## 🔹 Why Study Atomic Structure?

In Chapter 3, we learned atoms and molecules make up everything. But:

- Why are atoms of different elements different?
- Are atoms really indivisible as Dalton said?

Scientists discovered that atoms are not indivisible — they are made of even smaller particles called sub-atomic particles.

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# 4.1 Charged Particles in Matter

🧪 **Activity 4.1:**

- Comb dry hair → it attracts bits of paper
- Rub glass rod with silk → bring near balloon → balloon gets attracted

✅ Conclusion: Rubbing creates electric charge → things get charged → Matter contains charged particles!

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## 🔹 Sub-atomic Particles

1. **Electron ($e^-$)**
   📘 Discovered by: J.J. Thomson
   📘 Charge: –1 unit
   📘 Mass: negligible (1/2000 of a hydrogen atom)
2. **Proton ($p^+$)**
   📘 Discovered by: E. Goldstein (via canal rays)
   📘 Charge: +1 unit
   📘 Mass: 1 atomic mass unit (u)
3. **Neutron (n)**
   📘 Discovered by: James Chadwick
   📘 Charge: 0
   📘 Mass: ≈ 1 u

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# 4.2 Atomic Models

Let's understand how scientists explained atom's structure using models.

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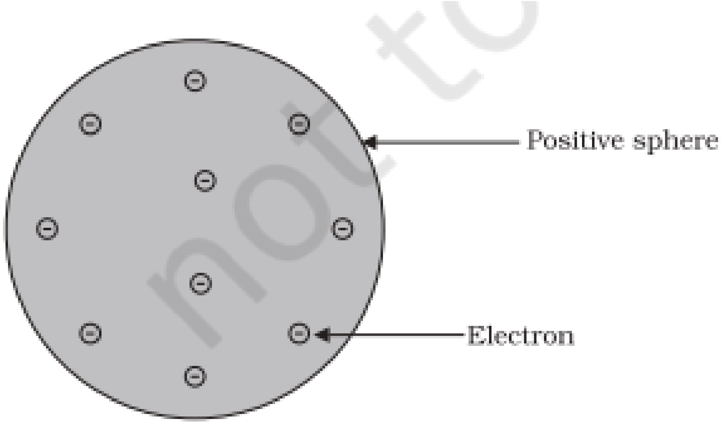
### 🔸 4.2.1 Thomson's Model of Atom

📘 **Definition**: J.J. Thomson proposed that atoms are a sphere of positive charge with negatively charged electrons embedded like seeds in watermelon.

🟩 Key points:

- Atom is positively charged with embedded electrons
- Charges are equal & opposite → Atom is neutral




📷 *Fig. 4.1 – Thomson's Plum Pudding Model*

🛑 **Limitation:**
Couldn't explain results of other experiments, like Rutherford's.

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### 🔶 4.2.2 Rutherford's Alpha-Particle Scattering Experiment

🧪 **What He Did:**

- Bombarded thin gold foil with fast-moving α-particles ($He^{2+}$)
- Observed paths of particles

📘 **Observations:**

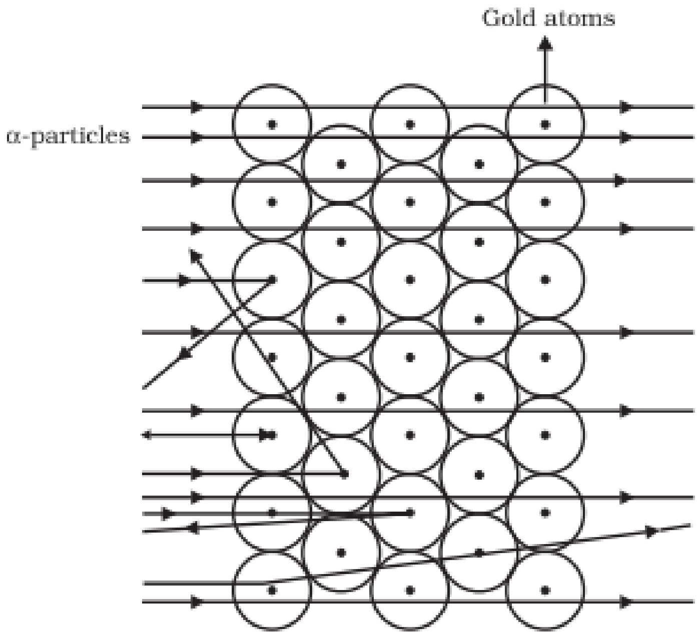
1. Most α-particles passed straight
2. Some deflected slightly
3. A few bounced back!

🟩 **Conclusion:**

- Most of atom is empty space
- Positive charge & mass concentrated in a tiny central part → nucleus
- Electrons revolve around nucleus

📘 **Rutherford's Model:**

- Atom has a dense nucleus (positive)
- Electrons revolve in circular orbits
- Nucleus is very small




📷 *Fig. 4.2 – Rutherford's gold foil experiment*

🛑 **Limitation:**

According to classical physics, electrons moving in circles should lose energy and spiral into nucleus → atom would collapse. But atoms are stable → this model couldn't explain that.

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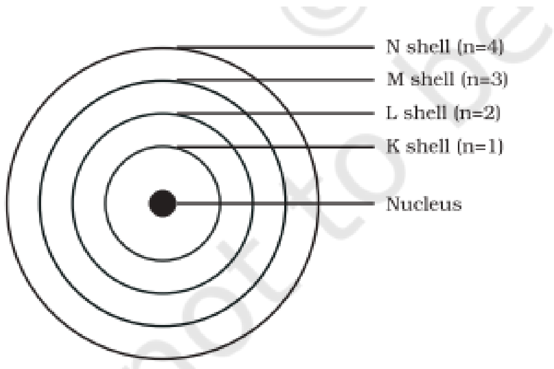
## 🔶 4.2.3 Bohr's Model of Atom

📘 Proposed by: Neils Bohr (to fix Rutherford's model)

🟩 **Key Postulates:**

1. Electrons move in fixed energy levels (orbits/shells)
2. As long as electrons stay in these orbits, they do not lose energy
3. Each orbit has fixed energy and capacity




📷 *Fig. 4.3 – Energy levels around nucleus*

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## 🔶 4.2.4 Neutrons

📘 Neutron → Neutral particle in nucleus
Discovered by: James Chadwick (1932)
Mass ≈ 1 u
Neutrons + Protons = nucleons
All atoms (except hydrogen-1) have neutrons

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## 4.3 Electron Distribution in Shells

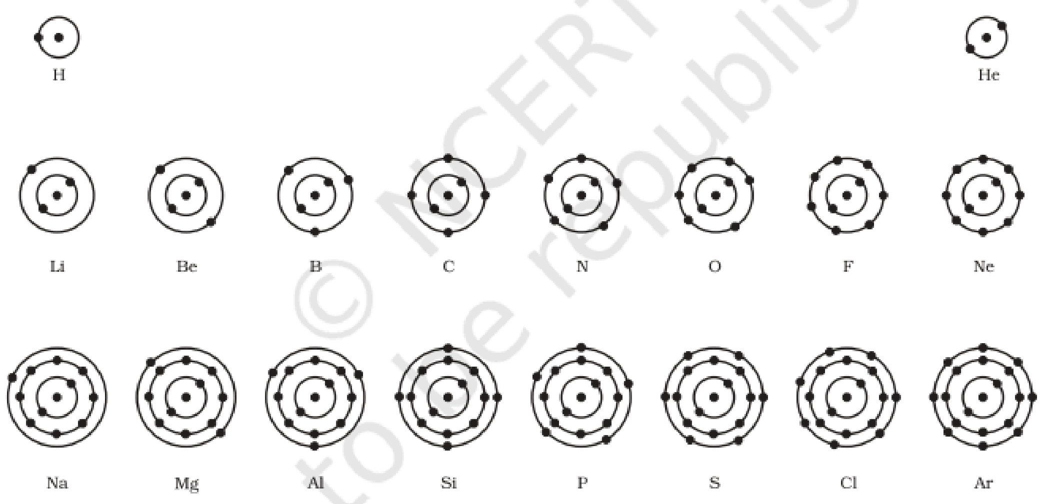
🟩 Rule for Electron Distribution: (Bohr-Bury scheme)

1. Max electrons in shell = $2n^2$
   (n = shell number: K = 1, L = 2...)

📘 **Capacity:**

- K → 2 ($2 \times 1^2$)
- L → 8 ($2 \times 2^2$)
- M → 18 ($2 \times 3^2$)
- N → 32 ($2 \times 4^2$)

1. Outer shell can have max 8 electrons
2. Inner shells must be filled first




📷 *Fig. 4.4 – Atomic structures of first 18 elements*

📘 Table 4.1 – Atomic number, shells, valency for first 18 elements

🧪 **Activity 4.2:**
Make a model showing electronic configurations of 1–18 elements

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## 4.4 Valency

📘 Valency:
The combining capacity of an atom (how many electrons it can lose/gain/share to complete its

outermost shell).

🟩 **Rules:**

- If outer shell has < 4 electrons → valency = number of electrons
- If outer shell has > 4 electrons → valency = 8 − number of electrons
- Completely filled shell (like He, Ne, Ar) → valency = 0

📘 Examples:

- Na (2,8,1) → Valency = 1
- O (2,6) → Gains 2 electrons → Valency = 2
- F (2,7) → Gains 1 → Valency = 1
- Mg (2,8,2) → Loses 2 → Valency = 2

📘 Table 4.1 → Column of valencies

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## 4.5 Atomic Number and Mass Number

🔸 **Atomic Number (Z):**
📘 Definition: Number of protons in an atom's nucleus

🟩 Z = number of protons = number of electrons (for neutral atom)

Example: Z of carbon = 6

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🔸 **Mass Number (A):**
📘 Definition: Total number of protons + neutrons in an atom

🟩 A = number of protons + number of neutrons

Notation:
A
Z X → where X = element symbol

Example:
Carbon → $^{12}_{6}C$ → 6 protons + 6 neutrons = 12

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## 4.6 Isotopes and Isobars

🔸 **Isotopes**

📘 **Definition:**
Atoms of the same element with same atomic number but different mass number.

🟩 Examples:

- $^{1}_{1}H$ (protium), $^{2}_{1}H$ (deuterium), $^{3}_{1}H$ (tritium)
- $^{12}_{6}C$ & $^{14}_{6}C$
- $^{35}_{17}Cl$ & $^{37}_{17}Cl$

📘 **Properties:**

- Chemically same
- Physically different
- Used in medicine & nuclear reactors

🧠 Average atomic mass of Cl = (35×75% + 37×25%) / 100 = 35.5 u

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🔶 **Isobars**

📘 **Definition:**
Atoms of different elements with same mass number but different atomic number

🟩 Example:
$^{40}_{20}Ca$ and $^{40}_{18}Ar$ → different elements, same mass