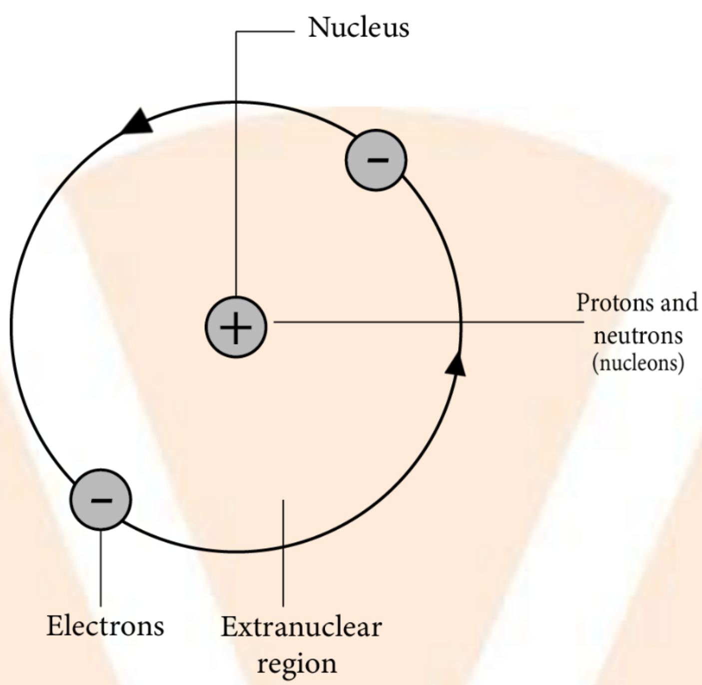
# Revision Notes for Class 9 Science

# Chapter 4 – Structure of the Atom

**Atoms**

- Atoms are the basic units of matter and the defining structure of elements.
- It consists of three basic particles, i.e., protons, electrons, and neutrons, that build the structure of an atom.
- Protons are positively charged particles and were founded by E. Goldstein.
- Electrons are negatively charged particles and were founded by J.J. Thomson.
- Neutrons have no charge and were founded by Chadwick.
- The nucleus is in the centre of the atom and contains protons and neutrons.
- The outer region of the atom which holds electrons in orbit around the nucleus is known as shell/energy level/orbits.
- These shells are further divided into subshells.
- The electrons present in the outermost shell of an atom are known as the valence electrons.

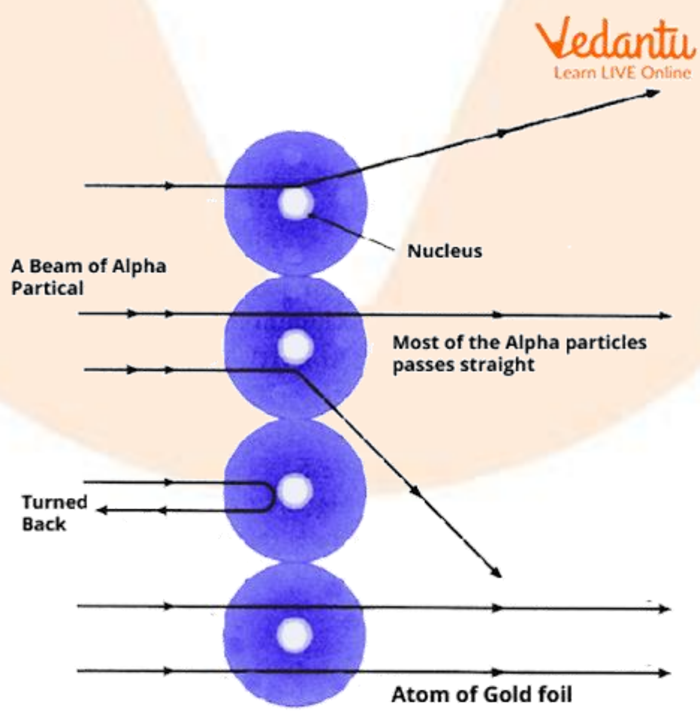



**Thomson's Model of an Atom**

- **Proposed by J.J. Thomson in 1897**: He discovered the electron and suggested a new model of the atom.
- **Plum Pudding Model**: Thomson compared the atom to a pudding or a sphere of positive charge, with negatively charged electrons embedded in it, like plums in a pudding.
- **No Nucleus**: According to this model, there was no central nucleus. The positive and negative charges were spread throughout the atom.
- **Neutral Atom**: The overall charge of the atom was neutral, as the positive charge balanced out the negative charge of electrons.
- **Failed Model**: This model was later rejected after Ernest Rutherford discovered the nucleus.

## Rutherford's Model of an Atom

- **Proposed by Ernest Rutherford in 1911**: Based on his famous gold foil experiment.
- **Nucleus**: He discovered that the atom has a small, dense, positively charged centre called the nucleus.
- **Electrons Orbit**: Electrons revolve around the nucleus in circular paths, similar to planets orbiting the sun.
- **Mostly Empty Space**: Most of the atom's volume is empty space, where electrons move.
- **Positive Charge in Nucleus**: The nucleus contains positively charged protons, which hold most of the atom's mass.
- **Failed to Explain Stability**: While this model explained the atom's structure, it couldn't explain why electrons don't spiral into the nucleus due to attraction, leading to the development of newer models.



**Rutherford's Experiment: Scattering of α-particles by a gold foil**

This scattering experiment was crucial in proposing the nuclear model of the atom.

1. **Conducted by Rutherford in 1909**: Rutherford and his team used a thin sheet of gold foil to observe how alpha (α) particles behave when they hit it.

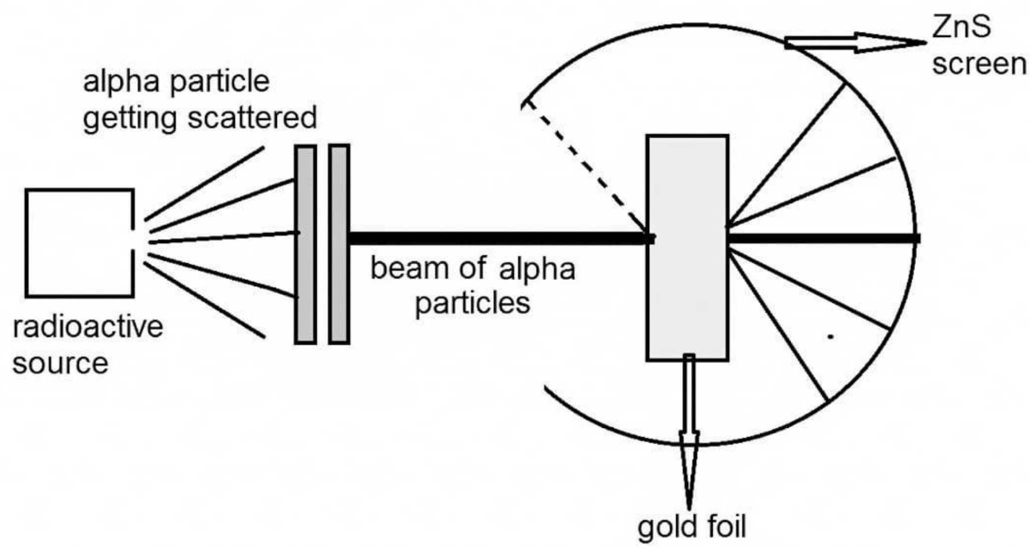
2. **Alpha Particles**: These are positively charged particles (helium nuclei) emitted from a radioactive source.

3. **Most Passed Through**: A majority of the α-particles passed straight through the gold foil without deflection, indicating that most of the atom is empty space.

4. **Some Deflected**: A few α-particles were deflected at small angles, suggesting that they encountered a small, positively charged region.

5. **Few Rebounded**: A very small number of particles bounced back, indicating they hit a dense, positively charged centre (nucleus).

6. **Discovery of the Nucleus**: This experiment led to the conclusion that the atom has a small, dense nucleus where most of its positive charge and mass are concentrated.



## Drawbacks of Rutherford's model of the atom

1. **Electrons in Orbit**: According to Rutherford's model, electrons revolve around the nucleus, which should cause them to lose energy and collapse into the nucleus.

2. **Problem with Classical Physics**: Classical physics predicts that electrons in motion would radiate energy and eventually spiral into the nucleus, leading to atom instability.

3. **Bohr's Explanation**: Niels Bohr solved this by proposing that electrons can only exist in specific, quantized orbits without radiating energy.

4. **Stable Energy Levels**: In Bohr's model, as long as electrons remain in these specific orbits (energy levels), they don't lose energy and the atom stays stable.

5. **Quantum Mechanics**: Modern atomic theory based on quantum mechanics explains that electrons have stable energy states, which prevents them from collapsing into the nucleus.

## Bohr's Model of Atom

- **Proposed by Niels Bohr in 1913**: Bohr modified Rutherford's model to explain atomic stability and the line spectra of elements.

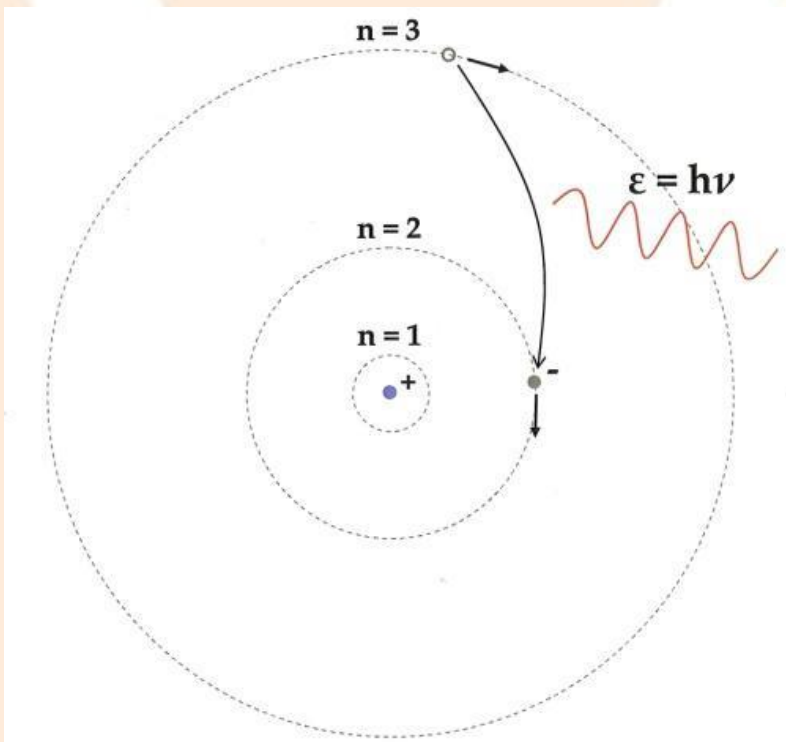
- **Quantized Orbits**: Electrons revolve around the nucleus in specific, fixed orbits or energy levels without radiating energy. These orbits are called quantized orbits.

- **Energy Levels**: The energy of an electron is fixed for each orbit. The closer the orbit to the nucleus, the lower the energy, and vice versa.

- **Electron Transitions**: Electrons can jump between energy levels by absorbing or emitting specific amounts of energy (quanta). When an electron jumps from a higher to a lower energy level, it emits energy in the form of light.

- **Explains Atomic Spectra**: Bohr's model successfully explained the discrete spectral lines seen in the hydrogen atom's emission and absorption spectra.

- **Stable Orbits**: Electrons in a particular orbit do not lose energy, which explains the stability of atoms.




**Neutrons:**

1. **Discovery in 1932**: J. Chadwick discovered a new subatomic particle called the neutron.

2. **No Charge**: Neutrons have no electric charge, unlike protons, which are positively charged.

3. **Mass of a Neutron**: The mass of a neutron is nearly equal to that of a proton.

4. **Presence in Nucleus**: Neutrons are found in the nucleus of all atoms except hydrogen, which has only one proton and no neutron.

5. **Representation**: Neutrons are represented by the symbol 'n'.

6. **Mass of Atom**: The mass of an atom is determined by the sum of the masses of the protons and neutrons in its nucleus.

**How are Electrons Distributed in Different Orbits (Shells)?**

Here's a simplified distribution of electrons in an atom based on Bohr and Bury's rules:

- **Electron Distribution Rules**: Bohr and Bury suggested rules for how electrons are arranged in different orbits or energy levels in an atom.

- **Maximum Electrons in a Shell**: The maximum number of electrons in a shell is calculated using the formula $2n^2$, where $n$ is the orbit number (1, 2, 3...).

- **Examples of Shells**:

  - **K-shell (first orbit)**: Can hold a maximum of 2 electrons ($2 \times 1^2 = 2$).

  - **L-shell (second orbit)**: Can hold up to 8 electrons ($2 \times 2^2 = 8$).

  - **M-shell (third orbit)**: Can hold up to 18 electrons ($2 \times 3^2 = 18$).

  - **N-shell (fourth orbit)**: Can hold up to 32 electrons ($2 \times 4^2 = 32$), and so on.

- **Outermost Shell Rule**: The outermost shell can hold a maximum of **8 electrons**, regardless of the shell's total capacity.

- **Stepwise Filling of Shells**: Electrons fill the shells in a step-by-step manner, meaning inner shells must be filled first before electrons occupy the outer shells.

- **Atomic Structure Example**: The electron configuration of the first 18 elements follows these rules, with their atomic structures shown in diagrams or tables.

**Valency:**

1. **Definition**: Valency is the ability of an atom to combine with other atoms, determined by the number of electrons in the outermost shell.
2. **Valency Rule**: It depends on the number of electrons in the outermost shell:
   - If the outer shell has fewer than 4 electrons, valency = number of outer electrons.
   - If the outer shell has more than 4 electrons, valency = 8 - number of outer electrons.
3. **Stable Configuration**: Atoms try to achieve 8 electrons (octet rule) in their outermost shell by gaining, losing, or sharing electrons.

**Schematic Atomic Structure of the First Eighteen Elements:**

The atomic structure for each of the first 18 elements is based on the arrangement of electrons in shells.

1. **Hydrogen (H)**: Atomic number 1; Valency = 1; Electron distribution: 1 (K-shell).
2. **Helium (He)**: Atomic number 2; Valency = 0; Electron distribution: 2 (K-shell).
3. **Lithium (Li)**: Atomic number 3; Valency = 1; Electron distribution: 2, 1 (K, L).
4. **Beryllium (Be)**: Atomic number 4; Valency = 2; Electron distribution: 2, 2 (K, L).
5. **Boron (B)**: Atomic number 5; Valency = 3; Electron distribution: 2, 3 (K, L).

6. **Carbon (C)**: Atomic number 6; Valency = 4; Electron distribution: 2, 4 (K, L).
7. **Nitrogen (N)**: Atomic number 7; Valency = 3; Electron distribution: 2, 5 (K, L).
8. **Oxygen (O)**: Atomic number 8; Valency = 2; Electron distribution: 2, 6 (K, L).
9. **Fluorine (F)**: Atomic number 9; Valency = 1; Electron distribution: 2, 7 (K, L).
10. **Neon (Ne)**: Atomic number 10; Valency = 0; Electron distribution: 2, 8 (K, L).
11. **Sodium (Na)**: Atomic number 11; Valency = 1; Electron distribution: 2, 8, 1 (K, L, M).
12. **Magnesium (Mg)**: Atomic number 12; Valency = 2; Electron distribution: 2, 8, 2 (K, L, M).
13. **Aluminium (Al)**: Atomic number 13; Valency = 3; Electron distribution: 2, 8, 3 (K, L, M).
14. **Silicon (Si)**: Atomic number 14; Valency = 4; Electron distribution: 2, 8, 4 (K, L, M).
15. **Phosphorus (P)**: Atomic number 15; Valency = 3; Electron distribution: 2, 8, 5 (K, L, M).
16. **Sulfur (S)**: Atomic number 16; Valency = 2; Electron distribution: 2, 8, 6 (K, L, M).
17. **Chlorine (Cl)**: Atomic number 17; Valency = 1; Electron distribution: 2, 8, 7 (K, L, M).
18. **Argon (Ar)**: Atomic number 18; Valency = 0; Electron distribution: 2, 8, 8 (K, L, M).

- The electronic configuration of an element is the representation of the arrangement of electrons distributed among the orbital shells and subshells. The valence electrons are the determining factor for the unique chemistry of the elements:

### Atomic Number

- The atomic number of an element is the same as the number of protons in the nucleus of its atom. As atoms are electrically neutral, an atom contains as many electrons as it has protons. The atomic number is denoted by Z.

### Mass Number

- The mass number of an atom is equal to the number of nucleons in its nucleus. Nucleons are the collective term for protons and neutrons. The mass number is denoted by A.
- In the notation of an atom, the atomic number is written as a subscript on the left of the element symbol and the mass number is written as a superscript on the left of the element symbol.

### Isotopes

- Isotopes can be defined as elements that possess the same number of protons and electrons, but a different number of neutrons.
- For example, Protium, deuterium, and tritium are the isotopes of hydrogen. They each have one single proton Z=1 and a single electron but differ in the number of their neutrons. Hydrogen has no neutrons, deuterium has one, and tritium has two neutrons.

### Isobars and Isotones

- Isobar is an element that differs in chemical property but has the same physical property which means isobars are those elements that have a different atomic number but the same mass number.

- For example, Calcium and chlorine are isobars since both have a mass number of 40 but calcium has an atomic number of 20 and chlorine has an atomic number of 17.

- Isotones are elements that have the same number of neutrons but different atomic numbers.

- For example, Chlorine with atomic number 37 and potassium with atomic number 39 is isotones because both chlorine and potassium have the same number of neutrons i.e., 20.