# *Chapter 4 – Structure of the Atom*

## *1. Introduction to Chapter 4: Structure of the Atom*

***Atoms are the basic building blocks of all matter around us. Even though they are extremely small and invisible to the naked eye, they form everything — from a tiny grain of sand to giant planets. This chapter takes us into the journey of how scientists gradually discovered the structure of the atom.***

◇ ***Why study the structure of atom?***

- ***To understand how elements combine to form compounds.***
- ***To explain chemical reactions and bonding.***
- ***To know how charges inside atoms are arranged.***

## *2. Early Atomic Theories*

### *Ancient Idea – Democritus (400 BC)*

***The earliest concept of the atom came from Greek philosopher Democritus. He wondered what would happen if matter was divided again and again into smaller pieces. He believed that there must be a stage where matter could not be divided further.***

- ***He named these smallest particles "atomos" (meaning indivisible).***
- ***According to him, atoms were eternal, indestructible, and different in shape and size for different substances.***
- ***However, this idea was purely philosophical. There was no experimental evidence.***

☞ ***Though not scientific, Democritus planted the seed for the idea of atoms.***

## 3. Dalton's Atomic Theory (1808)

***Much later, in the early 19th century, John Dalton, an English chemist, gave the first scientific theory of atoms.***

***His ideas were based on experiments with gases and chemical reactions. Dalton proposed that:***

1. ***All matter is made of atoms, which are tiny and indivisible.***
2. ***Atoms of a given element are identical in mass, size, and properties. For example, all oxygen atoms are exactly alike.***
3. ***Atoms of different elements are different in mass and properties. For example, hydrogen atoms differ from nitrogen atoms.***
4. ***Atoms combine in fixed whole-number ratios to form compounds. For instance, water ($H_2O$) is always made of 2 hydrogen atoms and 1 oxygen atom.***
5. ***Atoms are neither created nor destroyed in a chemical reaction; only rearranged.***

***Importance: Dalton's theory explained the laws of chemical combination (Law of conservation of mass, constant proportions, multiple proportions). It became the foundation of modern chemistry.***

### ❌ Limitations:

- ***It said atoms are indivisible, but later discovery of electrons, protons, and neutrons proved atoms have internal structure.***
- ***It couldn't explain why atoms of the same element can differ (as in isotopes).***
- ***It couldn't explain why atoms of different elements sometimes have the same mass number (isobars).***

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## 4. Discovery of Subatomic Particles (Late 19th–20th Century)

***Dalton's theory was very successful, but with advancements in experiments, scientists discovered that atoms are not indivisible. They have smaller particles inside them.***

- ***Electron (J.J. Thomson, 1897): Negatively charged particle.***
- ***Proton (E. Goldstein, 1886, refined by Rutherford): Positively charged particle.***
- ***Neutron (James Chadwick, 1932): Neutral particle.***

***These discoveries proved that Dalton's theory needed modification and gave rise to new atomic models (Thomson, Rutherford, Bohr).***

## 5. *Thomson's Model of Atom (Plum Pudding Model)*

- ***After the discovery of the electron (1897), J.J. Thomson proposed the first model of the atom.***

- ***According to him:***
    1. ***An atom is a positively charged sphere.***
    2. ***Electrons (negative charges) are embedded in it, like seeds in a watermelon or dry fruits in pudding.***
    3. ***The positive and negative charges balance each other, so the atom is neutral.***

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### *Significance:*

- ***First model to describe the internal structure of the atom.***
- ***Explained how atoms remain neutral even with charged particles.***

### *Limitations:*

- ***Could not explain the position and stability of electrons.***
- ***Failed when Rutherford's experiment showed that atoms have a dense nucleus.***

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## 6. *Rutherford's α-Particle Scattering Experiment*

### *The Experiment:*

- ***Scientist: Ernest Rutherford, along with Geiger and Marsden.***
- ***Setup:***
    - ***A very thin gold foil (about 1000 atoms thick) was bombarded with α-particles (positively charged, heavy particles).***
    - ***A fluorescent screen (zinc sulphide) surrounded the foil to detect where the α-particles went.***

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### *Observations:*

1. ***Most α-particles passed straight through → most of the atom is empty space.***
2. ***Some α-particles deflected at small angles → positive charge is not spread everywhere, but concentrated.***

3. ***A very few (1 in 12000) bounced back → there is a dense, positively charged core inside the atom.***

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### *Conclusions:*

***From these observations, Rutherford concluded:***

1. ***Atom has a tiny, dense, positively charged region at the center → called the nucleus.***
2. ***Nucleus carries almost the entire mass of the atom.***
3. ***Electrons revolve around the nucleus, and the rest of the atom is mostly empty space.***

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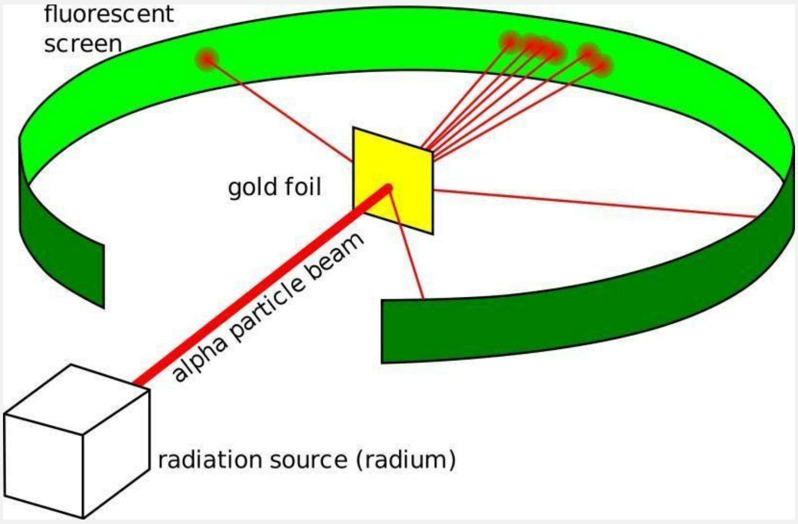
☑ ***In short: Rutherford's experiment showed that atoms are not solid spheres (like Thomson said) but have a small nucleus at the center and electrons moving around it, with large empty space in between.***

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## *7. Bohr's Model of Atom:*

### *Background:*

***Rutherford's model explained the nucleus but could not answer why electrons don't fall into the nucleus (as moving charges should lose energy). To solve this, Niels Bohr proposed a new model.***

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### *Main Postulates of Bohr's Model:*

1. ***Electrons revolve around the nucleus in fixed circular paths called orbits or shells.***

2. ***Each orbit has a fixed energy, so they are called energy levels (K, L, M, N… or n=1, 2, 3, 4…).***
3. ***While revolving in a fixed orbit, an electron does not lose energy.***
4. ***Electrons can jump from one orbit to another:***
    - ***Absorbing energy → move to a higher energy level.***
    - ***Releasing energy → move to a lower energy level.***
5. ***Energy is absorbed or emitted in the form of radiation (light/heat) during these jumps.***

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## *Importance:*

- ***Explained the stability of atoms (why electrons don't spiral into the nucleus).***
- ***Helped in understanding atomic spectra (lines of colors emitted by atoms like hydrogen).***

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## *Limitations:*

- ***Could not explain atoms with many electrons (multi-electron systems).***
- ***Works well mainly for hydrogen atom.***

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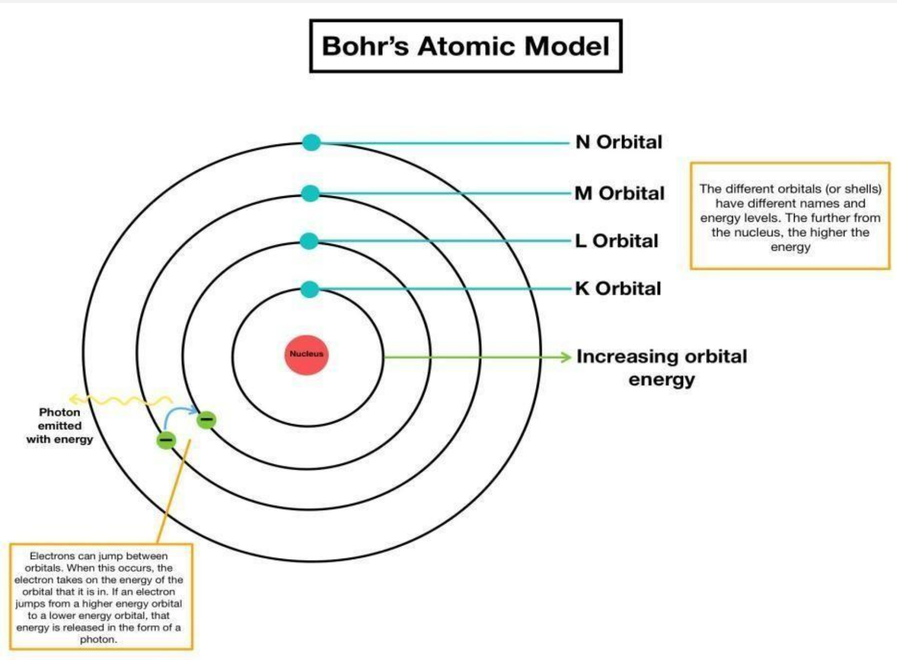
☑ ***In short:*** ***Bohr's model says electrons move in fixed orbits with fixed energies, and atoms emit/absorb energy only when electrons jump between these orbits.***

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## 7. Neutrons

### *Discovery*

- ***Discovered by: James Chadwick in 1932.***
- ***Neutrons are subatomic particles found in the nucleus of an atom.***

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### *Key Features:*

1. ***Charge: Neutrons are neutral, i.e., they carry no electric charge.***
2. ***Mass: Almost equal to that of protons (≈ 1 atomic mass unit).***
3. ***Location: Present inside the nucleus along with protons.***
4. ***Symbol: Usually represented by $n^0$.***

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### *Importance of Neutrons:*

- ***Neutrons contribute to the mass of the atom.***
- ***They provide stability to the nucleus by reducing the electrostatic repulsion between protons.***
- ***Neutrons play a role in nuclear reactions and radioactivity.***

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### *Quick Note:*

- ***Atoms of the same element may have different numbers of neutrons → called isotopes.***
- ***Example: Carbon-12 (6 neutrons), Carbon-14 (8 neutrons).***

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## 8. Distribution of Electrons (Bohr–Bury Rule)

***After Bohr proposed that electrons revolve in fixed orbits (energy levels) around the nucleus, scientists needed a way to arrange electrons in these shells for any atom. The Bohr–Bury Rule provides a simple method to do this.***

### *Bohr–Bury Rule (Octet Rule):*

1. ***Electrons are arranged in shells around the nucleus in the order K, L, M, N… or n = 1, 2, 3, 4…***
2. ***Maximum number of electrons in a shell is given by the formula: Maximum electrons in a shell = $2n^2$ where n = shell number (1 for K, 2 for L, etc.)***

### *Example:*

- ***K-shell (n=1): $2 \times 1^2 = 2$ electrons***

- ***L-shell (n=2): $2 \times 2^2 = 8$ electrons***
- ***M-shell (n=3): $2 \times 3^2 = 18$ electrons***
- ***N-shell (n=4): $2 \times 4^2 = 32$ electrons***

3. ***Electrons fill lower energy levels first before moving to higher levels.***

4. ***Outer shell electrons (valence electrons) determine chemical properties of the element.***

### *Example of Electron Distribution:*

- ***Sodium (Atomic number = 11) → K-shell: 2, L-shell: 8, M-shell: 1 → 2, 8, 1***
- ***Chlorine (Atomic number = 17) → K-shell: 2, L-shell: 8, M-shell: 7 → 2, 8, 7***

### *Summary:*

***The Bohr–Bury Rule helps systematically arrange electrons in shells, determine valence electrons, and predict chemical behavior of elements.***

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## *9. Atomic Number, Mass Number & Isotopes*

### *1. Atomic Number (Z):*

- ***The atomic number of an element is the number of protons in the nucleus of an atom.***
- ***In a neutral atom, number of electrons = number of protons.***
- ***Symbol: Z***
- ***Importance:*** ○ ***Determines the identity of the element.*** ○ ***Determines the position of the element in the periodic table.***
- ***Example:***
  ○ ***Carbon (C): Z = 6 (6 protons)*** ○ ***Oxygen (O): Z = 8 (8 protons)***

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### *2. Mass Number (A):*

- ***The mass number of an atom is the total number of protons and neutrons in its nucleus.***
- ***Symbol: A***
- ***Formula: A = Number of Protons + Number of Neutrons*** • ***Example:***
  ○ ***Carbon-12 → 6 protons + 6 neutrons → A = 12*** ○ ***Oxygen-16 → 8 protons + 8 neutrons → A = 16***

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### 3. Relation Between Atomic Number, Mass Number, and Neutrons:

***Number of Neutrons = Mass Number (A) − Atomic Number (Z)***

- ***Example:***
    - ***Carbon-14 → A = 14, Z = 6 → Neutrons = 14 − 6 = 8***

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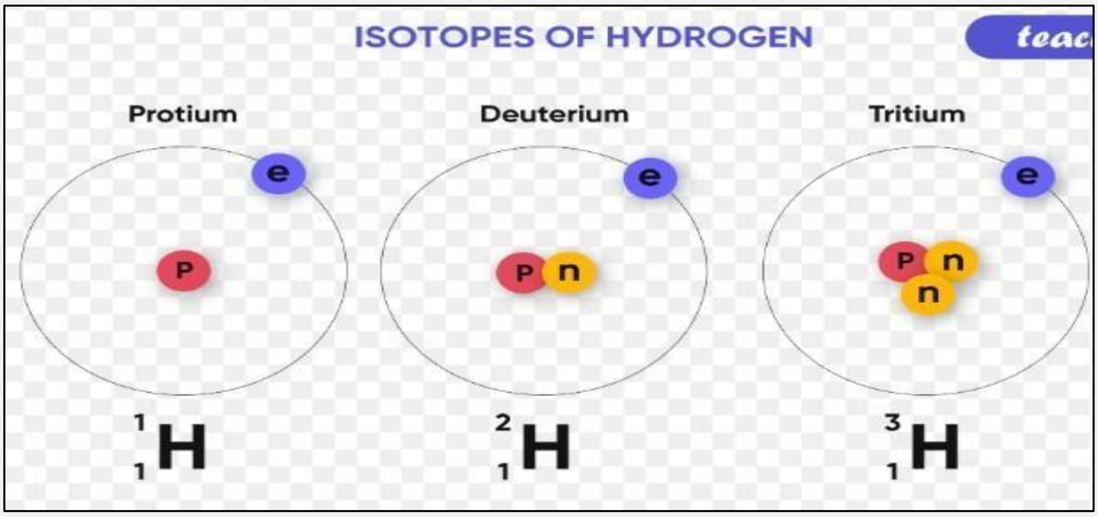
### 4. Isotopes:

- ***Atoms of the same element having same number of protons (Z) but different number of neutrons (A) are called isotopes.***
- ***Properties:***
    - ***Same chemical properties***
    - ***Different physical properties***
- ***Example: Hydrogen isotopes:***

1. ***Protium → 1 proton, 0 neutrons → Mass number = 1***
2. ***Deuterium → 1 proton, 1 neutron → Mass number = 2***
3. ***Tritium → 1 proton, 2 neutrons → Mass number = 3***

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### 5. Importance of Isotopes:

- ***Used in medicine (radioactive isotopes for diagnosis and treatment)***
- ***Used in carbon dating (Carbon-14)***
- ***Used in nuclear energy***

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### Summary:

- ***Atomic number (Z) = Number of protons***
- ***Mass number (A) = Number of protons + neutrons***
- ***Neutrons = A − Z***
- ***Isotopes = Same Z, different A***

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## *10. Valency*

### *1. Definition:*

***Valency of an element is the combining capacity of an atom, i.e., the number of electrons an atom can gain, lose, or share to form a stable compound.***

### *2. How Valency is Determined:*

- ***Atoms combine to complete their outermost shell (octet rule).***
- ***Valency is usually equal to the number of electrons needed to complete the outer shell.***
- ***Example:***
  - ***Hydrogen (H): 1 electron in outer shell → needs 1 electron → Valency = 1*** ○ ***Oxygen (O): 6 electrons in outer shell → needs 2 electrons → Valency = 2***

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### *3. Types of Valency:*

1. ***Electro-valency (Ionic Compounds):***
   - ***Formed by transfer of electrons.***
   - ***Example: Sodium (Na) loses 1 electron → $Na^{+}$ → Valency = 1***
2. ***Covalency (Covalent Compounds):***
   - ***Formed by sharing of electrons.***
   - ***Example: Hydrogen shares 1 electron with another hydrogen → $H_2$ → Valency = 1***

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### *4. Special Notes:*

- ***Some elements show variable valency.***
  - ***Example: Iron (Fe) → $Fe^{2+}$ (valency 2) or $Fe^{3+}$ (valency 3)***
- ***Noble gases generally have valency 0 because their outer shell is complete.***

### *Summary:*

- ***Valency shows how atoms combine to form compounds.***
- ***Determined by number of electrons to complete the outer shell.***
- ***Helps predict the chemical formula of compounds.***

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## *11. Summary Table*

| ***Concept*** | ***Key Points*** |
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| ***Subatomic Particles*** | ***Electron (-), Proton (+), Neutron (0)*** |
| ***Models*** | ***Dalton → Thomson → Rutherford → Bohr*** |
| ***Bohr–Bury Rule*** | ***Max electrons: $2n^2$*** |
| ***Atomic Number (Z)*** | ***No. of protons*** |
| ***Mass Number (A)*** | ***Protons + Neutrons*** |
| ***Isotopes*** | ***Same Z, diff A*** |
| ***Isobars*** | ***Same A, diff Z*** |
| ***Valency*** | ***Combining capacity*** |