECK YOURSELF

C3 How many radial nodes does a 4p orbital have?

C4 What is the maximum number of electrons with n = 3?

*Answers at the end of the chapter.*

SECTION H

# Filling orbitals

H

*Aufbau, Pauli and Hund.*

- **Aufbau principle:** orbitals fill in order of increasing energy, given by the **(n + l) rule** — lower n + l first; if equal, lower n first. 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 have all four quantum numbers the same → at most two electrons per orbital, with opposite spins.
- **Hund's rule:** in degenerate orbitals, electrons occupy separate orbitals with parallel spins before pairing.
- **Exceptions:** Cr is [Ar] $3d^5\ 4s^1$ and Cu is [Ar] $3d^{10}\ 4s^1$ — half-filled and filled subshells are extra stable because of **symmetry and larger exchange energy** (more pairs of parallel-spin electrons can exchange positions).

**SOLVED EXAMPLE 11** — EXCHANGE ENERGY

Why is $3d^5\ 4s^1$ (Cr) more stable than $3d^4\ 4s^2$?

SOLUTION

1. In $3d^5$ all five electrons have parallel spins: the number of exchange pairs is 5 × 4 / 2 = 10.
2. In $3d^4$ there are only 4 × 3 / 2 = 6 exchange pairs.
3. More exchanges → more stabilisation; the half-filled set is also symmetrical.

ANSWER $3d^5\ 4s^1$ gains extra exchange energy and symmetry.

- **Magnetic behaviour** from configuration: unpaired electrons → paramagnetic; μ = √[n(n + 2)] BM.

SECTION I

# Advanced problems

I

*Each needs two ideas from the chapter.*

### SOLVED EXAMPLE 12 — ENERGY TO WAVELENGTH

How much energy is needed to excite an electron in hydrogen from n = 1 to n = 2, and what wavelength of light does this?

SOLUTION

1. $\Delta E = 2.18 \times 10^{-18} (1 - 1/4) = 1.635 \times 10^{-18}$ J.
2. $\lambda = hc/\Delta E = (6.626 \times 10^{-34} \times 3.0 \times 10^{8}) / 1.635 \times 10^{-18}$.

ANSWER $\Delta E = 1.64 \times 10^{-18}$ J; $\lambda \approx 1.22 \times 10^{-7}$ m = 122 nm (ultraviolet, first Lyman line).

### SOLVED EXAMPLE 13 — RATIO OF RADII AND ENERGIES

Find the ratio of (a) the radius of the third orbit of $Li^{2+}$ to the first orbit of H, and (b) the energy of the second orbit of $He^{+}$ to the first orbit of H.

SOLUTION

1. (a) $r \propto n^2/Z$: (9/3) / (1/1) = 3.
2. (b) $E \propto Z^2/n^2$: (4/4) / (1/1) = 1.

ANSWER (a) 3 : 1. (b) 1 : 1 — the n = 2 level of $He^{+}$ has the same energy as the n = 1 level of H.

### SOLVED EXAMPLE 14 — WORK FUNCTION IN ELECTRON-VOLTS

The work function of sodium is 2.3 eV. Will light of wavelength 600 nm eject electrons? (1 eV = $1.602 \times 10^{-19}$ J)

SOLUTION

1. Photon energy = $hc/\lambda = (6.626 \times 10^{-34} \times 3.0 \times 10^{8}) / (6.00 \times 10^{-7}) = 3.31 \times 10^{-19}$ J.
2. In eV: $3.31 \times 10^{-19} / 1.602 \times 10^{-19} = 2.07$ eV.
3. 2.07 eV < 2.3 eV.

ANSWER No — 600 nm photons carry too little energy; no electrons are emitted, however bright the light.

**SOLVED EXAMPLE 15** — ELECTRONS WITH A GIVEN QUANTUM NUMBER

How many electrons in an atom can have n = 4 and $m_s = -½$? How many can have n = 3, l = 1?

SOLUTION

1. n = 4: 16 orbitals → 32 electrons, half with $m_s = -½$.
2. n = 3, l = 1: the 3p subshell, 3 orbitals.

ANSWER 16 electrons; 6 electrons.

SECTION J

# Exam trap gallery

J

*Where marks are lost in this chapter.*

| ✗ TRAP 01 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Brighter light gives faster photoelectrons. | Brighter light gives more electrons; their energy depends only on frequency. |
| WHY Each photon ejects one electron. | |

| ✗ TRAP 02 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Bohr's model works for He. | Only for one-electron species: H, $He^{+}$, $Li^{2+}$. |
| WHY Electron–electron repulsion is ignored. | |

| ✗ TRAP 03 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Energy levels get farther apart at higher n. | They get closer together ($E \propto -1/n^2$). |
| WHY Lines crowd toward the series limit. | |

| ✗ TRAP 04 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Balmer series lies in the ultraviolet. | Balmer is visible; Lyman is ultraviolet. |
| WHY $n_1 = 2$ for Balmer. | |

| ✗ TRAP 05 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| The 2p orbital has one radial node. | 2p: 0 radial nodes, 1 angular node. |
| WHY Radial nodes = n − l − 1. | |

| ✗ TRAP 06 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| l can equal n. | l runs from 0 to n − 1. |
| WHY Hence no 1p or 2d orbital. | |

| ✗ TRAP 07 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| 4s fills after 3d because 3 < 4. | 4s (n + l = 4) fills before 3d (n + l = 5). |
| **WHY** Use the (n + l) rule. | |

| ✗ TRAP 08 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Cu is [Ar] $3d^9\ 4s^2$. | Cu is [Ar] $3d^{10}\ 4s^1$. |
| **WHY** A filled d subshell is extra stable. | |

| ✗ TRAP 09 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| de Broglie waves matter for cricket balls too. | λ is about $10^{-34}$ m — undetectable. |
| **WHY** Only tiny masses show measurable λ. | |

| ✗ TRAP 10 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Number of lines from n = 4 to 1 is 3. | Up to 6 lines in a sample (n(n − 1)/2). |
| **WHY** Every pair of levels gives one line. | |

| ✗ TRAP 11 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Ionisation energy of H is 13.6 J. | 13.6 eV per atom = $2.18 \times 10^{-18}$ J = 1312 kJ $mol^{-1}$. |
| **WHY** Watch the units. | |

| ✗ TRAP 12 · WHAT STUDENTS WRITE | ✓ WHAT IS ACTUALLY CORRECT |
|---|---|
| Total energy and kinetic energy of an orbiting electron are equal. | KE = -E and PE = 2E. |
| **WHY** Virial relation for a Coulomb orbit. | |

**MEMORY HOOK**

"Radius $n^2/Z$, energy $Z^2/n^2$, speed Z/n."

"Lyman UV, Balmer visible, rest infrared."

"Radial n − l − 1, angular l."

"n + l small fills first."

SECTION K

# Practice MCQs

K

*Levels: Easy, Moderate, Tough.*

**Q01** The charge on an electron was measured by: [EASY]

(a) Thomson (b) Millikan
(c) Rutherford (d) Chadwick

**Q02** The neutron was discovered by: [EASY]

(a) Rutherford (b) Chadwick
(c) Bohr (d) Goldstein

**Q03** Radius of the first Bohr orbit of hydrogen: [EASY]

(a) 52.9 pm (b) 26.45 pm
(c) 105.8 pm (d) 0.529 pm

**Q04** The Balmer series of hydrogen lies in the: [EASY]

(a) UV region (b) visible region
(c) IR region (d) X-ray region

**Q05** Number of orbitals in the n = 3 shell: [EASY]

(a) 3 (b) 6
(c) 9 (d) 18

**Q06** Shape of a p orbital: [EASY]

(a) spherical (b) dumbbell
(c) cloverleaf (d) ring

**Q07** Which set of quantum numbers is not allowed? [EASY]

(a) n=1, l=0, $m_l$=0 (b) n=2, l=1, $m_l$=−1
(c) n=2, l=2, $m_l$=0 (d) n=3, l=2, $m_l$=+2

**Q08** Energy of the electron in the second orbit of H: [MODERATE]

(a) −13.6 eV (b) −3.4 eV
(c) −1.51 eV (d) −6.8 eV

**Q09** Number of radial nodes in 3s: [MODERATE]

(a) 0 (b) 1
(c) 2 (d) 3

**Q10** Electronic configuration of Cr (Z = 24): [MODERATE]

(a) [Ar] $3d^4\ 4s^2$ (b) [Ar] $3d^5\ 4s^1$
(c) [Ar] $3d^6$ (d) [Ar] $4s^2\ 4p^4$

**Q11** Wavelength of the Hα line (n = 3 → 2): [MODERATE]

(a) 486 nm (b) 656 nm
(c) 434 nm (d) 122 nm

**Q12** In the photoelectric effect, kinetic energy of electrons depends on: [MODERATE]

(a) intensity (b) frequency
(c) exposure time (d) area of metal

**Q13** Maximum number of lines when electrons fall from n = 6 to n = 1: [MODERATE]

(a) 5 (b) 10
(c) 15 (d) 21

**Q14** The orbital with n = 4, l = 2 is: [MODERATE]

(a) 4s (b) 4p
(c) 4d (d) 4f

**Q15** Which orbital fills first? [MODERATE]

(a) 3d (b) 4s
(c) 4p (d) both 3d and 4s together

**Q16** de Broglie wavelength is inversely proportional to: [MODERATE]

(a) mass only (b) momentum
(c) energy squared (d) frequency

**Q17** Energy of the first orbit of $He^+$: [MODERATE]

(a) -13.6 eV (b) -27.2 eV
(c) -54.4 eV (d) -122.4 eV

**Q18** Number of angular nodes in a 4f orbital: [MODERATE]

(a) 1 (b) 2
(c) 3 (d) 4

**Q19** Total number of electrons with l = 1 in phosphorus (Z = 15): [TOUGH]

(a) 6 (b) 9
(c) 12 (d) 3

**Q20** Which transition in $He^+$ gives the same wavelength as 2 → 1 in H? [TOUGH]

(a) 4 → 2 (b) 3 → 1
(c) 2 → 1 (d) 6 → 3

**Q21** Ionisation energy of H is 13.6 eV. That of $Li^{2+}$ is: [TOUGH]

(a) 40.8 eV (b) 122.4 eV
(c) 27.2 eV (d) 54.4 eV

**Q22** Minimum uncertainty in velocity of an electron with $\Delta x$ = 0.1 nm is about: [TOUGH]

(a) $5.8 \times 10^5$ m $s^{-1}$ (b) $5.8 \times 10^2$ m $s^{-1}$
(c) $5.8 \times 10^8$ m $s^{-1}$ (d) 5.8 m $s^{-1}$

**Q23** Orbital angular momentum of a d electron is: [TOUGH]

(a) $\sqrt{2}\ h/2\pi$ (b) $\sqrt{6}\ h/2\pi$
(c) $\sqrt{12}\ h/2\pi$ (d) $2\ h/2\pi$

**Q24** Ratio of radii of the third orbit of $Li^{2+}$ and the first orbit of H: [TOUGH]

(a) 1 : 3 (b) 3 : 1
(c) 9 : 1 (d) 1 : 9

**Q25** Shortest wavelength in the Lyman series of H: [TOUGH]

(a) 91.2 nm (b) 121.6 nm
(c) 656 nm (d) 364.6 nm

## Answer key with explanations

Q01 **(b) Millikan** — Oil-drop experiment.

Q02 **(b) Chadwick** — 1932.

Q03 **(a) 52.9 pm** — $r_1$ = 52.9 pm.

Q04 **(b) visible region** — $n_1 = 2$.

Q05 **(c) 9** — $n^2$.

Q06 **(b) dumbbell** — Two lobes.

Q07 **(c) n=2, l=2, $m_l$=0** — $l < n$.

Q08 **(b) -3.4 eV** — -13.6/4.

Q09 **(c) 2** — n - l - 1 = 2.

Q10 **(b) [Ar] $3d^5\ 4s^1$** — Half-filled d.

Q11 **(b) 656 nm** — Red line.

Q12 **(b) frequency** — KE = $h(\nu - \nu_0)$.

Q13 **(c) 15** — 6 × 5/2.

Q14 **(c) 4d** — l = 2 → d.

Q15 **(b) 4s** — n + l = 4 vs 5.

Q16 **(b) momentum** — $\lambda = h/p$.

Q17 **(c) -54.4 eV** — $Z^2 = 4$.

Q18 **(c) 3** — l = 3.

Q19 **(b) 9** — $2p^6 + 3p^3$.

Q20 **(a) 4 → 2** — $Z^2(1/n_1^2 - 1/n_2^2)$ equal.

Q21 **(b) 122.4 eV** — × $Z^2$ = 9.

Q22 **(a) 5.8 × $10^5$ m $s^{-1}$** — $h/(4\pi m\Delta x)$.

Q23 **(b) $\sqrt{6}\ h/2\pi$** — $\sqrt{[l(l+1)]}$ with l = 2.

Q24 **(b) 3 : 1** — $n^2/Z$ = 9/3.

Q25 **(a) 91.2 nm** — Series limit: $1/\lambda = R_h$.

## Check-yourself answers

C1 Balmer.

C2 Six.

C3 Two (4 – 1 – 1).

C4 18 ($2n^2$).

QUICK REVISION

# One-page revision

*Read this page the night before the exam.*

## Particles and models

- e = 1.602 × $10^{-19}$ C; $m_e$ = 9.109 × $10^{-31}$ kg
- Rutherford: tiny dense nucleus
- Isotopes (Z), isobars (A), isotones (n)
- Rutherford fails: radiation, spectra

## Light

- c = νλ; E = hν = hc/λ
- Photoelectric: hν = $h\nu_0$ + ½$mv^2$
- Intensity → number; frequency → KE
- Visible 400–750 nm

## Bohr

- $r_n$ = 52.9 $n^2/Z$ pm
- $E_n$ = −13.6 $Z^2/n^2$ eV = −2.18 × $10^{-18}$ $Z^2/n^2$ J
- mvr = nh/2π
- One-electron species only

## Spectrum

- $1/\lambda = R_h Z^2 (1/n_1^2 - 1/n_2^2)$
- Lyman UV, Balmer visible, Paschen IR
- Lines = n(n − 1)/2
- Hα 656 nm; Lyman limit 91.2 nm

## Waves and uncertainty

- λ = h/mv
- Δx · mΔv ≥ h/4π
- 2πr = nλ in Bohr orbits
- Large objects: negligible

## Quantum numbers

- n, l (0 to n−1), $m_l$ (2l+1), $m_s$ ±½
- Radial n − l − 1; angular l
- (n + l) rule; Pauli; Hund
- Cr $3d^5 4s^1$, Cu $3d^{10} 4s^1$ — exchange energy

END OF NOTES · RAMGOVINDYADAV.COM