Fundamentals of Chemistry and Matter

Chapter 1: Atomic Structure – The Building Blocks of Matter

Welcome to the amazing world of atoms! Everything around us, from the air we breathe to the solid ground we stand on, is made up of tiny particles called atoms. Understanding atoms is the first step to understanding chemistry.

1.1 Dalton's Atomic Theory

In the early 19th century, an English scientist named John Dalton proposed one of the first comprehensive theories about atoms. His theory, though later modified, laid the foundation for modern atomic theory.

Postulates of Dalton's Atomic Theory:

1. **Matter is composed of indivisible atoms:** All matter is made of extremely small particles called atoms, which cannot be broken down into simpler substances.
2. **Atoms of the same element are identical:** Atoms of a given element are identical in size, mass, and other properties.
3. **Atoms of different elements differ:** Atoms of different elements differ in size, mass, and other properties.
4. **Atoms combine in simple whole-number ratios:** Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
5. **Atoms are rearranged in chemical reactions:** Atoms are neither created nor destroyed in a chemical reaction; they are merely rearranged.

Limitations of Dalton's Atomic Theory:

- **Atoms are divisible:** We now know that atoms are not indivisible; they are made up of smaller sub-atomic particles (electrons, protons, neutrons).
- **Atoms of the same element are not always identical:** The discovery of isotopes showed that atoms of the same element can have different masses.

1.2 Atomic Structure – Sub-particles

By the end of the 19th century and early 20th century, scientists discovered that the atom was not a solid, indivisible sphere but contained even smaller particles.

- **Discovery of Electrons (J.J. Thomson, 1897):**
    - Using a **cathode ray tube** (a vacuum tube with electrodes), Thomson observed that when a high voltage was applied, a beam of "cathode rays" was produced.
    - These rays were deflected by both electric and magnetic fields, indicating they were negatively charged particles.

    - He called these particles **electrons** and determined their charge-to-mass ratio. This experiment showed that atoms could be divided.
- **Discovery of Protons (Goldstein, 1886 & Rutherford, 1919):**
    - Earlier, Eugen Goldstein observed "canal rays" or "anode rays" in a modified cathode ray tube. These rays traveled in the opposite direction to cathode rays and were positively charged.
    - Later, Ernest Rutherford, building on his gold foil experiment (discussed next), proposed that the positively charged particle in the nucleus was a **proton**. He performed experiments showing that alpha particles could dislodge these positive particles from nitrogen nuclei.
- **Discovery of Neutrons (James Chadwick, 1932):**
    - Rutherford had predicted the existence of a neutral particle in the nucleus to account for the "missing mass" of atoms.
    - Chadwick bombarded beryllium with alpha particles and observed the emission of highly penetrating, uncharged radiation. He identified these as **neutrons**, particles with a mass similar to protons but no electric charge.

Summary of Sub-atomic Particles:

Particle	Symbol	Charge	Relative Mass	Location
Electron	e^-	-1	1/1836 (very small)	Outside nucleus (electron cloud)
Proton	p^+	+1	1	Inside nucleus
Neutron	n^0	0	1	Inside nucleus

1.3 Models of the Atom

As new discoveries were made, scientists developed different models to explain the structure of the atom.

- **Thomson's "Plum Pudding" Model (1904):**
    - After discovering the electron, Thomson proposed that the atom was a sphere of uniformly distributed positive charge, with negatively charged electrons embedded within it, like plums in a pudding or raisins in a cake.
    - This model explained the overall neutrality of the atom but did not account for later experimental results.
- **Rutherford's Atomic Model (Nuclear Model) (1911):**
    - **The Gold Foil Experiment:** Rutherford and his students (Geiger and Marsden) bombarded a thin gold foil with positively charged **alpha particles**.
        - **Observations:**
            - Most alpha particles passed straight through the foil, undeflected.
            - A small fraction were deflected at large angles.
            - A very few (about 1 in 8000) were deflected back towards the source.
        - **Conclusions:**
            - The atom is mostly empty space (explains why most passed through).
            - There is a tiny, dense, positively charged center called the **nucleus** (explains large angle deflections and backscattering). The positive charge repelled the positive alpha particles.
            - Electrons orbit this nucleus.

    - **Limitations of Rutherford's Model:**
        - It couldn't explain why electrons, which are negatively charged and constantly accelerating around the positive nucleus, don't lose energy and spiral into the nucleus, causing the atom to collapse.
        - It couldn't explain the stability of atoms or the discrete atomic emission spectra (why atoms only emit light of specific colors).
- **Bohr's Model of the Atom (Planetary Model) (1913):**
    - Niels Bohr, building on Rutherford's model and Planck's quantum theory, proposed a more refined model.
    - **Postulates:**
        1. Electrons orbit the nucleus in specific, fixed paths called **orbits** or **energy levels** (shells).
        2. Each orbit has a definite, fixed amount of energy. Electrons in a particular orbit do not lose energy.
        3. Electrons can only move from one allowed energy level to another. When an electron moves from a higher energy level to a lower one, it emits energy in the form of light (a photon). When it absorbs energy, it jumps to a higher energy level.
        4. The energy levels are quantized, meaning only specific energy values are allowed.
    - **Limitations of Bohr's Model:**
        - It successfully explained the spectrum of hydrogen but failed for atoms with more than one electron.
        - It didn't account for the splitting of spectral lines in a magnetic field (Zeeman effect).
        - It still treated electrons as particles orbiting the nucleus in a fixed path, which is not entirely accurate according to modern quantum mechanics.

1.4 Key Atomic Concepts

- **Atomic Number (Z):**
    - The atomic number is the **number of protons** in the nucleus of an atom.
    - It is unique for each element and defines the element. For a neutral atom, the atomic number is also equal to the number of electrons.
    - Example: Hydrogen (Z=1), Carbon (Z=6), Oxygen (Z=8).
- **Mass Number (A):**
    - The mass number is the total **number of protons and neutrons** in the nucleus of an atom.
    - Mass Number (A) = Number of Protons (Z) + Number of Neutrons (n^o)
    - Example: An atom with 6 protons and 6 neutrons has a mass number of 12.
- **Isotopes:**
    - Isotopes are atoms of the **same element** (same atomic number, Z) but with **different numbers of neutrons** (and therefore different mass numbers, A).
    - They have the same chemical properties (due to the same number of electrons) but slightly different physical properties (due to different masses).

    - Example:
        - Hydrogen (^1H, Protium): 1 proton, 0 neutrons
        - Deuterium (^2H or D): 1 proton, 1 neutron
        - Tritium (^3H or T): 1 proton, 2 neutrons
    - Carbon-12 (^{12}C) and Carbon-14 (^{14}C) are also isotopes of carbon.
- **Relative Atomic Mass (Ar):**
    - The relative atomic mass of an element is the **average mass of an atom of the element**, taking into account the abundance of its isotopes, compared to 1/12th the mass of a carbon-12 atom.
    - Since most elements exist as a mixture of isotopes, the relative atomic mass is usually not a whole number.
    - Calculated as: Ar = Σ (isotopic mass × % abundance) / 100

1.5 Electronic Structure of the Atom

The arrangement of electrons around the nucleus is crucial as it determines an atom's chemical behavior.

- **Nature of Light and Atomic Emission Spectra:**
    - When atoms are excited (e.g., by heat or electricity), their electrons jump to higher energy levels.
    - These excited electrons are unstable and quickly fall back to lower energy levels.
    - As they fall, they emit the excess energy as light. Since the energy levels are specific and discrete (quantized), only specific wavelengths (colors) of light are emitted, forming a unique **atomic emission spectrum** for each element (like a fingerprint). This phenomenon supports Bohr's model.
- **The Four Quantum Numbers (Simplified):**
    - To describe the exact location and energy of an electron in an atom, a set of four quantum numbers is used. For secondary school, we simplify these concepts:
        1. **Principal Quantum Number (n):** Describes the **main energy level** or shell. n = 1, 2, 3, \ldots (corresponding to K, L, M shells, respectively). Higher 'n' means higher energy and further from the nucleus.
        2. **Azimuthal/Angular Momentum Quantum Number (l):** Describes the **shape of the orbital** and sub-energy level (subshell) within a main energy level. Values are 0, 1, 2, 3, \ldots, (n-1), corresponding to s, p, d, f subshells.
            - l=0: s orbital (spherical)
            - l=1: p orbital (dumbbell-shaped)
        3. **Magnetic Quantum Number (m_l):** Describes the **orientation of the orbital in space**. For a given 'l', m_l can be -l, …., 0, ….., +l. For example, a p subshell (l=1) has 3 orbitals (p_x, p_y, p_z).
        4. **Spin Quantum Number (m_s):** Describes the **intrinsic angular momentum (spin) of an electron**. Electrons can spin in one of two directions, denoted as +1/2 or -1/2 (often represented as up ⬆️ or down ⬇️).
- **Arrangement of Electrons in an Atom (Electronic Configuration):**
    - Electrons fill orbitals according to specific rules:
        1. **Aufbau Principle:** Electrons fill the lowest energy orbitals first. (e.g., 1s before 2s, 2s before 2p).
        2. **Pauli Exclusion Principle:** An atomic orbital can hold a maximum of **two electrons**, and these two electrons must have opposite spins (e.g., ⬆️ ⬇️).

3. **Hund's Rule of Maximum Multiplicity:** When filling degenerate orbitals (orbitals of the same energy, like the three p orbitals), electrons prefer to occupy separate orbitals with parallel spins before pairing up. (e.g., px ⬆️, py ⬆️, pz ⬆️ before px ⬆️ ⬇️).

- **Electronic Configuration Notation:**
  - Shows the distribution of electrons in shells and subshells (e.g., 1s^2 2s^2 2p^6).
  - **Orbital Diagram:** Uses boxes or lines to represent orbitals and arrows to represent electrons.
- **Examples:**
  - **Hydrogen (Z=1):** 1s^1 (⬆️)
  - **Helium (Z=2):** 1s^2 (⬆️ ⬇️)
  - **Carbon (Z=6):** 1s^2 2s^2 2p^2 (⬆️ ⬇️ ⬆️ ⬇️ ⬆️ ⬆️ __)
  - **Oxygen (Z=8):** 1s^2 2s^2 2p^4 (⬆️ ⬇️ ⬆️ ⬇️ ⬆️ ⬇️ ⬆️ ⬆️)
  - **Sodium (Z=11):** 1s^2 2s^2 2p^6 3s^1

1.6 The Periodic Table

The Periodic Table is a systematic arrangement of elements based on their atomic number and electron configuration. It allows us to predict the properties of elements.

- **Structure:** Elements are arranged in **periods** (horizontal rows, indicating the number of electron shells) and **groups** (vertical columns, indicating similar chemical properties due to similar valence electron configurations).
- **Rare Gases (Noble Gases):**
  - These are elements in **Group 18** (or Group 0) of the periodic table (Helium, Neon, Argon, Krypton, Xenon, Radon).
  - They have a **stable electronic configuration** (full outermost electron shells: 2 electrons for Helium, 8 electrons for others).
  - This stable configuration makes them very unreactive, which is why they are called "noble" or "inert" gases.

Chapter 2: Chemical Combinations – How Atoms Interact

Atoms rarely exist alone, except for noble gases. They tend to combine with other atoms to form molecules and compounds. Why? To achieve a more stable electronic configuration, usually a full outermost shell (like the noble gases). This is often referred to as the **octet rule** (except for hydrogen and helium, which aim for a duplet).

2.1 Introduction to Chemical Bonding

Chemical bonds are the forces that hold atoms together in molecules and compounds. They form through the rearrangement of valence (outermost shell) electrons.

2.2 Electrovalent (Ionic) Combination

- **Definition:** Electrovalent (or ionic) bonding involves the **complete transfer of one or more electrons** from one atom to another. This typically occurs between a metal atom (which tends to lose electrons) and a non-metal atom (which tends to gain electrons).

- **Formation:**
  - When a metal atom loses electrons, it forms a positively charged ion called a **cation**.
  - When a non-metal atom gains electrons, it forms a negatively charged ion called an **anion**.
  - The strong electrostatic force of attraction between the oppositely charged ions forms the ionic bond.
- **Example: Formation of Sodium Chloride (NaCl)**
  - Sodium (Na): 1s^2 2s^2 2p^6 3s^1 (1 valence electron)
  - Chlorine (Cl): 1s^2 2s^2 2p^6 3s^2 3p^5 (7 valence electrons)
  - Na loses 1 electron to become Na^+ (1s^2 2s^2 2p^6, stable octet).
  - Cl gains 1 electron to become Cl^- (1s^2 2s^2 2p^6 3s^2 3p^6, stable octet).
  - Na^+ and Cl^- are held together by strong electrostatic forces.
- **Properties of Ionic Compounds:**
  - **High Melting and Boiling Points:** Due to strong electrostatic forces requiring a lot of energy to break.
  - **Hard and Brittle:** Crystalline structure.
  - **Good Electrical Conductors in Molten or Aqueous State:** The ions are free to move and carry charge. They do not conduct electricity in the solid state because ions are fixed in a lattice.
  - **Soluble in Polar Solvents:** Like water.

2.3 Ordinary Covalent Combination

- **Definition:** Covalent bonding involves the **sharing of one or more pairs of electrons** between two atoms. This typically occurs between two non-metal atoms.
- **Formation:** Both atoms contribute electrons to the shared pair(s), allowing both to achieve a stable noble gas configuration (usually an octet or duplet).
- **Types of Covalent Bonds:**
  - **Single Covalent Bond:** One shared pair of electrons (e.g., H-H in H_2).
  - **Double Covalent Bond:** Two shared pairs of electrons (e.g., O=O in O_2).
  - **Triple Covalent Bond:** Three shared pairs of electrons (e.g., N≡N in N_2).
- **Example: Formation of Water (H_2O)**
  - Oxygen needs 2 electrons to complete its octet.
  - Each Hydrogen needs 1 electron to complete its duplet.
  - Oxygen shares one electron with each hydrogen atom, forming two O-H single bonds.
- **Properties of Covalent Compounds:**
  - **Low Melting and Boiling Points:** Due to weak intermolecular forces (forces between molecules), which require less energy to overcome. (The covalent bonds *within* the molecule are strong, but the bonds *between* molecules are weak).
  - **Poor Electrical Conductors:** No free moving ions or electrons to carry charge.
  - **Varying Solubility:** Depends on polarity; polar covalent compounds can dissolve in polar solvents.

2.4 Coordinate Covalent Combination (Dative Bonding)

- **Definition:** A coordinate covalent bond (or dative bond) is a special type of covalent bond where **one atom contributes both of the shared electrons** to the bond. The other atom provides an empty orbital to accept the electron pair.
- **Formation:** An atom with a lone pair of electrons (unshared pair) donates it to an atom that needs two electrons to complete its valency and has an empty orbital.
- **Representation:** Represented by an arrow (→) pointing from the donor atom to the acceptor atom.
- **Example: Formation of Ammonium Ion (NH_4^+)**
    - Ammonia (NH_3) has a lone pair on the nitrogen atom.
    - A hydrogen ion (H^+) has an empty 1s orbital.
    - The nitrogen in NH_3 donates its lone pair to H^+, forming a coordinate bond.

2.5 Other Binding Forces

Besides ionic and covalent bonds (which are *intramolecular* - within a molecule), there are also weaker *intermolecular* forces (between molecules):

- **Metallic Bonding (Brief):** In metals, valence electrons are delocalized and form a "sea of electrons" that holds the positively charged metal ions together. This explains why metals are good conductors of heat and electricity.
- **Van der Waals Forces (Brief):** Weak attractive forces between molecules, arising from temporary fluctuations in electron distribution. They are responsible for the existence of liquids and solids of non-polar substances (e.g., noble gases can be liquefied at very low temperatures).
- **Hydrogen Bonding (Brief):** A special type of strong dipole-dipole intermolecular force that occurs when hydrogen is bonded to a highly electronegative atom (like Oxygen, Nitrogen, or Fluorine). It plays a vital role in properties of water, DNA structure, etc.

2.6 Modifications of Dalton's Atomic Theory

Our understanding of atoms has evolved significantly since Dalton's time. Dalton's theory was revolutionary but incomplete. Here's how it was modified:

- **Atoms are Divisible:** The discovery of electrons, protons, and neutrons proved that atoms are not indivisible but are composed of sub-atomic particles.
- **Atoms of the Same Element Can Have Different Masses:** The discovery of isotopes showed that atoms of the same element can have different numbers of neutrons, leading to different mass numbers. For example, Dalton said all carbon atoms are identical, but we know about Carbon-12 and Carbon-14.
- **Atoms Can Be Converted into Other Atoms (Nuclear Reactions):** While atoms are conserved in ordinary chemical reactions, nuclear reactions (like radioactive decay or nuclear fusion/fission) can change one element into another.

Despite these modifications, Dalton's fundamental idea that elements are made of distinct atoms remains a cornerstone of chemistry.

Chapter 3: Kinetic Theory of Matter – The Dynamic World

Have you ever wondered why ice melts, or why a gas fills up any container it's in? The answers lie in the **Kinetic Theory of Matter**, which states that all matter is made up of tiny particles (atoms or molecules) that are in constant, random motion.

3.1 Introduction to Kinetic Theory

The kinetic theory is a model used to explain the properties of matter in terms of the motion of its particles.

Basic Postulates of Kinetic Theory (especially for gases):

1. Matter is composed of a large number of very small particles (atoms or molecules).
2. These particles are in constant, random motion.
3. There are empty spaces between the particles.
4. Particles collide with each other and with the walls of their container.
5. Collisions are perfectly elastic (no net loss of kinetic energy).
6. The average kinetic energy of the particles is directly proportional to the absolute temperature (in Kelvin).
7. Forces of attraction between particles are negligible (especially for gases).

3.2 States of Matter

The kinetic theory helps explain the distinct properties of solids, liquids, and gases.

Property	Solids	Liquids	Gases
Particle arrangement	Closely packed, orderly lattice	Closely packed, disordered	Far apart, random arrangement
Particle movement	Vibrate about fixed positions	Slide past each other	Move randomly and rapidly
Interparticle forces	Very strong	Stronger than gases, weaker than solids	Very weak, almost negligible
Shape	Fixed shape	No fixed shape (takes shape of container)	No fixed shape (takes shape of container)
Volume	Fixed volume	Fixed volume	No fixed volume (fills container)
Compressibility	Very difficult to compress	Difficult to compress	Easily compressed
Density	High	High (generally slightly less than solids)	Low
Fluidity	Non-fluid	Fluid (can flow)	Fluid (can flow)

3.3 Change of State

Matter can change from one state to another by gaining or losing energy (usually heat). These changes involve changes in the arrangement and movement of particles, not changes in the particles themselves.

- **Melting:** Solid ➡️ Liquid (absorbs energy, e.g., ice melting). Particles gain enough kinetic energy to overcome fixed positions and slide past each other.

- **Freezing:** Liquid ➡️ Solid (releases energy, e.g., water freezing). Particles lose kinetic energy and settle into fixed positions.
- **Boiling/Vaporization:** Liquid ➡️ Gas (absorbs energy, e.g., water boiling). Particles gain enough kinetic energy to completely overcome intermolecular forces and move freely.
- **Condensation:** Gas ➡️ Liquid (releases energy, e.g., steam condensing). Particles lose kinetic energy and forces of attraction become strong enough to bring them closer.
- **Sublimation:** Solid ➡️ Gas directly (absorbs energy, e.g., dry ice). Particles gain enough energy to escape solid state directly to gas.
- **Deposition:** Gas ➡️ Solid directly (releases energy, e.g., frost formation). Particles lose energy and settle directly into solid state.

3.4 Phenomena Supporting Kinetic Theory

Several everyday observations provide evidence for the kinetic theory:

- **Diffusion:** The spreading out of particles from a region of higher concentration to a region of lower concentration, due to their random motion.
    - **Gases:** Smell of perfume spreading across a room.
    - **Liquids:** Ink spreading in water.
    - Diffusion is fastest in gases, slower in liquids, and very slow/negligible in solids, supporting the idea of varying particle motion and spacing.
- **Brownian Motion:** The random, zigzag movement of tiny particles (like pollen grains or smoke particles) suspended in a fluid (liquid or gas), caused by the collision of invisible, fast-moving fluid particles with the suspended particles. This directly demonstrates the constant, random motion of fluid particles.
- **Evaporation:** The process by which liquid turns into gas *below* its boiling point. Faster-moving particles at the surface of the liquid escape into the gas phase. This explains why puddles dry up even when the water isn't boiling.

Chapter 4: Gas Laws – The Behavior of Gases

Gases exhibit unique behaviors due to the large distances between their particles and their constant, random motion. Gas laws describe how gases respond to changes in pressure, volume, and temperature.

4.1 Introduction to Gases

- **Properties of Gases:** Gases have no fixed shape or volume, are highly compressible, expand to fill their container, and have very low densities compared to liquids and solids.
- **Kinetic Theory of Gases (Ideal Gas Assumptions):** The gas laws are based on the concept of an "ideal gas," which is a theoretical gas whose particles:
    - Are in constant, random, straight-line motion.
    - Have negligible volume compared to the volume of the container.
    - Have no intermolecular forces (no attraction or repulsion).
    - Undergo perfectly elastic collisions.
    - The average kinetic energy is proportional to absolute temperature.
    - Real gases behave like ideal gases at high temperatures and low pressures.

4.2 Pressure Exerted by Gases

- **Definition:** Gas pressure is defined as the force exerted by the gas particles per unit area on the walls of their container. It results from the continuous collisions of gas particles with the container walls.
- **Units of Pressure:**
    - Pascal (Pa) or kilopascal (kPa) - SI unit.
    - Atmosphere (atm).
    - Millimeters of Mercury (mmHg) or Torricelli (Torr).
    - Bar.
- **Standard Atmospheric Pressure (STP):** 1 { atm} = 760 { mmHg} = 101325 { Pa} ~ 101.3 { kPa}.

4.3 Boyle's Law

- **Statement:** For a fixed mass of gas at **constant temperature**, the volume of the gas is inversely proportional to its pressure.
- **Mathematical Representation:** P ↔ {1}/{V} PV = {constant} Therefore, for two different states of the same gas: P1V1 = P2V2 Where P1, V1 are initial pressure and volume, and P2, V2 are final pressure and volume.
- **Graphical Representation:** A plot of P vs. V gives a curve (hyperbola). A plot of P vs. 1/V gives a straight line passing through the origin.
- **Explanation (Kinetic Theory):** If you decrease the volume, particles hit the walls more frequently, leading to higher pressure.

4.4 Charles' Law

- **Statement:** For a fixed mass of gas at **constant pressure**, the volume of the gas is directly proportional to its absolute temperature (temperature in Kelvin).
- **Importance of Kelvin Temperature Scale:** All gas law calculations involving temperature must use the Kelvin scale. K = {°C} + 273.15 (often rounded to 273 for calculations).
- **Mathematical Representation:** V ↔ T {V}/{T} = {constant} Therefore: {V1}/{T1} = {V2}/{T2}
- **Graphical Representation:** A plot of V vs. T (Kelvin) gives a straight line passing through the origin.
- **Explanation (Kinetic Theory):** Increasing temperature increases the kinetic energy of particles, causing them to move faster and collide more forcefully and frequently. To maintain constant pressure, the volume must expand.

4.5 General Gas Equation (Combined Gas Law)

- This law combines Boyle's and Charles' Laws.
- **Statement:** For a fixed mass of gas, the ratio of the product of its pressure and volume to its absolute temperature is constant.
- **Mathematical Representation:** {PV}/{T} = {constant} Therefore: {P1V1}/{T1} = {P2V2}/{T2} This equation is very useful for solving problems where all three variables (P, V, T) change.

4.6 Dalton's Law of Partial Pressures

- **Statement:** For a mixture of non-reacting gases, the **total pressure** exerted by the mixture is equal to the sum of the **partial pressures** of the individual gases.
    - **Partial Pressure:** The pressure that each gas would exert if it alone occupied the entire volume of the container at the same temperature.
- **Mathematical Representation:** P_{\text{total}} = P1 + P2 + P3 + ….
- **Application:** Often used when collecting gases over water, where the total pressure collected includes the partial pressure of water vapor. P{gas} = P{total} - P{water vapor} (Vapor pressure of water depends on temperature).

4.7 Ideal Gas Equation

- This equation relates pressure, volume, temperature, and the number of moles of a gas.
- **Mathematical Representation:** PV = nRT Where:
    - P = Pressure
    - V = Volume
    - n = Number of moles of gas
    - R = Ideal Gas Constant (8.314 { J mol}^{-1} { K}^{-1} or 0.0821{ L atm mol}^{-1} { K}^{-1}, depending on units of P and V)
    - T = Absolute temperature (in Kelvin)
- This equation is very powerful as it directly relates the amount of gas (moles) to its macroscopic properties.

4.8 Gay-Lussac's Law of Combining Volumes

- **Statement:** When gases react together, they do so in volumes that bear a simple whole-number ratio to one another, and to the volumes of any gaseous products (if measured at the same temperature and pressure).
- **Example:** 2{H}2({g}) + {O}2({g}) ➡️ 2{H}2{O}({g}) 2 volumes of hydrogen react with 1 volume of oxygen to produce 2 volumes of steam. The ratio of volumes is 2:1:2, a simple whole-number ratio.

4.9 Avogadro's Hypothesis and Moles

- **Avogadro's Hypothesis (or Law):** Equal volumes of all gases, at the same temperature and pressure, contain the same number of molecules (or moles).
    - This means that if you have 1 L of oxygen and 1 L of nitrogen at the same temperature and pressure, they contain the same number of gas molecules.
- **The Mole Concept:**
    - A **mole (mol)** is the SI unit for the amount of substance. It is defined as the amount of substance that contains as many elementary entities (atoms, molecules, ions, etc.) as there are atoms in exactly 12 grams of carbon-12.
    - This number is **Avogadro's Number (N_A)**: 6.022 \times 10^{23} particles/mol.
- **Molar Volume of a Gas:**
    - At Standard Temperature and Pressure (STP: 0 {°C} or 273.15 { K} and 1 { atm}), one mole of any ideal gas occupies a volume of **22.4 { L}**.

    - At Room Temperature and Pressure (RTP: 25 {°C} or 298 { K} and 1 { atm}), one mole of any ideal gas occupies a volume of **24 { L}**.
- **Calculations:**
    - Number of moles (n) = Mass (m) / Molar Mass (M)
    - Number of moles (n) = Volume of gas / Molar Volume (at STP/RTP)
    - Number of moles (n) = Number of particles / Avogadro's Number

4.10 Relative Vapour Density of Gases

- **Definition:** The relative vapour density (RVD) of a gas is the ratio of the density of a gas to the density of hydrogen gas, both measured under the same conditions of temperature and pressure.
- **Mathematical Relationship to Molar Mass:** Due to Avogadro's Law, RVD is also approximately equal to half of the molar mass of the gas. {RVD} = {Density of gas} {{Density of H}2} = {Molar Mass of gas}}/ {2 × {Molar Mass of H (2 for H}_2)} {RVD} = {Molar Mass of gas}/{2} This relationship is useful for determining the molar mass of an unknown gas if its vapour density is known.

4.11 Graham's Law of Diffusion of Gases

- **Statement:** The rate of diffusion (or effusion) of a gas is inversely proportional to the square root of its molar mass (or density), at constant temperature and pressure.
    - **Diffusion:** Spreading of gas particles through another gas.
    - **Effusion:** Escape of gas particles through a tiny hole.
- **Mathematical Representation:**

- Mathematical Representation:

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{D_2}{D_1}}$$

Where:

- Rate = rate of diffusion/effusion
- M = molar mass
- D = density

- **Explanation:** Lighter gas particles move faster and thus diffuse/effuse more quickly than heavier ones.
- **Application:** Used to separate isotopes or gases with different molar masses.

Glossary of Terms

- **Atom:** The smallest unit of an element that retains the chemical properties of that element.
- **Electron:** A negatively charged sub-atomic particle found orbiting the nucleus.
- **Proton:** A positively charged sub-atomic particle found in the nucleus.
- **Neutron:** An uncharged sub-atomic particle found in the nucleus.
- **Nucleus:** The dense, positively charged center of an atom, containing protons and neutrons.
- **Atomic Number (Z):** The number of protons in an atom's nucleus, defining the element.
- **Mass Number (A):** The total number of protons and neutrons in an atom's nucleus.
- **Isotopes:** Atoms of the same element with different numbers of neutrons.
- **Relative Atomic Mass:** The average mass of an element's atoms compared to 1/12th the mass of a carbon-12 atom.
- **Electronic Configuration:** The arrangement of electrons in the energy levels (shells and subshells) of an atom.
- **Valence Electrons:** Electrons in the outermost shell of an atom, involved in chemical bonding.
- **Octet Rule:** Tendency of atoms to gain, lose, or share electrons to achieve 8 valence electrons.
- **Ionic Bond:** Chemical bond formed by the complete transfer of electrons, creating oppositely charged ions.
- **Covalent Bond:** Chemical bond formed by the sharing of electrons between atoms.
- **Coordinate Covalent Bond (Dative Bond):** A covalent bond where one atom donates both shared electrons.
- **Kinetic Theory of Matter:** Model stating that matter consists of tiny particles in constant motion.
- **Diffusion:** Spreading of particles from high to low concentration.
- **Brownian Motion:** Random movement of particles in a fluid caused by collisions with fluid molecules.
- **Evaporation:** Liquid changing to gas below its boiling point.
- **Pressure:** Force per unit area exerted by gas particles on container walls.
- **Boyle's Law:** PV = {constant} (at constant T).
- **Charles' Law:** V/T = {constant} (at constant P, T in Kelvin).
- **General Gas Equation:** P1V1/T1 = P2V2/T2.
- **Dalton's Law of Partial Pressures:** P{total} = P1 + P2 + …..
- **Ideal Gas Equation:** PV = nRT.
- **Avogadro's Hypothesis:** Equal volumes of gases at same T and P contain equal numbers of molecules.
- **Mole:** SI unit for amount of substance (6.022 × 10^{23} particles).
- **Molar Volume:** Volume occupied by one mole of a gas at specified conditions (e.g., 22.4 L at STP).
- **Relative Vapour Density (RVD):** Ratio of gas density to hydrogen density; also half its molar mass.

- **Graham's Law of Diffusion:** Rate of diffusion is inversely proportional to the square root of molar mass.

Practice Questions

(Note: Answers are not provided here, but you would typically include them at the end of a physical ebook.)

Chapter 1 & 2 Questions:

1. State two postulates of Dalton's atomic theory that have since been modified.
2. Describe the key observations and conclusions from Rutherford's gold foil experiment.
3. Explain why Bohr's model of the atom was an improvement over Rutherford's model.
4. Define atomic number and mass number. How many protons, neutrons, and electrons are in a neutral atom of ^{35}\text{Cl}?
5. What are isotopes? Give an example of two isotopes for the same element.
6. Write the full electronic configuration for oxygen (Z=8) and sodium (Z=11).
7. Explain the difference between an ionic bond and a covalent bond, providing an example for each.
8. List three properties of ionic compounds and explain why they have these properties.
9. What is a coordinate covalent bond? Give an example.
10. Why do atoms combine to form chemical bonds?

Chapter 3 & 4 Questions:

1. State three key postulates of the Kinetic Theory of Matter.
2. Using the kinetic theory, explain why gases are easily compressible while liquids are not.
3. Describe the particle arrangement and movement in solids, liquids, and gases.
4. What is diffusion? Give an example for both gases and liquids.
5. State Boyle's Law and write its mathematical formula. If the pressure of a gas is doubled at constant temperature, what happens to its volume?
6. A gas occupies 2.0 L at 27 \text{°C}. What volume will it occupy at 127 \text{°C} if the pressure remains constant? (Remember to convert to Kelvin!)
7. State Dalton's Law of Partial Pressures. A container has 0.5 atm of oxygen and 0.3 atm of nitrogen. What is the total pressure?
8. Calculate the volume occupied by 0.5 mol of methane gas (CH_4) at STP.
9. Which gas will diffuse faster: Helium (He, Molar Mass = 4 g/mol) or Neon (Ne, Molar Mass = 20 g/mol)? Explain your answer using Graham's Law.
10. Explain the meaning of "absolute zero" in relation to Charles' Law.