Chapter 3: Atomic Structure

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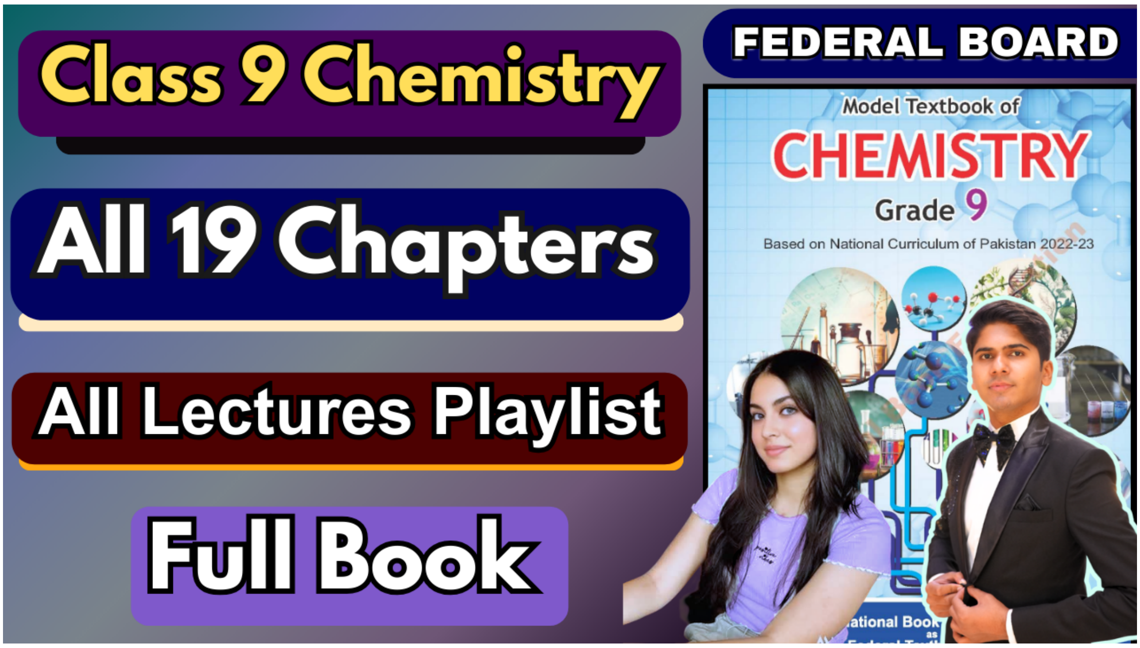

3.1. Introduction to Atomic Models:

Scientists developed different atomic models over time. These models help us understand **how atoms behave** and **their structure**. Every new model improved our knowledge about **atomic structure**.

Dalton's Atomic Model (1803)

John Dalton (British Chemist) gave the **first scientific theory** of matter called **Dalton's Atomic Theory**.

Main Postulates (Important Points):

- Matter is made of tiny particles called **atoms** that **cannot be divided**.
- Atoms of the **same element are identical** in: Mass, Size, Volume
- Atoms **combine**, **separate**, or **rearrange** in **simple ratios** to form compounds during chemical reactions.
- Atoms cannot be created or destroyed.

Importance of Dalton's Model:

- Explained many chemical laws like: **Law of Multiple Proportions** and **Law of Conservation of Mass**
- Helped in understanding **chemical combinations**.

Rutherford's Atomic Model (1911)

Ernest Rutherford performed an experiment to understand the **structure of atom**.

Rutherford's Experiment:

- Used **alpha particles** (He^{2+} nuclei) from polonium.
- These were bombarded at a **very thin gold foil** (0.00004 cm thick).
- Expected particles to pass straight through the foil.
- Observed that most alpha particles passed straight through, but a small fraction were deflected (bounced back at large angles).

Rutherford's Conclusions/Postulates:

- **Atom is mostly empty space** (because most particles passed through).
- **Nucleus** (The centre of the atom) is a tiny, dense, positively charged part (because some positive alpha particles were deflected).
- The size of the nucleus is very small compared to the size of the atom.
- Electrons orbit the nucleus (like planets orbit the sun) $\rightarrow$ **Planetary Model**.

Defects of Rutherford's Model (Why it was Incomplete):

1. According to Classical Physics, a moving, charged particle (like the orbiting electron) should constantly lose energy by emitting radiation. If the electron loses energy, its orbit size should decrease, causing it to spiral into the nucleus. This would make the atom unstable, which is incorrect.
2. The model could not explain why atoms emit light only at certain, discrete wavelengths (atomic spectra).

Bohr's Atomic Model (1913)

Niels Bohr (Danish physicist) proposed this model to correct Rutherford's defects by applying **Quantum Theory**.

Postulates of Bohr's Model:

- **Energy is Quantized:** Electrons orbit the nucleus only in certain fixed paths called **orbits** or **shells** or **energy levels**.
- Each shell is associated with a fixed, specific amount of energy and is represented by the letter 'n' (n=1, 2, 3, etc.) or K, L, M, N...
- **Electron Stability:** As long as the electron stays in its orbit, it does **NOT** gain or lose energy. (This fixed the stability problem of Rutherford's model).
- **Energy Absorption/Emission:** An electron only gains or loses energy when it moves from one shell to another.

  - **Excitation (Absorption):** Electron **absorbs** energy and moves from a lower shell to a higher shell.
  - **De-excitation (Emission):** Electron **emits** energy (as light) and moves from a higher shell to a lower shell.
- The amount of energy released or absorbed is given by the equation: $E = h\nu$ (where E is energy, h is Planck's constant, and ν is the frequency of light).

Success of Bohr's Model:

- Explained the **stability** of the atom.
- Successfully explained the **atomic spectrum of Hydrogen** (emission lines).

Defects of Bohr's Model:

- Only worked for **Hydrogen** and hydrogen-like ions (single electron systems like He^{+}).
- It could **not** explain the spectra of atoms with more than one electron (multi-electron atoms).
- It could **not** explain the splitting of spectral lines in the presence of a magnetic field (**Zeeman effect**) or an electric field (**Stark effect**).

3.2. Sub-Atomic Particles

The particles that make up an atom. The three main sub-atomic particles are: **electrons**, **protons**, and **neutrons**.

- **Electron (e^{-}):** Discovered by **J.J. Thomson**. Carries a **negative** charge.
- **Proton (p^{+}):** Discovered by **Ernest Rutherford**. Carries a **positive** charge.
- **Neutron (n):** Discovered by **James Chadwick**. Has **no charge** (neutral).

Location and Mass:

- **Protons and Neutrons** are located in the **nucleus** (collectively called **nucleons**). They account for almost all the mass of the atom.
- **Electrons** revolve around the nucleus in shells. They have negligible mass compared to protons and neutrons.

Atomic Number (Z)

- Defined as the **number of protons** in the nucleus of an atom.
- For a neutral atom, Atomic Number = Number of Protons = Number of Electrons.
- The atomic number determines the **chemical identity** of an element.

Mass Number (A)

- Defined as the **total number of protons and neutrons** in the nucleus.
- Mass Number (A) $=$ Number of Protons (Z) $+$ Number of Neutrons
- Number of Neutrons $=$ Mass Number (A) $-$ Atomic Number (Z)

Isotopes

Atoms of the **same element** that have the **same atomic number (Z)** but a **different mass number (A)**. This difference is due to the difference in the **number of neutrons**.

Examples of Isotopes:

- **Hydrogen Isotopes:**
  1. **Protium** (^{1}H): 1 proton, 0 neutrons (Most common)
  2. **Deuterium** (^{2}H): 1 proton, 1 neutron
  3. **Tritium** (^{3}H): 1 proton, 2 neutrons (Radioactive)
- **Chlorine Isotopes:** ^{35}Cl and ^{37}Cl
- **Uranium Isotopes:** ^{235}U and ^{238}U

Uses of Isotopes:

- **Carbon-14 (^{14}C)**: Used for **radioactive dating** of very old fossils and archaeological artifacts.
- **Iodine-131 (^{131}I)**: Used to treat **thyroid cancer** and diagnose thyroid gland disorders.
- **Cobalt-60 (^{60}Co)**: Used for **radiotherapy** (cancer treatment).
- **Uranium-235 (^{235}U)**: Used as a **fuel in nuclear power plants**.

3.3. Electron Energy Levels

The concept of electron energy levels was first introduced by Niels Bohr.

Shells (Principal Energy Levels)

- Represented by the Principal Quantum Number (**n**).
- $n = 1$ is the lowest energy level (closest to the nucleus).
- $n = 1 \to$ K shell; $n = 2 \to$ L shell; $n = 3 \to$ M shell; $n = 4 \to$ N shell.
- **Maximum number of electrons** in a shell is given by the formula $\mathbf{2n^2}$.
  - K shell ($n = 1$): $2(1)^2 = 2$ electrons
  - L shell ($n = 2$): $2(2)^2 = 8$ electrons

- M shell ($n = 3$): $2(3)^2 = 18$ electrons
- N shell ($n = 4$): $2(4)^2 = 32$ electrons

Sub-shells

Energy levels (shells) are further divided into sub-shells.

- Represented by the Azimuthal Quantum Number (**l**).
- The number of sub-shells in a shell is equal to the value of **n**.
- The sub-shells are named $\mathbf{s}$, $\mathbf{p}$, $\mathbf{d}$, and $\mathbf{f}$.
- $\mathbf{s}$ sub-shell: $l = 0$; $\mathbf{p}$ sub-shell: $l = 1$; $\mathbf{d}$ sub-shell: $l = 2$; $\mathbf{f}$ sub-shell: $l = 3$.

Orbitals

Sub-shells are composed of orbitals. An orbital is a **region of space** where there is a **high probability** (90%) of finding an electron.

- **s sub-shell:** 1 orbital (spherical shape) $\rightarrow$ max 2 electrons
- **p sub-shell:** 3 orbitals (p_x, p_y, p_z) (dumbbell shape) $\rightarrow$ max 6 electrons
- **d sub-shell:** 5 orbitals $\rightarrow$ max 10 electrons
- **f sub-shell:** 7 orbitals $\rightarrow$ max 14 electrons

Maximum number of electrons in any orbital is **2**.

3.4. Electronic Configuration

The arrangement of electrons in the shells and sub-shells of an atom.

Rules for Electronic Configuration

1. Aufbau Principle (Building-up Principle)

According to this principle, electrons fill the **lowest energy sub-shell that is available first**. This means electron will fill first 1s, then 2s, then 2p and so on.

$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow \ldots$

The increasing order of energy of sub-shells (according to the **n + l rule**) is: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p$

Example for Oxygen (Z $= 8$): $1s^2, 2s^2, 2p^4$

Symbols for Atoms and Ions

The symbol for an atom represents the element. It consists of one or two-letters, the mass number as a left superscript, the atomic number as a left subscript, and the charge as a right superscript. For example,

$${}^{24}_{12}\text{Mg}^{2+}$$

The atomic number is often omitted. This diagram shows a symbol for magnesium “Mg” which stands for magnesium. The number to the upper left of the symbol is the mass number, which is 24. The number to the upper right of the symbol is the charge which is positive 2. The number to the lower left of the symbol is the atomic number which is 12.

Planck Constant (Additional Information)

Planck’s constant, denoted by the symbol ’h’, is a fundamental constant in physics, specifically within quantum mechanics. It relates the energy of a photon (a quantum of light) to its frequency: $E = h\nu$. It is a small number: h $\approx 6.626 \times 10^{-34}$ J $\cdot$ s. Its discovery by Max Planck was foundational to the development of quantum theory, which revolutionized the understanding of energy at the atomic and subatomic levels.


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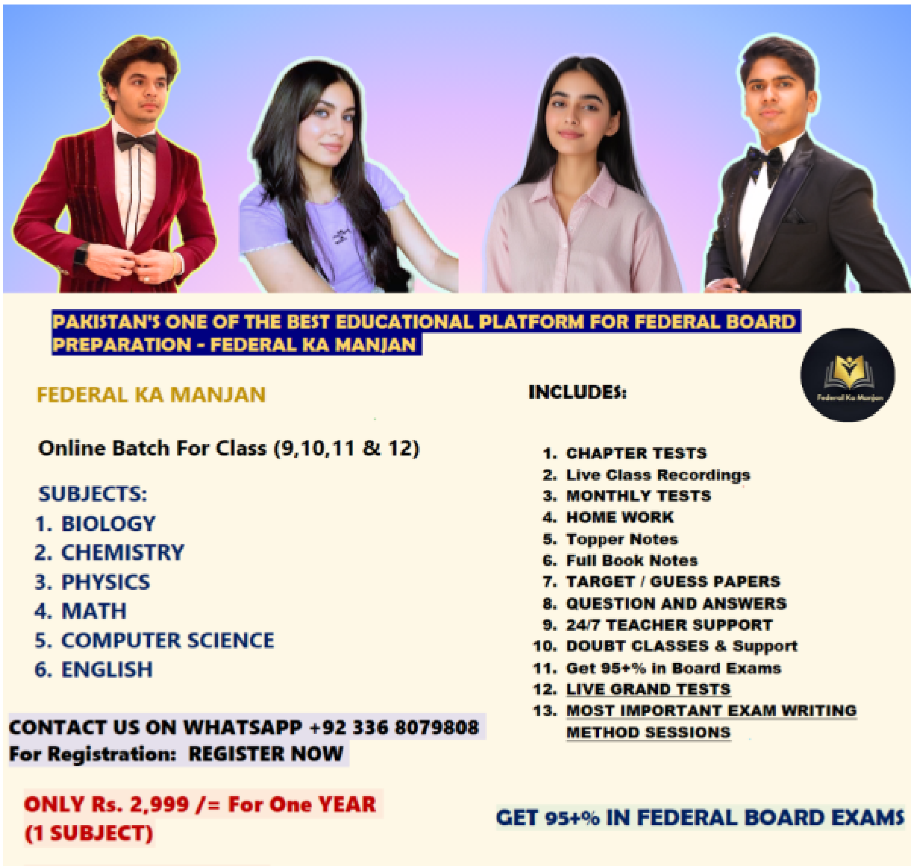
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