Unit 3

STRUCTURE OF THE ATOM

3.1 Historical Development of the Atomic Theories of Matter

Introduction to Atomic Theory

The concept of the atom has evolved over centuries, from ancient philosophical ideas to modern scientific theories. Understanding the historical development of atomic theory helps us appreciate how our current understanding of matter has been shaped.

Early Greek Philosophers and Atomic Theory

1. **Empedocles and the Four Elements (5th Century BC)**
    - **Empedocles** proposed that all matter was composed of four fundamental elements: earth, air, water, and fire.
    - **Plato** and **Aristotle** supported this idea, with Aristotle adding that these elements could be differentiated by properties like hot vs. cold and wet vs. dry.
2. **Democritus and Atomism (460-370 BC)**
    - **Democritus**, a Greek philosopher, expanded on the idea of atoms. He proposed that everything is composed of indivisible and indestructible particles called atoms.
    - The term **atom** comes from the Greek word *atomos*, meaning "indivisible".
    - Democritus believed that atoms were the smallest units of matter and that they existed in empty space, or void. This idea included:
        - Atoms are indivisible and indestructible.
        - There is void between atoms allowing movement.
        - Atoms are solid and homogeneous.
        - Atoms differ in size, shape, and weight.
3. **Limitations of Early Greek Theories**
    - Despite the profound insights, early Greek philosophers lacked experimental evidence. Their theories were based on reasoning rather than experimentation and observation. Their ideas were largely ignored for centuries.

Rebirth of Atomic Theory in the 17th and 18th Centuries

1. **Antoine Lavoisier and the Law of Conservation of Mass (1789)**
    - Lavoisier's experiments showed that mass is conserved in chemical reactions, laying the groundwork for modern chemistry. His work demonstrated that matter is neither created nor destroyed in chemical reactions.
2. **John Dalton's Atomic Theory (Early 19th Century)**
    - Dalton built upon the ideas of early Greek philosophers and introduced a modern atomic theory. His key points included:

- Each element is composed of atoms.
- Atoms of a given element are identical in mass and properties.
- Atoms combine in simple, whole-number ratios to form compounds.
- Chemical reactions involve the rearrangement of atoms.

3.2 Fundamental Laws of Chemical Reactions

1. The Law of Conservation of Mass

- **Definition:** The Law of Conservation of Mass states that in a closed system, the total mass remains constant during a chemical reaction. This means that mass cannot be created or destroyed.
- **Origin:** Proposed by Antoine Lavoisier in 1789, this law was fundamental in establishing the principles of modern chemistry.
- **Application:** In any chemical reaction, the mass of the reactants equals the mass of the products.

  Example: Burning wood seems to lose mass, but the mass of the ash, gases, and remaining charcoal combined equals the original mass of the wood and oxygen.

- **Experiment:** To demonstrate this law, measure the mass of reactants before and after a reaction. In the reaction of silver nitrate and hydrochloric acid, the mass of the system remains unchanged before and after the reaction.

  Procedure Summary:

  1. React silver nitrate with hydrochloric acid.
  2. Weigh the system before and after the reaction to show that mass remains constant.

2. The Law of Definite Proportions

- **Definition:** The Law of Definite Proportions, also known as Proust's Law, states that a chemical compound contains the same elements in the same proportion by mass, regardless of the amount of the compound.
- **Example:** Water (H_2O) always consists of 2 grams of hydrogen for every 16 grams of oxygen, regardless of the source or amount of water.

  Calculation: For 25 g of water, the ratio of hydrogen to oxygen by weight remains 1:8.

3. The Law of Multiple Proportions

- **Definition:** The Law of Multiple Proportions states that when two elements form more than one compound, the ratios of the masses of one element that combine with a fixed mass of the other element are in small whole numbers.

- **Example:** Carbon monoxide (CO) and carbon dioxide (CO_2) show that for the same amount of carbon, the mass of oxygen in CO_2 is twice that in CO.

Conclusion

The historical development of atomic theory shows the progress from philosophical ideas to scientifically validated theories. Understanding the laws of chemical reactions helps in explaining and predicting chemical behavior, providing a foundation for further study in chemistry.

3.1 Atomic Theory

Definition of a Scientific Theory

A scientific theory is a comprehensive explanation of a natural phenomenon based on repeated testing and verification using the scientific method. It provides a framework for understanding and interpreting facts about the natural world. Atomic Theory explains the nature and behavior of atoms, the fundamental building blocks of matter.

Historical Development of Atomic Theory

The concept of atoms dates back to ancient Greece, where philosophers like Democritus proposed that matter is made up of indivisible particles called atoms. However, it was not until the early 19th century that the atomic theory was scientifically developed and tested.

3.1.1 Dalton’s Atomic Theory

John Dalton’s Contribution

In 1803, John Dalton, an English chemist and physicist, introduced his atomic theory based on experimental evidence. His theory provided a more accurate description of atoms and their interactions.

Key Postulates of Dalton’s Atomic Theory:

1. **Elements are composed of tiny, indivisible particles called atoms.**
2. **Atoms of the same element are identical in size, mass, and other properties.**
3. **Atoms of different elements differ in size, mass, and properties.**
4. **Atoms cannot be created or destroyed in chemical reactions.**
5. **Atoms combine in simple, whole-number ratios to form compounds.**

Dalton’s theory supported the Law of Definite Proportions, which states that a chemical compound always contains the same elements in the same proportions by mass.

Symbols Proposed by Dalton

Dalton also introduced symbols for elements and compounds, though these symbols were complex. Later, Berzelius proposed a simpler system using letters, which became more widely accepted.

Successes and Limitations

- **Successes:** Dalton's theory accurately described how elements combine to form compounds.
- **Limitations:** Dalton's idea that atoms are indivisible was later proven incorrect. Modern science has shown that atoms consist of subatomic particles—protons, neutrons, and electrons.

3.1.2 Modern Atomic Theory

Advancements in Atomic Theory

The understanding of atoms has evolved significantly since Dalton's time. Key contributions from various scientists have refined our knowledge:

1. **Eugene Goldstein (1886):** Predicted the existence of protons in the atom.
2. **J.J. Thomson (1897):** Discovered the electron and its charge-to-mass ratio, proposing the 'Plum Pudding' model.
3. **Robert Millikan (1909):** Determined the charge and mass of the electron.
4. **Ernest Rutherford (1911):** Developed the planetary model of the atom, showing a nucleus surrounded by electrons.
5. **Niels Bohr (1913):** Proposed that electrons orbit the nucleus in fixed energy levels.
6. **James Chadwick (1932):** Discovered the neutron, another subatomic particle in the nucleus.

Modern Atomic Theory Postulates

The modern atomic theory incorporates these advancements:

1. **Elements are made of small particles called atoms.**
2. **Atoms cannot be created or destroyed during ordinary chemical reactions.**
3. **Atoms of the same element have the same atomic number but may vary in mass due to isotopes.**
4. **Atoms of different elements are distinct.**
5. **Atoms combine in whole-number ratios to form compounds.**

3.2 Structure of the Atom

Basic Structure of an Atom

An atom consists of three main subatomic particles:

1. **Protons:** Positively charged particles located in the nucleus.
2. **Neutrons:** Neutral particles (no charge) also located in the nucleus.
3. **Electrons:** Negatively charged particles orbiting the nucleus in electron shells or energy levels.

Atomic Models

1. Dalton's Model: Atoms were considered indivisible particles.

2. Thomson's Plum Pudding Model: Proposed that atoms are a positive sphere with embedded electrons.

3. Rutherford's Model: Suggested that atoms have a dense nucleus with electrons orbiting around it.

4. Bohr's Model: Introduced fixed orbits for electrons around the nucleus.

5. Quantum Mechanical Model: Describes electrons in probabilistic regions called orbitals rather than fixed orbits.

3.3 Atomic Number and Mass Number

Atomic Number

The atomic number (Z) is the number of protons in the nucleus of an atom. It defines the element and determines its position on the periodic table. For example, the atomic number of carbon is 6, meaning it has 6 protons.

Mass Number

The mass number (A) is the total number of protons and neutrons in the nucleus of an atom. It is calculated as:

Mass Number=Number of Protons+Number of Neutrons

Isotopes

Isotopes are variants of an element with the same number of protons but different numbers of neutrons. They have the same atomic number but different mass numbers. For example, carbon-12 and carbon-14 are isotopes of carbon with mass numbers of 12 and 14, respectively.

Example Calculation

For a carbon atom with 6 protons and 6 neutrons:

- **Atomic Number:** 6
- **Mass Number:** 6 (protons) + 6 (neutrons) = 12

Understanding the atomic number and mass number is crucial for identifying elements and their isotopes, which have important applications in chemistry and various scientific fields.

3.4 Discoveries of the Fundamental Subatomic Particles and the Atomic Nucleus

Following Dalton's Atomic Theory, scientists delved deeper into the nature of the atom. This section covers the discoveries of fundamental subatomic particles, including electrons, protons, neutrons, and the nucleus.

3.4.1 Discovery of the Proton

Historical Context and Experiment:

The concept of positively charged particles in atoms was first proposed by Eugen Goldstein in 1886. Goldstein conducted experiments using a cathode ray tube with a perforated cathode. He observed rays that passed through the perforations and were not deflected by the cathode. These rays were named "anode rays" or "canal rays."

Understanding Anode Rays:

Anode rays are produced when an electric discharge ionizes gas in the tube, creating positive ions. These positive ions move towards the perforated cathode and produce a glow on the glass of the discharge tube. Goldstein found that the particles in these rays had the highest charge-to-mass ratio when hydrogen gas was used. Rutherford later identified these particles as protons, confirming that they are present in all matter and are the only positively charged particles in an atom.

Properties of Anode Rays:

1. **Travel in Straight Lines:** Anode rays move in straight paths.
2. **Material Particles:** They consist of material particles.
3. **Deflection:** They are deflected in electric and magnetic fields opposite to cathode rays.
4. **Nature of Rays:** Their charge-to-mass ratio depends on the gas in the tube.
5. **Positive Ions:** Anode rays are positively charged ions.

3.4.2 Discovery of the Electron

The Crookes Discharge Tube:

In 1855, Heinrich Geissler developed the mercury pump, which, when improved by Sir William Crookes, led to the discovery of cathode rays. Crookes' Discharge Tube, or Cathode Ray Tube, includes a glass tube with two electrodes connected to a high-voltage source. The resulting discharge creates cathode rays, which are visible as a greenish glow.

J.J. Thomson's Experiments:

J.J. Thomson discovered the electron in 1897. His experiments included:

1. **Path of Cathode Rays:** Thomson showed that cathode rays travel in straight lines and cast shadows of objects placed in their path.
2. **Particle Nature:** Thomson placed a paddle wheel in the path of cathode rays and observed its rotation, confirming that cathode rays are composed of particles with mass.
3. **Charge of Cathode Rays:** Thomson used electric and magnetic fields to show that cathode rays are negatively charged, calculating the charge-to-mass ratio of electrons as 1.76 x 10^8 C/g.

Properties of Cathode Rays:

- **Straight-Line Travel:** They travel in straight lines and cast shadows.
- **Mechanical Effect:** They can move a paddle wheel, indicating they have mass.
- **Charge-to-Mass Ratio:** The ratio is constant regardless of the gas in the tube.
- **Deflection:** They bend towards positive plates in electric fields and north poles in magnetic fields.
- **Ionization:** They ionize gases, causing fluorescence in the tube.
- **Penetrating Power:** They can pass through thin metal foils.
- **X-Ray Production:** They can produce X-rays when they strike certain metals.

3.4.3 Discovery of the Nucleus

Rutherford's Experiment:

In 1920, Ernest Rutherford challenged Thomson's atomic model. Using α-particles and a thin gold foil, Rutherford observed that while most α-particles passed through the foil, some were deflected at large angles. This experiment suggested that the atom has a dense, positively charged nucleus. Rutherford proposed that protons and neutrons reside in the nucleus, while electrons orbit around it.

Explanation for Nuclear Stability:

Rutherford also addressed why the nucleus doesn't disintegrate despite proton repulsion. He hypothesized the presence of neutrons, which neutralize proton repulsion. For atomic stability, he explained that electrons remain in orbit due to a balance between electrostatic attraction to the nucleus and centrifugal force.

3.5 Composition of an Atom and the Isotopes

Atoms are made up of subatomic particles—electrons, protons, and neutrons. These particles determine an atom's mass, charge, and overall properties. In this section, we'll explore each particle and discuss isotopes.

3.5.1 Electrons, Protons, and Neutrons

Electrons

- **Description**: Electrons are fundamental particles with a negative electric charge. They are extremely small, with a mass about 1/2000 of a proton or neutron, roughly 0.00054897 amu (atomic mass unit) or 9×10^{-31} kg.
- **Charge**: -1 (negative)
- **Location**: Outside the nucleus of the atom.
- **Role**: Electrons balance the positive charge of protons in an atom, making the atom electrically neutral. Since the number of electrons equals the number of protons in a neutral atom, their charges cancel out.

Protons

- **Description**: Protons are positively charged particles found in the nucleus of an atom. They have a mass of 1.0073 amu or 1.67×10^{-27} kg.
- **Charge**: +1 (positive)
- **Location**: Inside the nucleus.
- **Role**: Protons contribute significantly to the mass of an atom and determine the element's identity (atomic number).

Neutrons

- **Description**: Neutrons are neutral particles (no charge) also found in the nucleus. They have a mass slightly greater than protons, about 1.0087 amu or 1.67×10^{-27} kg.
- **Charge**: 0 (no charge)
- **Location**: Inside the nucleus.
- **Role**: Neutrons contribute to the atom's mass and help stabilize the nucleus. They vary in number within isotopes of the same element.

Table 3.1 Physical Properties of Subatomic Particles

Particle	Symbol	Charge (coulombs)	Mass (kg)	Relative Charge	Mass (amu)	Location
Proton	P^{+}	$+1.59 \times 10^{-19}$	1.673×10^{-27} kg	+1	1.0073	Inside nucleus
Neutron	n^{0}	0	1.675×10^{-27} kg	0	1.0087	Inside nucleus
Electron	e^{-}	-1.59×10^{-19}	9×10^{-31} kg	-1	0.00054897	Outside nucleus

3.5.2 Atomic Number and Mass Number

Atomic Number (Z)

- **Definition**: The number of protons in an atom's nucleus. It uniquely identifies an element.
- **Example**: An atom with 6 protons is carbon. Therefore, the atomic number of carbon is 6.
- **Significance**: The atomic number also tells us the number of electrons in a neutral atom.

Mass Number (A)

- **Definition**: The total number of protons and neutrons in an atom's nucleus.
- **Formula**: Mass Number = Number of Protons + Number of Neutrons
- **Example**: For helium with 2 protons and 2 neutrons, the mass number is 2+2=4.

Nuclear Symbols

- **Format**: The mass number (A) is written as a superscript and the atomic number (Z) as a subscript to the left of the element symbol.
- **Example**: For carbon-12: ${}^{12}_{6}\mathrm{C}$

3.5.3 Atomic Mass and Isotopes

Isotopes

- **Definition**: Atoms of the same element with different numbers of neutrons.
- **Example**: Carbon has isotopes such as carbon-12 (6 protons, 6 neutrons) and carbon-13 (6 protons, 7 neutrons).

- **Properties**: Isotopes of the same element have similar chemical properties but different masses.

Stability of Isotopes

- **Stable Isotopes**: Have a balanced number of protons and neutrons. Example: Carbon-12 and Carbon-13 are stable.
- **Radioactive Isotopes**: Have an imbalance in neutron-to-proton ratio and decay over time. Example: Carbon-14 is radioactive and used in dating fossils.

Atomic Mass

- **Definition**: The weighted average mass of all isotopes of an element, based on their natural abundance.
- **Calculation**: Average Atomic Mass$\frac{=(\%\ Isotope\ 1\times\ Mass\ of\ Isotope\ 1)+(\%\ Isotope\ 2\times\ Mass\ of\ Isotope\ 2)+\dots}{100}$

Examples

1. **Lithium**:
   - **Isotope with 3 Neutrons**: Atomic Number = 3, Mass Number = 6.
   - **Isotope with 4 Neutrons**: Atomic Number = 3, Mass Number = 7.
2. **Average Atomic Mass Calculation**:
   - **Lithium**: Li-6 (7% abundance), Li-7 (93% abundance).

     Average Atomic Mass = (7×0.07)+(6×0.93)=6.93 amu.

   - **Boron**: B-10 (20% abundance), B-11 (80% abundance).

     Average Atomic Mass = (10×0.20)+(11×0.80)=10.8 amu.

   - **Neon**: Ne-20 (90.92% abundance), Ne-21 (0.30% abundance), Ne-22 (8.85% abundance).

     Average Atomic Mass = (19.99×0.9092)+(20.99×0.003)+(21.99×0.0885)=20.17 amu.

3.5.4 Main Energy Levels

Understanding Main Energy Levels:

The concept of energy levels, also known as principal energy levels or shells, describes how electrons are arranged around the nucleus of an atom. This idea was first proposed by Niels Bohr in 1913. Each energy level represents a specific distance from the nucleus where electrons are likely to be found.

- **Principal Quantum Number (n):** Each energy level is denoted by a principal quantum number nnn, which is a positive integer (1, 2, 3, 4, ...). These numbers can also be represented by letters (K, L, M, N, etc.).
- **Energy of Levels:** The energy associated with these levels increases as you move farther from the nucleus. Thus, the order of increasing energy is: K < L < M < N or 1 < 2 < 3 < 4.

Activity 3.15

1. **Discovery of Energy Levels:**
   - Niels Bohr discovered the concept of energy levels in 1913. Energy levels are also known as principal energy levels or shells.
2. **Electron Dispersion in Bohr's Model:**
   - In Bohr's atomic model, electrons arrange themselves in distinct orbits or shells around the nucleus. Electrons disperse themselves in these orbits to achieve a stable configuration.
3. **Electron Arrangement According to Bohr:**
   - Electrons are arranged in concentric orbits around the nucleus. Each orbit can hold a specific number of electrons, following the rule that electrons fill the lowest energy levels first before occupying higher ones.

3.5.5 Electronic Configuration on Main Shells

Electronic Configuration:

Electronic configuration describes how electrons are distributed among the various energy levels or shells in an atom. The distribution follows specific rules:

- **Formula for Maximum Electrons per Shell:** The maximum number of electrons that each shell can hold is given by $2n^2$, where n is the principal quantum number.
  - K shell (n=1): 2×1^2= electrons
  - L shell (n=2): $2\times2^2=8$ electrons
  - M shell (n=3): $2\times3^2=18$ electrons
  - N shell (n=4): $2\times4^2=32$ electrons

Activity

1. **Seating Arrangement Analogy:**
   - Think of the seating arrangement where the number of seats in each row increases. This is similar to how electrons fill energy levels, starting from the closest (K shell) and moving to higher levels (L, M, N, etc.).
2. **Arranging Electrons:**
   - For an atom with 20 electrons, start filling the K shell with 2 electrons, L shell with 8 electrons, and then M shell with the remaining 10 electrons. Thus, the configuration is 2, 8, 10.

Example :

- **Calcium (Atomic Number 20):** The electronic configuration is 2, 8, 8, 2.

Example :

- **Argon (Atomic Number 18):** The electronic configuration is 2, 8, 8.

3.5.6 Valence Electrons

Valence Electrons:

Valence electrons are the electrons located in the outermost shell of an atom. They are the electrons most involved in chemical reactions and determine the chemical properties of the element.

Example 3.8:

- **Silicon (Atomic Number 14):** Electronic configuration is 2, 8, 4. The valence electrons are in the M shell.

Example 3.9:

- **Argon (Atomic Number 18):** Electronic configuration is 2, 8, 8. The valence shell is the third shell with 8 electrons.

Summary

- **Energy Levels:** The arrangement of electrons in an atom is based on discrete energy levels, denoted by principal quantum numbers (n) or letters (K, L, M, N).
- **Electronic Configuration:** Describes the distribution of electrons among these energy levels, starting from the lowest energy level and moving outward.
- **Valence Electrons:** The electrons in the outermost shell that determine an element's chemical properties and reactivity.