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CUET · CHEMISTRY · CLASS XI · CODE 306

# Structure of Atom

CUET unit: Structure of Atom

By **UniDrill** · NCERT-grounded study material

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## 📌 Snapshot

- Traces the journey from Dalton's indivisible atom to the modern quantum-mechanical model via Thomson's plum-pudding atom, Rutherford's nuclear atom, Bohr's quantised orbits, de Broglie's matter waves, and Schrödinger's wave equation.
- Establishes the three sub-atomic particles — electron (Thomson 1897, Millikan 1909), proton (canal-ray, Goldstein/Rutherford), neutron (Chadwick 1932) — with their charges, masses and the experiments that discovered them.
- Develops the wave–particle duality of electromagnetic radiation through Planck's quantum theory ($E = h\nu$), the photoelectric effect (Einstein 1905), and the hydrogen line spectrum (Rydberg/Bohr).
- Introduces de Broglie's dual nature of matter ($\lambda = h/mv$) and Heisenberg's uncertainty principle ($\Delta x \cdot \Delta p \geqslant h/4\pi$) as the conceptual bridge to quantum mechanics.
- Concludes with the four quantum numbers (n, l, m?, m?), shapes of s/p/d orbitals, the (n+l) Aufbau rule, Pauli's exclusion principle, Hund's rule of maximum multiplicity, and the Cr/Cu exceptions.
- High-yield CUET chapter — about 10–12 MCQs per attempt, mostly on quantum numbers, Bohr formulae for H-like ions, hydrogen spectral series, de Broglie/photoelectric numericals and electronic configurations.

## 📖 Detailed Notes

### 2.1 Core concepts

**Discovery of the electron.** J.J. Thomson (1897) constructed a cathode ray discharge tube and measured the charge-to-mass ratio of the cathode-ray particles by balancing mutually perpendicular electric and magnetic fields. The result, $e/m_e = 1.758820 \times 10^{11}$ C $kg^{-1}$, was independent of both the cathode material and the residual gas, proving that the negatively charged particle now called the electron is a universal constituent of all atoms. Cathode rays travel in straight lines in field-free space, originate at the cathode, and are deflected toward the positive plate in an electric field (NCERT §2.1.1–2.1.2, pp. 30–31; Figs. 2.1, 2.2). R.A. Millikan's oil-drop experiment (1906–14) measured the charge on each electron by balancing gravity against an applied electric field on charged oil droplets; the charge was always an integer multiple of e = 1.6022 × $10^{-19}$ C, and combining with $e/m_e$ gave the electron mass $m_e$ = 9.1094 × $10^{-31}$ kg (NCERT §2.1.3, p. 31).

**Proton and neutron.** Modified discharge tubes with perforated cathodes produced canal rays moving toward the cathode — these were positive ions whose smallest, lightest member came from hydrogen and was identified as the proton (charge +1.602 × $10^{-19}$ C; mass 1.672 × $10^{-27}$ kg). The neutron was discovered by James Chadwick (1932) by bombarding beryllium with $\alpha$-particles; the resulting penetrating neutral radiation consisted of particles slightly heavier than protons ($m_n$ = 1.674 × $10^{-27}$ kg, q = 0) (NCERT §2.1.4 + Table 2.1, p. 32–33).

**Thomson model (1898).** An atom is a sphere of radius ~$10^{-10}$ m of uniformly distributed positive charge with electrons embedded like seeds in a watermelon. Total charge is zero, mass is uniformly distributed. The model explained electrical neutrality but failed later scattering tests (NCERT §2.2.1, p. 33; Fig. 2.4).

**Rutherford's nuclear atom (1911).** Geiger and Marsden bombarded a very thin gold foil (~100 nm) with $\alpha$-particles. Most $\alpha$-particles passed undeflected; a few were deflected through small angles; and roughly one in 20,000 bounced almost straight back. Rutherford concluded that (i) most of the atom is empty space, (ii) the entire positive charge and almost all the mass are concentrated in a tiny nucleus of radius ~$10^{-15}$ m, and (iii) electrons revolve around the nucleus in orbits like planets around the Sun (NCERT §2.2.2, p. 34; Fig. 2.5).

**Atomic and mass numbers.** The atomic number Z = number of protons in the nucleus = number of electrons in the neutral atom — Z defines the element. The mass number A = Z + n = total nucleons. The isotope notation is $^{A}_{Z}X$. **Isotopes** have the same Z but different A ($^{1}H$, $^{2}D$, $^{3}T$; $^{12}C$, $^{13}C$, $^{14}C$; $^{35}Cl$, $^{37}Cl$) and share chemistry because chemistry depends on the electron count. **Isobars** have the same A but different Z ($^{14}_{6}C$ and $^{14}_{7}N$) (NCERT §2.2.3–2.2.4, p. 35).

**Failure of the Rutherford model.** A revolving electron is undergoing centripetal acceleration; by Maxwell's electromagnetism an accelerating charge must continuously radiate, lose energy, and spiral into the nucleus in ~$10^{-8}$ s. The model also says nothing about the distribution or energies of electrons (NCERT §2.2.5, p. 36).

**EM radiation as waves.** Maxwell (1870) showed that oscillating charged particles produce mutually perpendicular electric and magnetic fields propagating as transverse electromagnetic waves at c = 3.0 × $10^{8}$ m $s^{-1}$ in vacuum, with no medium required. Wave parameters: frequency $\nu$ (Hz), wavelength $\lambda$ (m), c = $\nu\lambda$, wavenumber $\bar{\nu} = 1/\lambda$. The electromagnetic spectrum stretches from radio (~$10^{6}$ Hz) through microwave, IR, visible (380–780 nm, ~$10^{15}$ Hz), UV, X-ray to $\gamma$-ray (NCERT §2.3.1, pp. 37–39; Fig. 2.7).

**Planck's quantum theory.** Classical wave theory could not explain black-body radiation or the photoelectric effect. Planck (1900) proposed that energy is absorbed and emitted only in discrete quanta of E = h$\nu$, where h = 6.626 × $10^{-34}$ J s, so that allowed energies are E = 0, h$\nu$, 2h$\nu$, ⋯ nh$\nu$. The black-body intensity now varies smoothly with temperature and explains why hot objects glow red before orange/white (NCERT §2.3.2, pp. 40–41).

**Photoelectric effect.** Heinrich Hertz (1887) observed that UV light ejects electrons from a metal surface. Three classical puzzles emerged: (i) emission occurs only above a threshold frequency $\nu_0$; (ii) the kinetic energy of the ejected electron depends on frequency, not intensity; (iii) the number of electrons depends on intensity. Einstein (1905) used Planck's quanta to explain all three at once: $h\nu = h\nu_0 + \frac{1}{2}m_e v^2 = W_0 +$ KE, where $W_0 = h\nu_0$ is the work function. KE = $h(\nu - \nu_0)$. This earned Einstein the 1921 Nobel Prize (NCERT §2.3.2, pp. 41–43; Fig. 2.9; eq. 2.7).

**Atomic line spectra.** Incandescent solids emit a continuous spectrum, but excited atoms emit at discrete wavelengths producing a line spectrum — every element has a unique fingerprint (He was discovered spectroscopically in the Sun before being found on Earth). The hydrogen line spectrum is described by Rydberg's formula $\bar{\nu} = 109{,}677\ (1/n_1^2 - 1/n_2^2)\ cm^{-1}$, organised into series — **Lyman** ($n_1 = 1$, UV), **Balmer** ($n_1 = 2$, visible), **Paschen** ($n_1 = 3$, IR), **Brackett** ($n_1 = 4$, IR) and **Pfund** ($n_1 = 5$, far IR) (NCERT §2.3.3 + Table 2.3, pp. 44–46; Figs. 2.10–2.11).

**Bohr's model of the atom (1913).** Three postulates: (i) the electron moves only in fixed stationary circular orbits of definite radius and energy without radiating; (ii) energy is constant in an orbit but is absorbed (n_low → n_high) or emitted (n_high → n_low) only in transitions with frequency $\nu = \Delta E/h$; (iii) angular momentum is quantised: $m_e v r = n\, h/2\pi$. From these, the energy and radius of hydrogen-like (one-electron) species are $E_n = -2.18 \times 10^{-18}\ (Z^2/n^2)$ J atom$^{-1}$ and $r_n = 52.9\ (n^2/Z)$ pm. The Bohr radius $a_0 = 52.9$ pm is the radius for n = Z = 1. Negative energy means the bound electron is more stable than a free electron at rest (n = ∞, E = 0). Bohr's model elegantly reproduces the Rydberg constant from first principles (NCERT §2.4, eqs. 2.10–2.15, pp. 46–48).

**Limitations of Bohr.** The model fails to explain (i) the fine structure (doublets, triplets) of even hydrogen lines, (ii) the spectra of multi-electron atoms, (iii) the Zeeman effect (splitting in a magnetic field) and the Stark effect (splitting in an electric field), and (iv) the formation of chemical bonds — molecular geometry. It also violates the de Broglie/Heisenberg framework by giving the electron a definite trajectory (NCERT §2.4.2, p. 49).

**Dual nature of matter — de Broglie (1924).** Every moving particle has a wavelength $\lambda = h/mv = h/p$. The wave nature is significant only for microscopic objects: an electron at $v = 10^6$ m s$^{-1}$ has $\lambda \approx 7.3 \times 10^{-10}$ m (X-ray range) — diffractable by crystals (Davisson–Germer, 1927) — whereas a 1 kg cricket ball at 10 m s$^{-1}$ has $\lambda = 6.6 \times 10^{-35}$ m, far below any measurable scale (NCERT §2.5.1, eq. 2.22, p. 50).

**Heisenberg's uncertainty principle (1927).** It is impossible to determine simultaneously and exactly both the position and the momentum of a microscopic particle: $\Delta x \cdot \Delta p \geq h/4\pi$, equivalently $\Delta x \cdot \Delta v \geq h/(4\pi\, m)$. For an electron ($m_e \approx 9.1 \times 10^{-31}$ kg) the product is sizeable and forbids talk of "the path of the electron"; for a 1 mg dust particle it is utterly negligible. The principle invalidates Bohr-style fixed orbits and motivates the probabilistic orbital description (NCERT §2.5.2, eq. 2.23, pp. 51–52).

**Quantum-mechanical model.** Schrödinger's wave equation $\hat{H}\psi = E\psi$ yields the allowed energies E and wave functions $\psi$ (atomic orbitals) for the hydrogen atom. $|\psi|^2$ (Max Born interpretation) is the probability density of finding the electron at a point. For one-electron systems (H, $He^+$, $Li^{2+}$) orbital energy depends only on n; in multi-electron atoms it depends on both n and l because of electron-electron repulsion and shielding (NCERT §2.6, pp. 53–54).

**Quantum numbers.** Each orbital is specified by three quantum numbers and each electron by an additional spin quantum number. **n** (principal, 1, 2, 3 ⋯) sets shell, size and (mostly) energy. **l** (azimuthal, 0 to n−1) sets the subshell (0 = s, 1 = p, 2 = d, 3 = f, 4 = g ⋯) and orbital shape. **m**⍰ (magnetic, −l to +l, 2l+1 values) gives orientation. **m**⍰ (spin, +½ or −½, Uhlenbeck and Goudsmit 1925) is the intrinsic electron spin. The maximum number of orbitals in shell n is $n^2$ and the maximum number of electrons is $2n^2$ (NCERT §2.6.1, pp. 54–56).

**Shapes and nodes.** s-orbitals are spherically symmetric; an ns orbital has (n − 1) radial nodes. p-orbitals are dumbbell-shaped along x, y or z axes ($2p_x$, 2p_y, 2p_z); each has one angular nodal plane and (n − 2) radial nodes. d-orbitals come in five orientations — $d_{xy}$, $d_{yz}$, $d_{xz}$, $d_{x^2-y^2}$, $d_{z^2}$ — the first four have four-lobed shapes, while $d_{z^2}$ has two lobes plus a torus. Total nodes per orbital = n − 1; angular nodes = l; radial nodes = n − l − 1 (NCERT §2.6.2, pp. 57–59; Figs. 2.12–2.15).

**Energy ordering and Aufbau.** In hydrogen, orbitals of the same n are degenerate; in multi-electron atoms, energy increases with (n + l), and for tied (n + l), with smaller n first. The diagonal rule gives the filling order: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s ⋯ (NCERT §2.6.3–2.6.4, Table 2.5, pp. 60–62; Fig. 2.17). **Pauli's exclusion principle** — no two electrons can share all four quantum numbers (an orbital holds two electrons with opposite spins). **Hund's rule** — pairing in degenerate orbitals begins only after each has one electron with parallel spin. **Cr** (Z = 24) is [Ar] $3d^5 4s^1$ and **Cu** (Z = 29) is [Ar] $3d^{10} 4s^1$ because half-filled and filled subshells gain extra stability from symmetric distribution, low mutual shielding and large exchange energy (NCERT §2.6.4–2.6.6, pp. 62–65; Fig. 2.18).

## 2.2 Definitions to memorise

| Term | Definition | Page |
|---|---|---|
| Cathode rays | Stream of negatively charged particles (electrons) moving from cathode to anode in an evacuated discharge tube | 30 |
| Canal rays | Positively charged rays moving toward the cathode in a discharge tube with perforated cathode | 32 |
| Atomic number (Z) | Number of protons in the nucleus; equals the number of electrons in a neutral atom | 35 |

| Term | Definition | Page |
|---|---|---|
| Mass number (A) | Total number of nucleons (protons + neutrons) in the nucleus | 35 |
| Isotopes | Atoms of the same element with same Z but different A ($^{1}H$, $^{2}D$, $^{3}T$) | 35 |
| Isobars | Atoms with same A but different Z ($^{14}C$ and $^{14}N$) | 35 |
| Quantum | The smallest discrete packet of energy emitted or absorbed as EM radiation; $E = h\nu$ | 41 |
| Photon | A quantum of light, treated as a particle of energy $h\nu$ and momentum $h/\lambda$ | 42 |
| Work function ($W_0$) | Minimum energy $h\nu_0$ required to eject an electron from a metal surface | 42 |
| Threshold frequency ($\nu_0$) | Minimum frequency below which photoelectric effect does not occur | 41 |
| Black body | Ideal body that absorbs and emits radiation of all frequencies uniformly | 40 |
| Continuous spectrum | Spectrum showing all wavelengths without gaps (e.g., white light through a prism) | 44 |
| Line spectrum | Atomic emission/absorption spectrum showing discrete wavelengths characteristic of each element | 44 |
| Rydberg constant | $R\_H = 109{,}677\ cm^{-1}$ in the formula $\bar{\nu} = R\_H(1/n_1^2 - 1/n_2^2)$ | 45 |
| Stationary state | Bohr orbit in which an electron has constant energy without radiating | 46 |
| Bohr radius ($a_0$) | Radius of the first Bohr orbit of hydrogen, 52.9 pm | 47 |
| de Broglie wavelength | $\lambda = h/mv = h/p$; matter wave associated with a moving particle | 50 |
| Heisenberg uncertainty | $\Delta x \cdot \Delta p \geq h/4\pi$; position and momentum cannot both be known exactly | 51 |
| Atomic orbital | One-electron wave function $\psi$(n, l, m?); $\|\psi\|^2$ gives probability density | 54 |
| Principal quantum number (n) | 1, 2, 3 ⋯ — sets shell, size and energy | 55 |
| Azimuthal quantum number (l) | 0 to n−1 — defines subshell shape (s, p, d, f) | 55 |
| Magnetic quantum number (m?) | −l to +l, 2l+1 values — orientation in space | 55 |
| Spin quantum number (m?) | +½ or −½ — intrinsic spin of the electron | 56 |
| Degenerate orbitals | | 60 |

| Term | Definition | Page |
|---|---|---|
| | Orbitals of identical energy (same subshell in H; same (n+l) in multi-electron atoms) | |
| Aufbau principle | Orbitals fill in order of increasing energy ((n+l) rule) | 62 |
| Pauli exclusion principle | No two electrons in an atom can share all four quantum numbers | 62 |
| Hund's rule | In degenerate orbitals, pairing begins only after each is singly occupied | 62 |
| Node | Region where probability density $\|\psi\|^2$ is zero; total nodes = n − 1 | 57, 59 |

## 2.3 Diagrams / processes to remember

CUET routinely tests which figure depicts which phenomenon. The discovery sequence begins with **Fig. 2.1** (cathode ray discharge tube — note the perforated anode used to make a beam), **Fig. 2.2** (Thomson's $e/m_e$ set-up showing electric and magnetic field plates with deflection points A, B, C), and **Fig. 2.3** (Millikan's oil-drop apparatus with the atomiser, condenser plates and microscope viewer). **Fig. 2.4** is Thomson's plum-pudding sphere and **Fig. 2.5** the Geiger–Marsden gold-foil scattering set-up that overturned it — almost all $\alpha$-particles pass straight through; one in 20,000 rebounds. **Fig. 2.6** sketches Rutherford's miniature solar system.

The wave–particle figures begin with **Fig. 2.7** — the full EM spectrum from radio (long $\lambda$, low $\nu$) through visible (380–780 nm) to $\gamma$-rays (short $\lambda$, high $\nu$) — and **Fig. 2.8** (wave parameters: amplitude, wavelength and frequency). **Fig. 2.9** is the classic photoelectric apparatus: a clean metal cathode irradiated by tunable-frequency UV; you should be able to label threshold frequency, work function and the linear KE-vs-$\nu$ Einstein plot. **Fig. 2.10** contrasts emission and absorption line spectra of atomic hydrogen, and **Fig. 2.11** maps Lyman, Balmer and Paschen transitions onto the H energy-level diagram — particularly useful for series-identification MCQs.

The orbital figures come next. **Fig. 2.12** shows $\psi$(r) and $\psi^2$(r) plots for 1s and 2s — the 2s function crosses zero once, creating a radial node. **Fig. 2.13** gives boundary-surface diagrams for 1s and 2s as spheres. **Fig. 2.14** shows $2p_x$, $2p_y$, $2p_z$ dumbbells along the coordinate axes. **Fig. 2.15** displays the five 3d orbitals — four "cloverleaf" lobes ($d_{xy}$, $d_{yz}$, $d_{xz}$, $d_{x^2-y^2}$) and the distinct $d_{z^2}$ with axial lobes plus an equatorial torus. **Fig. 2.16** compares the energy-level diagrams of hydrogen (degenerate within n) versus multi-electron atoms ((n+l) ordering). **Fig. 2.17** is the famous (n+l) diagonal mnemonic for the Aufbau filling order. **Fig. 2.18** illustrates the six exchange pairs in a $d^5$ configuration that account for the extra stability of half-filled subshells in Cr and Mn.

Three procedural workflows worth memorising: (1) the photoelectric pipeline — given metal work function $W_0$ and incident frequency $\nu$, compute KE = $h(\nu - \nu_0)$ and $v_{max}$ from ½ $m_e v^2$ (NCERT Problem 2.8, p. 43); (2) the spectral-line pipeline — given $n_1$, $n_2$, use Rydberg $\bar{\nu} = 109{,}677(1/n_1^2 - 1/n_2^2)\ cm^{-1}$, then $\lambda = 1/\bar{\nu}$; (3) the H-like ion pipeline — E⍰ = $-2.18 \times 10^{-18} (Z^2/n^2)$ J, r⍰ = $52.9 (n^2/Z)$ pm, and ionisation energy = $E_\infty - E_n$.

## 2.4 Common confusions / NTA trap points

- **Orbit vs orbital** — a Bohr orbit is a fixed circular trajectory; an orbital is a 3-D probability region described by $\psi$. Don't treat them as synonyms (p. 56).
- **Filling order vs removal order** — 4s fills before 3d in the Aufbau order, but in transition-metal cations 4s electrons are removed first because once 3d is occupied, 3d sinks below 4s in energy. So $Fe^{2+}$ is [Ar] $3d^6$, not [Ar] $3d^4\ 4s^2$.
- **Cr and Cu exceptions** — half-filled ($3d^5$) and filled ($3d^{10}$) subshells are extra-stable; valence configurations are $3d^5\ 4s^1$ and $3d^{10}\ 4s^1$, not $3d^4\ 4s^2$ and $3d^9\ 4s^2$ (p. 64).
- **Photoelectric KE depends on frequency, NOT intensity** — intensity merely changes the number of electrons ejected (p. 42). Common distractor.
- **Number of nodes** — total nodes = $n - 1$; angular nodes = $l$; radial nodes = $n - l - 1$. Don't confuse "nodes" (where $\psi = 0$) with "nodal planes" (specific angular nodes).
- **Bohr energy formula** — E⍰ ∝ $-Z^2/n^2$ ($Z^2$ in numerator, $n^2$ in denominator). NTA loves to swap them in distractors.
- **Isotopes vs isobars vs isotones** — isotopes share Z, isobars share A, isotones share neutron number n (not in CUET syllabus but tested as a distractor).
- **de Broglie $\lambda$ for macroscopic objects** — though $\lambda = h/mv$ applies to all objects, it is unobservably small for everyday bodies. Wave nature manifests only for very small m.
- **Schrödinger's $\psi$ has no physical meaning by itself** — only $|\psi|^2$ is observable (Born interpretation, p. 54).
- **(n+l) tie-break** — when two orbitals have the same (n+l), the smaller n is filled first; thus 3d (n+l=5) is below 4p (n+l=5) only because 3 < 4 → no, here 3 < 4 so 3d first. But 4s (n+l=4) sits below 3d (n+l=5). Read the rule carefully.
- **He emission lines** — Helium was discovered in the Sun's spectrum before on Earth; it is not a hydrogen-like ion ($He^+$ is), and only $He^+$ obeys Bohr exactly.
- **Spin quantum number values** — +½ and −½, not +1 and −1 (a common slip).

## 2.5 Key reactions & formulas

| Reaction / Formula | Conditions / Notes | NCERT page |
|---|---|---|
| $e/m_e = 1.758820 \times 10^{11}\ C\ kg^{-1}$ | Thomson cathode-ray experiment | 31 |
| $e = -1.6022 \times 10^{-19}$ C; $m_e = 9.1094 \times 10^{-31}$ kg | Millikan + Thomson combined | 31 |
| $c = \nu\lambda$; $\bar{\nu} = 1/\lambda$ | Wave–parameter relations, $c = 3.0 \times 10^8\ m\ s^{-1}$ | 38 |

| Reaction / Formula | Conditions / Notes | NCERT page |
|---|---|---|
| $E = h\nu = hc/\lambda$ | Planck quantum, $h = 6.626 \times 10^{-34}$ J s | 41 |
| $\frac{1}{2} m_e v^2 = h\nu - h\nu_0$ | Einstein photoelectric equation | 42 |
| $\bar{\nu} = 109,677(1/n_1^2 - 1/n_2^2)$ $cm^{-1}$ | Rydberg formula for H | 45 |
| Lyman $n_1$=1 UV; Balmer $n_1$=2 visible; Paschen $n_1$=3 IR | Hydrogen spectral series | 45 |
| $m_e v r = n h/2\pi$ | Bohr quantisation of angular momentum | 46 |
| $E_n = -2.18 \times 10^{-18} (Z^2/n^2)$ J $atom^{-1}$ | Bohr energy, H-like one-electron species | 47 |
| $r_n = 52.9 (n^2/Z)$ pm | Bohr radius, H-like ions; $a_0$ = 52.9 pm | 48 |
| $\Delta E = E_2 - E_1 = h\nu$ | Bohr frequency rule for emission/ absorption | 46 |
| $\lambda = h/(m v) = h/p$ | de Broglie wavelength | 50 |
| $\Delta x \cdot \Delta p \geq h/4\pi$ | Heisenberg uncertainty | 51 |
| Number of orbitals in shell n = $n^2$ | From quantum numbers | 56 |
| Maximum electrons in shell n = $2n^2$ | Pauli + spin | 56 |
| Subshell counts: s(1), p(3), d(5), f(7) orbitals | 2l+1 values of $m_l$ | 56 |
| Total nodes = n − 1; radial = n − l − 1; angular = l | Orbital shape | 59 |
| (n + l) Aufbau rule | Lower (n+l) → lower energy; tie-break with smaller n | 62 |
| Cr [Ar] $3d^5 4s^1$; Cu [Ar] $3d^{10} 4s^1$ | Half/fully-filled stability exceptions | 64 |

## 🎯 Practice MCQs

**Q1.** **The charge-to-mass ratio ($e/m_e$) of the electron, as determined by J.J. Thomson, is**

**A.** $1.602 \times 10^{-19}$ C $kg^{-1}$

**B.** $9.109 \times 10^{-31}$ C $kg^{-1}$

**C.** $1.758820 \times 10^{11}$ C $kg^{-1}$

**D.** $6.626 \times 10^{-34}$ C $kg^{-1}$

**Q2.** **Which experiment first proved that an atom contains a tiny, dense, positively charged nucleus?**

**A.** Thomson's cathode-ray experiment

**B.** Millikan's oil-drop experiment

**C.** Rutherford's $\alpha$-particle scattering by gold foil

**D.** Chadwick's beryllium bombardment

**Q3.** **Which pair represents isobars?**

**A.** $^{1}H$ and $^{2}D$

**B.** $^{14}_{6}C$ and $^{14}_{7}N$

**C.** $^{35}_{17}Cl$ and $^{37}_{17}Cl$

**D.** $^{12}_{6}C$ and $^{13}_{6}C$



# 📊 PYQ Alignment

Structure of Atom is among the highest-yielding chapters in CUET Chemistry — typically 10–12 MCQs per attempt across CUET 2023–25. The most-tested themes are: quantum numbers (permitted sets, total orbitals/electrons for given n); Bohr energy/radius for H-like ions ($He^{+}$, $Li^{2+}$, $Be^{3+}$); the photoelectric effect with KE = $h(\nu - \nu_0)$ numericals; hydrogen spectral series (Lyman, Balmer, Paschen identifications); de Broglie wavelength numericals; isotopes vs isobars; and electronic configurations including the Cr/Cu exceptions and transition-metal cation removal order. Assertion–reason and statement-based formats appear often, especially around Heisenberg's uncertainty and the limitations of Bohr's model.

For chapter-wise PYQ practice and consolidated mock tests on these themes, see [/pyq/chemistry](/pyq/chemistry).

## CUETL 2023 — Actual PYQs from this chapter

**Q.1 (CUET 2023)** In case of phosphorus, $PCl_3$ and $PCl_5$ are possible while nitrogen forms $NCl_3$ and not $NCl_5$. This is due to:

- A) Nitrogen is a gas while phosphorus is solid at room temperature. B) Nitrogen does not have vacant d-orbitals in its valence shell. C) Electronegativity of nitrogen is higher than phosphorus. D) Nitrogen atom is smaller than phosphorus atom.

Tests: Structure of Atom Answer: Not in extracted key