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Structure of Atom Class 11 Notes | CBSE Chemistry Chapter 2

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Structure of Atom Class 11 Notes | CBSE Chemistry

Chapter 2

Structure of Atom is Chapter 2 of CBSE Class 11 Chemistry — the chapter that finally explains *what* an atom is actually made of and *why* electrons sit where they do. From the discovery of the electron to the strange quantum world of orbitals, this chapter rebuilds your idea of matter from the ground up. Master it and the rest of physical chemistry, chemical bonding, and the periodic table suddenly click into place.

By the end of these notes you will be able to describe the discovery of subatomic particles, compare Thomson, Rutherford and Bohr models, apply Planck's quantum theory and the photoelectric effect, calculate energies and radii in the hydrogen atom, write quantum numbers and electronic configurations, and dodge the common mistakes examiners love. This is a high-weightage chapter carrying roughly 6–8 marks in boards and the foundation for Classification of Elements, Chemical Bonding, and almost all of physical chemistry.

Table of Contents

- Key Concepts — subatomic particles, atomic models, quantum theory, hydrogen spectrum, orbitals, configuration
 - Weightage in Board & Entrance Exams
 - Important Definitions
 - Solved Examples
 - Important Questions for Board Exams
 - Quick Revision Points
-

Key Concepts

1. Discovery of Subatomic Particles

Atoms were once thought to be indivisible (Dalton). Experiments with electric discharge through gases at low pressure proved they contain even smaller particles — **electrons, protons and neutrons**.

The Electron — Cathode Ray Discharge Tube

When a high voltage is applied across electrodes in a gas at low pressure, rays travel from the cathode to the anode — **cathode rays**. These rays are streams of negatively charged particles called **electrons**, and their properties do not depend on the gas or electrode material.

- J. J. Thomson measured the **charge-to-mass ratio**: $e/m_e = 1.758 \times 10^{11} \text{ C/kg}$.
- R. A. Millikan's **oil-drop experiment** found the charge on the electron: $e = -1.602 \times 10^{-19} \text{ C}$.

- Mass of electron $m_e = 9.109 \times 10^{-31}$ kg (about 1/1837 of a hydrogen atom).

The Proton — Anode (Canal) Rays

Positively charged rays travelling towards the cathode (anode rays) were heaviest for hydrogen gas, identifying the **proton**. Charge = $+1.602 \times 10^{-19}$ C; mass = 1.672×10^{-27} kg.

The Neutron

James Chadwick (1932) bombarded beryllium with α -particles and discovered the **neutron** — a neutral particle of mass 1.675×10^{-27} kg, slightly heavier than a proton.

2. Atomic Number, Mass Number, Isotopes and Isobars

Atomic number (Z) = number of protons = number of electrons in a neutral atom. **Mass number (A)** = number of protons + neutrons. A nuclide is written as ${}^A_Z\text{X}$.

- **Isotopes:** same Z, different A (same protons, different neutrons) — e.g. ${}^1\text{H}$, ${}^2\text{H}$, ${}^3\text{H}$.
 - **Isobars:** different Z, same A — e.g. ${}^{14}_6\text{C}$ and ${}^{14}_7\text{N}$.
 - **Isotones:** same number of neutrons — e.g. ${}^{14}_6\text{C}$ and ${}^{16}_8\text{O}$.
-

3. Thomson and Rutherford Models

Thomson Model (Plum-Pudding / Watermelon Model)

Thomson pictured the atom as a uniform sphere of positive charge with electrons embedded in it like seeds in a watermelon. It explained overall neutrality but failed to explain the results of the gold-foil experiment.

Rutherford's α -Scattering Experiment

A thin gold foil was bombarded with α -particles. Most passed straight through, a few deflected, and very few (1 in 20,000) bounced back.

- Most of the atom is **empty space**.
- The entire positive charge and almost all mass are concentrated in a tiny dense **nucleus**.
- Electrons revolve around the nucleus like planets around the sun.

[DIAGRAM: α -particles fired at gold foil — most go straight, a few bend, very few rebound from the central nucleus.]

Drawback: A revolving electron is accelerating, so by classical theory it should radiate energy, spiral inward, and the atom should collapse — which does not happen. It also could not explain line spectra.

4. Electromagnetic Radiation

Light is an **electromagnetic wave** with oscillating electric and magnetic fields perpendicular to each other and to the direction of travel. Key wave terms:

- **Wavelength (λ):** distance between two consecutive crests.
 - **Frequency (ν):** number of waves per second; SI unit hertz (Hz).
 - **Wavenumber ($\bar{\nu}$):** number of waves per unit length, $\bar{\nu} = 1/\lambda$.
 - **Speed:** $c = \nu\lambda = 3 \times 10^8$ m/s in vacuum.
-

5. Planck's Quantum Theory

Classical physics could not explain black-body radiation. Max Planck proposed that energy is emitted or absorbed not continuously but in discrete packets called **quanta** (a quantum of light is a **photon**).

$$E = h\nu = hc/\lambda$$

- Planck's constant $h = 6.626 \times 10^{-34}$ J·s.
 - Energy of n photons: $E = nh\nu$.
 - Energy is directly proportional to frequency and inversely proportional to wavelength.
-

6. Photoelectric Effect

When light of suitable frequency falls on a metal surface, electrons are ejected — the **photoelectric effect** (explained by Einstein).

- Electrons are emitted only if frequency $\geq$ a minimum **threshold frequency (ν_0)**, however intense the light.
- The minimum energy needed to eject an electron is the **work function ($W_0 = h\nu_0$)**.
- Kinetic energy of the ejected electron: $\frac{1}{2}mv^2 = h\nu - h\nu_0$.
- Increasing intensity increases the *number* of electrons, not their kinetic energy.

The photoelectric effect proved that light behaves as a stream of particles (photons) — its **particle nature**.

7. Atomic Spectra and the Hydrogen Spectrum

An **emission spectrum** is produced when atoms emit radiation; an **absorption spectrum** when they absorb it. Atomic spectra are **line spectra** — discrete coloured lines, not a continuous band — acting as a fingerprint of each element.

The wavelengths of the hydrogen spectrum lines are given by the **Rydberg formula**:

$$\bar{\nu} = 1/\lambda = R_H (1/n_1^2 - 1/n_2^2), \text{ where } R_H = 109677 \text{ cm}^{-1} \text{ and } n_2 > n_1.$$

Series	n_1	Region
Lyman	1	Ultraviolet
Balmer	2	Visible
Paschen	3	Infrared
Brackett	4	Infrared
Pfund	5	Infrared

8. Bohr's Model of the Hydrogen Atom

Niels Bohr combined Rutherford's nuclear model with Planck's quantum idea to explain the hydrogen spectrum.

Postulates

- Electrons revolve only in certain fixed circular paths called **stationary orbits** without radiating energy.
- Angular momentum is quantised: $mvr = nh/2\pi$ ($n = 1, 2, 3 \dots$).
- Energy is absorbed or emitted only when an electron jumps between orbits: $\Delta E = E_2 - E_1 = h\nu$.

Key Results for Hydrogen-like Species

- Radius:** $r_n = 0.529 \times n^2/Z \text{ \AA}$ (52.9 pm for $n = 1$ in H).
- Energy:** $E_n = -13.6 \times Z^2/n^2 \text{ eV/atom} = -2.18 \times 10^{-18} \times Z^2/n^2 \text{ J/atom}$.
- Velocity:** $v_n = 2.18 \times 10^6 \times Z/n \text{ m/s}$.

The negative energy means the electron is bound to the nucleus; energy increases (becomes less negative) as n increases.

Limitations of Bohr's Model

- Works only for one-electron (hydrogen-like) species; fails for multi-electron atoms.
- Could not explain the fine structure of spectral lines, the Zeeman effect (magnetic field) or the Stark effect (electric field).
- Violated the Heisenberg uncertainty principle by fixing both orbit and velocity.

9. Dual Nature of Matter — de Broglie

Louis de Broglie proposed that matter, like light, has a **dual (wave and particle) nature**. Every moving particle has an associated wavelength:

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

The wavelength is significant only for very small particles like electrons; for large objects it is negligibly small. This idea was confirmed by electron diffraction (Davisson–Germer).

10. Heisenberg's Uncertainty Principle

It is impossible to determine simultaneously and exactly both the position and the momentum (or velocity) of a microscopic particle like an electron.

$$\Delta x \cdot \Delta p \geq h/4\pi, \text{ or } \Delta x \cdot \Delta v \geq h/4\pi m$$

This is why the idea of a fixed Bohr orbit is wrong; we can only speak of the **probability** of finding an electron in a region — leading to the concept of an orbital.

11. Quantum Mechanical Model and Quantum Numbers

Schrödinger's wave equation describes the electron as a wave. Its solutions, the **wave functions (ψ)**, give the allowed energy states. ψ^2 gives the probability of finding the electron — the region of highest probability is an **orbital**.

Each electron in an atom is described by four **quantum numbers**:

Quantum Number	Symbol	Tells us	Values
Principal	n	Size and energy of shell	1, 2, 3 ...
Azimuthal	l	Subshell and shape (s,p,d,f)	0 to $(n-1)$
Magnetic	m_l	Orientation of orbital	$-l$ to $+l$
Spin	m_s	Spin of electron	$+\frac{1}{2}$ or $-\frac{1}{2}$

- Maximum electrons in a shell = $2n^2$.
- Number of orbitals in a shell = n^2 ; in a subshell = $(2l + 1)$.
- $l = 0, 1, 2, 3$ corresponds to s, p, d, f subshells.

12. Shapes of Orbitals

- **s-orbital**: spherical, non-directional (1 orientation).
- **p-orbital**: dumb-bell shaped along x, y, z axes (3 orientations), with a node at the nucleus.
- **d-orbital**: double dumb-bell / cloverleaf shaped (5 orientations).

[DIAGRAM: spherical s-orbital; three dumb-bell p-orbitals along x, y, z axes.]

- **Nodes**: regions of zero probability. Radial nodes = $n - l - 1$; angular nodes = l ; total nodes = $n - 1$.

13. Rules for Filling Orbitals — Electronic Configuration

- **Aufbau principle:** electrons fill orbitals of lowest energy first (order by increasing $n + l$; for equal $n + l$, lower n fills first).
- **Pauli exclusion principle:** no two electrons can have all four quantum numbers identical; an orbital holds at most 2 electrons with opposite spins.
- **Hund's rule of maximum multiplicity:** electrons fill degenerate orbitals singly first, all with parallel spin, before pairing.

Filling order: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s ...

Extra stability of half-filled and fully-filled subshells explains exceptions like Cr ($[\text{Ar}]3d^54s^1$) and Cu ($[\text{Ar}]3d^{10}4s^1$), where electrons rearrange for symmetry and exchange-energy stability.

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	6–8 marks	Quantum numbers, Bohr model, electronic configuration, photoelectric effect
JEE Main / Advanced	1–2 questions	Bohr energy/radius, de Broglie, uncertainty principle, Rydberg formula
NEET	2–3 questions	Quantum numbers, electronic configuration, hydrogen spectrum, photons

[TABLE: Question-type split — VSA (1 mark): definitions, quantum number values; SA (2–3 marks): Bohr/de Broglie/photoelectric numericals, configuration; LA (5 marks): postulates of Bohr's model, derivation of energy/radius, quantum-number based questions.]

Important Definitions

Term	Definition
Atomic number (Z)	Number of protons in the nucleus of an atom
Mass number (A)	Total number of protons and neutrons in the nucleus
Isotopes	Atoms of the same element with same Z but different mass number A
Quantum (photon)	Smallest discrete packet of energy of radiation: $E = h\nu$
Threshold frequency (ν_0)	Minimum frequency of light needed to eject an electron from a metal
Work function (W_0)	Minimum energy required to remove an electron from a metal surface: $W_0 = h\nu_0$
Orbit	A fixed circular path (Bohr) in which an electron revolves around the nucleus
Orbital	The three-dimensional region of highest probability of finding an electron
de Broglie wavelength	Wavelength associated with a moving particle: $\lambda = h/mv$
Uncertainty principle	Position and momentum of an electron cannot be known exactly together: $\Delta x \cdot \Delta p \geq h/4\pi$
Aufbau principle	Orbitals are filled in order of increasing energy (lowest first)
Hund's rule	Degenerate orbitals are singly filled with parallel spins before pairing

Solved Examples

Example 1

Calculate the energy of one photon of light of wavelength 4000 Å. ($h = 6.626 \times 10^{-34}$ J·s, $c = 3 \times 10^8$ m/s)

Answer: $E = hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^8)/(4000 \times 10^{-10}) = \mathbf{4.97 \times 10^{-19} \text{ J}}$.

Example 2

Calculate the radius of the second Bohr orbit of the hydrogen atom.

Answer: $r_n = 0.529 \times n^2/Z \text{ Å} = 0.529 \times 4/1 = \mathbf{2.116 \text{ Å (211.6 pm)}}$.

Example 3

Find the energy of the electron in the third orbit ($n = 3$) of the hydrogen atom in eV.

Answer: $E_n = -13.6/n^2 \text{ eV} = -13.6/9 = \mathbf{-1.51 \text{ eV}}$.

Example 4

Calculate the de Broglie wavelength of an electron (mass 9.1×10^{-31} kg) moving with a velocity of 2.2×10^6 m/s.

Answer: $\lambda = h/mv = (6.626 \times 10^{-34})/(9.1 \times 10^{-31} \times 2.2 \times 10^6) = 3.3 \times 10^{-10} \text{ m (3.3 \AA)}$.

Example 5

For the first line of the Balmer series of hydrogen, find $\bar{\nu}$ for the transition $n_2 = 3 \rightarrow n_1 = 2$. ($R_H = 109677 \text{ cm}^{-1}$)

Answer: $\bar{\nu} = R_H(1/2^2 - 1/3^2) = 109677(1/4 - 1/9) = 109677 \times 5/36 = 15233 \text{ cm}^{-1}$ ($\lambda \approx 656 \text{ nm}$).

Example 6

Write the electronic configuration of chromium ($Z = 24$) and explain the anomaly.

Answer: Expected $[\text{Ar}]3d^44s^2$, but the actual configuration is **$[\text{Ar}]3d^54s^1$** . The half-filled $3d^5$ configuration is extra stable because of symmetrical distribution and maximum exchange energy.

Important Questions for Board Exams

1-Mark Questions (VSA)

1. What is the maximum number of electrons that can be accommodated in a shell with principal quantum number n ?
2. Define threshold frequency.
3. How many spherical (radial) nodes are present in a $3s$ orbital?
4. Which quantum number determines the shape of an orbital?
5. Write the relation between de Broglie wavelength and momentum.

2–3-Mark Questions (SA)

1. State Heisenberg's uncertainty principle and give its mathematical expression. Why is it not significant for large objects?
2. Explain the main features of the photoelectric effect that classical physics could not explain.
3. State the Aufbau principle, Pauli exclusion principle and Hund's rule with one example each.
4. Why does the actual electronic configuration of copper differ from the expected one?

5-Mark Questions (LA)

1. State the postulates of Bohr's model of the hydrogen atom and give its limitations.
2. Describe Rutherford's α -scattering experiment, its observations and conclusions, and state its drawbacks.

3. Explain the four quantum numbers, their permissible values and significance, and use them to describe the electrons of a 2p subshell.
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Quick Revision Points

- Electron: $e/m_e = 1.758 \times 10^{11}$ C/kg (Thomson), charge -1.602×10^{-19} C (Millikan)
 - Atomic number Z = protons; mass number A = protons + neutrons
 - Isotopes: same Z ; isobars: same A ; isotones: same neutrons
 - Rutherford: tiny dense nucleus, mostly empty space; drawback — spiralling electron
 - Planck: $E = h\nu = hc/\lambda$; $h = 6.626 \times 10^{-34}$ J·s
 - Photoelectric: $\frac{1}{2}mv^2 = h\nu - h\nu_0$; intensity raises number, not KE
 - Rydberg: $\bar{\nu} = R_H(1/n_1^2 - 1/n_2^2)$; Lyman (UV), Balmer (visible), Paschen (IR)
 - Bohr: $r_n = 0.529 n^2/Z$ Å; $E_n = -13.6 Z^2/n^2$ eV; $mvr = nh/2\pi$
 - de Broglie: $\lambda = h/mv$; Heisenberg: $\Delta x \cdot \Delta p \geq h/4\pi$
 - Quantum numbers: n (shell), l (shape), m_l (orientation), m_s (spin); max $e^- = 2n^2$
 - Orbitals: s spherical, p dumb-bell, d double dumb-bell; total nodes = $n - 1$
 - Filling: Aufbau, Pauli, Hund; half/fully-filled subshells (Cr, Cu) are extra stable
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Next Chapter: Chapter 3 — Classification of Elements and Periodicity in Properties



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