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Science

SIERRA INNOVATIONS

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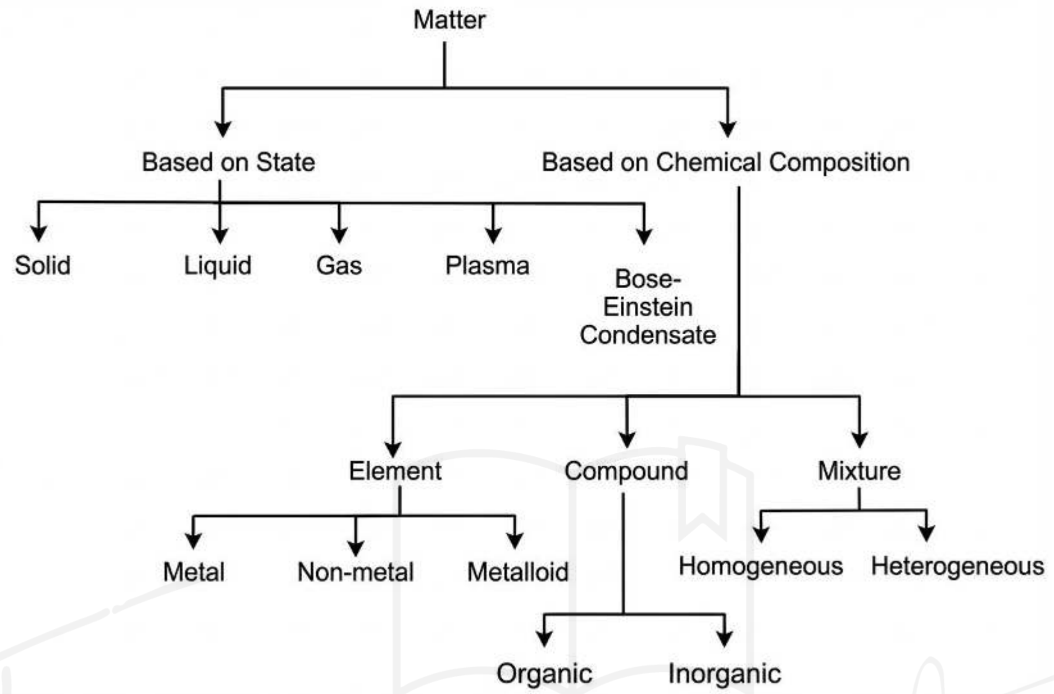
1 CHAPTER

Matter and Nature of Matter

Matter: A substance that has mass, occupies space (volume), possesses physical resistance and inertia, and can be perceived through the sense organs is called matter.



Types of Matter Based on Physical State:

1. **Solid:** The state of matter that has a definite shape, volume, and mass is called a solid. The particles are held together by strong intermolecular forces, due to which they are very closely packed. Solids have high density and generally high melting points. **Examples:** Ice, wood, and metals.
   a. **Crystalline Solids**: These have a definite shape and a sharp melting point. They are anisotropic in nature. Example: Quartz.
   b. **Amorphous Solids:** These have an irregular shape and no definite melting point. They are isotropic in nature.
2. **Liquid:** The state of matter that has no definite shape but has a definite volume is called a liquid. The intermolecular forces between their particles are weaker than those in solids. Therefore, they can flow and take the shape of the container. **Examples:** Water, oil, and mercury.
3. **Gas:** The state of matter that has neither definite shape nor definite volume is called a gas. The intermolecular forces between their particles are extremely weak, and their kinetic energy is very high. They can be compressed easily. **Examples:** Oxygen, nitrogen, and carbon dioxide.
4. **Plasma:** The state of matter in which particles are highly energetic and exist in an excited state in the form of ionized gas is called plasma. **Examples:** The Sun, stars, neon bulbs, and fluorescent tubes.
5. **Bose–Einstein Condensate (BEC):** This state is formed at extremely low temperatures. In this state, particles combine and behave as a single quantum entity. It is a very low-energy state of matter. It is used in superconductivity and quantum simulation research. This phenomenon was predicted in 1924 by Albert Einstein based on the quantum theories of the Indian physicist Satyendra Nath Bose.

Gas Laws

1. **Boyle's Law (Pressure–Volume Relationship):** At constant temperature, the pressure of a gas is inversely proportional to its volume. **$P_1V_1 = P_2V_2$**
   Example: If a balloon is squeezed, its volume decreases and the pressure inside increases.
2. **Charles's Law (Temperature–Volume Relationship):_** At constant pressure, the volume of a gas is directly proportional to its absolute temperature. **$T_1/V_1 = T_2/V_2$**
   Example: When a balloon is heated, the gas inside expands and the balloon inflates.
3. **Gay-Lussac's Law (Pressure–Temperature Relationship):** At constant volume, the pressure of a fixed mass of gas is directly proportional to its absolute temperature.
   $P \propto T$
4. **Avogadro's Law (Volume–Mole Relationship):_** At the same temperature and pressure, equal volumes of gases contain an equal number of molecules.
   $V \propto n$ (where *n* = number of moles of the gas)
   Example*:* If the amount of gas in a balloon is doubled while keeping temperature and pressure constant, its volume will also double.
5. **Ideal Gas Law:** This is the combined form of Boyle's, Charles's, and Avogadro's laws. It expresses the relationship among all the variables simultaneously.
   $PV = nRT$
   Where:
   P = pressure (atm, Pa)
   V = volume (L, m^3)
   n = number of moles of the gas
   R = universal gas constant (8.314 $J \cdot mol^{-1} \cdot K^{-1}$)
   T = temperature (K)
   - ✓ One mole of an ideal gas occupies a volume of 22.4 liters at 273 K and 1 atmospheric pressure (atm). Boltzmann constant = 1.38×10^{-23} $J \cdot K^{-1}$.
6. **Graham's Law of Diffusion/Effusion:** Graham's law states that the rate of diffusion or effusion of a gas is inversely proportional to the square root of its molar mass.
   $r_1/r_2 = \sqrt{(M_2/M_1)}$
   Where: r_1, r_2 = rates of diffusion/effusion of gas 1 and gas 2
   M_1, M_2 = molar masses of gas 1 and gas 2
   - ✓ Lighter gases diffuse faster than heavier gases.
7. **Dalton's Law of Partial Pressures:** Dalton's law states that the total pressure exerted by a mixture of non-reacting gases is equal to the sum of the partial pressures of the individual gases.
   $P_{total} = P_1 + P_2 + P_3 + \cdots + P_n$
   Where:
   P_{total} = total pressure of the gas mixture
   P_1, P_2, P_3, …, P_n = partial pressures of the gases present in the mixture.

Types of Matter Based on Chemical Composition:

1. **Element:** An element is a substance made up of only one type of atom. It cannot be broken down into simpler substances or components by chemical means. **Examples:** Iron, chlorine, gold, silver, and sulfur.
2. **Compound:** A compound is a substance formed by the chemical combination of two or more elements in a fixed proportion. **Example:** Water (H_2O) → Formed from hydrogen and oxygen, but its physical and chemical properties are different from those of its constituent elements.
3. **Mixture:** A mixture is a substance formed by mixing two or more elements and/or compounds in any proportion. Its components can be separated by physical or mechanical methods.
   a. **Homogeneous Mixture:** A mixture in which the composition and properties are uniform throughout. **Examples:** Salt–water solution, sugar–water solution.
   b. **Heterogeneous Mixture:** A mixture in which the composition and properties are not uniform throughout. **Example:** Suspension of chalk or sand in water.

Methods of Separation of Mixtures:

1. **Filtration:** In this method, the mixture is passed through a filter to separate the solid component. It is not used for separating two liquids.
   - ✓ **Principle:** Solid particles cannot pass through the filter paper, whereas the liquid passes through.
   - ✓ **Example:** Separating sand from water.
2. **Evaporation:** The process by which a liquid changes into vapor. This method is used to separate a volatile component.
   - ✓ **Principle:** The volatile component evaporates, leaving the solid component behind.
   - ✓ **Example:** Obtaining salt from salt water.
3. **Centrifugation:** Components of different densities are separated by rapid spinning.
   - ✓ **Principle:** Heavier particles move outward, while lighter particles remain toward the center.
   - ✓ **Example:** Separating cream from milk.
4. **Sublimation:** Some solid substances directly convert into gas on heating without passing through the liquid state.
   - ✓ **Principle:** One component vaporizes and escapes, while the other remains behind.
   - ✓ **Example:** Separating iodine from a mixture.

   Note: Sublimation can be used to separate a mixture of ammonium chloride and common salt.
5. **Chromatography:** The word "chroma" means color. This method is based on the different adsorption capacities of the components of a mixture. It was discovered by Tswett.
   - ✓ **Principle:** Different substances move at different speeds over a medium.
   - ✓ **Example:** Separating colors in ink.

   Note: Column chromatography can also be used to separate impurities in milk.
6. **Simple Distillation:** If two liquids are miscible and have a sufficient difference in their boiling points, they can be separated by this method.
   - ✓ **Principle:** The liquid with the lower boiling point vaporizes first and then condenses.
   - ✓ **Example:** Obtaining pure water from a mixture of alcohol and water.

Solution:

- ➢ A homogeneous mixture of two or more substances is called a solution. **Examples:** Sharbat, lemonade, salt–water solution.
- ➢ **Solvent:** The component that dissolves another substance in itself. It is present in a larger amount.
- ➢ **Solute:** The component that is dissolved in the solvent. It is present in a smaller amount.
- ➢ **Example:** In a sugar–water solution, sugar is the solute and water is the solvent.

Classification Based on the Physical States of Solvent and Solute

Type of Solution	Physical State of Solvent	Physical State of Solute	Example
Solid Solution	Solid	Solid	Alloys (e.g., Brass)
Liquid Solution	Liquid	Gas	Carbonated beverages
Liquid Solution	Liquid	Liquid	Alcohol in water
Liquid Solution	Liquid	Solid	Sugar in water; Tincture of iodine
Gaseous Solution	Gas	Gas	Air (Oxygen in Nitrogen)

Suspension:

- ➢ A suspension is a heterogeneous mixture in which the solute particles are dispersed throughout the medium. The size of the dispersed particles is approximately 10^{-8} to 10^{-4} cm or larger.
- ➢ The particles of a suspension can be seen with the naked eye due to their relatively large size (more than 10 cm).
- ➢ **Example:** Milk of magnesia is a suspension of magnesium hydroxide, $Mg(OH)_2$, in water.

Colloidal Solution:

- ➢ A colloidal solution is a mixture that is actually heterogeneous but appears homogeneous because the particles are uniformly distributed throughout the medium.

- **Examples:** Milk, blood, ink.
- The color of colloidal solutions is due to the scattering of light. The color depends on the wavelength of light scattered by the dispersed particles.
  - ✓ The path of a beam of light becomes visible in a colloidal solution due to the relatively larger size of the dispersed particles (Tyndall effect).
- According to the Hardy–Schulze rule, the amount of electrolyte required to coagulate a fixed amount of a colloidal solution depends upon the valency of the coagulating ion.

Classification of Colloids

Dispersed Phase	Dispersion Medium	Type of Colloid	Example
Gas	Liquid	Aerosol	Fog, Mist
Gas	Solid	Aerosol	Smoke
Liquid	Gas	Foam	Whipped cream
Liquid	Liquid	Emulsion	Milk
Liquid	Solid	Sol	Paint, Ink
Solid	Gas	Solid Foam	Pumice stone, Sponge
Solid	Liquid	Gel	Jelly
Solid	Solid	Solid Sol	Alloys, Gemstones

Tyndall Effect:

- When light is passed through a colloidal sol, the colloidal particles scatter the light, making the path of the light beam visible. This phenomenon is called the Tyndall effect.
- **Examples:** The blue color of water and the sky, fog, and smoke.

Brownian Motion:

- This phenomenon was first observed in 1827 by the botanist Robert Brown, when he noticed the random motion of pollen grains in water under a microscope.
- Brownian motion refers to the irregular, zigzag movement shown by microscopic particles suspended in a liquid or gas.

Emulsion:

- An emulsion is a liquid colloidal system in which fine droplets of one liquid are dispersed in another liquid. When two immiscible or partially miscible liquids are mixed and shaken together, an emulsion is formed. This process is called emulsification.
- **Examples:** Milk, butter, cream, egg yolk, paint, cough syrup, face cream, insecticides, etc.

Do You Know?

- Thomas Andrews was an Irish chemist and physicist, known for his pioneering work on the behavior of gases.

2 CHAPTER

Atoms and Atomic Structure

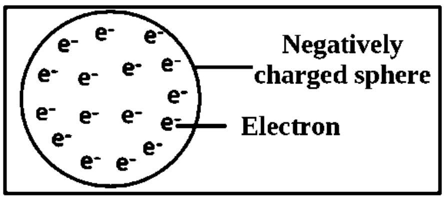
Atom: The term "atom" is derived from the Greek word *a-tomio*, which means indivisible. It is the smallest particle of matter and participates in chemical reactions.

Dalton's Atomic Theory:

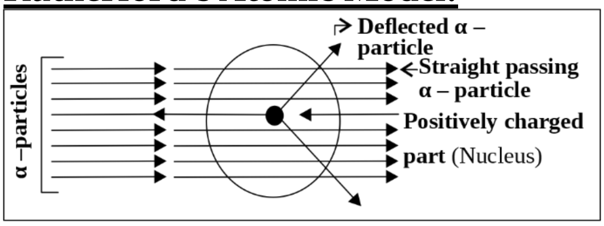
- In 1808, John Dalton proposed the atomic theory of matter, known as Dalton's Atomic Theory. He is regarded as the father of modern atomic theory.
- Dalton provided a physical explanation of how compounds are formed by the combination of two or more different types of atoms.
- This theory is based on the laws of chemical combination. Matter is composed of indivisible, smallest particles called atoms. Atoms of the same element have identical masses and chemical properties, whereas atoms of different elements have different masses and chemical properties.

Thomson's Atomic Model:

- This model was proposed by J. J. Thomson in 1898. According to him, "An atom is like a solid sphere in which positive charge is uniformly distributed, and negatively charged electrons are embedded within it like seeds in a watermelon. The mass of the atom is evenly distributed throughout the sphere."
- This model is also known as the **Plum Pudding Model** or **Watermelon Model**.
- **Limitation:** It could not explain the internal structure and stability of the atom.

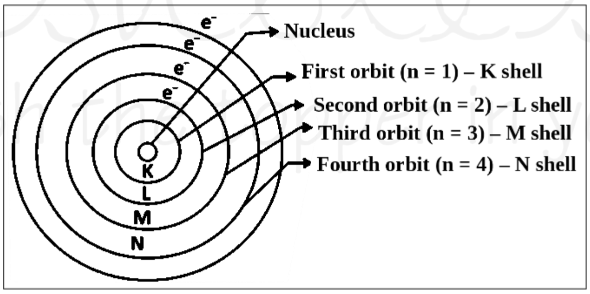


Rutherford's Atomic Model:



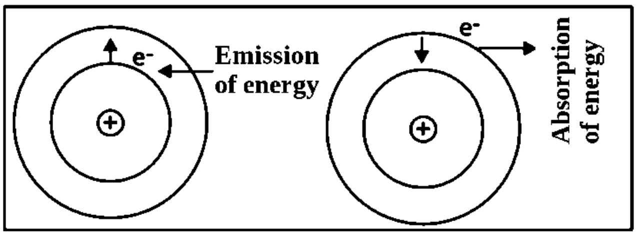
- This model was proposed by Ernest Rutherford in 1911. He conducted the gold foil experiment, in which alpha particles were bombarded onto a thin sheet of gold. Based on the observations and conclusions of this experiment, he proposed the atomic model described below:
  - ✓ At the center of the atom, there is a small, dense, and positively charged nucleus. Electrons revolve around the nucleus in circular orbits.
  - ✓ The entire mass of the atom is concentrated in the nucleus. The nucleus is composed of protons and neutrons. Electrons are held around the nucleus by electrostatic forces of attraction.
- **Limitation:** This model could not explain the stability of the atom or the discrete energy levels of electrons.

Bohr's Atomic Model:

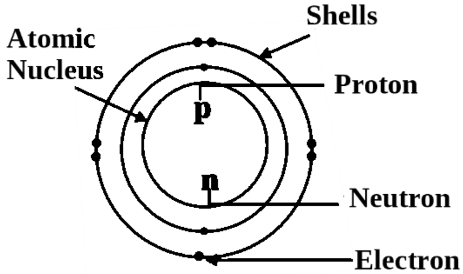


- In 1912–13, Niels Bohr proposed a new atomic model based on quantum theory, which overcame the limitations of Rutherford's model. The main points of the hydrogen atomic model are as follows:
  - ✓ In a hydrogen atom, electrons revolve around the nucleus in circular orbits with fixed radius and fixed energy. These orbits are called shells and are represented by K, L, M, N, O, etc. While revolving in a particular orbit, the energy of the electron does not change.

- ✓ However, when an electron moves from a higher to a lower orbit or from a lower to a higher orbit, energy is emitted or absorbed.
- ✓ The angular momentum (mvr) of an electron in an orbit is equal to **h/2π** or an integral multiple of it, where **h** is Planck's constant.



- ➢ Through this model, we can understand how electrons are distributed in the shells around the nucleus:
- ✓ The maximum number of electrons in a shell is given by **$2n^2$**, where *n* is the principal quantum number (shell number).
- ✓ The outermost shell can contain a maximum of 8 electrons.
- ✓ Electrons first occupy the subshell with the lowest energy level. An electron does not enter the next shell until the lower-energy shell is completely filled.



Subatomic Particles of the Atom:

Name of Particle	Discovery / Discoverer	Charge and Charge Value	Mass (kg)	Position and Role / Description
Electron	J. J. Thomson (1897), Cathode Ray Experiment	Negative (–1), –1.602 × 10^{-19} coulomb	9.11 × 10^{-31}	Revolves around the nucleus in orbitals; participates in chemical bonding and electrical conductivity
Proton	Ernest Rutherford (1917), Gold Foil Experiment	Positive (+1), +1.602 × 10^{-19} coulomb	1.67 × 10^{-27}	Located inside the nucleus; determines atomic number and identity of the element
Neutron	James Chadwick (1932)	Neutral (0), 0 coulomb	1.67 × 10^{-27}	Present inside the nucleus; maintains nuclear stability and reduces repulsion between protons
Positron	Carl Anderson (1932)	Positive (+1), equal in magnitude to electron	Equal to electron	Antiparticle of electron; does not revolve around the nucleus like an electron
Neutrino	Predicted by Pauli (1930), Confirmed by Clyde Cowan & Frederick Reines (1956)	Neutral (0), 0	Extremely small	Emitted during radioactive decay; can pass through matter easily
Graviton	Hypothetical particle	Neutral (0), 0	—	Considered the carrier of gravitational force according to quantum physics

Properties of the Atom:

1. **Atomic Number (Z):** The number of protons present in the nucleus of an atom is called the atomic number. It is denoted by **Z. Atomic Number = Number of protons in nucleus = Number of electrons in a neutral atom**
2. **Atomic Mass:** The mass of a single atom of an element measured relative to 1/12th of the mass of a carbon-12 atom is called atomic mass. It is expressed in **amu (atomic mass unit).**
3. **Mass Number (A):** The total number of protons and neutrons present in the nucleus of an atom is called the mass number. **Mass Number = Number of protons + Number of neutrons**

Mass number

Element

Atomic number

Types of Atomic Species

1. **Isotopes:** Atoms of the same element that have the **same atomic number** but **different mass numbers** are called isotopes. They have identical chemical properties but different physical properties. **Example:** Hydrogen isotopes – Protium (^1H), Deuterium (^2H), Tritium (^3H)

2. **Isobars:** Atoms of **different elements** that have the **same mass number** but **different atomic numbers** are called isobars. They have similar physical properties but different chemical properties. **Example:** Carbon-14 (C-14) and Nitrogen-14 (N-14)

3. **Isotones:** Atoms that have the **same number of neutrons** but **different atomic numbers and mass numbers** are called isotones. **Example: C-14 (6p + 8n) and O-16 (8p + 8n)** all the atmons have to electrons.

4. **Isoelectronic:** Atoms, ions, or molecules that have the **same number of electrons** but belong to **different elements or ions** are called isoelectronic species. **Example:** O^{2-}, F^-, Ne, Na^+ → all have 10 electrons

Molecule:

- A molecule is the **smallest unit of a substance** that retains its chemical properties.
- It consists of **two or more atoms chemically bonded** together by chemical bonds.
  Mole Concept
  Example: Water (H_2O) – Two hydrogen atoms bonded to one oxygen atom
- **Molecular Mass:** The sum of the atomic masses of all atoms in a molecule. Expressed in **amu (atomic mass unit).**
  Molecular Mass = ∑ (Atomic mass of each atom × Number of atoms)
  Example: Water (H_2O)
  - ✓ Atomic mass of hydrogen = 1 amu
  - ✓ Atomic mass of oxygen = 16 amu
  - ✓ Molecular mass = (2 × 1) + 16 = 18 amu

Mole Concept:

- In 1896, Wilhelm Ostwald introduced the term *mole,* which means heap or collection. One mole is defined as the amount of a substance whose mass in grams is equal to its atomic or molecular mass. The mole is the SI unit for measuring the amount of substance.
  Number of particles = (Given mass / Molar mass) × NA
- Where: NA = Avogadro's number = 6.022 × 10^{23} (the number of atoms present in exactly 12 grams of carbon-12).

Electronic Configuration:

1. **Orbital:** An orbital is the three-dimensional region around the nucleus of an atom where the probability of finding a moving electron is maximum. An orbital can hold a maximum of 2 electrons.
2. **Shell:** A shell is the circular path around the nucleus in which electrons revolve. These are represented by K, L, M, N, etc., or by n = 1, 2, 3, 4, etc.
3. **Subshell:** A subshell is a specific region within a shell where electrons move. It is represented by s, p, d, and f.

Subshell	Shape	Number of Orbitals	Maximum Electrons
s	Spherical, symmetrical	1	2
p	Dumbbell-shaped	3	6
d	Double dumbbell (clover-shaped)	5	10
f	Complex shape	7	14

Shell (n)	Subshells Present	Number of Orbitals (n^2)	Maximum Electrons ($2n^2$)
K (n=1)	1 (s)	1	2
L (n=2)	2 (s, p)	1 + 3 = 4	8
M (n=3)	3 (s, p, d)	1 + 3 + 5 = 9	18
N (n=4)	4 (s, p, d, f)	1 + 3 + 5 + 7 = 16	32
O (n=5)	4 (s, p, d, f)*	1 + 3 + 5 + 7 = 16	50
P (n=6)	3 (s, p, d)*	1 + 3 + 5 = 9	18
Q (n=7)	2 (s, p)*	1 + 3 = 4	8

- **Electronic Configuration:** Electronic configuration is the arrangement of electrons of an element in subshells (s, p, d, f) in order of increasing energy.

Principles for Filling Electrons in Orbitals:

1. **Aufbau Principle:** Electrons first occupy the orbitals with the lowest energy and then progressively move to higher energy levels.
2. **Increasing order of energy of subshells:** 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p
3. **Pauli Exclusion Principle:** An orbital can contain a maximum of two electrons, and no two electrons in an atom can have the same set of all four quantum numbers. Therefore, if two electrons are present in an orbital, they must have opposite spins.
4. **Hund's Rule of Maximum Multiplicity:** In p, d, and f subshells, electrons occupy the orbitals singly first, with parallel spins, before pairing occurs. This results in minimum energy and maximum stability.

Quantum Numbers:

- Quantum numbers provide information about the position and energy of an electron in an atom.
- There are four types:
  - ✓ **Principal Quantum Number (n):** Indicates the energy level of the electron and its average distance from the nucleus.
  - ✓ **Azimuthal (Angular Momentum) Quantum Number (l):** Describes the shape of the subshell (orbital). The value of *l* ranges from 0 to (n − 1).
  - ✓ **Magnetic Quantum Number (m_l):** Determines the orientation of the orbital in space. Its value depends on *l* and ranges from −l to 0 to +l.
  - ✓ **Spin Quantum Number (m_s):** Indicates the direction of spin of the electron. Its value is ±1/2.

Do You Know?

- The electron is a member of the lepton family of particles. Its mass is approximately **1/1836** of the mass of a proton.
- Carbon (Z = 6) → Electronic configuration: **$1s^2\ 2s^2\ 2p^2$** → 4 electrons in the L shell.
- Oxygen (Z = 8) has **6 valence electrons**, all present in the L shell (n = 2).
- In decreasing order of the number of valence electrons:
  F > O > C > Li (7, 6, 4, 1 respectively).
- Sulfur (S) exists in the **octatomic form (S_8)**.
- Magnesium (Mg): Electronic configuration – **2, 8, 2**; Atomic number – **12**.
- Fluorine (F) has **7 valence electrons**, which participate in chemical bonding.
- Aluminium (Al) has **3 valence electrons**.
- Nitrogen (Z = 7) has the electronic configuration **$1s^2\ 2s^2\ 2p^3$**. In the formation of the N_2 molecule, two nitrogen atoms share three pairs of electrons to form a **triple bond**. The triple bond in N_2 is formed by the overlap of three 2p orbitals from each nitrogen atom. The L shell of nitrogen participates in the formation of the N_2 molecule.
- Copper (Cu): Atomic number **29**; Electronic configuration – **[Ar] $3d^{10}\ 4s^1$**.
- Scandium has **13 isotopes**, but only one stable isotope occurs naturally.
- Cesium (Cs) has many isotopes, but it does not have the maximum number of isotopes among all elements.

3
CHAPTER

Physical and Chemical Changes

Physical Change:

- A physical change occurs when a substance changes its form or state without any change in its chemical composition. This change is generally reversible and depends on temperature and pressure.

Physical Changes of Matter:

1. **Fusion (Melting):** The process in which a solid changes into a liquid is called fusion or melting.
2. **Melting Point:** The temperature at which a solid melts into a liquid at normal atmospheric pressure is called its melting point.
3. **Sublimation:** This process applies to solids that, on heating, directly change from solid to gas without passing through the liquid state, and on cooling, change directly from gas back to solid. Such solids are called sublimable substances.
   Example: Camphor.
4. **Deposition:**The reverse process of sublimation, in which a gas directly changes into a solid without first becoming a liquid.
5. **Vaporization:** In this process, a liquid changes rapidly into gas upon heating. The conversion of a liquid into vapor at a temperature below its boiling point is called evaporation.
6. **Boiling Point:** The temperature at which a liquid starts boiling at normal atmospheric pressure is called its boiling point.
7. **Condensation:** The process in which a gas changes into a liquid.
8. **Freezing:** The process in which a liquid changes into a solid.

Chemical Change:

- A chemical change is a process in which a substance undergoes a change in its **chemical composition to form one or more new substances.**
- This change is generally **irreversible** and **involves the breaking and formation of chemical bonds**.
- **Examples:** Rusting of iron, burning of wood, cooking of food, photosynthesis, formation of a precipitate, fermentation.

Laws of Chemical Combination:

1. **Law of Conservation of Mass:** Proposed in 1789 by Antoine Lavoisier. According to this law, matter can neither be created nor destroyed. In a chemical reaction, **the total mass of the reactants is equal to the total mass of the products.**
2. **Law of Definite Proportions:** Proposed in 1799 by Joseph Proust. According to this law, the elements present in a pure compound are always combined in a fixed ratio by mass. It is also called the Law of Constant Composition.
3. **Law of Multiple Proportions:** Proposed in 1803 by John Dalton. According to this law, if two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other are in the ratio of simple whole numbers.

Chemical Formula:

A chemical formula is the symbolic representation of the composition of a compound. To write the chemical formula of a compound, the symbols of elements, their valencies, and the charges of ions are required.

Chemical Equation:

- A chemical equation is the symbolic representation of a chemical reaction using the symbols of elements or the formulas of compounds.
- **Reactants:** Substances or compounds that take part in a reaction. They are written on the left-hand side of the equation, separated by a plus (+) sign.

- **Products:** Substances or compounds formed after the reaction. They are written on the right-hand side of the equation, separated by a plus (+) sign.
- The arrow points toward the products and indicates the direction of the reaction.
- **Representation of Physical States:** (s) – solid; (l) – liquid; (g) – gas; (aq) – aqueous solution.

Balancing of Chemical Equations:

- It is important to balance chemical equations to satisfy the Law of Conservation of Mass.
- The hit-and-trial method is a common and straightforward technique used to balance chemical equations. The objective is to ensure that the number of atoms of each element is equal on both sides of the equation.
- **Rules:**
  - ✓ The number of atoms of each element in the reactants must be equal to the number of atoms in the products.
  - ✓ The physical states of reactants and products should be indicated in parentheses along with their chemical formulas.
  - ✓ An equation in which the enthalpy change (heat energy) is also written along with the reaction is called a thermochemical equation.
  - ✓ An equation in which reactants and products are represented in the form of atoms or ions (cations/anions) is called an ionic equation.

Chemical Reaction:

- A chemical reaction is a process in which two or more reactant substances combine to form one or more new substances (products).
- The results of a chemical reaction may include:
  - ✓ Change in color
  - ✓ Change in physical state
  - ✓ Emission or absorption of energy

Types of Chemical Reactions:

Type	Description	Examples
Combination Reaction	When two or more reactants combine to form a single product. $A + B \rightarrow AB$	➢ $CaO + H_2O \rightarrow Ca(OH)_2$ – Calcium oxide reacts with water to form slaked lime. ➢ $Mg + O_2 \rightarrow 2MgO$ – Magnesium oxide forms on burning magnesium in oxygen (also an oxidation reaction).
Decomposition Reaction	A chemical reaction in which a single compound breaks down into two or more simpler products. $AB \rightarrow A + B$	$CaCO_3 \rightarrow CaO + CO_2$ – On heating calcium carbonate, calcium oxide and CO_2 gas are formed..
Thermal Decomposition	A decomposition reaction carried out by supplying heat.	➢ $FeSO_4 \rightarrow Fe_2O_3 + SO_2 + SO_3$: – On heating ferrous sulfate, solid Fe_2O_3 forms and SO_2 (sulfur dioxide) and SO_3 (sulfur trioxide) gases are released ➢ $2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$ – Thermal decomposition of lead nitrate produces NO_2 and O_2 gases..
Photochemical Decomposition	A decomposition reaction that occurs in the presence of sunlight.	➢ $2AgCl \rightarrow 2Ag + Cl_2$ (in sunlight) ➢ $2AgBr(s) \rightarrow$ (in sunlight). $2Ag(s) + Br_2(g)$
Electrolytic Decomposition	A decomposition reaction carried out by passing electric current.	➢ $2H_2O \rightarrow 2H_2 + O_2$ ➢ $2NaCl + 2H_2O \rightarrow 2NaOH + Cl_2 + H_2$ (Chlor-alkali process; NaCl is used).

Displacement Reaction	A chemical reaction in which one element or group replaces another element or group from a compound. A + BC → AC + B	➢ $Pb + CuCl_2 \rightarrow PbCl_2 + Cu$ – Lead displaces copper. ➢ $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$ – Hydrogen gas is evolved. ➢ $FeCl_3 + Zn \rightarrow FeCl_2 + ZnCl_2$: – Zn displaces Fe^{3+} ions forming Zn^{2+} ($ZnCl_2$) and Fe^{2+} ions remain; the yellow color of ferric chloride disappears.
Double Displacement Reaction	A reaction occurring between ionic compounds involving exchange of ions. AB + CD → AD + CB	➢ $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4\downarrow + 2NaCl$ – White precipitate ($BaSO_4$) forms. ➢ $Pb(NO_3)_2 + 2KI \rightarrow PbI_2\downarrow + 2KNO_3$ – Yellow precipitate forms. ➢ $3BaCl_2(aq) + Al_2(SO_4)_3(aq) \rightarrow 3BaSO_4(s) + 2AlCl_3(aq)$ – ($BaSO_4$) forms a white insoluble precipitate; ($AlCl_3$) remains in solution
Combustion Reaction	When a substance reacts with oxygen producing heat and light. $C_xH_Y + O_2 \rightarrow CO_2 + H_2O$ + Heat	➢ $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + Energy – Combustion of methane. ➢ $C + O_2 \rightarrow CO_2$ – In sufficient oxygen, hydrocarbons undergo complete combustion to form CO_2 and water.
Neutralization Reaction	Acid and base neutralize each other to form salt and water. Acid + Base → Salt + Water (H_2O)	➢ $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$ The reaction between an acid and a base produces salt and water. ➢ $Zn + H_2SO4 \rightarrow ZnSO_4 + H_2$ Hydrogen gas is released. (Hydrogen gas burns with a "pop" sound when brought near a flame.)
Precipitation Reaction	Occurs when the cation of one aqueous reactant combines with the anion of another to form an insoluble ionic solid.	➢ $AgNO_3 + NaCl \rightarrow AgCl\downarrow + NaNO_3$: $Ag^+ + Cl^- \rightarrow AgCl$, white solid; turns grey on exposure to light). ➢ $Ca(OH)_2(aq)+CO_2(g) \rightarrow CaCO_3(s)+H_2O(l)$: – When carbon dioxide gas is passed through lime water, the lime water turns milky.
Reversible Reaction	A reaction that occurs in both directions (the reactants combine to form products, and the products recombine to form reactants) . = $A \rightleftharpoons B$	➢ $N_2 + 3H_2 \rightleftharpoons 2NH_3$
Irreversible Reaction	A reaction that proceeds only in one direction. A → B	Burning of wood
Exothermic Reaction	A reaction that releases heat. $A + B \rightarrow C + \Delta E$	$CaO + H_2O \rightarrow Ca(OH)_2$ – Heat is released.
Endothermic Reaction	A reaction that absorbs heat. $A + B \rightarrow C - \Delta E$	$CaCO_3 \rightarrow CaO + CO_2$ – Heat is absorbed.
Oxidation Reaction	A chemical reaction in which a substance gains oxygen or any electronegative element, or loses hydrogen or electrons, showing an increase in oxidation number.	➢ $2Cu + O_2 \rightarrow 2CuO$ – Cu is oxidized. ➢ $2Mg+O2 \rightarrow 2MgO$ – Magnesium is oxidized.

Reduction Reaction	A chemical reaction in which a substance gains hydrogen or electropositive element or electrons, or loses oxygen or electronegative element, showing a decrease in oxidation number.	➢ $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$: – CO acts as reducing agent. ➢ $CuSO_4 + Mg \rightarrow MgSO_4 + Cu$:– Mg acts as reducing agent.
Redox Reaction	Oxidation and reduction occur simultaneously; involves transfer of electrons between substances.	➢ $CuO + H_2 \rightarrow Cu + H_2O$ ➢ $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$ – Al is reducing agent, $Fe_2O_3$is oxidizing agent.
Test Reaction	Reactions used to identify chemical substances.	➢ Starch + I_2 → Blue-black color (Iodine test). **Example:** When iodine is added to a paste of rice flour, the paste turns blue-black in color. ➢ $Na_2CO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$ – CO_2 turns lime water milky.
Gas Evolution Reaction	A reaction in which a gas is produced.	➢ $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$: – Hydrogen gas forms. ➢ CH_3COOH (vinegar) + $NaHCO_3$ (baking soda) → $CH_3COONa + CO_2 + H_2O$ – CO_2 inflates a balloon.

Catalyst:

➢ The concept of a catalyst was introduced by **Jöns Jakob Berzelius**. A catalyst is a substance that increases or decreases the rate of a chemical reaction without itself undergoing permanent chemical change.

Factors Affecting the Rate of a Chemical Reaction:

1. **Concentration of Reactants:** At constant temperature, increasing the concentration of reactants increases the rate of reaction.
2. **Temperature:** Increasing the temperature increases the rate of reaction.
3. **Effect of Temperature Rise:** For many reactions, a rise of 10°C approximately doubles the reaction rate.
4. **Surface Area of Reactants:** A greater surface area of reactant particles increases the rate of reaction.

Rate of Chemical Reaction:

➢ The rate of a chemical reaction is defined as the change in concentration of reactants or products per unit time.

➢ **Order of Reaction:** The sum of the powers (exponents) of the concentrations of reactants as expressed in the rate law is called the order of the reaction.

➢ **Zero-Order Reaction:** A reaction in which the rate is independent of the concentration of reactants (i.e., proportional to concentration^0) is called a zero-order reaction.

➢ **First-Order Reaction:** A reaction in which the rate is directly proportional to the first power of the concentration of a reactant is called a first-order reaction.

➢ **Second-Order Reaction:** A reaction in which the rate is proportional to the square of the concentration of a reactant is called a second-order reaction.

➢ **Third-Order Reaction:** A reaction in which the rate depends on the cube (third power) of the concentration of a reactant is called a third-order reaction.

Do You Know?

➢ Crystals of ferrous sulfate ($FeSO_4 \cdot 7H_2O$) lose water when heated, and their color changes from light green to white.

➢ **Chlor-Alkali Process:** Electrolysis of aqueous sodium chloride (brine) produces Cl_2, H_2, and NaOH. In this process, sodium chloride solution is decomposed by electricity to form sodium hydroxide (alkali), chlorine gas, and hydrogen gas.

- When a substance combines with oxygen, it is said to be oxidized. Oxidation is generally an exothermic process.
- If a substance gains oxygen during a reaction, it is said to be oxidized.
- **Sandmeyer Reaction:** An organic reaction in which an aryl diazonium salt is converted into an aryl halide using a copper(I) halide catalyst is called the Sandmeyer reaction.

Examples of Chemical Reactions in Daily Life:

- Burning of coal (formation of carbon dioxide from carbon and oxygen) – Combustion reaction.
- Electrochemical cell – Indirect redox reaction.
- Electrolysis of water – Decomposition reaction.
- Explosion of firecrackers – Combustion reaction.
- Burning of natural gas – An exothermic reaction.
- Storing food in airtight containers helps to slow down oxidation – Oxidation reaction.
- Conversion of ethanol into acetic acid – Oxidation reaction.
- Formation of water from hydrogen and oxygen – Combination reaction.
- Adding lemon juice to milk causes it to curdle, forming solid curds – Precipitation reaction.
- When slaked lime reacts with carbon dioxide present in the atmosphere, calcium carbonate is formed.
- Change in color (tarnishing) of jewelry – Oxidation process.

4 CHAPTER

Chemical Bond

Oxidation Number:

- The oxidation number of an atom is defined as the charge that an atom appears to have when it forms an ionic bond with atoms of different elements.
- **Rules for Determining Oxidation Number:**
- In a free (uncombined) element, the oxidation number is **0**.
- The oxidation number of a monoatomic ion is equal to its charge.
- The oxidation number of **hydrogen (H)** is generally **+1**, but in metal hydrides it is **–1**.
- **The oxidation number of oxygen (O)** is generally **–2**, but in peroxides it is **–1**.
- **Alkali metals (Group 1)** always have an oxidation number of **+1**.
- **Alkaline earth metals (Group 2)** always have an oxidation number of **+2**.
- **Halogens (Cl, Br, I)** generally have an oxidation number of **–1**, but they can show positive oxidation states when combined with oxygen or more electronegative elements.
- **Sum Rule:** In a neutral compound, the sum of the oxidation numbers of all atoms is equal to zero. In a polyatomic ion, the sum of oxidation numbers is equal to the net charge of the ion.

Valency:

- The electrons present in the outermost shell of an atom are called **valence electrons**, and the outermost shell is called the **valence shell**.
- The combining capacity of an element with other elements is called its **valency**.
- The electrons present in the outermost shell (valence shell) of an atom are known as **valence electrons**.

Some Examples:

Element	Atomic Number	Electronic Configuration	Valency
Hydrogen (H)	1	1	1
Oxygen (O)	8	2,6	2
Nitrogen (N)	7	2,5	3
Carbon (C)	6	2,4	4
Sodium (Na)	11	2,8,1	1
Magnesium (Mg)	12	2,8,2	2
Boron (B)	5	2,3	3
Fluorine (F)	9	2,7	1
Calcium (Ca)	20	2,8,8,2	2
Aluminium (Al)	13	2,8,3	3
Lithium (Li)	3	2,1	1
Carbon (C)	6	2,4	4

Ion:

- When an atom loses or gains electrons, it acquires an electric charge. This electrically charged form of an atom is called an **ion**.
- **Cation:** If an atom loses one or more electrons, it acquires a positive charge. In a cation, the number of protons is greater than the number of electrons. **Examples:** Hydrogen (H^+), Calcium (Ca^{2+}), Potassium (K^+), Ammonium (NH_4^+).
- **Anion:** If an atom gains one or more electrons, it acquires a negative charge. In an anion, the number of electrons is greater than the number of protons. **Examples:** Chloride (Cl^-), Bromide (Br^-), Iodide (I^-), Fluoride (F^-), Cyanide ion (CN^-).

Do You Know?

- **The common valency of silver (Ag) is +1, which means it can lose one electron to form a +1 charged cation (Ag^+).**

• Silver (I) – **Argentous** → Ag^+
• Silver (II) – **Argentic** → Ag^{2+}

Valency Theory:

- In 1916, **Kossel and Lewis** proposed the electronic theory of chemical bonding. According to this theory, atoms combine in such a way that they transfer or share their valence electrons with other atoms to attain a stable noble gas configuration of **eight electrons in their outermost shell (octet rule).**
- Atoms achieve stable configuration in the following ways:
  - ✓ **By losing electrons (forming cations)** → **Metals**
  - ✓ **By gaining electrons (forming anions) Non-metals**
  - ✓ **By sharing electrons** → Covalent compounds

Chemical Bonding:

- The force of attraction between two or more atoms that holds them together to form a molecule or compound is called a **chemical bond**.

1. **Ionic or Electrovalent Bond:** An ionic (electrovalent) bond is formed by the transfer of electrons from one atom to another, usually between dissimilar atoms. The atom that loses electrons is called the electron donor, and the atom that gains electrons is called the electron acceptor.
   Example: Sodium Chloride (NaCl): Na (metal) loses one electron to form Na^+, and Cl (non-metal) gains one electron to form Cl^- → Na^+Cl^-.
   - ✓ Compounds formed by this bond are called **ionic or electrovalent compounds**.

Compound	Melting Point
NaCl (Sodium Chloride)	1074 K
$CaCl_2$ (Calcium Chloride)	1045 K
CaO (Calcium Oxide)	2850 K

- ✓ **Characteristics of Ionic (Electrovalent) Compounds:**
  - They are generally crystalline solids with a definite shape and are hard and brittle.
  - They remain solid due to strong electrostatic forces of attraction between oppositely charged ions.
  - They have high melting and boiling points because of strong electrostatic forces between ions.
  - They are soluble in polar solvents such as water but insoluble in non-polar solvents like benzene, acetone, and kerosene.
  - They conduct electricity in the molten state or in aqueous solution because the ions are free to move in these states.

2. **Covalent Bond:**
   - ✓ In 1916, **G. N. Lewis** proposed that atoms share electrons to complete their outermost shell. Later, **Langmuir** refined this concept and introduced the term **covalent bond** in 1919.
   - ✓ **When** electrons are shared between two atoms having similar or nearly equal electronegativities, the bond formed is called a covalent bond.
     Example: Water (H_2O): Oxygen shares electrons with two hydrogen atoms.

Types of Covalent Bonds:

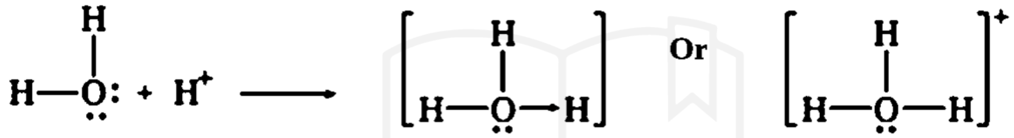
Type	Description	Bond Representation	Examples
Single Bond	One electron pair (2 electrons) is shared between atoms.	A—B	H_2, Cl_2, CH_4
Double Bond	Two electron pairs (4 electrons) are shared between atoms.	A═B	O_2, CO_2, C_2H_4
Triple Bond	Three electron pairs (6 electrons) are shared between atoms.	A≡B	N_2, C_2H_2, CO
Polar Covalent Bond	Unequal sharing of electrons due to electronegativity difference.	A—B (δ^+—δ^-)	H_2O, HCl, NH_3
Nonpolar Covalent Bond	Equal sharing of electrons between identical atoms.	A—A	O_2, N_2, CH_4

✓ **Properties or Characteristics of Covalent Compounds:**

- Covalent compounds generally exist as gases or liquids. Some are soft solids. Diamond, silica (SiO_2), and carborundum (SiC) are notable exceptions.
- They have relatively low melting and boiling points because the intermolecular forces between their molecules are weak.
- They are generally insoluble in water but soluble in non-polar solvents such as benzene and carbon tetrachloride. However, some compounds like sucrose and alcohol are soluble in water due to hydrogen bonding.
- Covalent compounds are generally poor conductors of electricity in both solid and molten states. However, graphite is an exception because it conducts electricity due to the presence of free electrons.
- Covalent bonds are strong and directional. Therefore, covalent molecules have definite shapes and geometries.
- The reactions of covalent compounds are generally complex and slow.

3. **Coordinate or Dative Covalent Bond:** A coordinate (or dative covalent) bond is a special type of covalent bond in which the shared pair of electrons is contributed entirely by one atom.

✓ The arrow always points toward the atom that accepts the electron pair (electron acceptor).



Hydronium ion (H_3O^+)

4. **Hydrogen Bond:** A hydrogen bond is a special type of attractive intermolecular force that arises due to dipole–dipole interaction between a hydrogen atom bonded to a highly electronegative atom and another electronegative atom possessing a lone pair of electrons. The electronegative atom (such as N, O, or F) forms a hydrogen bond with the hydrogen atom.

Examples: Water (H_2O), ammonia (NH_3), hydrogen fluoride (HF).

➢ The strength of hydrogen bonding is determined by the Coulombic interaction between the lone pair of electrons on the electronegative atom of one molecule and the hydrogen atom of another molecule.

Types of Hydrogen Bonds:

Type	Description	Examples	Effects
Intermolecular Hydrogen Bond	Hydrogen bonding between different molecules.	H_2O, NH_3, HF, DNA (A–T, G–C base pairs)	Increases boiling point, melting point, solubility, and surface tension; stabilizes DNA structure.
Intramolecular Hydrogen Bond	Hydrogen bonding within the same molecule.	Ortho-nitrophenol, Proteins (α-helix, β-sheet), Salicylaldehyde	Decreases solubility in water; stabilizes molecular structure; affects physical properties.

Van der Waals Force

The **Van der Waals force** is a weak **intermolecular attractive force** that arises due to temporary or permanent **dipoles** between molecules.

It is not a chemical bond (such as a **covalent** or **ionic bond**), but rather a **physical force**.

Type	Description	Strength	Examples
London Dispersion Forces (Instantaneous Dipole–Induced Dipole Forces)	When unequal movement of electrons in a molecule creates a temporary dipole, which induces a dipole in a nearby molecule.	Weakest	He, Ne, Ar, CH_4, Cl_2, I_2
Dipole–Dipole Attraction (Keesom Forces)	Attraction between polar molecules having permanent dipoles.	Moderate	HCl, SO_2, CH_3Cl
Dipole–Induced Dipole Forces (Debye Forces)	A polar molecule induces a temporary dipole in a nearby non-polar molecule.	Weak	O_2 in H_2O, CCl_4 in HCl

Hybridization:

- Hybridization is the process in which two or more atomic orbitals of nearly equal energy within the same atom mix together to form new hybrid orbitals. These hybrid orbitals have identical energy, shape, and properties, and they participate in bond formation in molecules.

Types of Hybridization:

Hybridization Type	Orbitals Involved	Bond Angle	Molecular Geometry	Examples
sp	1s + 1p	180°	Linear	$BeCl_2$, CO_2, C_2H_2 (Ethyne), acetylene (C_2H_2)
sp^2	1s + 2p	120°	Trigonal Planar	BF_3, C_2H_4 (Ethene), SO_3
sp^3	1s + 3p	109.5°	Tetrahedral	CH_4, NH_3, H_2O
sp^3d	1s + 3p + 1d	90°, 120°	Trigonal Bipyramidal	PCl_5, PF_5
sp^3d^2	1s + 3p + 2d	90°	Octahedral	SF_6, $[Ni(CN)_4]^{2-}$