ATOMIC STRUCTURE

Presented by

Mr. S. S. Chivate

M.Sc. (Chem) NET, B.Ed.

Head, Department of Chemistry and Agrochemicals & Fertilizers

Shri Shivaji College of Arts, Commerce & Science, Kandhar

Affiliated to S.R.T.M. University, Nanded.

Atomic Models

1. Dalton's Atomic Model **(1803)**

Atoms are **tiny, solid, indivisible spheres**

All atoms of the same element are identical

Atoms combine in simple whole-number ratios

Limitation: Didn't explain subatomic particles

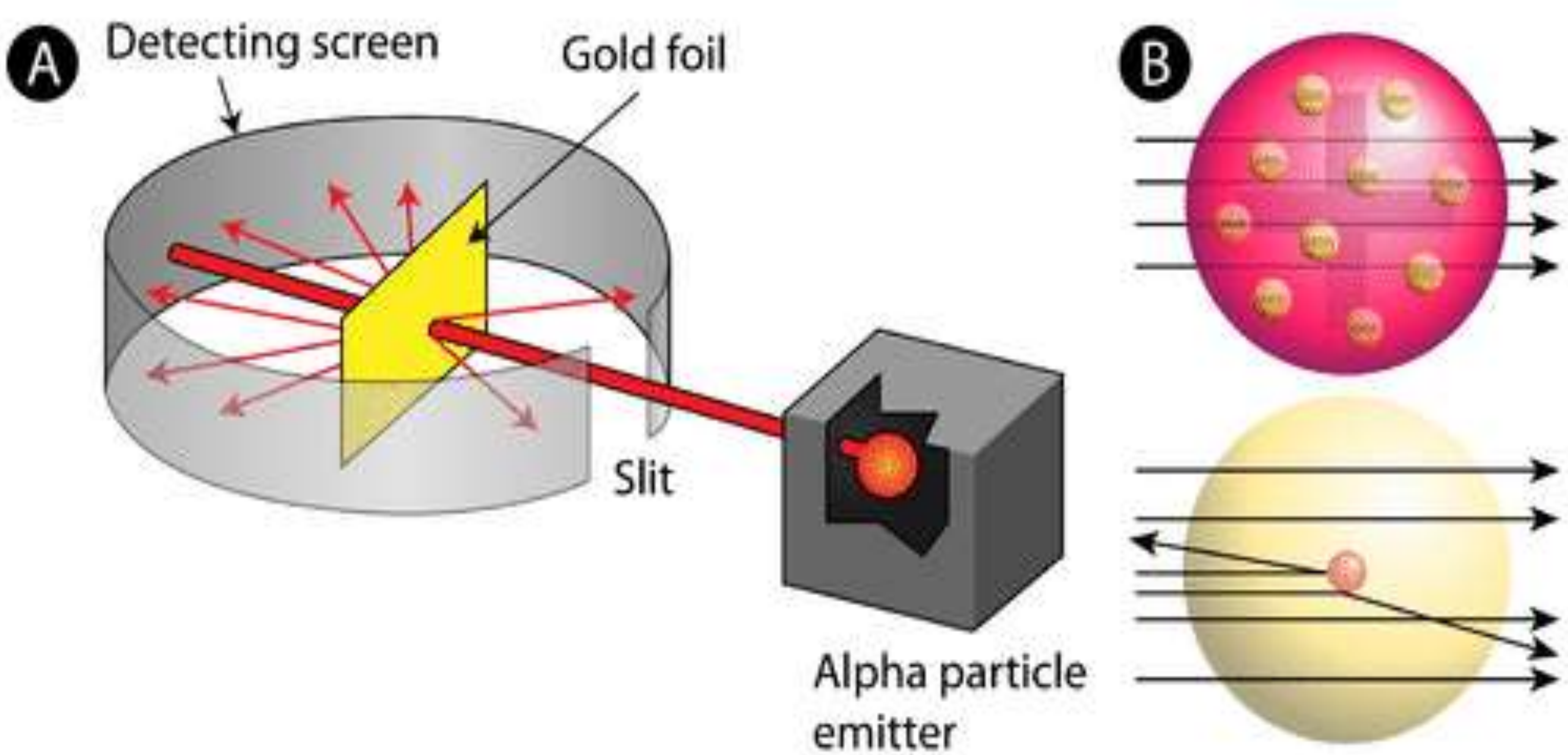
2. Thomson's Plum Pudding Model **(1897)**

Atom is a **positively charged sphere.**

Electrons are embedded inside it like "plums in pudding."

Limitation: No nucleus.

3. Rutherford’s Nuclear Model (1911)



Rutherford’s Alpha Particle Scattering Experiment

Observations

1.Most α-particles passed straight through the gold foil without any deflection.

2.Some α-particles were deflected through small angles.

3.Very few α-particles were deflected back (rebounded).

Conclusions

1.Most of the atom is empty space (because most particles passed straight through).

2.Positive charge and most of the mass are concentrated in a very small region called the **nucleus**.

3.The **nucleus is small, dense, and positively charged**.

4.Electrons revolve around the nucleus.

Limitations

- Could not explain the **stability of electrons** around the nucleus.
- Failed to explain **atomic spectra**.

4. Bohr’s Model (1913)

Postulates :-

1) Electrons move around the nucleus in fixed circular paths called orbits or energy levels.
2) While moving in a permitted orbit, an electron does not lose or gain energy.
3) Each orbit has a fixed energy($E= h\gamma$) and these energies increase as the distance from the nucleus increases.
4) Electrons can revolve only in those orbit whose angular momentum is an integral multiple of $h/2\pi$.

Merits

1. Successfully explains the **line spectrum of hydrogen** and hydrogen-like atoms.
2. Explains **atomic stability** (electrons do not spiral into nucleus).
3. Provides correct values for:
    - Radius of orbit
    - Energy of electron
4. Gives theoretical basis for **Rydberg equation**.
5. Explains **ionization energy** of hydrogen atom.

Demerits

Fails to explain spectra of **multi-electron atoms**.

2. Cannot explain **fine structure** of spectral lines.
3. Does not explain **Zeeman effect** (magnetic field).
4. Does not explain **Stark effect** (electric field).
5. Violates **Heisenberg's uncertty principle**.
6. Does not consider **wave nature of electrons**.

Quantum Numbers

1. Principal Quantum Number (n)
2. Azimuthal Quantum Number (l)
3. Magnetic Quantum Number (m□)
4. Spin Quantum Number (m□)

1. Principal Quantum Number (n)

- Values: 1, 2, 3, …
- Indicates energy level (shell) and size of orbital
- Number of orbitals in a shell $= n^2$
- Maximum electrons in a shell $= \mathbf{2n^2}$

2. Azimuthal Quantum Number (l)

- Values: 0 to (n − 1)
- Determines shape of orbital
- l = 0 (s), l = 1 (p), l = 2 (d), l = 3 (f)

3. Magnetic Quantum Number (m□)

- Values: −l to +l (including zero)
- Determines the orientation of orbitals in space
- Number of orbitals = 2l + 1
- Example: For l = 1 (p-subshell) → m□ = −1, 0, +1 → 3 orbitals

4. Spin Quantum Number (m□)

- Symbol: m□
- Values: +½ or −½
- Significance:

  Describes the spin of the electron

  Explains magnetic properties of atoms

1) Aufbau Principle

- Electrons are filled in orbitals in order of **increasing energy**.

2) Hund's Rule of Maximum Multiplicity

- In a given subshell, electrons occupy degenerate orbitals singly first before pairing.

Pauli's Exclusion Principle

- No two electrons in an atom can have the same set of all four quantum numbers.

Thank you!