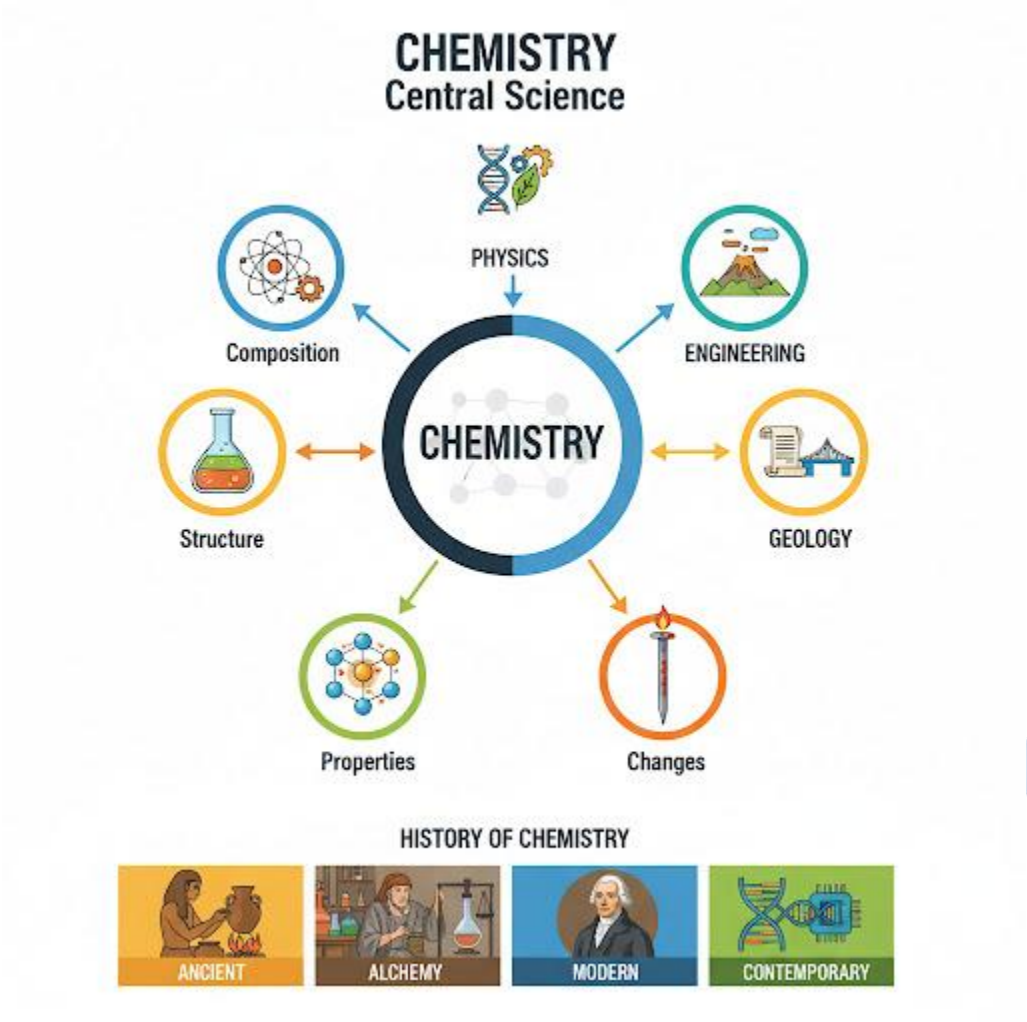
SS1 CHEMISTRY LESSON NOTES

FIRST TERM

1. Introduction to Chemistry – Definition, history, branches, and importance.
2. Chemical Industries and Laboratory Apparatus – Common apparatus and their uses.
3. Nature of Matter – States, properties, changes of state.
4. Separation Techniques – Filtration, distillation, sublimation, chromatography.
5. **Midterm Examination**
6. Atomic Structure – Structure of the atom, subatomic particles.
7. **Midterm Break**
8. Periodic Table – Groups, periods, and periodicity.
9. Chemical Combination – Laws of chemical combination.
10. Acids, Bases, and Salts – Definitions, examples, and properties, pH scale, neutralization.
11. Chemical Symbols, Formulae, and Equations – Writing symbols, formulae, balancing equations.
12. Examination
13. Closing

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WEEK 1 – INTRODUCTION TO CHEMISTRY



Definition of Chemistry

Chemistry is the branch of science that studies the composition, structure, properties, and changes of matter.

- **Composition** refers to the kinds of elements or compounds a substance is made of.
- **Structure** refers to how atoms or molecules are arranged in a substance.
- **Properties** include physical properties (colour, boiling point, density) and chemical properties (reactivity, flammability).
- **Changes** include physical changes (melting, dissolving) and chemical changes (rusting, burning, fermentation).

Key Point: Chemistry is often called the “central science” because it connects physics, biology, geology, medicine, engineering, and many other sciences.

History of Chemistry

The development of chemistry can be divided into four main historical periods:

1. Ancient Chemistry (Practical Beginnings)

- Early humans learned to use fire, cook food, extract oils, make pottery, and use natural dyes.
- Egyptians mastered techniques such as glassmaking, papyrus processing, mummification, and extraction of metals like gold and copper.
- The Chinese developed gunpowder from saltpetre, charcoal, and sulphur.
- These early practices were based on observation and trial-and-error, not scientific theory.

2. Alchemy (Middle Ages, 500 – 1600 AD)

- Originated in Egypt, Arabia, and later spread to Europe.
- Alchemists sought to transform base metals like lead into gold (transmutation) and discover the "elixir of life" for immortality.
- Although their goals were mystical, alchemists discovered many laboratory techniques: distillation, filtration, crystallization, and use of apparatus like alembics.
- Famous alchemists include Jabir ibn Hayyan (Geber), who is considered the father of chemistry in the Arab world.

3. The Age of Modern Chemistry (1600s – 1800s)

- Robert Boyle (1627–1691) emphasized experimentation and introduced Boyle’s Law on gases.

- Antoine Lavoisier (1743–1794), called the "Father of Modern Chemistry," introduced the Law of Conservation of Mass, named oxygen and hydrogen, and helped develop the metric system.
- John Dalton proposed the Atomic Theory (1803), explaining that matter is made of indivisible particles called atoms.

4. Contemporary Chemistry (1900s – Present)

- Development of the Periodic Table by Dmitri Mendeleev (1869) and later refinements.
- Discovery of subatomic particles (electrons, protons, neutrons) and the structure of the atom.
- Advancement in organic chemistry, biochemistry, materials science, and nuclear chemistry.
- Chemistry now impacts industries such as pharmaceuticals, agriculture, electronics, cosmetics, and environmental management.

Branches of Chemistry

Chemistry is vast and can be divided into major and applied branches.

Major Branches

1. **Organic Chemistry** – Study of carbon-containing compounds (except carbonates and simple oxides). Examples: fuels, plastics, pharmaceuticals.
2. **Inorganic Chemistry** – Study of non-carbon compounds such as salts, metals, and minerals.
3. **Physical Chemistry** – Study of how matter behaves on a molecular and atomic level, and how chemical reactions occur, using principles of physics.
4. **Analytical Chemistry** – Concerned with identifying substances (qualitative analysis) and determining their quantities (quantitative analysis).
5. **Biochemistry** – Study of chemical processes within living organisms (e.g., enzymes, DNA, metabolism).

Applied / Specialized Branches

1. **Environmental Chemistry** – Studies chemical processes in the environment and pollution control.
2. **Industrial Chemistry** – Application of chemical knowledge in manufacturing processes.
3. **Pharmaceutical Chemistry** – Design and production of drugs.
4. **Nuclear Chemistry** – Study of radioactive substances and nuclear reactions.
5. **Forensic Chemistry** – Application of chemistry in criminal investigations.

Importance of Chemistry

Chemistry plays a central role in everyday life, industry, and scientific development.

1. Daily Life

- Production of soap, detergents, and cosmetics.
- Cooking processes (baking, boiling, frying).
- Preservation of food using chemical additives.

2. Industry

- Petrochemical industries produce fuels, plastics, synthetic fibres.
- Agricultural chemicals (fertilizers, pesticides) improve crop yield.
- Cement, glass, paint, textile, and paper industries all depend on chemical processes.

3. Medicine and Health

- Discovery and manufacture of drugs, vaccines, antiseptics.
- Diagnostic techniques (blood tests, imaging using contrast chemicals).

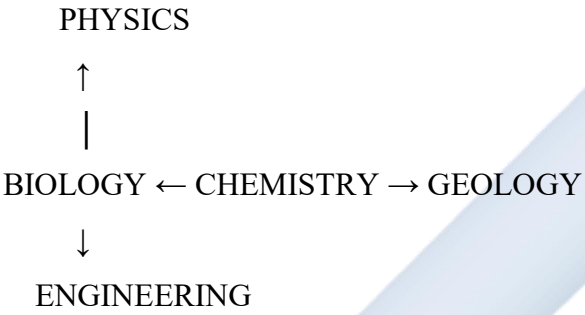
4. Environmental Management

- Water purification, waste management, pollution control.
- Understanding climate change and developing renewable energy technologies.

5. Scientific and Technological Advancement

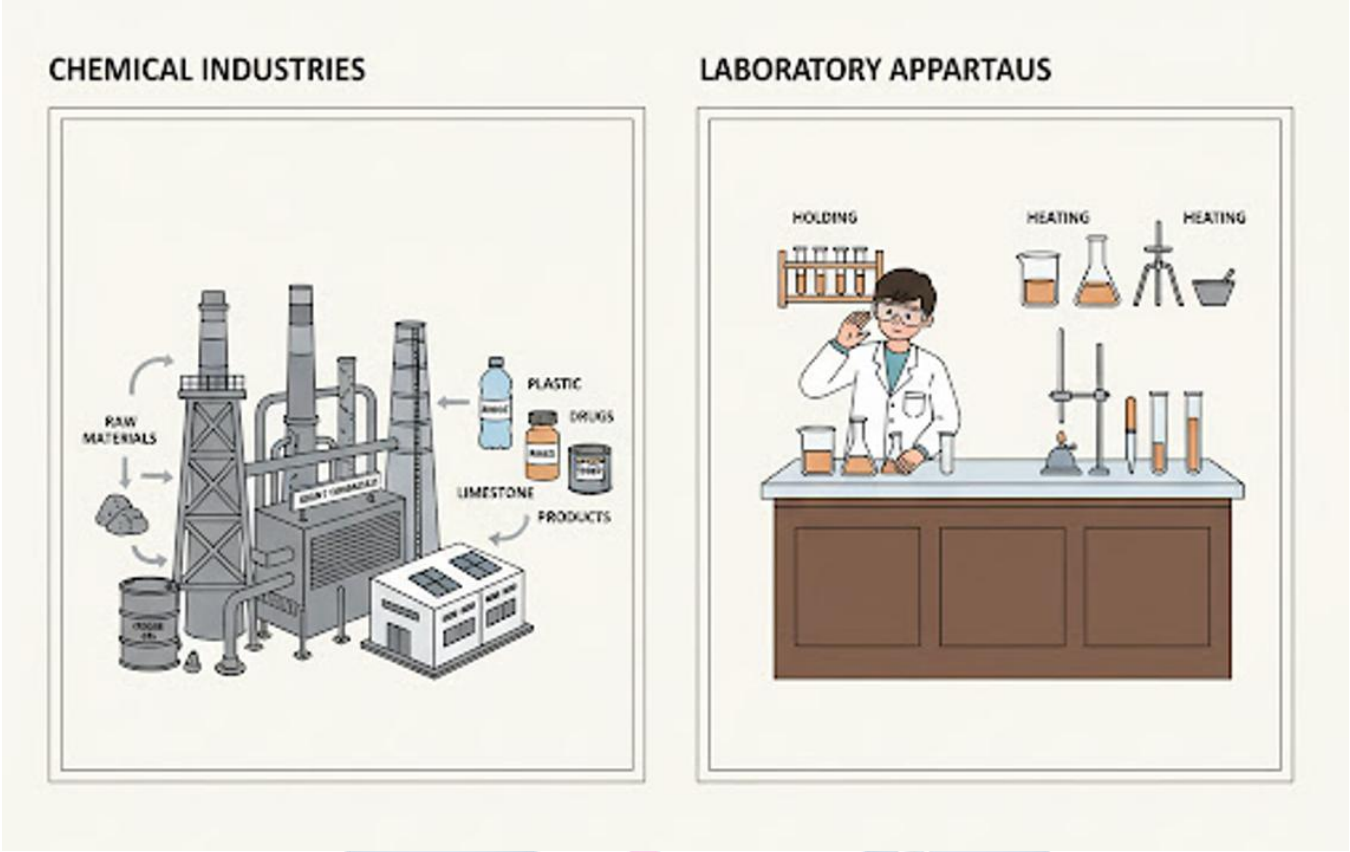
- Development of new materials (nanomaterials, superconductors).
- Space exploration (rocket fuels, life support systems).

A flow chart showing the **relationship of chemistry** with other sciences:



This shows chemistry as the “central science.”

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WEEK 2 – CHEMICAL INDUSTRIES AND LABORATORY APPARATUS

Chemical Industries (Overview)

Meaning of Chemical Industries

Chemical industries are industries that use chemical substances at one stage or another to produce other chemicals or finished goods.

- **Raw materials**: Substances used as starting materials in production.
- **Products**: Finished goods from an industry, which may also serve as raw materials for another industry.

Types of Raw Materials

1. **Natural raw materials** – Obtained directly from nature without major alteration.
   Examples: air, water, coal, limestone, petroleum, iron ore, sulphur deposits.
2. **Synthetic raw materials** – Man-made materials produced for further manufacturing.
   Examples: synthetic fibres, plastics, palm kernel oil for soaps, thread for textiles.

Classification of Chemical Industries

1. **Heavy Chemicals** – Needed in large quantities; purity is not the top priority.
   Examples: tetraoxosulphate (VI) acid (H_2SO_4), sodium hydroxide, cement, palm kernel oil.
2. **Fine Chemicals** – Required in small quantities; produced in very pure form.
   Examples: sodium chloride (NaCl), dyes, preservatives, analytical reagents.

Divisions According to Products

- **Glass industries** – Use sand, clay, sodium carbonate, and colourants to make glass.
- **Ceramics industries** – Use clay and fine chemicals to make ceramics.
- **Pharmaceutical industries** – Use organic, inorganic, and synthetic chemicals to produce drugs.
- **Metallurgical industries** – Extract and process metals from ores.
- **Fertilizer industries** – Produce nitrogen, phosphorus, potassium compounds for agriculture.
- **Paint industries** – Use pigments and inorganic chemicals to make paints.

Importance of Chemical Industries

- Provide raw materials for other industries.
- Generate income and employment.
- Serve as research centres and educational resources.

Disadvantages of Chemical Industries

- Cause environmental pollution.
- Can lead to rural-urban migration.

- Capital intensive to establish.

Laboratory Apparatus – Common Apparatus and Their Uses

Laboratory work in chemistry requires various apparatus for measuring, holding, heating, mixing, and observing substances. These apparatus can be broadly grouped according to their functions.

1. Apparatus for Holding and Storing

Apparatus	Description	Uses
Beaker	Cylindrical container with a pouring lip; comes in various sizes (50 mL – 1 L).	Holding, mixing, heating liquids, rough measurements.
Conical Flask (Erlenmeyer flask)	Conical body, narrow neck.	Mixing solutions, titration, heating without spillage.
Test Tube	Small cylindrical glass tube with rounded bottom.	Holding small samples, heating small amounts of substances.
Flat-bottom Flask	Round body, flat base.	Heating liquids evenly.
Round-bottom Flask	Round body, no flat base; used with support.	Uniform heating, distillation, reactions under heat.
Reagent Bottle	Glass or plastic bottle, often amber-coloured.	Storing chemicals (liquid or solid).
Measuring Cylinder (Graduated Cylinder)	Tall cylindrical vessel with graduations.	Measuring accurate volumes of liquids.

2. Apparatus for Heating

Apparatus	Description	Uses

Bunsen Burner	Gas burner with adjustable air holes.	Heating, sterilizing, combustion in the lab.
Spirit Lamp	Small container with wick, burns ethanol.	Gentle heating where gas is unavailable.
Tripod Stand	Three-legged stand.	Supports apparatus over heat source.
Wire Gauze	Metal mesh, often with ceramic centre.	Placed on tripod to support glassware and distribute heat evenly.
Crucible	Small ceramic pot with lid.	Heating substances to very high temperatures.
Evaporating Dish	Shallow ceramic or porcelain dish.	Evaporating liquids to leave solid residues.

3. Apparatus for Measuring

Apparatus	Description	Uses
Pipette	Long narrow tube calibrated to deliver a specific volume.	Measuring and transferring exact volumes of liquids.
Burette	Long graduated tube with a tap at the bottom.	Accurate measurement and delivery of liquid in titrations.
Volumetric Flask	Flat-bottomed flask with a long neck and calibration mark.	Preparing solutions of exact concentration.
Stopwatch	Time-measuring device.	Measuring time in reactions.
Thermometer	Graduated glass tube containing mercury or alcohol.	Measuring temperature.

4. Apparatus for Separation and Purification

Apparatus	Description	Uses
Separating Funnel	Pear-shaped glass with a stopcock at the bottom.	Separating immiscible liquids (e.g., oil and water).

Filter Funnel	Cone-shaped apparatus.	Directing liquid into containers and for filtration with filter paper.
Distillation Flask	Round-bottom flask with a side arm.	Separating mixtures by distillation.
Condenser (Liebig condenser)	Glass tube with outer water jacket.	Cooling vapour back to liquid in distillation.
Desiccator	Airtight container with drying agent.	Keeping substances dry or removing moisture.

5. Apparatus for Support and Safety

Apparatus	Description	Uses
Retort Stand with Clamp	Vertical rod fixed to a base with adjustable clamps.	Holding apparatus in place during experiments.
Tongs	Metal grip tool.	Handling hot crucibles and glassware.
Test Tube Holder	Wooden or metal clip.	Holding test tubes during heating.
Fume Cupboard	Enclosed ventilated space.	Conducting experiments that produce harmful gases.
Goggles and Lab Coat	Safety gear.	Protecting eyes and body.

Practical Safety Tips for Students:

- Always wear safety goggles and a lab coat.
- Do not heat a closed container.
- Handle acids and alkalis with care.
- Always read labels on reagent bottles before use.
- Never taste or directly smell chemicals; waft gently if necessary.

Evaluation

1. Define chemical industries.
2. Differentiate between natural and synthetic raw materials with two examples each.
3. Mention two differences between heavy chemicals and fine chemicals.
4. List any four divisions of chemical industries according to products and give one example of raw material for each.
5. State two importance and two disadvantages of chemical industries.
6. Name five laboratory apparatus used for measuring and state one use of each.
7. Differentiate between a beaker and a conical flask in terms of shape and function.
8. What is the function of a condenser in distillation?
9. Give two safety rules to follow when using laboratory apparatus.

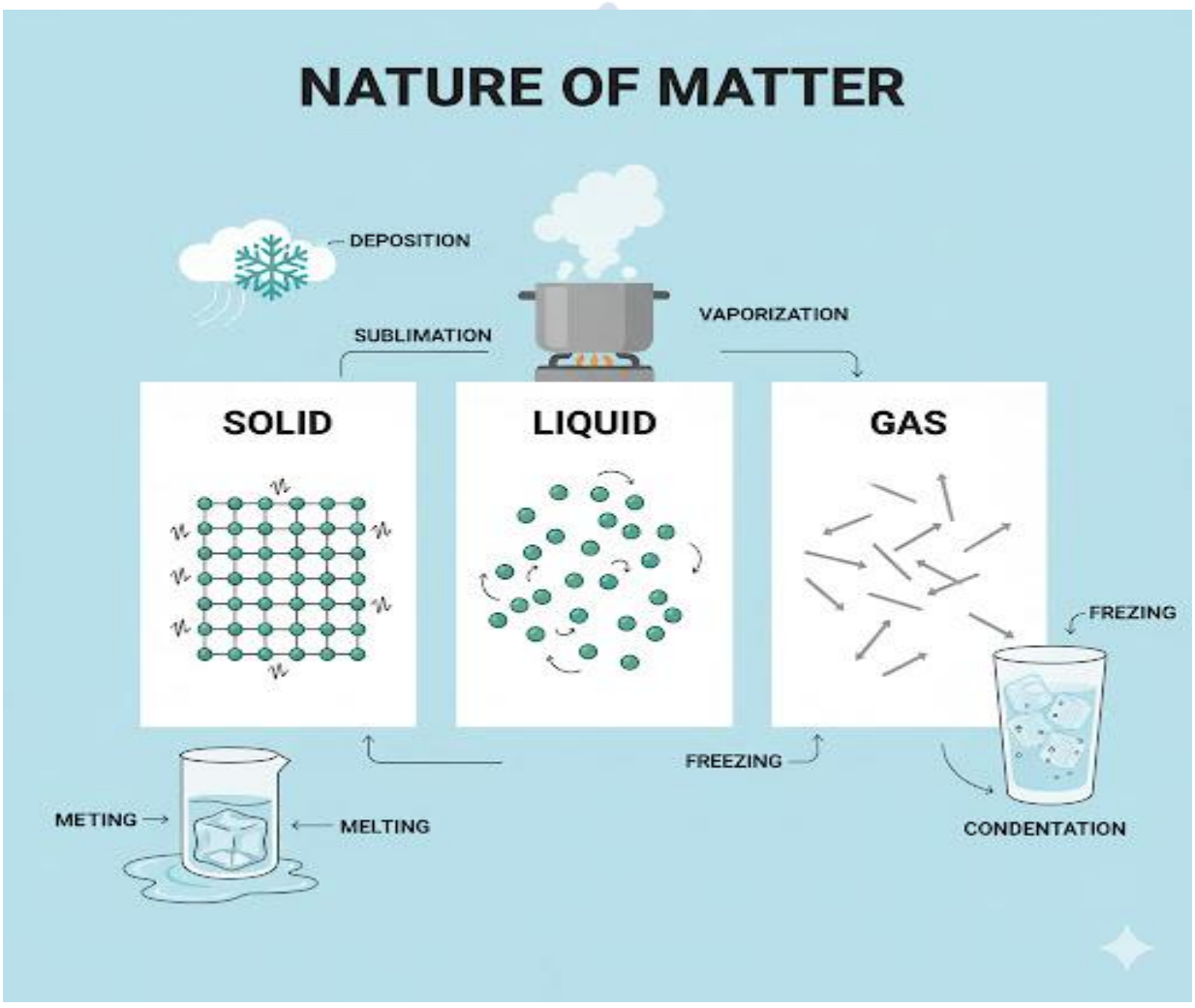
Assignment

1. Draw and neatly label **five apparatus** used for measuring and heating substances in the laboratory.
2. List **five Nigerian chemical industries** and state their main products.

WEEK 3 - NATURE OF MATTER – STATES, PROPERTIES, AND CHANGES OF STATE

Meaning of Matter

Matter is anything that has mass and occupies space. All physical objects—whether visible (like a table) or invisible (like air)—are made of matter. Matter is composed of tiny particles called atoms and molecules that are in constant motion.



States of Matter

Matter exists in three primary states: solid, liquid, and gas. In advanced science, two more states—plasma and Bose–Einstein condensate—are recognized, but at the secondary school level, we focus on the main three.

(a) Solid

- **Particle arrangement:** Particles are closely packed in a fixed, orderly arrangement.
- **Movement:** Particles vibrate in fixed positions.
- **Shape and volume:** Has a definite shape and definite volume.
- **Compressibility:** Cannot be easily compressed due to little space between particles.
- **Examples:** Wood, metal, stone, ice.

(b) Liquid

- **Particle arrangement:** Particles are close together but not in a fixed arrangement.
- **Movement:** Particles can slide over each other, allowing liquids to flow.
- **Shape and volume:** Has definite volume but no definite shape—it takes the shape of its container.
- **Compressibility:** Almost incompressible.
- **Examples:** Water, oil, alcohol, mercury.

(c) Gas

- **Particle arrangement:** Particles are far apart and not in a fixed position.
- **Movement:** Move freely at high speed in all directions.
- **Shape and volume:** Has neither definite shape nor definite volume—it expands to fill its container.
- **Compressibility:** Highly compressible due to large spaces between particles.
- **Examples:** Air, oxygen, carbon dioxide, nitrogen.

Properties of the States of Matter

Property	Solid	Liquid	Gas

Shape	Definite	Indefinite	Indefinite
Volume	Definite	Definite	Indefinite
Arrangement of particles	Fixed, closely packed	Close, random	Far apart, random
Movement of particles	Vibrate in place	Slide over each other	Move freely in all directions
Density	High	Medium	Low
Compressibility	Very low	Very low	High
Energy of particles	Lowest	Moderate	Highest

Changes of State

Changes of state occur when matter is transformed from one state to another by **heating** or **cooling**. These processes involve energy changes but **do not alter the chemical composition** of the substance.

(a) Melting

- **Definition:** Change from solid to liquid upon heating.
- **Example:** Ice melting into water.
- **Energy change:** Heat energy is absorbed to overcome the forces holding particles in fixed positions.

(b) Freezing (Solidification)

- **Definition:** Change from liquid to solid upon cooling.
- **Example:** Water freezing into ice.
- **Energy change:** Heat energy is released as particles slow down and become fixed.

(c) Vaporization (Boiling/Evaporation)

- **Boiling:** Rapid change from liquid to gas at a specific temperature (boiling point).
- **Evaporation:** Slow change from liquid to gas at the surface, at any temperature.
- **Example:** Water boiling at 100°C; drying of wet clothes.

(d) Condensation

- **Definition:** Change from gas to liquid upon cooling.
- **Example:** Water vapour forming droplets on a cold surface.

(e) Sublimation

- **Definition:** Direct change from solid to gas without passing through the liquid state.
- **Examples:** Iodine crystals, dry ice (solid CO_2), naphthalene balls.

(f) Deposition

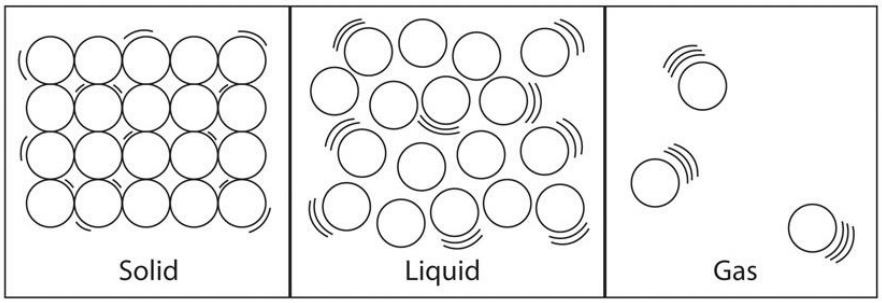
- **Definition:** Change from gas directly to solid without becoming liquid.
- **Example:** Formation of frost.

Particle Theory of Matter

The **Kinetic Theory of Matter** explains that:

1. All matter is made up of tiny particles (atoms or molecules).
2. Particles are in constant motion.
3. The higher the temperature, the faster the particles move.
4. The state of matter depends on the spacing and movement of the particles.

Arrangement of particles in solids, liquids, and gases



Common Misconceptions

- *Misconception:* Gases have no mass.

  Correction: Gases have mass, though they are light and invisible.
- *Misconception:* Melting and dissolving are the same.

  Correction: Melting changes state due to heat; dissolving is mixing with a solvent.

Practical Examples

- Melting: Ice in a glass on a hot day.
- Boiling: Water boiling for cooking.
- Evaporation: Sweat drying from the skin.
- Condensation: Fogging of mirrors after a hot bath.
- Sublimation: Disappearance of mothballs in clothes.

Evaluation

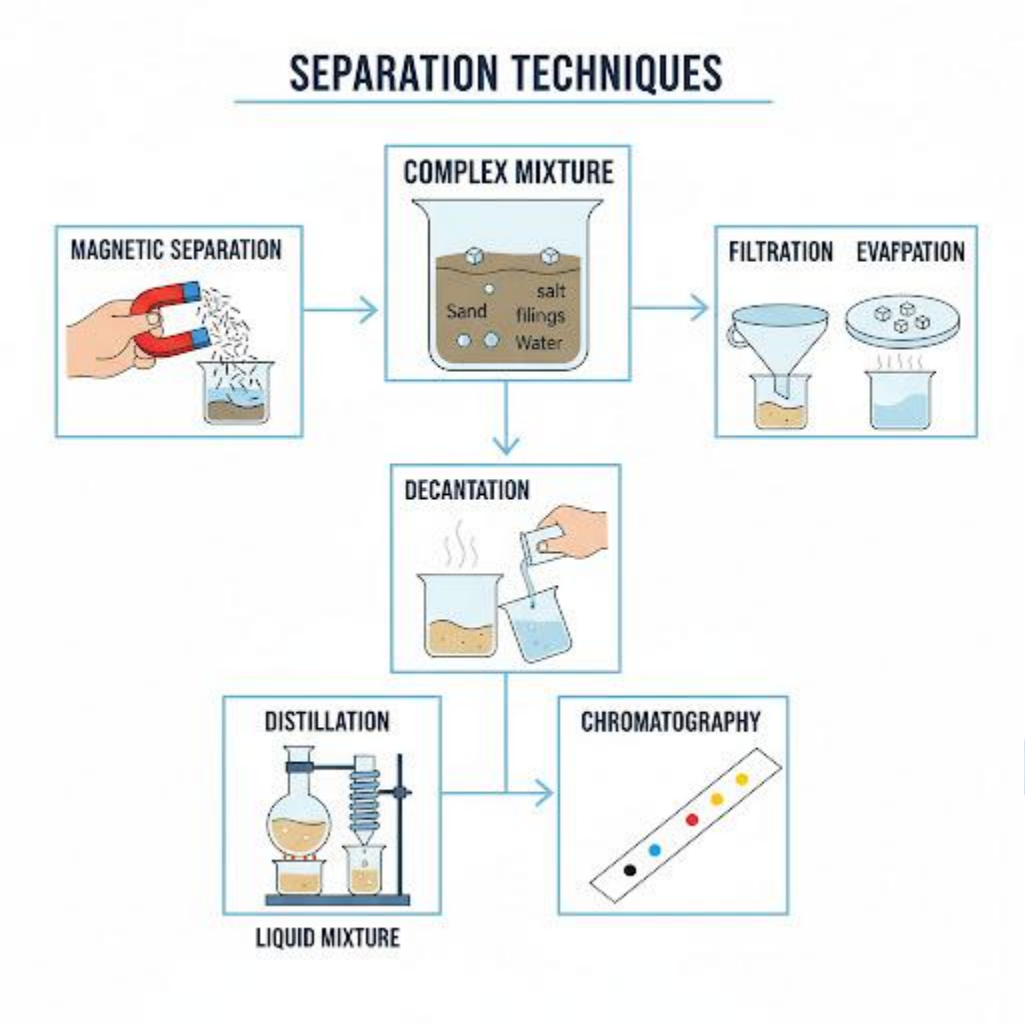
1. Define matter and list its three main states.
2. Differentiate between solids, liquids, and gases using at least four properties.
3. Explain the difference between boiling and evaporation.
4. Give two examples each of:
    - Melting
    - Sublimation
    - Condensation
5. State the particle arrangement and motion in gases.

Assignment

1. With the aid of a diagram, show the particle arrangement in solids, liquids, and gases.
2. State five differences between evaporation and boiling.
3. Write short notes on the following changes of state: freezing, deposition, melting.

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WEEK 4 – SEPARATION TECHNIQUES



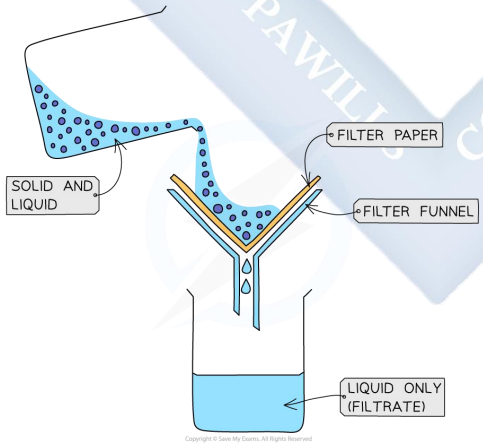
In chemistry, separation techniques are methods used to divide mixtures into their individual components, based on differences in **physical or chemical properties** such as particle size, density, solubility, boiling point, or magnetic properties. The choice of a suitable technique depends on:

- The nature of the mixture (solid–solid, solid–liquid, liquid–liquid, gas mixtures).
- The properties of the components.
- The level of purity required.

Separation method	What it separates
Chromatography	Compounds in a solution with same properties
Filtration	Solids or group of solids and liquids in a mixture
Evaporation	Solids that cannot decompose when heated in a solution
Crystallisation	Dissolved solids in a solution
Simple Distillation	Liquids in a solution
Fractional Distillation	Mixture of miscible (dissolved) liquids
Separating Funnel	Immiscible (undissolved) liquids
Sublimation	Substances that sublime from two substances
Magnetic Attraction	Magnetic substances from non-magnetic ones

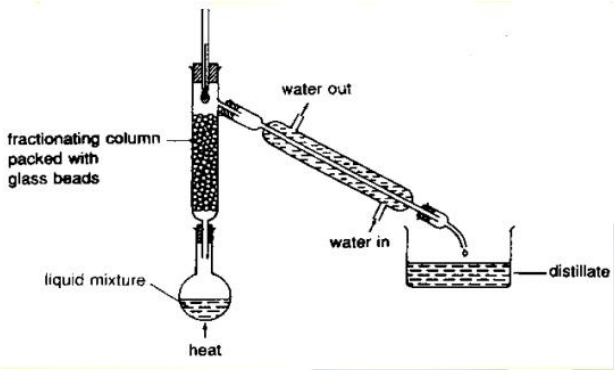
1. Filtration

- **Definition:** A process of separating an insoluble solid from a liquid using a porous barrier (filter paper).
- **Principle:** Based on particle size — solid particles are too large to pass through the filter pores.
- **Example:** Separating sand from water.

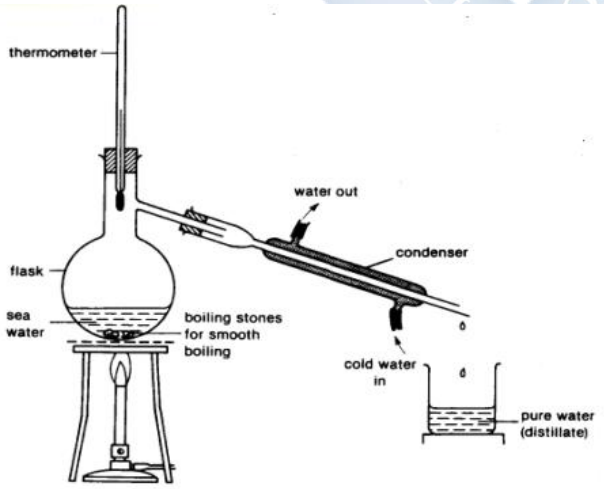


2. Distillation

- **Definition:** Separation of a liquid from a mixture using differences in boiling points.
- **Principle:** Components with lower boiling points vaporize first, then condense back to liquid.
- **Example:** Purifying water from saltwater.



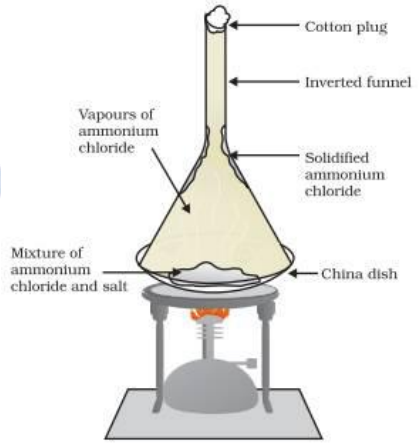
3. Fractional Distillation



- **Definition:** A refined form of distillation used to separate a mixture of liquids with close boiling points.
- **Principle:** Utilizes a fractionating column to allow repeated condensation and vaporization for better separation.
- **Example:** Separating ethanol from water, crude oil refining.

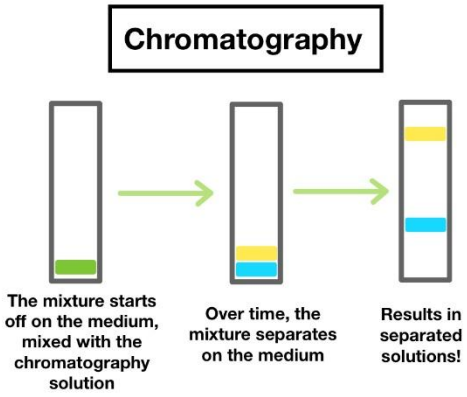
4. Sublimation

- **Definition:** Direct conversion of a solid to gas without passing through the liquid phase.
- **Principle:** Based on differences in sublimation tendencies between substances.
- **Example:** Separating iodine crystals from a mixture.



5. Chromatography

- **Definition:** Technique for separating components of a mixture based on their movement through a medium.
- **Principle:** Components move at different rates due to differences in solubility and adsorption.
- **Example:** Separating plant pigments on paper.



6. Centrifugation

- **Definition:** Separation of mixtures using centrifugal force by spinning at high speed.
- **Principle:** Denser particles move outward faster than less dense ones.
- **Example:** Separating plasma from blood cells.



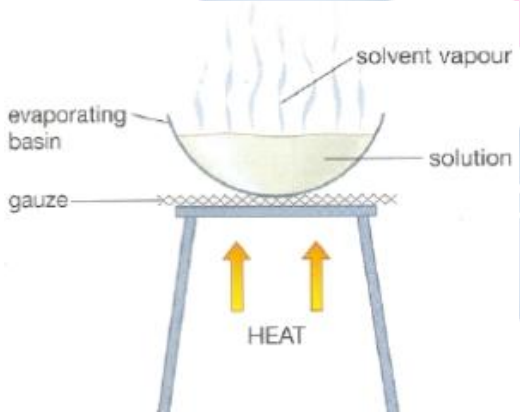
7. Decantation

- **Definition:** Pouring off a liquid to separate it from sediment or an immiscible liquid.
- **Principle:** Relies on density differences between substances.
- **Example:** Removing water from settled sand.



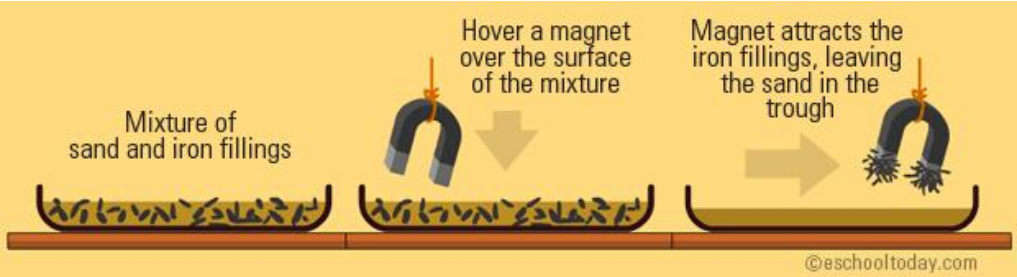
8. Evaporation

- **Definition:** Removal of a liquid by heating until only the solid residue remains.
- **Principle:** Components with lower boiling points evaporate, leaving non-volatile solids behind.
- **Example:** Obtaining salt from seawater.



9. Magnetic Separation

- **Definition:** Use of magnets to separate magnetic materials from non-magnetic ones.
- **Principle:** Based on magnetic properties of substances.



- **Example:** Separating iron filings from sand.

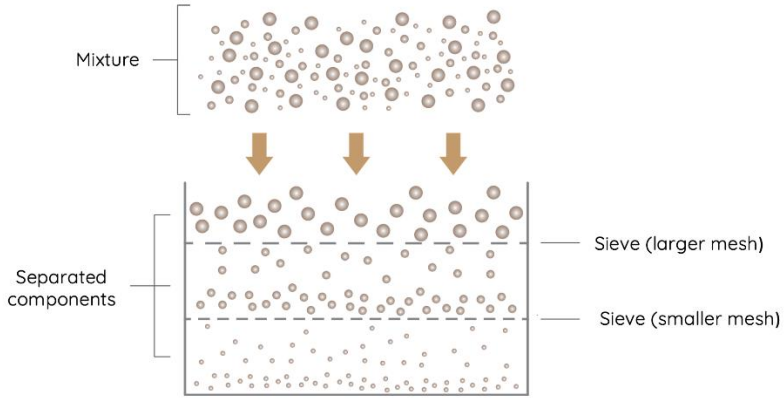
10. Winnowing

- **Definition:** Winnowing is a separation technique used to separate lighter particles (such as husks or chaff) from heavier particles (such as grains or seeds) using the action of wind or air currents.
- **Principle:** The method works on the difference in weight (density) between the two substances — lighter materials are blown away by air, while heavier materials fall straight down.
- **Example:** Separating paddy rice grains from husk after threshing.



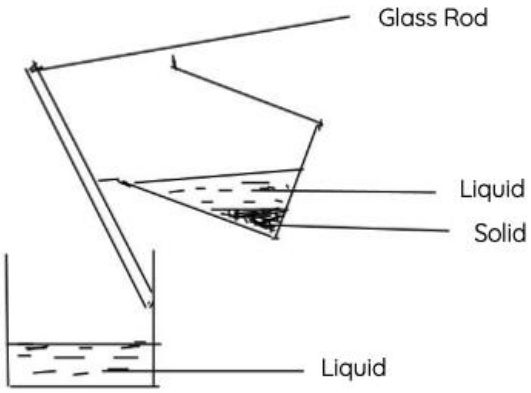
11. Sieving

- **Definition:** Sieving is a separation technique used to separate **particles of different sizes** by passing a mixture through a **sieve** or mesh that allows smaller particles to pass through while retaining the larger ones.
- **Principle:** The method works based on the **difference in particle size**.
- **Example:** Separating wheat flour from bran using a fine mesh sieve.



12. Sedimentation

- **Definition:** Sedimentation is a separation technique in which **insoluble, heavier particles** in a liquid settle at the bottom under the influence of **gravity**.
- **Principle:** The method is based on the **difference in density** between solid particles and the liquid.
- **Example:** Allowing muddy water to stand so that sand and mud settle at the bottom.



EVALUATION

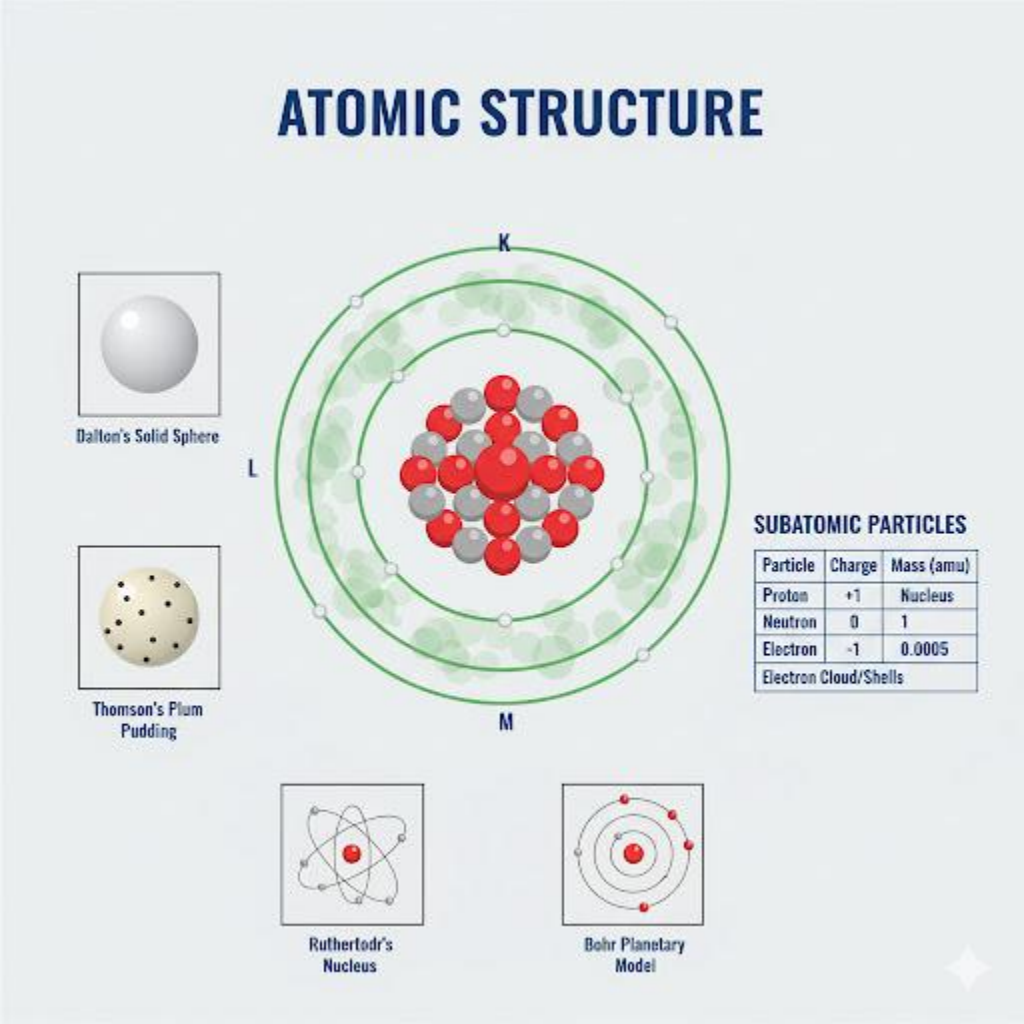
1. Define separation techniques and state three factors that determine the choice of method.
2. Distinguish between simple distillation and fractional distillation.
3. Give two examples of substances that can be separated by sublimation.
4. Explain the principle behind chromatography.
5. Name three separation techniques based on density differences.

ASSIGNMENT

- Draw and label the apparatus for filtration, simple distillation, chromatography, and separating funnel method.
- Research two industrial applications of fractional distillation and write short notes on them.

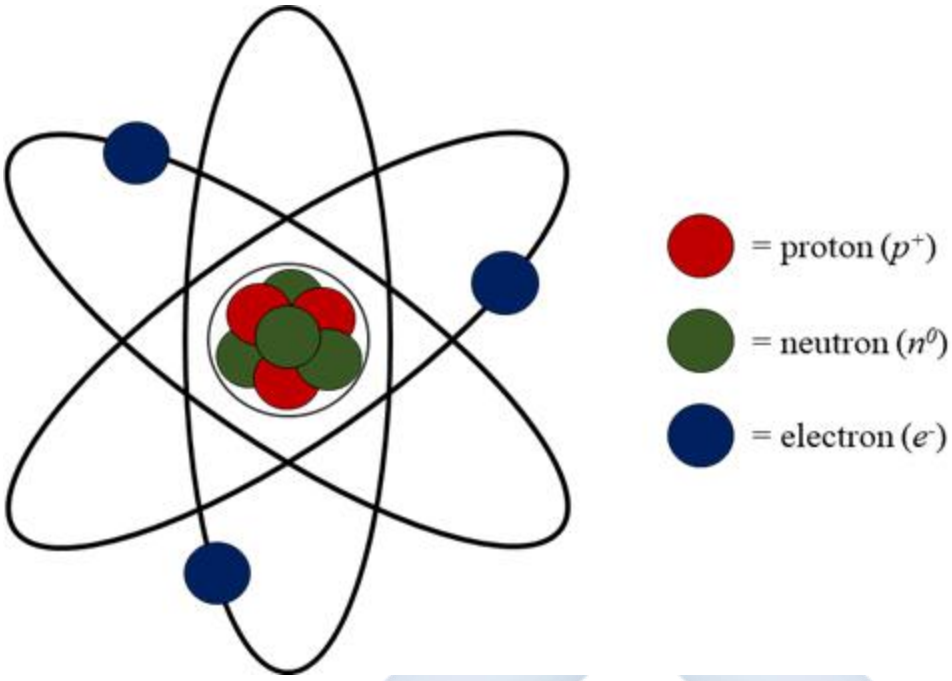
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WEEK 6 – ATOMIC STRUCTURE – STRUCTURE OF THE ATOM, SUBATOMIC PARTICLES



Meaning of Atom

- An **atom** is the smallest particle of an element that takes part in chemical reactions and retains the chemical properties of that element.
- Atoms are extremely small: e.g., radius ≈ 10^{-10} m.



Historical Development of the Atomic Model

1. **Dalton's Atomic Theory (1803)**
    - Atoms are indivisible, indestructible particles.
    - Atoms of the same element are identical; atoms of different elements differ.
    - Atoms combine in simple ratios to form compounds.
    - Limitation: Failed to explain subatomic particles and isotopes.
2. **Thomson's Model (1897)** – Plum pudding model
    - Discovered the **electron**.
    - Proposed that an atom is a sphere of positive charge with electrons embedded.
    - Limitation: Could not explain the stability of atoms.
3. **Rutherford's Model (1911)** – Nuclear model
    - Gold foil experiment showed atoms are mostly empty space.
    - Nucleus at the center, containing positive charges (protons).

- Electrons revolve around nucleus like planets.
- Limitation: Could not explain why electrons don't collapse into the nucleus.

4. **Bohr's Model (1913)** – Planetary model
    - Electrons revolve in fixed orbits (energy levels/shells).
    - Energy is absorbed or released when electrons jump between shells.
    - Limitation: Could not fully explain multi-electron atoms.
5. **Modern Atomic Theory (Quantum Mechanical Model)**
    - Developed by Schrödinger, Heisenberg, etc.
    - Electrons are in regions called **orbitals**, not fixed paths.
    - Atoms consist of a dense nucleus (protons + neutrons) with electrons in shells/subshells.

Subatomic Particles

Particle	Symbol	Charge	Relative Mass	Location in Atom
Proton	p^+	+1	1 a.m.u. ($\approx 1.67 \times 10^{-27}$ kg)	In nucleus
Neutron	n^0	0	1 a.m.u. ($\approx 1.67 \times 10^{-27}$ kg)	In nucleus
Electron	e^-	–1	1/1836 a.m.u. ($\approx 9.11 \times 10^{-31}$ kg)	Orbiting nucleus in shells

- **Protons**: Determine the identity of an element (Atomic number).
- **Neutrons**: Provide stability in the nucleus.
- **Electrons**: Involved in chemical bonding and reactions.

Atomic Number and Mass Number

- **Atomic Number (Z):** Number of protons in the nucleus.

- **Mass Number (A):** Number of protons + neutrons.
- **Neutron number (N):** N = A – Z.

Example: Carbon-12 has Z = 6, A = 12 → Protons = 6, Neutrons = 6, Electrons = 6.

Isotopes

- Atoms of the same element (same Z) but different mass numbers (A).
- Example: ^{12}C, ^{13}C, ^{14}C are isotopes of carbon.
- They have similar chemical properties but different physical properties.

Atomic Structure Representation

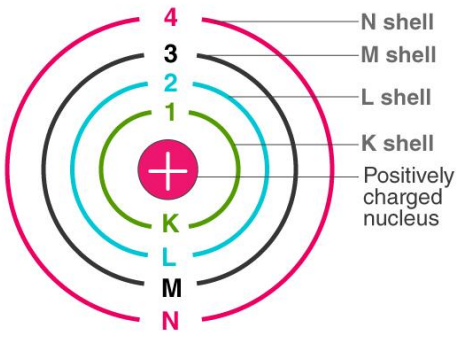
- Electrons are arranged in **shells** (K, L, M, N) according to the $2n^2$ rule:
  - K shell → 2 electrons
  - L shell → 8 electrons
  - M shell → 18 electrons
  - N shell → 32 electrons

Example:

- Sodium (Z = 11) → 2, 8, 1
- Oxygen (Z = 8) → 2, 6

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BOHR'S MODEL OF AN ATOM



Evaluation

1. Define an atom.
2. State three limitations of Dalton's Atomic Theory.
3. Mention the scientist who discovered the electron.
4. Distinguish between **atomic number** and **mass number**.
5. Write the electronic configuration of:
    - Magnesium (Z = 12)
    - Chlorine (Z = 17)
6. Fill in the table:

Element	Atomic Number	Mass Number	Protons	Neutrons	Electrons
Sodium	11	23	?	?	?
Oxygen	8	16	?	?	?

Assignment

1. List and explain **five differences** between Bohr's model and Rutherford's model of the atom.
2. Write the electronic configurations of the following elements:

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- (a) Nitrogen (Z = 7)
- (b) Potassium (Z = 19)
- (c) Sulphur (Z = 16)

3. Neon (Z = 10, A = 20). Determine the number of protons, neutrons, and electrons.
4. Explain briefly the significance of isotopes in:
    - Medicine
    - Agriculture
    - Industry

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WEEK 8: PERIODIC TABLE – GROUPS, PERIODS, AND PERIODICITY

The periodic table is one of the most important tools in Chemistry. It is a systematic arrangement of all the known elements in rows and columns in such a way that elements with similar properties are placed in the same vertical column (group). This arrangement is based on **increasing atomic number**, which is the number of protons in the nucleus of an atom. The periodic table not only helps chemists to classify elements but also enables them to predict the properties of elements and their compounds.

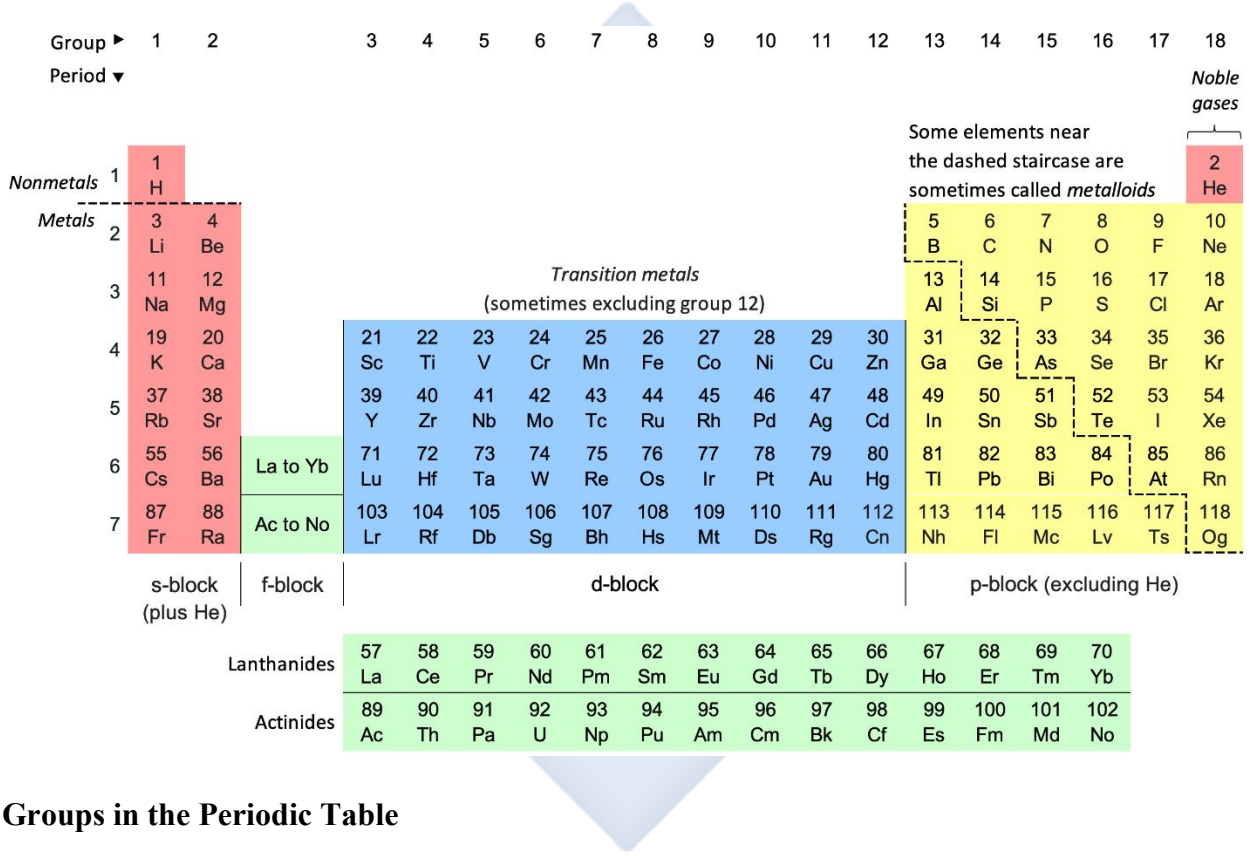
Historical Development of the Periodic Table

- **Johann Döbereiner (1829):** Proposed the Law of Triads. He noticed that elements with similar chemical properties could be grouped in threes (triads). The atomic mass of the middle element was approximately the average of the other two. Example: Lithium (Li), Sodium (Na), and Potassium (K).
- **John Newlands (1864):** Proposed the Law of Octaves. He observed that when elements were arranged in increasing order of atomic mass, every eighth element had properties similar to the first. This was not universally applicable.
- **Dmitri Mendeleev (1869):** Developed the first widely accepted periodic table. He arranged elements in order of increasing atomic mass but placed elements with similar properties under the same group. Importantly, he left gaps for undiscovered elements and correctly predicted their properties (e.g., Germanium).
- **Henry Moseley (1913):** Through X-ray experiments, Moseley showed that the atomic number, not atomic mass, is the correct basis for arrangement. This solved the anomalies in Mendeleev's table and gave rise to the **modern periodic table**.

Structure of the Modern Periodic Table

- The periodic table is arranged in **7 horizontal rows (Periods)** and **18 vertical columns (Groups)**.

- Elements are arranged in order of increasing atomic number.
- The periodic table is divided into **blocks** based on electron configuration:
    - **s-block:** Groups 1 and 2 (alkali and alkaline earth metals).
    - **p-block:** Groups 13–18 (includes halogens and noble gases).
    - **d-block:** Transition metals (Groups 3–12).
    - **f-block:** Lanthanides and actinides, usually placed separately at the bottom.



Groups in the Periodic Table

A **group** is a vertical column of elements in the periodic table. Elements in the same group have the same number of **valence electrons**, which explains their similar chemical properties.

- **Group 1 (Alkali metals):** Li, Na, K, Rb, Cs, Fr. Very reactive, soft metals, low ionization energies, form +1 cations, react vigorously with water to produce hydrogen gas and alkaline hydroxides.
- **Group 2 (Alkaline earth metals):** Be, Mg, Ca, Sr, Ba, Ra. Reactive metals but less than Group 1, form +2 cations, react with acids to form hydrogen gas.

- **Groups 3–12 (Transition metals):** Have partially filled d-orbitals, show variable oxidation states, form colored compounds, many act as catalysts (e.g., Fe, Ni, Pt).
- **Group 17 (Halogens):** F, Cl, Br, I, At. Very reactive non-metals, form -1 anions, exist as diatomic molecules, strong oxidizing agents.
- **Group 18 (Noble gases):** He, Ne, Ar, Kr, Xe, Rn. Inert gases, stable due to full valence shell configuration, very low chemical reactivity, used in lighting and welding.

Periods in the Periodic Table

A **period** is a horizontal row of elements in the periodic table. There are **7 periods**.

- In a period, the atomic number increases from left to right.
- All elements in the same period have the **same number of electron shells** but different numbers of valence electrons.
- Across a period, properties change progressively:
    - Metallic character decreases from left to right.
    - Non-metallic character increases from left to right.
    - Electronegativity and ionization energy increase.
    - Atomic radius decreases because of increased nuclear charge pulling electrons closer.

Periodicity in Properties of Elements

Periodicity is the repeating pattern of properties of elements across different periods and groups of the periodic table. These periodic trends arise due to changes in atomic structure (mainly nuclear charge and electron configuration).

- **Atomic Radius:** Distance from nucleus to the outermost electron.
    - Decreases across a period (due to increasing nuclear charge).

    - Increases down a group (due to addition of electron shells).
- **Ionization Energy:** Energy required to remove the outermost electron.
    - Increases across a period (greater nuclear attraction).
    - Decreases down a group (outer electrons farther from nucleus, easier to remove).
- **Electronegativity:** Ability of an atom to attract shared electrons.
    - Increases across a period.
    - Decreases down a group.
- **Electron Affinity:** Energy change when an atom gains an electron.
    - Generally increases across a period.
    - Decreases down a group.
- **Metallic Character:** Tendency of an element to lose electrons.
    - Decreases across a period.
    - Increases down a group.

EVALUATION

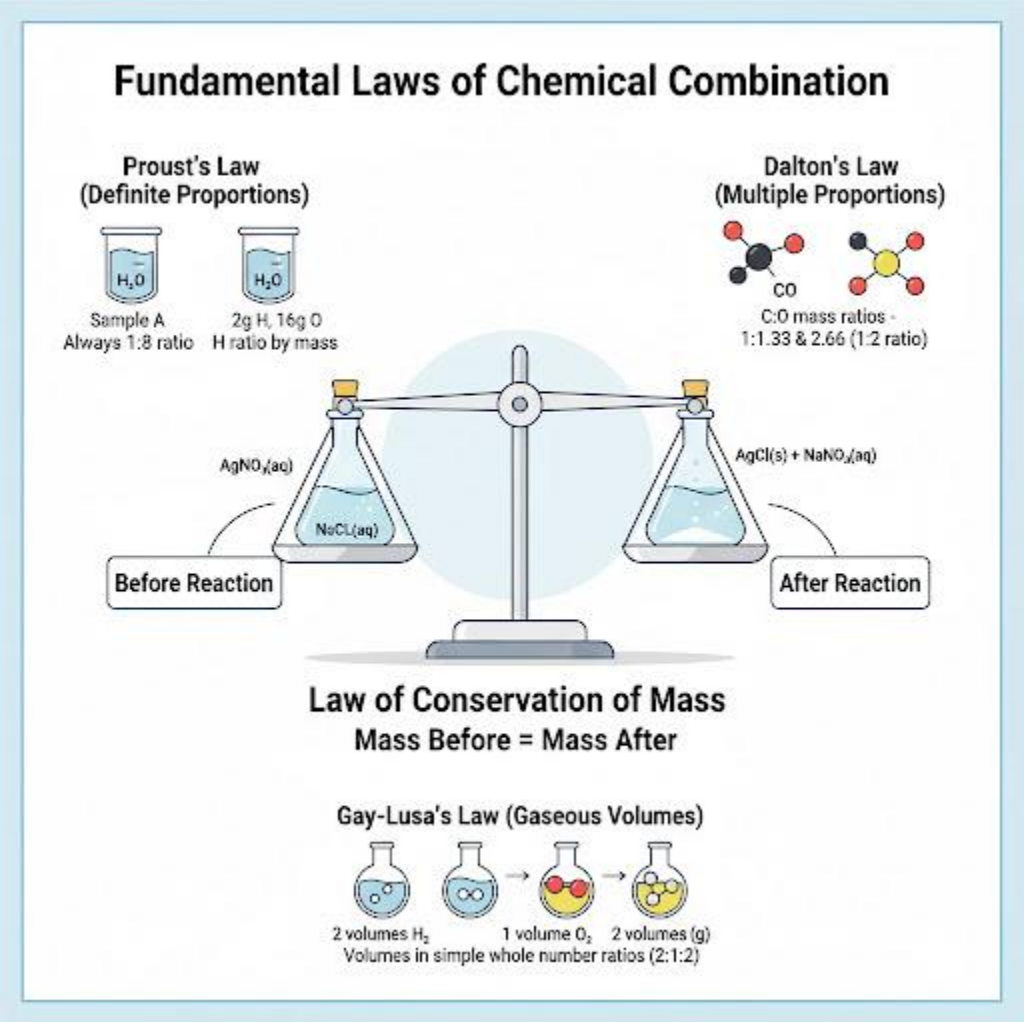
1. Differentiate between a group and a period in the periodic table.
2. State two major defects of Mendeleev's periodic table and explain how the modern periodic table corrected them.
3. Explain why fluorine is more electronegative than chlorine.
4. Describe the trend in ionization energy across a period and down a group.
5. Compare the chemical reactivities of Group 1 and Group 17 elements.

ASSIGNMENT

1. Draw a simplified periodic table showing at least the first 20 elements. Indicate clearly the groups and periods.
2. Discuss in detail the periodic trends of atomic radius, ionization energy, and metallic character with examples.
3. Explain the similarities and differences between Group 1 (alkali metals) and Group 2 (alkaline earth metals).

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WEEK 9: CHEMICAL COMBINATION – LAWS OF CHEMICAL COMBINATION



Meaning of Chemical Combination

- Chemical combination is the process by which two or more substances (elements or compounds) unite to form a new substance with new properties.
- The substances that react are called **reactants**, while the new substance formed is called the **product**.
- Chemical combination differs from **physical combination** (e.g., sand + salt, iron + sulphur mixture) because:

- In chemical combination, the substances lose their individual properties and form a new substance (e.g., iron + sulphur → iron(II) sulphide).
- In physical combination, the components retain their properties and can usually be separated by physical means.

- Example:
  - **Physical**: Mixture of iron filings and sulphur can be separated by magnet.
  - **Chemical**: Heating iron filings and sulphur gives iron(II) sulphide (FeS) – a compound with entirely new properties.

This process follows certain **natural rules or principles** called the **Laws of Chemical Combination**, which were discovered by early chemists through experiments and form the foundation of modern chemistry.

The Laws of Chemical Combination

(a) Law of Conservation of Mass

- **Proposed by:** Antoine Lavoisier (1789) – often called *the Father of Modern Chemistry*.
- **Statement:** *Matter is neither created nor destroyed in the course of a chemical reaction. The total mass of the reactants equals the total mass of the products.*
- **Explanation:** During a chemical reaction, atoms are simply rearranged, not lost or created. This means the overall mass remains constant.
- **Equation Example:**

$2H_2 + O_2 \longrightarrow 2H_2O$

  - Mass of reactants = (4 g H_2 + 32 g O_2) = 36 g
  - Mass of products = 36 g H_2O
    → Mass is conserved.

- **Experimental Verification:**

- When mercuric oxide (HgO) is heated, it decomposes into mercury (Hg) and oxygen (O_2). Careful weighing before and after reaction shows no change in total mass.

- **Application:** Forms the basis of **stoichiometric calculations** in chemistry (mole concept, chemical equations).

(b) Law of Definite Proportions (Law of Constant Composition)

- **Proposed by:** Joseph Proust (1799).
- **Statement:** *A given compound always contains the same elements combined in the same fixed proportion by mass, regardless of the method of preparation or the source.*
- **Explanation:** Water from a river, rain, or laboratory experiment always contains hydrogen and oxygen in the ratio **1:8 by mass**.
- **Example:**
  - Water (H_2O):
    - Atomic mass of H = 1, O = 16
    - 2 atoms of H = 2 × 1 = 2
    - 1 atom of O = 16
    - Ratio of H: O = 2:16 = **1:8** (constant always).
- **Application:** Used to determine chemical formulae of compounds.

(c) Law of Multiple Proportions

- **Proposed by:** John Dalton (1803).
- **Statement:** *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 simple whole-number ratios.*

- **Explanation:** Same elements can combine in different ways to form different compounds, but their combining masses remain in simple ratios.
- **Example (Carbon + Oxygen):**
    - **Carbon monoxide (CO):** 12 g C combines with 16 g O.
    - **Carbon dioxide (CO_2):** 12 g C combines with 32 g O.
    - Ratio of O in the two compounds = 16:32 = 1:2 (simple whole number).
- **Application:** Supports Dalton's Atomic Theory (atoms combine in whole numbers, not fractions).

(d) Law of Reciprocal Proportions

- **Proposed by:** Jeremias Richter (1792).
- **Statement:** *If element A combines separately with element B and with element C, then the ratio in which B and C combine with each other is either the same or a simple multiple of the ratio in which they combine with A.*
- **Example:**
    - H combines with O to form water (H_2O).
    - H combines with Cl to form HCl.
    - O and Cl combine to form compounds like Cl_2O.
      → Their masses follow simple ratios that are consistent with the law.
- **Application:** Helps in establishing chemical relationships among different compounds.

(e) Law of Combining Volumes (Gay-Lussac's Law)

- **Proposed by:** Joseph Gay-Lussac (1808).

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- **Statement:** *When gases react, they do so in volumes that bear a simple whole-number ratio to one another and to the volume of the gaseous products, provided all are measured at the same temperature and pressure.*
- **Example:**

$2H_2(g) + O2(g) \longrightarrow 2H_2O(g)$

  - Volumes: H_2: O_2: H_2O = 2: 1: 2 (simple whole numbers).

- **Application:** Foundation for Avogadro's hypothesis and the **molar volume concept** (22.4 dm^3 at STP).

Experimental Verifications (Summary of Classic Experiments)

1. **Lavoisier's Experiment:** Decomposition of mercuric oxide to verify conservation of mass.
2. **Proust's Experiment:** Analysis of water from different sources always gave H:O = 1:8 by mass.
3. **Dalton's Example:** CO and CO_2 to verify multiple proportions.
4. **Gay-Lussac's Experiment:** Reaction of hydrogen and oxygen gases in a closed vessel to form water vapor in a simple volume ratio.

Importance of the Laws of Chemical Combination

- Provided strong experimental evidence for **Dalton's Atomic Theory**.
- Laid the foundation for the **mole concept** and **stoichiometry**.
- Enabled chemists to determine the **formulae of compounds** accurately.
- Guided the development of **modern atomic structure** and the **periodic table**.
- Essential for **quantitative chemistry** (calculating masses, moles, and volumes of substances).

- Still used today in industries (pharmaceuticals, fertilizers, metallurgy, etc.) where chemical reactions must be carried out in precise proportions.

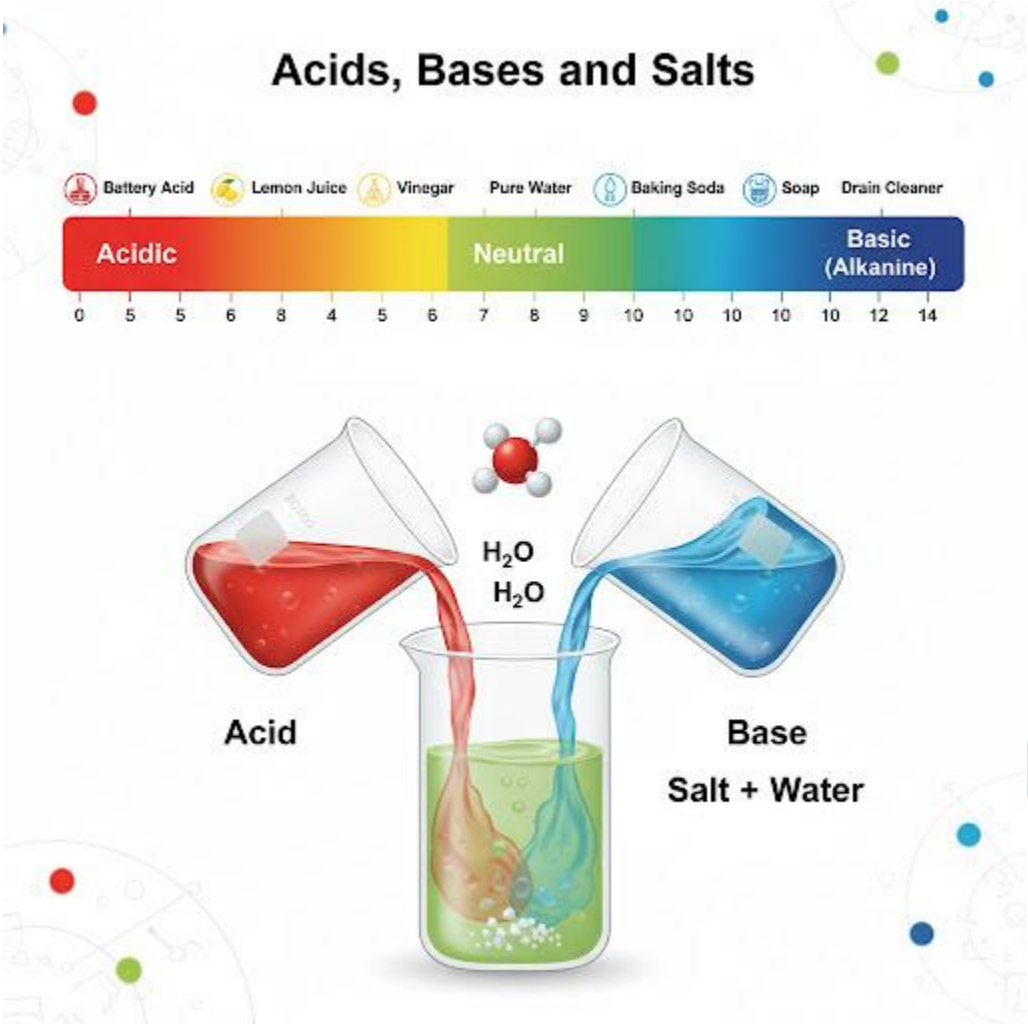
Evaluation:

1. Define chemical combination and differentiate it from physical combination, with examples.
2. State the Law of Conservation of Mass and verify it using an example.
3. Water from a laboratory and a river were analyzed. In both cases, hydrogen and oxygen were found in the ratio of 1:8 by mass. Which law does this illustrate? Explain.
4. Carbon and oxygen form two oxides. In one, 12 g of carbon combines with 16 g of oxygen, and in the other, 12 g of carbon combines with 32 g of oxygen. Show how this supports the Law of Multiple Proportions.
5. Using a balanced chemical equation, demonstrate Gay-Lussac's Law of Combining Volumes.
6. Mention three applications of the laws of chemical combination in modern chemistry.

Assignment:

1. Differentiate between the **Law of Definite Proportions** and the **Law of Multiple Proportions**, giving two examples each.
2. Hydrogen and chlorine combine to form hydrogen chloride (HCl). 2 dm^3 of hydrogen reacts with 2 dm^3 of chlorine to produce 4 dm^3 of HCl gas. Use this data to explain Gay-Lussac's Law of Combining Volumes.
3. Research and write a detailed account (one page) on the historical experiments of **Antoine Lavoisier** and how his work changed the course of chemistry.

WEEK 10: ACIDS, BASES AND SALTS



Definitions of Acids, Bases, and Salts

- **Acids:**
  Acids are substances that ionize in aqueous solution to produce hydrogen ions (H^+) or protons as the only positive ion. According to Arrhenius, an acid increases the concentration of H^+ ions in solution. Bronsted–Lowry defines an acid as a **proton donor**, while Lewis defines it as an **electron pair acceptor**.
    - Examples: HCl (Hydrochloric acid), H_2SO_4 (Sulphuric acid), HNO_3 (Nitric acid), CH_3COOH (Ethanoic acid).

- **Bases:**

  A base is any substance that reacts with an acid to neutralize it and form salt and water only. Arrhenius defines a base as a substance that increases the concentration of OH^- ions in aqueous solution. Bronsted–Lowry defines a base as a **proton acceptor**, while Lewis defines it as an **electron pair donor**.

  - Examples: NaOH (Sodium hydroxide), KOH (Potassium hydroxide), $Ca(OH)_2$ (Calcium hydroxide), NH_3 (Ammonia).

- **Alkalis:**

  Alkalis are soluble bases which, when dissolved in water, release OH^- ions. All alkalis are bases, but not all bases are alkalis.

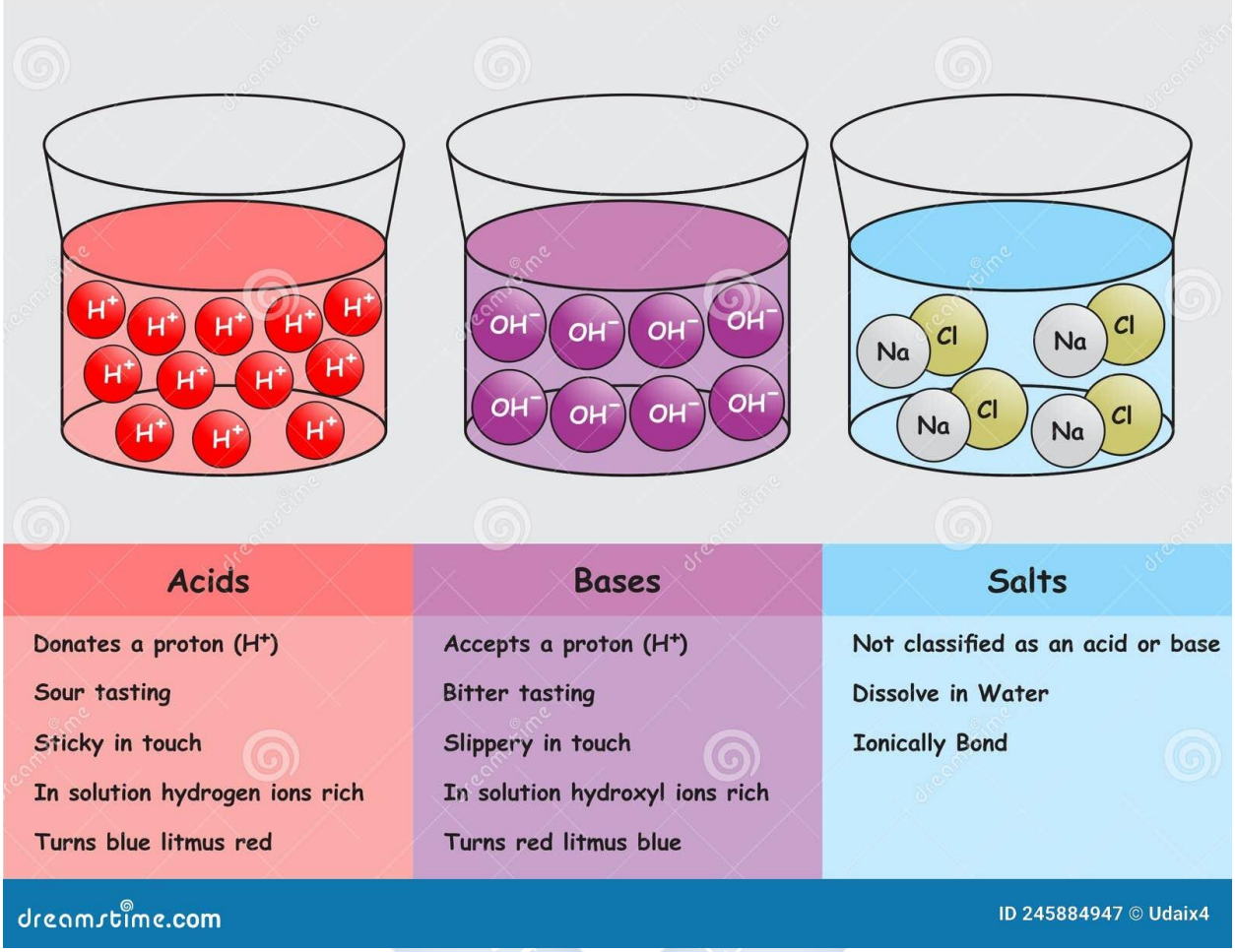
  - Examples: NaOH, KOH, NH_4OH.

- **Salts:**

  Salts are compounds formed when the hydrogen ion of an acid is wholly or partially replaced by a metal ion or ammonium ion. Salts can be normal, acidic, or basic, depending on the extent of replacement.

  - Examples: NaCl (Sodium chloride), K_2SO_4 (Potassium sulphate), NH_4NO_3 (Ammonium nitrate).

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Examples of Acids, Bases, and Salts in Everyday Life

- **Acids:**
    - Vinegar (contains ethanoic acid).
    - Citrus fruits like orange and lemon (contain citric acid).
    - Soft drinks (contain carbonic acid).
    - Battery acid (sulphuric acid).
- **Bases:**
    - Soap and detergents (contain sodium or potassium hydroxides).

- Toothpaste (contains mild alkalis to neutralize mouth acids).
- Limewater ($Ca(OH)_2$) used in construction.
- Baking soda ($NaHCO_3$).

- **Salts:**
  - Table salt (NaCl).
  - Fertilizers (e.g., ammonium nitrate NH_4NO_3, potassium chloride KCl).
  - Washing soda ($Na_2CO_3 \cdot 10H_2O$).
  - Epsom salt ($MgSO_4 \cdot 7H_2O$).

Properties of Acids, Bases, and Salts

(a) Properties of Acids

1. Taste sour (e.g., citric acid in lemon).
2. Turn blue litmus paper red.
3. Have pH less than 7.
4. Strong acids (like HCl, H_2SO_4) are corrosive to skin and metals.
5. React with metals to produce hydrogen gas (e.g., $2HCl + Zn \rightarrow ZnCl_2 + H_2$).
6. React with bases/alkalis to form salt and water (neutralization).
7. React with carbonates and bicarbonates to release carbon dioxide ($HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$).

(b) Properties of Bases

1. Taste bitter and feel soapy or slippery.
2. Turn red litmus paper blue.
3. Have pH greater than 7.

4. Strong alkalis (NaOH, KOH) are corrosive and can burn skin.
5. React with acids to form salt and water.
6. React with ammonium salts on heating to release ammonia gas ($NaOH + NH_4Cl \rightarrow NaCl + H_2O + NH_3$).
7. Some bases are insoluble in water (e.g., CuO, $Fe(OH)_3$).

(c) Properties of Salts

1. Many salts are crystalline solids.
2. They can be soluble (NaCl, KNO_3) or insoluble ($PbSO_4$, $BaSO_4$).
3. Neutral salts (NaCl) have no effect on litmus.
4. Acidic salts ($NaHSO_4$) turn blue litmus red.
5. Basic salts (Pb(OH)Cl) turn red litmus blue.
6. Conduct electricity when molten or in aqueous solution.

The pH Scale

- The pH scale is a logarithmic scale used to measure the concentration of hydrogen ions (H^+) in a solution, ranging from **0 to 14**.
- **pH < 7:** Acidic solution.
- **pH = 7:** Neutral solution (pure water).
- **pH > 7:** Basic/alkaline solution.

Indicators Used in pH Determination:

- Litmus (red in acid, blue in base).
- Methyl orange (red in acid, yellow in base).
- Phenolphthalein (colourless in acid, pink in base).

- Universal indicator (shows different colours corresponding to pH values).

Practical Importance of pH:

- Agriculture: pH of soil determines fertility. Acidic soils are treated with lime ($Ca(OH)_2$).
- Medicine: Stomach acid (pH 1.5 – 2.5) is neutralized by antacids when excessive.
- Biology: Blood has a pH of about 7.35 – 7.45. Any deviation is dangerous.

Neutralization Reaction

- A **neutralization reaction** is a chemical reaction between an acid and a base (or alkali) to produce a salt and water.
- General equation:

Acid + Base → Salt + Water

Examples:

1. $HCl + NaOH \rightarrow NaCl + H_2O$
2. $H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$
3. $HNO_3 + NH_4OH \rightarrow NH_4NO_3 + H_2O$

Characteristics of Neutralization:

- Exothermic reaction (releases heat).
- Can be used for titration experiments to determine concentration of solutions.

Applications of Neutralization:

1. **Medicine:** Antacids neutralize excess stomach acid.
2. **Agriculture:** Quicklime (CaO) or slaked lime ($Ca(OH)_2$) used to reduce soil acidity.
3. **Environmental Science:** Industrial effluents containing acids are neutralized before disposal.

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4. **Daily Life:** Baking soda neutralizes bee stings (acidic) while vinegar neutralizes wasp stings (alkaline).

EVALUATION

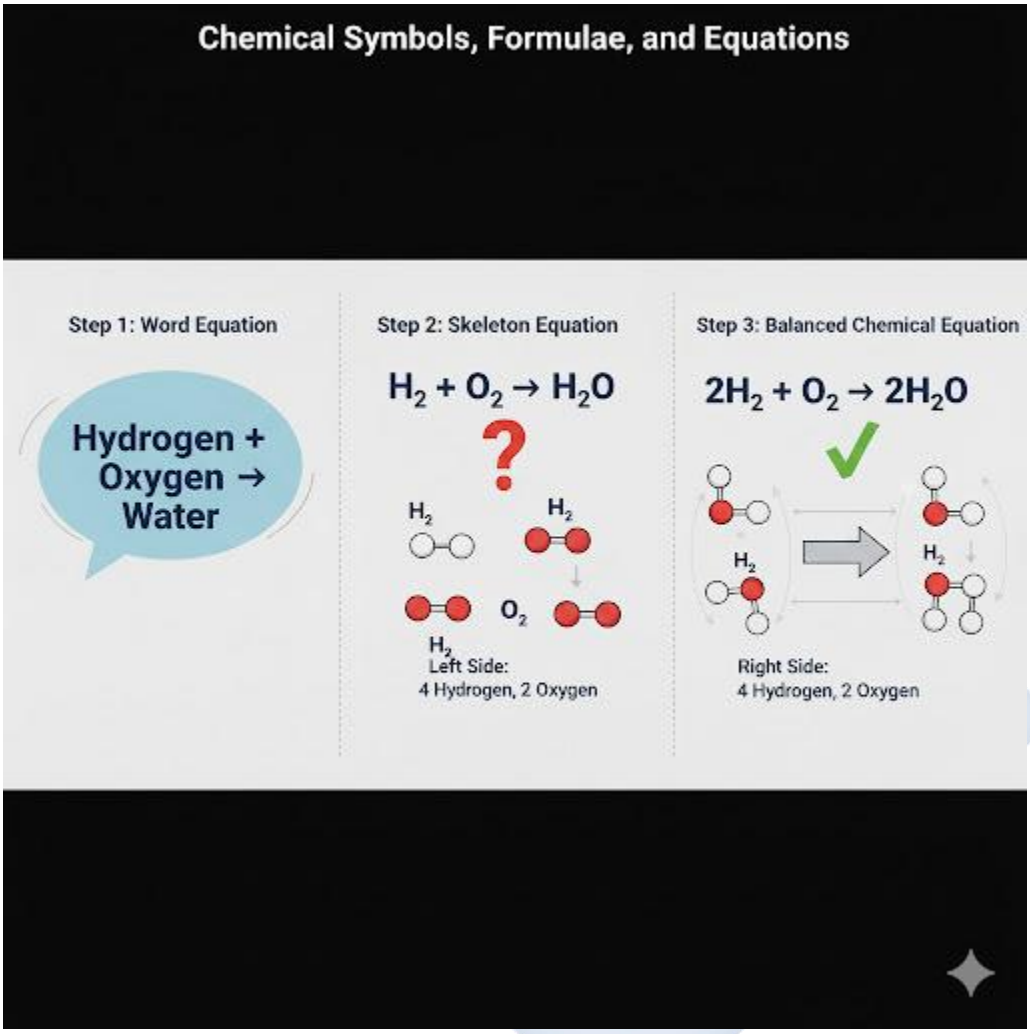
1. Define acid, base, alkali, and salt with examples.
2. Explain the differences between a base and an alkali.
3. List four properties of acids and four properties of bases.
4. State three real-life uses of neutralization reactions.
5. A solution has a pH of 4.5. Is it acidic, neutral, or alkaline? Give reasons.
6. Write balanced equations for the following neutralization reactions:
   (a) HCl with NaOH
   (b) H_2SO_4 with KOH
   (c) HNO_3 with NH_4OH

ASSIGNMENT

1. Provide three examples each of acids, bases, and salts found in:
   (a) the home,
   (b) industry,
   (c) agriculture.
2. Explain why strong acids and strong bases are considered dangerous substances in laboratories and industries.
3. Describe, with examples, how the pH scale is used in:
   (a) medicine,
   (b) agriculture,
   (c) environmental management.

WEEK 11: CHEMICAL SYMBOLS, FORMULAE, AND EQUATIONS

(Writing symbols, writing formulae, balancing chemical equations)



CHEMICAL SYMBOLS

- **Definition:**
  A chemical symbol is a universal shorthand representation of a chemical element. Each element has a unique symbol, typically one or two letters, based on the element's English or Latin name.
- **Importance of Chemical Symbols:**

1. **Universal Understanding:** Chemists worldwide use the same symbols regardless of language differences, ensuring clear communication.
2. **Space and Time Saving:** Instead of writing "oxygen," we simply write **O**; instead of "sodium," we write **Na**.
3. **Precision and Accuracy:** Symbols reduce ambiguity in chemical descriptions.
4. **Foundation for Formulae and Equations:** Without symbols, writing equations or molecular compositions would be cumbersome.

- **Examples of Chemical Symbols:**

Element Name	Symbol	Latin/Greek Name (if different)
Hydrogen	H	Hydrogenium
Sodium	Na	Natrium
Potassium	K	Kalium
Iron	Fe	Ferrum
Copper	Cu	Cuprum
Silver	Ag	Argentum
Lead	Pb	Plumbum
Tin	Sn	Stannum
Mercury	Hg	Hydrargyrum
Antimony	Sb	Stibium

- **Rules for Writing Chemical Symbols:**

1. The **first letter is always capitalized**, and the second (if present) is lowercase (e.g., Mg, Al).
2. No two elements share the same symbol.

3. Symbols may derive from **English names** (e.g., C for carbon) or **Latin names** (e.g., Au for gold, from *Aurum*).
4. For elements with similar first letters, other letters are used to avoid confusion (e.g., C for carbon, Cl for chlorine, Ca for calcium).

CHEMICAL FORMULAE

- **Definition:**
  A **chemical formula** expresses the composition of a chemical substance using symbols and numbers. The subscript in the formula indicates the number of atoms of each element present in one molecule of the compound.
- **Significance of Formulae:**
  - They indicate the **types of atoms** present in a compound.
  - They show the **ratio of elements** in a molecule.
  - They allow chemists to calculate **molar mass** and predict reaction quantities.
  - They are used in writing **balanced chemical equations**.

- **Types of Formulae:**

Type of Formula	Description	Example
Empirical Formula	Simplest whole-number ratio of atoms of elements in a compound.	CH_2O (for glucose)
Molecular Formula	Shows the actual number of each type of atom in a molecule.	$C_6H_{12}O_6$
Structural Formula	Diagrammatic representation showing how atoms are bonded.	H–O–H (for water)

Condensed Formula	Shows bonding groups in a molecule without drawing full structure.	CH_3CH_2OH (for ethanol)

- **Steps to Write Chemical Formulae:**
    1. Write down the symbols of the elements.
    2. Write their **valencies** (combining capacities).
    3. Cross-multiply valencies to determine subscripts.
    4. Write the subscripts (omit if "1").
    5. Simplify subscripts to the smallest whole numbers.

Examples:

- Sodium chloride: $Na^{1+} + Cl^{1-} \rightarrow NaCl$
- Calcium chloride: $Ca^{2+} + Cl^{1-} \rightarrow CaCl_2$
- Aluminium oxide: $Al^{3+} + O^{2-} \rightarrow Al_2O_3$

- **Common Valencies Table:**

Element	Symbol	Valency
Hydrogen	H	1
Sodium	Na	1
Potassium	K	1
Calcium	Ca	2
Magnesium	Mg	2
Zinc	Zn	2

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Aluminium	Al	3
Oxygen	O	2
Sulphur	S	2
Nitrogen	N	3
Chlorine	Cl	1
Iron (II)	Fe^{2+}	2
Iron (III)	Fe^{3+}	3

CHEMICAL EQUATIONS

- **Definition:**

  A **chemical equation** is a symbolic representation of a chemical reaction showing the reactants, products, and their molar relationships.

- **Types of Equations:**

  1. **Word Equation:** Written in words only.
     Example: Hydrogen + Oxygen → Water.
  2. **Skeleton (Unbalanced) Equation:** Symbols used but not balanced.
     Example: $H_2 + O_2 \rightarrow H_2O$.
  3. **Balanced Equation:** Shows equal numbers of each atom on both sides.
     Example: $2H_2 + O_2 \rightarrow 2H_2O$.

- **Importance of Writing Equations:**

  - Provides a concise way to describe chemical changes.
  - Shows quantitative relationships between reactants and products.
  - Ensures conservation of mass and atoms.

- Helps predict the amount of substances needed or produced.

BALANCING CHEMICAL EQUATIONS

- **Law of Conservation of Mass:**
  In a chemical reaction, matter is neither created nor destroyed; the total mass of reactants equals the total mass of products.
  - This law forms the basis of balancing equations.
- **General Steps:**

1. Write down correct symbols and formulae of all reactants and products.
2. Write the unbalanced equation.
3. List the number of each atom on both sides.
4. Adjust the **coefficients** (whole numbers placed before formulae) to balance atoms.
5. Never change subscripts in the formula; only adjust coefficients.
6. Check to ensure all atoms are balanced and coefficients are in the lowest ratio.

- **Worked Examples:**

1. **Combustion of methane:**
   $CH_4 + O_2 \rightarrow CO_2 + H_2O$
   Count atoms: C=1, H=4, O=2 (LHS) vs O=3 (RHS).
   Balanced: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
2. **Reaction of sodium with water:**
   $Na + H_2O \rightarrow NaOH + H_2$
   Balanced: $2Na + 2H_2O \rightarrow 2NaOH + H_2$
3. **Formation of ammonia (Haber process):**
   $N_2 + H_2 \rightarrow NH_3$
   Balanced: $N_2 + 3H_2 \rightarrow 2NH_3$

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4. **Combustion of ethane:**

   $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

   Balanced: $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

- **Tips for Balancing:**
  - Balance metals first, non-metals second, hydrogen next, and oxygen last.
  - If odd numbers occur, try multiplying all coefficients by 2.
  - Always simplify coefficients to the smallest whole numbers.

PRACTICE EQUATION TYPES

Reaction Type	**Example (Balanced Equation)**
Combination	$2Mg + O_2 \rightarrow 2MgO$
Decomposition	$2H_2O_2 \rightarrow 2H_2O + O_2$
Single Displacement	$Zn + 2HCl \rightarrow ZnCl_2 + H_2$
Double Displacement	$Na_2SO_4 + BaCl_2 \rightarrow 2NaCl + BaSO_4$
Combustion	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Evaluation:

1. Define:

   a) Chemical symbol

   b) Chemical formula

   c) Chemical equation

2. State **five reasons** why chemical equations are important.

3. Differentiate between **empirical** and **molecular** formulae with examples.

4. Write chemical formulae for:
    - Ammonium sulphate
    - Aluminium hydroxide
    - Calcium nitrate
5. Balance these equations:
    a) $H_2 + Cl_2 \rightarrow HCl$
    b) $Al + O_2 \rightarrow Al_2O_3$
    c) $Fe + HCl \rightarrow FeCl_2 + H_2$
    d) $NH_3 + O_2 \rightarrow NO + H_2O$

Assignment:

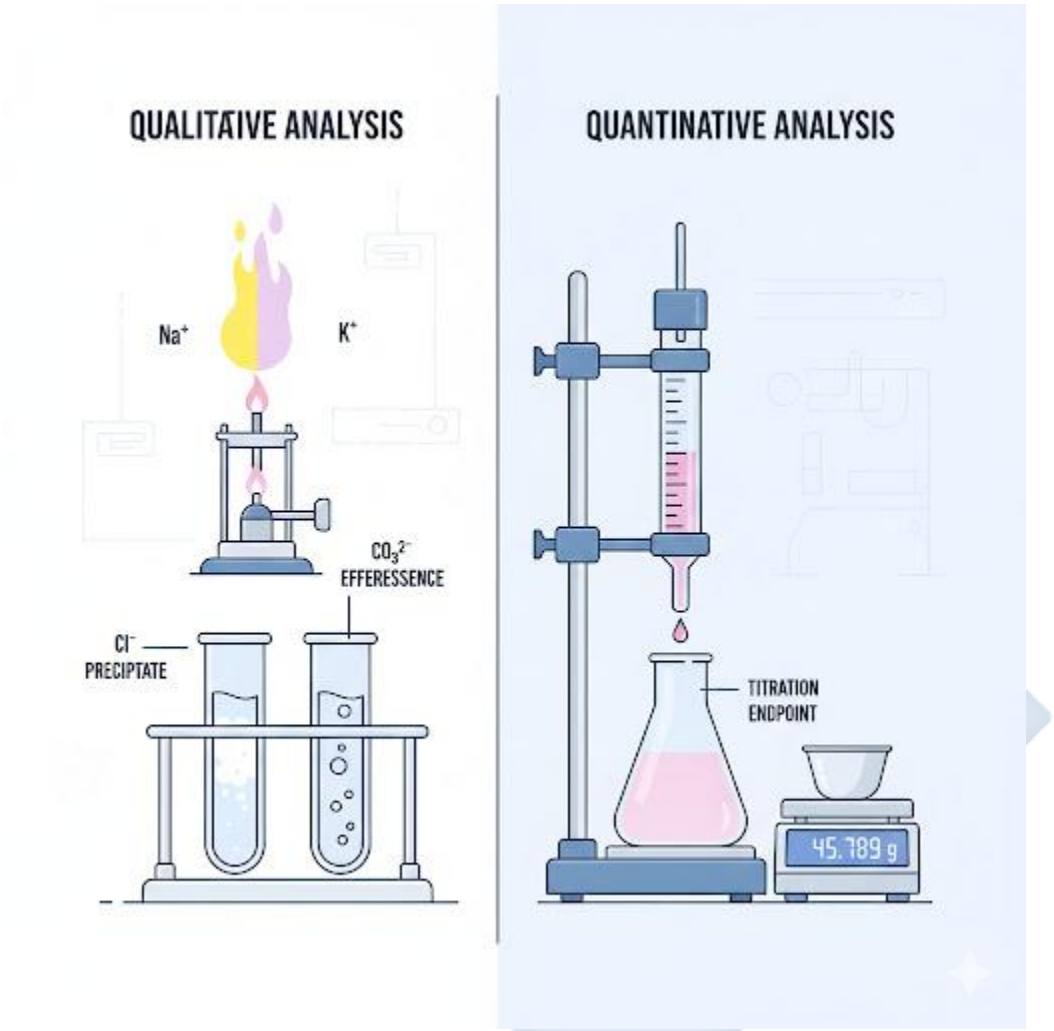
1. Write symbols for **20 elements**, including those with Latin origins.
2. Research and list **10 compounds**, their **chemical names**, and **molecular formulae**.
3. Write and balance equations for:
    a) Reaction between hydrogen sulphide and oxygen to form sulphur dioxide and water.
    b) Combustion of propane gas.
4. Explain, with an example, why changing subscripts in a formula is incorrect.
5. Prepare a **valency chart** of the first 20 elements of the periodic table.

SS1 CHEMISTRY LESSON NOTES

SECOND TERM

1. Quantitative and Qualitative Analysis (Introduction) – Basic concepts.
2. Mole Concept and Stoichiometry – Mole, molar mass, Avogadro's number.
3. Percentage Composition and Empirical Formula – Calculations.
4. Gas Laws – Boyle's law, Charles' law, ideal gas equation.
5. **Midterm Examination**
6. Structure of the Atom and Electronic Configuration – Bohr's model, configurations.
7. **Midterm break**
8. Ionic and Covalent Bonding – Formation, properties.
9. Oxidation and Reduction – Definition, examples, redox reactions.
10. Water and Solution – Properties, types of solutions, solubility.
11. Solubility and Crystallization – Saturated solutions, solubility curves.
12. Examination
13. Closing

WEEK 1: QUANTITATIVE AND QUALITATIVE ANALYSIS (INTRODUCTION) – BASIC CONCEPTS



Chemistry, being an experimental science, depends largely on **observation, measurement, and analysis**. The term **analysis** refers to a set of techniques used to identify substances (what they are) and to determine the amount present (how much is there). This forms the foundation of **analytical chemistry**.

Analytical chemistry is broadly divided into two main branches:

1. **Qualitative Analysis** – concerned with identifying the type or nature of substances.

2. **Quantitative Analysis** – concerned with measuring the exact amount of substances present.

Both are complementary: qualitative analysis first establishes the identity of a substance, and quantitative analysis then determines its amount.

QUALITATIVE ANALYSIS

Definition

Qualitative analysis is the type of analysis used to determine **the identity of elements, ions, functional groups, or compounds present in a sample**. It does not give a numerical value but answers the question: *"What is in the sample?"*

Main Features

- It is **descriptive** rather than numerical.
- It is usually the **first step** in analysis.
- Relies on **observation of changes** (colour change, precipitation, effervescence, odour, etc.).

Methods of Qualitative Analysis

1. **Observation of Physical Properties**
    - Colour (e.g., copper compounds are usually blue or green).
    - State of matter (gas, liquid, solid).
    - Odour (e.g., ammonia has a pungent smell).
    - Solubility in water or organic solvents.
2. **Chemical Tests**
    - **Acid Tests:** Reaction of carbonates with dilute acids producing effervescence due to CO_2.

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- **Alkali Tests:** Reaction of ammonium salts with alkalis releasing ammonia gas.

3. **Flame Tests**
    - Some metals produce characteristic colours in a non-luminous flame due to electronic transitions.
        - Sodium (Na^+): Yellow
        - Potassium (K^+): Lilac
        - Calcium (Ca^{2+}): Brick red
        - Copper (Cu^{2+}): Blue-green
4. **Precipitation Reactions**
    - Certain ions form insoluble salts with specific reagents.
        - Chloride ions (Cl^-) + silver nitrate ($AgNO_3$) → White precipitate of silver chloride.
        - Sulphate ions (SO_4^{2-}) + barium chloride ($BaCl_2$) → White precipitate of barium sulphate.
5. **Confirmatory Tests**
    - Used after preliminary results for greater accuracy.
    - Example: Distinguishing between carbonate and sulphite ions by adding dilute acid and then passing gas through limewater (sulphur dioxide decolourises acidified $KMnO_4$ while CO_2 does not).

QUANTITATIVE ANALYSIS

Definition

Quantitative analysis is the type of analysis used to determine **the exact amount, concentration, or numerical value** of a substance in a given sample. It answers the question: *"How much of it is present?"*

Main Features

- It is **numerical** and involves measurements.
- Requires **calculations** to determine concentration, mass, or volume.
- It is essential in quality control, medicine dosage, and industrial formulations.

Methods of Quantitative Analysis

1. **Volumetric Analysis**
    - Involves accurate measurement of volumes of solutions that react together.
    - Commonly used in **titrations** (acid-base, redox, precipitation).
    - Example: Determining the molarity of hydrochloric acid by titration against standard sodium hydroxide.
2. **Gravimetric Analysis**
    - Involves determination of a substance by **precipitation, drying, and weighing**.
    - Example: Estimating the percentage purity of sodium chloride by precipitating silver chloride using silver nitrate and weighing it.
3. **Instrumental/Modern Methods**
    - More advanced and highly accurate.
    - Examples:
        - Spectrophotometry (measures light absorbed by a solution).
        - Chromatography (separates mixtures into components).
        - Mass spectrometry (gives molecular structure and mass).
        - Atomic absorption spectroscopy (for metal analysis).

DIFFERENCES BETWEEN QUALITATIVE AND QUANTITATIVE ANALYSIS

Qualitative Analysis	Quantitative Analysis
Determines **what** is present in a substance	Determines **how much** is present
Deals with **identity/nature**	Deals with **numerical value/amount**
Involves **observation of colour, smell, flame test, precipitate**	Involves **measurements, weighing, titration, calculations**
Often **simple and preliminary**	Usually **more advanced and precise**
Example: Testing for chloride with $AgNO_3$	Example: Determining concentration of HCl via titration

IMPORTANCE OF CHEMICAL ANALYSIS

Chemical analysis is very important in science, industry, medicine, and everyday life:

1. **Medicine and Pharmacy**
    - Determination of drug composition and purity.
    - Ensures correct dosage of active ingredients.
    - Detects harmful impurities.
2. **Food Industry**
    - Testing for nutritional content (proteins, fats, carbohydrates).
    - Detection of food adulterants (e.g., presence of harmful chemicals like formalin in milk or fish).
    - Ensuring proper concentration of preservatives.
3. **Agriculture**
    - Analysis of soil samples to determine mineral content.
    - Quality check of fertilizers (NPK content).

- Detection of pesticide residues.

4. **Environmental Monitoring**
    - Analysis of air, water, and soil for pollutants.
    - Checking presence of heavy metals, toxic gases, or sewage contamination.
5. **Industrial Applications**
    - Quality control in manufacturing (paints, cement, beverages).
    - Analysis of raw materials and finished products.
    - Ensuring compliance with safety and international standards.

BASIC LABORATORY CONCEPTS IN ANALYSIS

- **Analyte:** The substance being tested or measured.
- **Reagent:** The chemical used to react with the analyte.
- **Indicator:** A dye or compound that shows the end-point of a reaction by changing colour (e.g., phenolphthalein, methyl orange).
- **Standard Solution:** A solution of precisely known concentration used in titration.
- **End Point:** The stage in titration where the indicator changes colour, signalling completion.
- **Equivalence Point:** The theoretical point where exactly equivalent amounts of reactants have reacted.
- **Accuracy:** Closeness of a measured value to the true value.
- **Precision:** The degree of reproducibility of repeated measurements.

EVALUATION

1. Define qualitative and quantitative analysis.
2. Mention and explain four methods of qualitative analysis with examples.
3. List and describe three major methods of quantitative analysis.
4. Differentiate between analyte, reagent, and indicator.
5. State five importance of chemical analysis in everyday life.
6. Why is qualitative analysis usually performed before quantitative analysis?

ASSIGNMENT

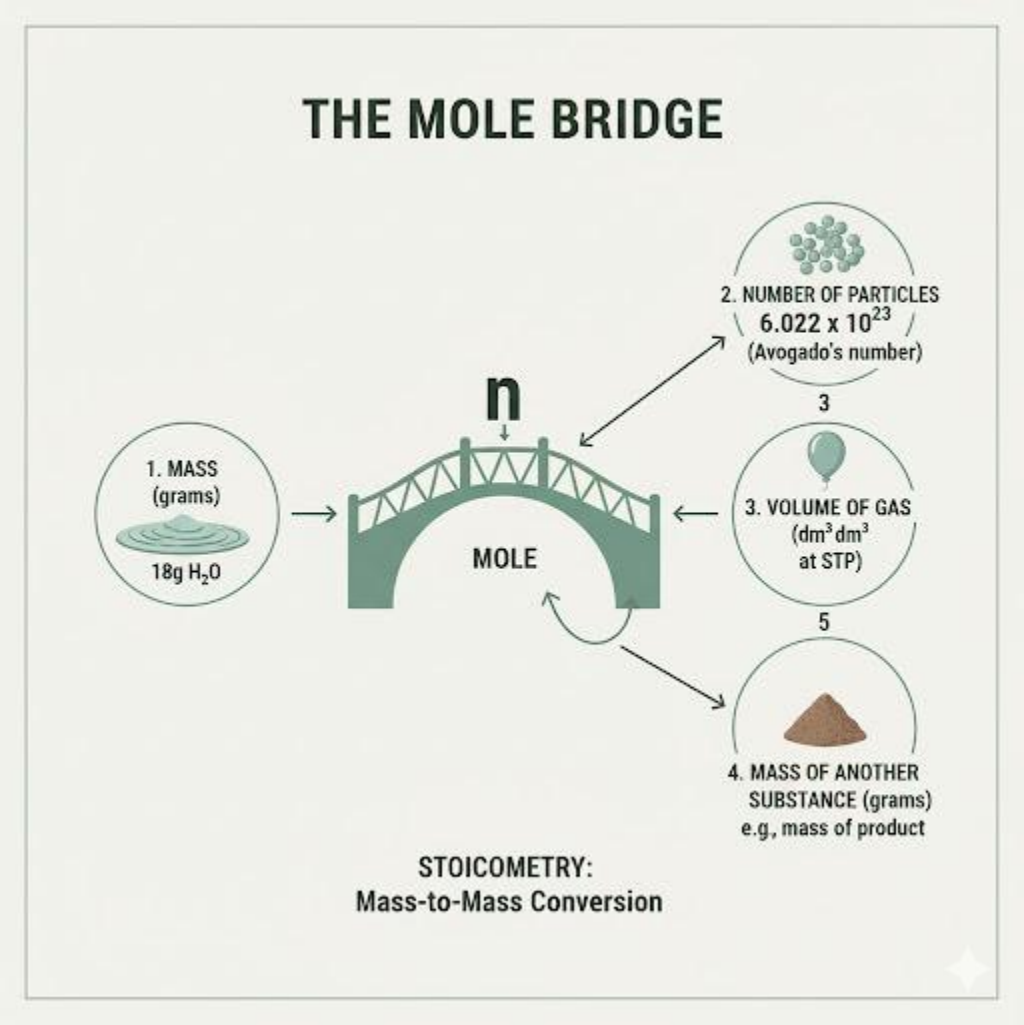
1. Write a comprehensive essay on the differences between qualitative and quantitative analysis, giving **at least five examples** each.
2. Discuss how chemical analysis can be applied in (a) medicine, (b) agriculture, and (c) environmental monitoring.
3. Explain in detail the principle of **volumetric analysis** and describe one common example (acid-base titration).

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WEEK 2: MOLE CONCEPT AND STOICHIOMETRY – MOLE, MOLAR MASS, AVOGADRO'S NUMBER

Lesson Objectives: By the end of this lesson, students should be able to:

1. Define a mole and state Avogadro's number.
2. Calculate molar mass of elements and compounds.
3. Interconvert between mass, moles, number of particles, and volume of gases at STP.
4. Perform stoichiometric calculations involving mass-mass and mass-volume relationships.
5. Identify the limiting reactant in a chemical reaction.



1. Introduction to the Mole Concept

Chemistry deals with tiny particles like atoms and molecules, which are impossible to count individually. We need a special "counting unit," just like we use a **dozen** for 12 items. In chemistry, this unit is the **mole (mol)**.

- **Definition:** A mole is the amount of a substance that contains as many elementary particles (atoms, molecules, ions, etc.) as there are atoms in **exactly 12 grams of Carbon-12**.
- This specific number of particles is known as **Avogadro's Number (N_a)**.

2. Avogadro's Number (N_a)

- **Value: $N_a = 6.022 \times 10^{23}$ particles per mole**
- **Meaning:**
  - 1 mole of Hydrogen **atoms** = 6.022×10^{23} H atoms.
  - 1 mole of Water **molecules** = 6.022×10^{23} H_2O molecules.
  - 1 mole of NaCl **formula units** = 6.022×10^{23} NaCl units.

🔑 **Important Note:** The type of "particle" we count depends on the substance:

- **Elements:** Atoms (e.g., Na, Fe) or molecules (e.g., O_2, H_2, Cl_2).
- **Covalent Compounds:** Molecules (e.g., H_2O, CO_2, NH_3).
- **Ionic Compounds:** Formula Units (e.g., NaCl, KNO_3).

3. Molar Mass

- **Definition:** The **mass of one mole** of a substance, expressed in **grams per mole (g/mol)**.
- It is numerically equal to the **Relative Atomic Mass (A_r)** or **Relative Molecular/Formula Mass (M_r)** from the periodic table.

Examples:

- A_r of Carbon (C) = 12 → Molar Mass of C = **12 g/mol**

- M_r of Water (H_2O) = (2×1) + 16 = 18 → Molar Mass = **18 g/mol**
- M_r of Sodium Chloride (NaCl) = 23 + 35.5 = 58.5 → Molar Mass = **58.5 g/mol**

☞ **Remember:** Always express molar mass in **g/mol**.

4. The Mole Relationship Formulas

To Find:	**Formula**	**Key Information**
Moles (n) from Mass	n = mass (g) / Molar Mass (g/mol)	
Number of Particles	Particles = n × N_a	N_a = 6.022 × 10^{23}
Moles from Particles	n = Particles / N_a	
Moles from Gas Volume at STP	n = Volume (dm^3) / 22.4	**STP:** 0°C (273 K) & 1 atm. 1 mole of **any** gas = 22.4 dm^3
Moles from Solution	n = Molarity (mol/dm^3) × Volume (dm^3)	n = M × V

5. Worked Examples

Example 1: Moles from Mass

How many moles are in 18 g of water?

- M_r of H_2O = 18 g/mol
- n = 18 g / 18 g/mol = 1 mol
  Answer: 1 mole of water.

Example 2: Particles from Mass

How many molecules are there in 54 g of water?

- M_r of H_2O = 18 g/mol
- Moles, n = 54 g / 18 g/mol = 3 mol

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- Number of molecules = 3 mol × 6.022 × 10^{23} molecules/mol = 1.807 × 10^{24} molecules
  Answer: 1.807 × 10^{24} molecules of water.

Example 3: Atoms from Mass

Calculate the number of atoms in 4 g of hydrogen gas (H_2).

- M_r of H_2 = 2 g/mol
- Moles = 4 g / 2 g/mol = 2 mol
- Number of H_2 **molecules** = 2 × 6.022 × 10^{23} = 1.204 × 10^{24}
- Each H_2 molecule has **2 atoms**.
- Total atoms = 1.204 × 10^{24} molecules × 2 atoms/molecule = 2.408 × 10^{24} atoms
  Answer: 2.408 × 10^{24} atoms.

Example 4: Gas Volume at STP

Find the volume occupied by 0.25 mol of oxygen gas at STP.

- Volume = n × 22.4 dm^3/mol
- Volume = 0.25 mol × 22.4 dm^3/mol = 5.6 dm^3
  Answer: 5.6 dm^3.

6. Stoichiometry (Quantitative Chemistry)

- **Definition:** Stoichiometry is the calculation of the quantities (masses, volumes, particles) of reactants and products in a chemical reaction based on the **balanced chemical equation**.
- **The Heart of Stoichiometry: The Mole Ratio**
  - The coefficients in a balanced equation give the **mole ratio** of reactants and products.
- **Steps in Stoichiometric Calculations:**
  1. Write a **balanced chemical equation**.

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2. Extract the **mole ratio** from the coefficients.
3. Convert the given quantity (mass, volume, particles) to **moles (n)**.
4. Use the mole ratio to find the moles of the unknown substance.
5. Convert the moles of the unknown back to the required quantity (mass, volume, particles).

7. Stoichiometry Worked Examples

Example 5 (Mass–Mass Relationship):

How many grams of water are produced when 8 g of hydrogen reacts completely with oxygen?

1. **Equation:** $2H_2 + O_2 \rightarrow 2H_2O$
2. **Mole Ratio:** 2 mol H_2 : 1 mol O_2 : 2 mol H_2O
3. **Moles of given:** Moles of H_2 = 8 g / 2 g/mol = 4 mol
4. **Mole Ratio to Unknown:**
   - 2 mol H_2 produces 2 mol H_2O → Ratio is 1:1.
   - So, **4 mol H_2 produces 4 mol H_2O**.
5. **Convert to Mass:** Mass of H_2O = 4 mol × 18 g/mol = 72 g
   Answer: 72 g of water.

Example 6 (Mass–Volume Relationship):

What volume of CO_2 is produced at STP when 50 g of $CaCO_3$ decomposes?

1. **Equation:** $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
2. **Mole Ratio:** 1 mol $CaCO_3$: 1 mol CO_2
3. **Moles of given:** Moles of $CaCO_3$ = 50 g / 100 g/mol = 0.5 mol
4. **Mole Ratio to Unknown:** Moles of CO_2 = **0.5 mol** (from 1:1 ratio)
5. **Convert to Volume:** Volume of CO_2 = 0.5 mol × 22.4 dm^3/mol = 11.2 dm^3
   Answer: 11.2 dm^3 of CO_2 at STP.

Example 7 (Limiting Reactant):

When 10 g of H_2 reacts with 80 g of O_2, what is the limiting reactant and what mass of water is formed?

1. **Equation:** $2H_2 + O_2 \rightarrow 2H_2O$
2. **Mole Ratio:** 2 mol H_2 : 1 mol O_2
3. **Moles of reactants:**
    - Moles of H_2 = 10 g / 2 g/mol = 5.0 mol
    - Moles of O_2 = 80 g / 32 g/mol = 2.5 mol
4. **Identify Limiting Reactant:**
    - The ratio requires 2 mol H_2 for every 1 mol O_2.
    - With 5.0 mol H_2, we need 5.0 / 2 = 2.5 mol of O_2 to react completely. We have exactly 2.5 mol of O_2.
    - **Therefore, there is no limiting reactant; both are used up completely.**
5. **Calculate Product:** Moles of H_2O produced = 5.0 mol (from 5 mol H_2 and the 1:1 H_2:H_2O ratio).
   Mass of H_2O = 5.0 mol × 18 g/mol = 90 g
   Answer: No limiting reactant. **90 g** of water is formed.

(Note: This is a special case where the reactants are in the exact perfect ratio. Most problems will have one reactant that limits the reaction.)

8. Common Misconceptions & Pitfalls

- **Molar Mass:** Thinking it has no units. It **must** be in **g/mol**.
- **Atoms vs. Molecules:** Confusing O (an atom) with O_2 (a molecule). Specify the particle type.

- **Avogadro's Number:** Forgetting it applies to ions and formula units, not just atoms and molecules.
- **Molar Volume:** Using 22.4 dm^3/mol for non-gases or at conditions that are **not STP**. (At room temperature, it's approximately 24 dm^3/mol).

9. Applications of Mole Concept & Stoichiometry

- **Industry:** Calculating raw materials for products like fertilizers, soap, and cement.
- **Medicine:** Preparing precise drug dosages with the correct mole ratios of active ingredients.
- **Environmental Science:** Modeling CO_2 emissions from combustion processes.
- **Laboratory:** Titrations and gravimetric analysis to determine unknown concentrations.
- **Everyday Life:** Understanding fuel efficiency and water treatment processes.

Evaluation / In-Class Exercises

1. Define a mole and state the value of Avogadro's number.
2. What is the molar mass of Na_2CO_3? (Na=23, C=12, O=16)
3. Calculate the number of moles in: (a) 9 g of water (b) 22 g of CO_2
4. How many molecules are in 18 g of ammonia (NH_3)?
5. Calculate the volume at STP occupied by 11 g of CO_2.
6. If 20 g of calcium reacts to form calcium oxide, calculate the mass of CaO formed. (Ar: Ca = 40, O = 16)
7. Explain three applications of stoichiometry.

Assignment

1. **27 g of aluminium reacts completely with oxygen.**
    - Write the balanced equation.

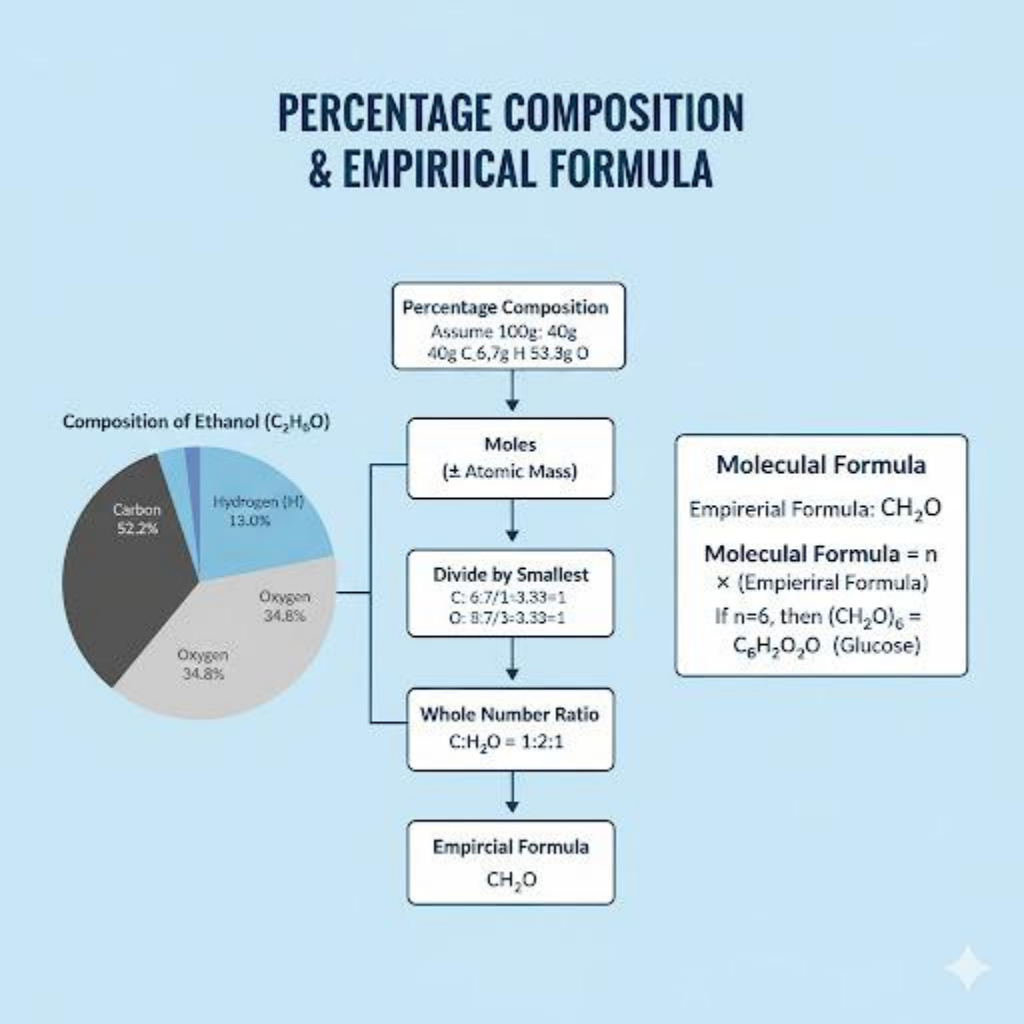
- Calculate: (a) Moles of Al that reacted. (b) Moles of O_2 required. (c) Mass of Al_2O_3 formed. (Ar: Al = 27, O = 16)

2. **A 50 g sample of calcium carbonate was heated until it decomposed.**
   - Calculate: (a) Moles of $CaCO_3$ that decomposed. (b) Volume of CO_2 produced at STP.

3. **Research Assignment:**
   Write a one-page note on Amedeo Avogadro. Explain how his hypothesis contributed to the development of the modern atomic theory and the mole concept.

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WEEK 3: PERCENTAGE COMPOSITION AND EMPIRICAL FORMULA – CALCULATIONS



PERCENTAGE COMPOSITION BY MASS

The **percentage composition of a compound** is the percentage by mass of each element present in the compound relative to the total mass of the compound. It gives us an idea of how much of each element is present in a compound, and is particularly useful in chemical analysis and industrial chemistry.

Mathematical Formula:

Percentage of an element=Total mass of the element in the formulaMolar mass of the compound ×100\text{Percentage of an element} = \frac{\text{Total mass of the element in the formula}}{\text{Molar mass of the compound}} \times 100

- **Molar mass (M_r)** = Sum of relative atomic masses of all the atoms in the formula.
- The numerator = (number of atoms of that element × relative atomic mass).

Worked Example 1

Calculate the percentage composition of carbon, hydrogen, and oxygen in **ethanol (C_2H_6O)**.

Step 1: Find molar mass.

- C = 12, H = 1, O = 16
- M_r = (2 × 12) + (6 × 1) + (16) = 24 + 6 + 16 = **46 g/mol**

Step 2: Find mass contribution of each element.

- Carbon: 2 × 12 = 24
- Hydrogen: 6 × 1 = 6
- Oxygen: 1 × 16 = 16

Step 3: Apply formula.

- %C = (24 ÷ 46) × 100 = **52.2%**
- %H = (6 ÷ 46) × 100 = **13.0%**
- %O = (16 ÷ 46) × 100 = **34.8%**

Final Answer: Ethanol contains 52.2% carbon, 13.0% hydrogen, and 34.8% oxygen.

Worked Example 2

Determine the percentage composition of sodium and chlorine in **NaCl**.

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Step 1: Molar mass = Na (23) + Cl (35.5) = **58.5 g/mol**

Step 2: Contributions: Na = 23, Cl = 35.5

Step 3:

- %Na = (23 ÷ 58.5) × 100 = **39.3%**
- %Cl = (35.5 ÷ 58.5) × 100 = **60.7%**

Worked Example 3

Calculate the percentage composition of Ca, C, and O in **$CaCO_3$**.

Step 1: M_r = Ca (40) + C (12) + (3 × 16) = 40 + 12 + 48 = **100 g/mol**

Step 2: Contributions: Ca = 40, C = 12, O = 48

Step 3:

- %Ca = (40 ÷ 100) × 100 = **40%**
- %C = (12 ÷ 100) × 100 = **12%**
- %O = (48 ÷ 100) × 100 = **48%**

Final Answer: $CaCO_3$ = 40% Ca, 12% C, 48% O.

Empirical Formula

The **empirical formula** of a compound shows the simplest whole-number ratio of atoms of each element in the compound. It does not necessarily show the actual number of atoms, but it reveals the smallest ratio.

STEPS TO DETERMINE EMPIRICAL FORMULA FROM PERCENTAGE COMPOSITION:

1. Assume 100 g of the compound (so percentage = grams).
2. Divide each element's mass by its relative atomic mass (Ar) to find moles.

3. Divide all mole values by the smallest mole obtained.
4. If ratios are not whole numbers, multiply all by a suitable factor (e.g., ×2, ×3) to make them whole.
5. Write the formula using the whole number ratios.

Worked Example 1

A compound contains 40% C, 6.7% H, and 53.3% O. Find its empirical formula.

Step 1: Assume 100 g.

- C = 40 g, H = 6.7 g, O = 53.3 g

Step 2: Convert to moles.

- C = 40 ÷ 12 = 3.33
- H = 6.7 ÷ 1 = 6.7
- O = 53.3 ÷ 16 = 3.33

Step 3: Divide by smallest (3.33).

- C = 3.33 ÷ 3.33 = 1
- H = 6.7 ÷ 3.33 = 2
- O = 3.33 ÷ 3.33 = 1

Step 4: Ratio = 1 : 2 : 1

Final Answer: Empirical formula = **CH_2O**

Worked Example 2

A substance is found to contain 27.3% C and 72.7% O. Find the empirical formula.

Step 1: Assume 100 g.

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- C = 27.3 g, O = 72.7 g

Step 2: Convert to moles.

- C = 27.3 ÷ 12 = 2.28
- O = 72.7 ÷ 16 = 4.54

Step 3: Divide by smallest (2.28).

- C = 2.28 ÷ 2.28 = 1
- O = 4.54 ÷ 2.28 ≈ 2

Final Answer: Empirical formula = **CO_2**

RELATIONSHIP BETWEEN EMPIRICAL AND MOLECULAR FORMULA

The **molecular formula** shows the actual number of atoms of each element in one molecule of the compound.

Relationship:

Molecular formula = (Empirical formula)n

Where:

$$n = \frac{\text{Molar mass of the compound}}{\text{Molar mass of empirical formula}}$$

Worked Example

The empirical formula of a compound is CH_2O. Its molar mass is 180 g/mol. Find its molecular formula.

Step 1: Find M_r of empirical formula CH_2O.
= 12 + (2 × 1) + 16 = **30**

Step 2: Calculate n.

$n = 180 \div 30 = 6$

Step 3: Multiply subscripts by 6.

$= (CH_2O)_6 = \mathbf{C_6H_{12}O_6}$

Final Answer: The molecular formula is **$C_6H_{12}O_6$ (Glucose)**.

Applications of Percentage Composition and Empirical Formula

1. **Pharmaceutical Industry:** Ensures accurate proportion of active ingredients in drugs.
2. **Food Industry:** Used in nutritional analysis (e.g., determining protein or carbohydrate percentage).
3. **Agriculture:** Helps in fertilizer production and analysis (e.g., %N in ammonium nitrate).
4. **Industrial Chemistry:** Ensures correct formulation of alloys, explosives, and fuels.
5. **Forensic Science:** Helps in identifying unknown chemical substances in investigations.
6. **Environmental Chemistry:** Used in pollution analysis and waste treatment.

Common Pitfalls in Solving Problems

- Forgetting to convert percentage to grams (by assuming 100 g).
- Stopping at fractional ratios instead of multiplying to obtain whole numbers.
- Mixing up empirical and molecular formulas.
- Using the wrong atomic masses.

EVALUATION

1. Define:
   (a) Percentage composition,

(b) Empirical formula,
(c) Molecular formula.

2. Distinguish between empirical and molecular formulas with one example each.
3. Calculate the percentage composition of nitrogen in NH_3.
4. A compound contains 36% carbon, 12% hydrogen, and 52% oxygen. Determine its empirical formula.
5. The empirical formula of a compound is CH. If its molar mass is 78 g/mol, determine its molecular formula.
6. State three real-life applications of empirical and molecular formulas.

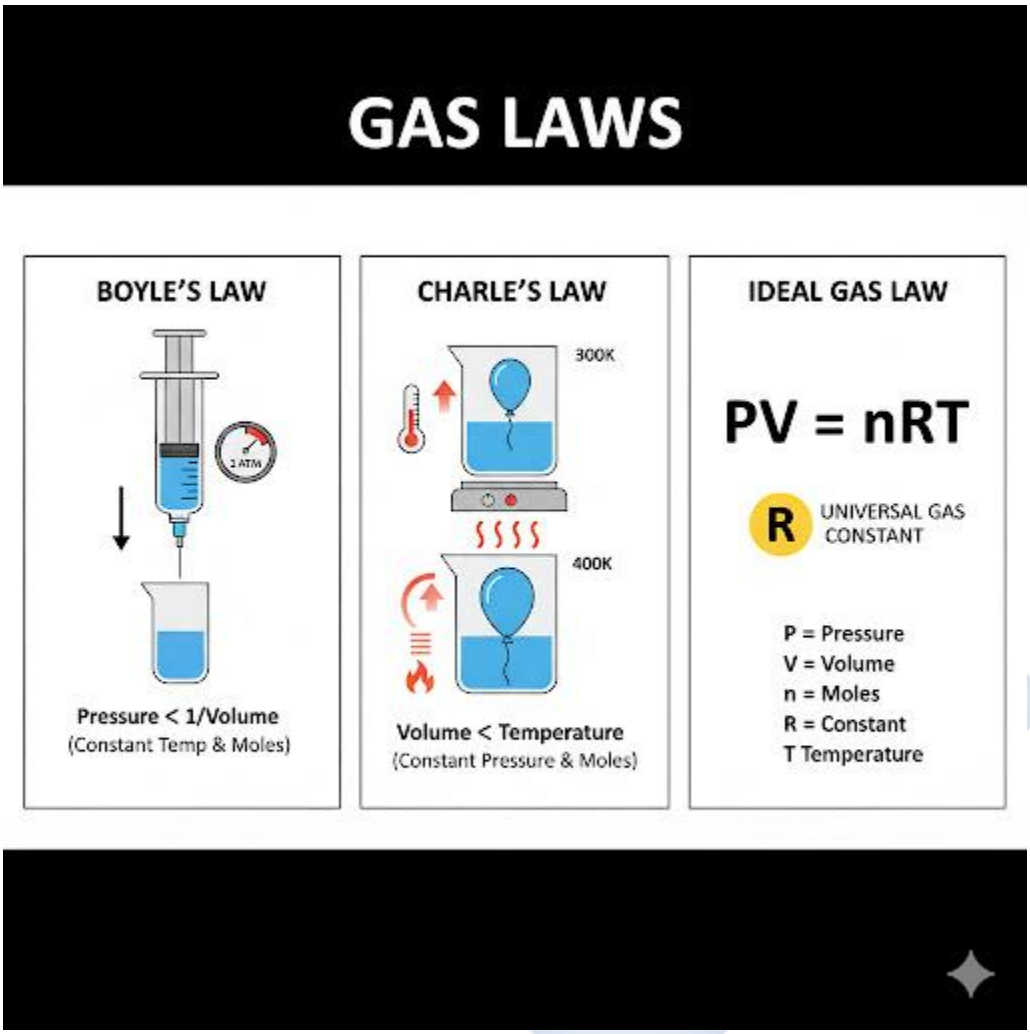
ASSIGNMENT

1. Calculate the percentage composition of hydrogen and oxygen in H_2O.
2. An organic compound contains 54.5% carbon, 9.1% hydrogen, and 36.4% oxygen. Find its empirical formula.
3. A hydrocarbon contains 85.7% carbon and 14.3% hydrogen. If its molar mass is 56 g/mol, determine:
(a) its empirical formula,
(b) its molecular formula.

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WEEK 4: GAS LAWS

(Boyle's Law, Charles' Law, and the Ideal Gas Equation)



INTRODUCTION TO GAS LAWS

Gases are one of the fundamental states of matter, consisting of particles (atoms or molecules) in **constant random motion**. Unlike solids and liquids, gases have **no definite shape or volume**, and their particles are widely spaced. Because gases can be easily compressed or expanded, scientists have studied how gases behave under various conditions, leading to the discovery of the **gas laws**.

These gas laws describe the mathematical relationships between:

- **Pressure (P):** The force exerted per unit area by gas molecules when they collide with the container walls.

  $P = \frac{\text{Force}}{\text{Area}}$ Unit: Pascal (Pa), where 1 Pa = 1 N/m².

- **Volume (V):** The amount of space occupied by the gas, measured in **litres (L)** or **cubic metres (m³)**.

- **Temperature (T):** A measure of the average kinetic energy of gas molecules, measured in **Kelvin (K)**. Always convert from Celsius:

  T(K) = T(°C) + 273

- **Number of moles (n):** Quantity of gas in terms of the number of particles, where 1 mole = 6.022×10^{23} molecules.

These measurable properties are **interdependent**: changing one affects the others. Boyle's Law, Charles' Law, and the Ideal Gas Equation summarize these relationships.

Boyle's Law

- **Statement:**
  At constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure.

Mathematically:

$P \propto \frac{1}{V}$ or $PV = k$

where k is a constant (for a fixed amount of gas at a constant temperature).

- **Equation for Calculations:**

$P_1V_1 = P_2V_2$

where:

  - P_1V_1 = initial pressure and volume

- P_2V_2 = final pressure and volume

- **Graphical Representations:**
  - **P vs V:** A hyperbolic curve; as pressure increases, volume decreases.
  - **P vs 1/V:** A straight line through the origin.

(Diagrams should be included when teaching: Label axes and indicate units clearly.)

- **Molecular Explanation:**
  - When a gas is compressed, the distance between molecules decreases.
  - Molecules collide more frequently with container walls, increasing pressure.
  - If volume doubles, pressure halves (and vice versa) as long as temperature remains constant.

- **Worked Example 1:**
  A gas occupies **4.0 L** at a pressure of **200 kPa**. Calculate its new volume if the pressure is increased to **400 kPa**, assuming temperature is constant.

  $P_1V_1 = P_2V_2$

  $200 \times 4.0 = 400 \times V_2$

  $800 = 400V_2 \Rightarrow V_2 = 2.0L$

Charles' Law

- **Statement:**
  At constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature.

Mathematically:

$$V \propto T \text{ or } \frac{V}{T} = k$$

- **Equation for Calculations:**

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

(Temperatures must always be in Kelvin.)

- **Graphical Representation:**

  A graph of **V vs T(K)** is a **straight line through the origin**.

  - If extended backward, the line meets the temperature axis at **–273°C (0K)**, the theoretical temperature where molecular motion stops.

- **Molecular Explanation:**

  - Increasing temperature increases molecular kinetic energy.
  - Molecules move faster, colliding harder and more often with the walls.
  - To keep pressure constant, volume must increase.

- **Worked Example 2:**

  A gas has a volume of **300 cm³** at **27°C (300K)**. What is its volume at **127°C (400K)** if pressure remains constant?

$$\frac{V_2}{T_2} = \frac{V_2}{T_2}$$

$$\frac{300}{300} = \frac{V_2}{400} \Rightarrow V_2 = 400 \text{ cm}^3$$

4. Ideal Gas Equation (General Gas Law)

By combining Boyle's and Charles' Laws, we get:

$$PV \propto T \Rightarrow PV = nRT$$

Where:

- P = pressure (Pa)
- V = volume (m^3)
- n = number of moles
- R = universal gas constant 8.314 $Jmol^{-1}K^{-1}$
- T = temperature (K)

- **Alternative Form for Comparing Two States of the Same Gas:**

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

- **Molecular Explanation of the Ideal Gas Model:**
  - Molecules are in **constant random motion**.
  - Collisions are **perfectly elastic** (no loss of energy).
  - Molecular volume is negligible compared to the container.
  - There are no intermolecular forces.
  - Kinetic energy is proportional to absolute temperature.
  - Real gases behave ideally at **low pressures and high temperatures**.

- **Worked Example 3:**
  Calculate the pressure exerted by **0.5 mol** of gas in a **10 L** container at **300 K**.
  Convert volume: 10L=0.010m310 L = 0.010 m^3.

Using PV = nRT:

$P = \frac{nRT}{V} = \frac{0.5 \times 8.314 \times 300}{0.010} = 124{,}710$ Pa (≈125 kPa)

Comparison of Gas Laws

Law	Constant Variable(s)	Mathematical Form	Relationship	Graph
Boyle's Law	Temperature, n	$P \propto {}^{1}/_{V}$ or PV = k	Inverse	Hyperbola
Charles' Law	Pressure, n	$V \propto T$ or ${}^{V}/_{T} = k$	Direct	Straight line
Ideal Gas Law	n constant	PV = nRT	Combines Boyle's & Charles'	3D relationship

Applications of Gas Laws

1. **Boyle's Law:**
    - Scuba diving: Explains why divers must ascend slowly to prevent nitrogen bubbles in the blood (decompression sickness).
    - Syringes and pumps: Sucking liquid in by reducing pressure.
    - Compressors: Increasing gas pressure by reducing volume.
2. **Charles' Law:**
    - Hot air balloons: Heating air increases volume, making the balloon rise.
    - Automotive engines: Explains expansion of gases during combustion.
3. **Ideal Gas Law:**
    - Calculating molar masses of gases.

- Predicting chemical reaction yields.
- Designing pressurized gas storage systems.

Evaluation:

1. Define **Boyle's Law**, **Charles' Law**, and **Ideal Gas Law**.
2. Write the mathematical expressions for each law.
3. Explain why temperature must be converted to Kelvin when applying gas laws.
4. A gas occupies **2.5 L** at a pressure of **100 kPa**. Calculate its new volume if the pressure is increased to **250 kPa**, temperature constant.
5. A balloon has a volume of **1.2 L** at **20°C**. What will be its volume at **80°C** (constant pressure)?
6. Calculate the pressure exerted by **2 moles** of gas in a **5 L** container at **350 K**.
7. Sketch **Boyle's Law** and **Charles' Law** graphs.

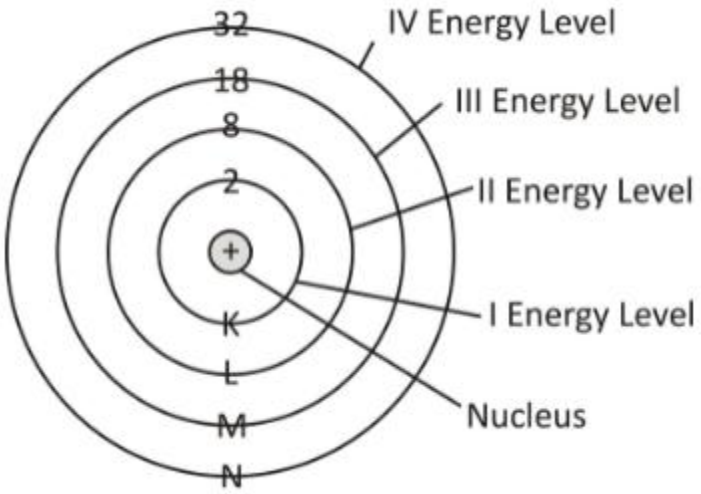
Assignment:

1. Derive the **Ideal Gas Equation** from Boyle's and Charles' Laws.
2. A cylinder contains nitrogen gas at **2.0×10^5 Pa** and **300 K**. If the cylinder's volume is **0.05 m^3**, calculate:
   - The number of moles of nitrogen gas.
   - The mass of nitrogen (N_2) in the cylinder.
3. Explain five key assumptions of the kinetic theory of gases.
4. Describe three ways gas laws are applied in everyday life.
5. Solve: A gas has a pressure of **150 kPa** and a volume of **10 L** at **300 K**. What is its pressure at **400 K** if volume remains constant?

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WEEK 6: STRUCTURE OF THE ATOM AND ELECTRONIC CONFIGURATION

Bohr's Model of the Atom; Energy Levels; Rules of Electronic Configuration; Applications in the Periodic Table



Historical Development Leading to Bohr's Model

Understanding Bohr's theory requires knowledge of earlier attempts:

- **Dalton's Atomic Theory (1803):**
  Proposed that atoms are indivisible, solid particles. Limitation → did not explain internal structure.
- **Thomson's Model (1897):**
  Plum pudding model: atom as a positively charged sphere with electrons embedded in it. Limitation → could not explain Rutherford's experiment.
- **Rutherford's Nuclear Model (1911):**
  Gold foil experiment: showed that atoms have a dense, positively charged nucleus with electrons orbiting around it. Limitation → could not explain why electrons did not spiral into the nucleus or why atoms emitted discrete spectra.
- **Bohr's Model (1913):**
  Improved Rutherford's model by introducing **quantized energy levels** for electrons.

Bohr's Postulates of the Atom

Bohr made the following assumptions:

1. **Electrons revolve around the nucleus in fixed paths (orbits/shells) without losing energy.**
    - These orbits are stable and are called **stationary states**.
2. **Each orbit has a fixed energy value.**
    - Lower orbits = lower energy; higher orbits = higher energy.
3. **Electrons can jump between orbits:**
    - When an electron absorbs energy, it jumps to a higher orbit (excited state).
    - When it falls back, it emits energy as light (photon).
    - This explains **atomic spectra**.
4. **Energy Difference Formula:**
   The energy emitted or absorbed when an electron changes orbit is given by:

   $\Delta E = E2 - E1 = h\nu$

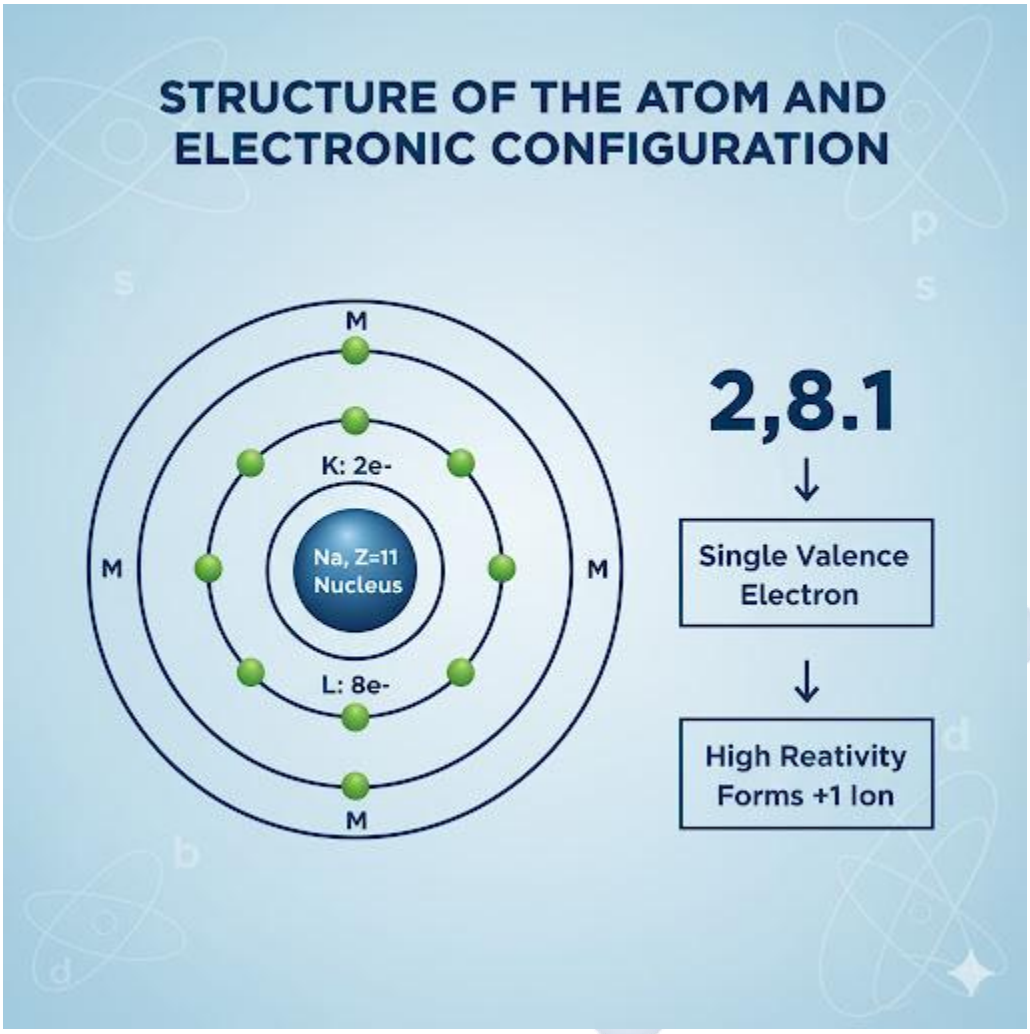
where h = Planck's constant (6.626×10^{-34} Js) and ν = frequency of radiation.

Energy Levels and Shells

- The nucleus is surrounded by concentric shells denoted as:
    - **K (n=1), L (n=2), M (n=3), N (n=4), O (n=5), P (n=6), Q (n=7).**
- Each shell can hold a maximum of **$2n^2$ electrons**, where n = principal quantum number.
    - K (n=1): $2(1)^2 = 2$
    - L (n=2): $2(2)^2 = 8$

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

- **Example:** Calcium (Z = 20) has 20 electrons → distribution is 2, 8, 8, 2.



Electronic Configuration of Atoms

Electronic configuration describes how electrons are arranged in shells and subshells.

- **Subshells:** Each energy level is divided into subshells (s, p, d, f).
  - s → 2 electrons
  - p → 6 electrons

- d → 10 electrons
- f → 14 electrons

- **Aufbau Principle:** Electrons occupy the lowest energy orbitals first before filling higher ones.
- **Pauli's Exclusion Principle:** An orbital can hold a maximum of two electrons, and they must spin in opposite directions.
- **Hund's Rule:** Electrons fill orbitals singly before pairing up in degenerate orbitals (orbitals of equal energy).
- **Order of Filling Orbitals (Aufbau order):**
  1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f → 5d → 6p → 7s → 5f → 6d → 7p

Examples of Electronic Configuration

- **Hydrogen (Z = 1):** $1s^1$ → Shell notation: 1
- **Carbon (Z = 6):** $1s^2\ 2s^2\ 2p^2$ → Shell notation: 2, 4
- **Sodium (Z = 11):** $1s^2\ 2s^2\ 2p^6\ 3s^1$ → Shell notation: 2, 8, 1
- **Chlorine (Z = 17):** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ → Shell notation: 2, 8, 7
- **Potassium (Z = 19):** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$ → Shell notation: 2, 8, 8, 1
- **Iron (Z = 26):** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

Importance of Electronic Configuration

- Explains the **valency** of elements:
  - Sodium (2,8,1) → valency = 1
  - Oxygen (2,6) → valency = 2
- Explains **chemical properties**:

- Elements with similar outer electron configurations belong to the same group.

- Explains **periodicity**: Position of elements in the periodic table is based on electronic configuration.
- Explains **stability of noble gases**: Neon (2,8) and Argon (2,8,8) are inert because their outer shells are completely filled.

Evaluation

1. State four differences between Rutherford's atomic model and Bohr's atomic model.
2. Using the formula $2n^2$, calculate the maximum number of electrons that can be in the M and N shells.
3. Draw the electronic configuration of:
   a. Oxygen (Z = 8)
   b. Magnesium (Z = 12)
   c. Potassium (Z = 19)
4. Write the full (subshell) electronic configuration of Aluminium (Z = 13).
5. Explain how Bohr's theory accounts for the line spectrum of hydrogen.

Assignment

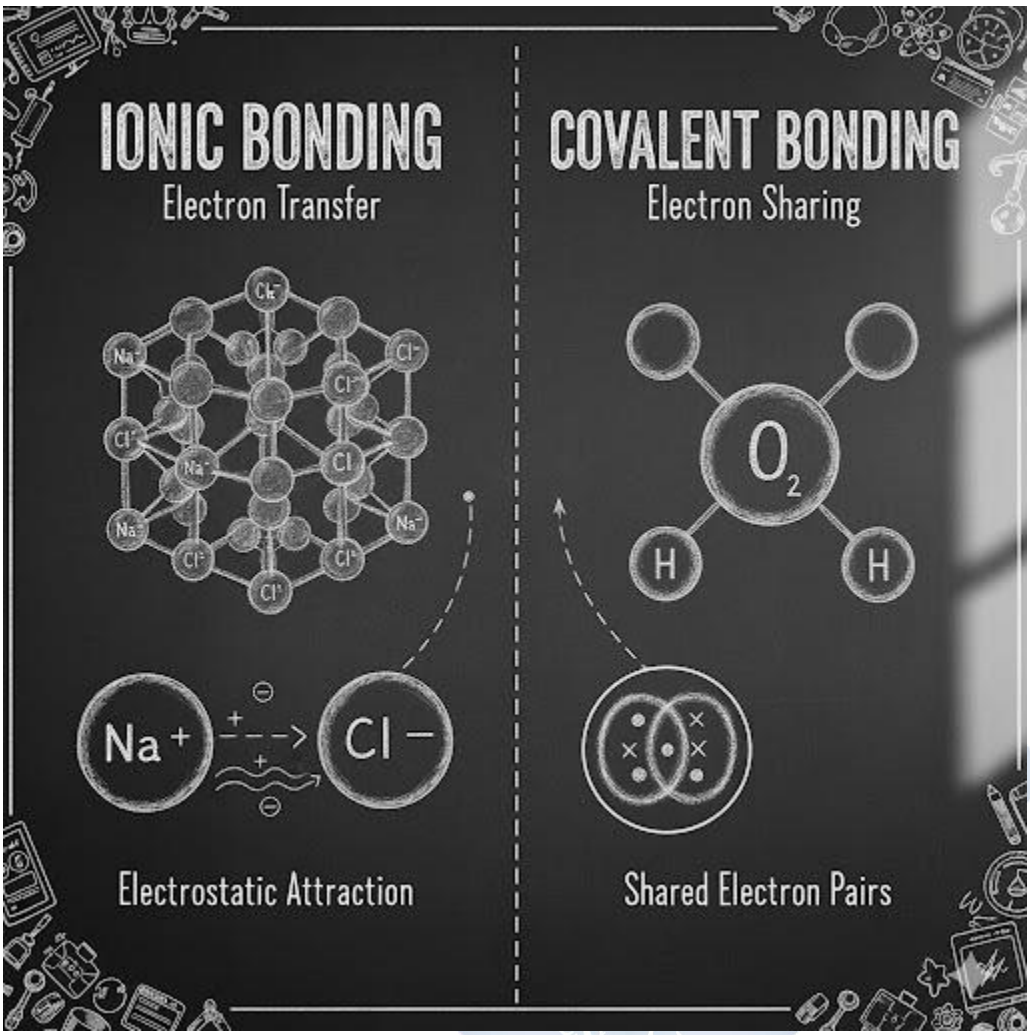
1. Write the **subshell** electronic configuration for:
   a. Sulphur (Z = 16)
   b. Phosphorus (Z = 15)
   c. Calcium (Z = 20)
   d. Copper (Z = 29)
   e. Zinc (Z = 30)
2. Explain, with examples, the relationship between electronic configuration and:
   - Position in the periodic table.

- Valency of elements.

3. Why are elements in **Group 1 (alkali metals)** very reactive while those in **Group 18 (noble gases)** are inert? Use electronic configuration to justify your answer.

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WEEK 8: IONIC AND COVALENT BONDING – FORMATION AND PROPERTIES



All elements, except the noble gases, are chemically unstable in their free atomic state because their outermost energy levels are not completely filled. In order to achieve stability, atoms combine with one another. This combination is called **chemical bonding**.

The stability attained by atoms after bonding is usually explained using the **octet rule**. According to the octet rule:

- Atoms tend to gain, lose, or share electrons until they have **eight electrons in their outermost shell**, just like the noble gases.
- Hydrogen and helium are exceptions because they only need **two electrons** to achieve stability (duplet rule).

There are different types of chemical bonds, but the most common and important are **ionic (electrovalent) bonding** and **covalent bonding**.

IONIC (ELECTROVALENT) BONDING

Definition

Ionic bonding is the type of chemical bonding formed when atoms **transfer electrons** from one atom (usually a metal) to another atom (usually a non-metal). The atom that loses electrons becomes a **positively charged ion (cation)**, while the atom that gains electrons becomes a **negatively charged ion (anion)**. The strong **electrostatic force of attraction** between the oppositely charged ions is called an **ionic bond**.

Steps in Ionic Bond Formation

1. One atom loses one or more electrons to attain a stable noble gas configuration.
2. Another atom gains these electrons to also attain stability.
3. Oppositely charged ions are formed.
4. A strong electrostatic attraction holds them together.

Conditions for Ionic Bond Formation

- Large difference in **electronegativity** (greater than 1.7) between the two atoms.
- Usually occurs between **Group 1–3 metals** and **Group 5–7 non-metals**.
- Metals have low ionization energy (lose electrons easily).
- Non-metals have high electron affinity (gain electrons easily).

Examples of Ionic Bond Formation

1. **Sodium Chloride (NaCl):**
    - Sodium (atomic number 11): 2,8,1 → loses 1 electron → Na^+.
    - Chlorine (atomic number 17): 2,8,7 → gains 1 electron → Cl^-.

- Result: Na^+ and Cl^- ions are held together by electrostatic attraction in a giant ionic lattice.

$Na \rightarrow Na^+ + e^- Na \rightarrow Na^+ + e^-$

$Cl + e^- \rightarrow Cl^- Cl + e^- \rightarrow Cl^-$

$Na^+ + Cl^- \rightarrow NaClNa^+ + Cl^- \rightarrow NaCl$

2. **Magnesium Oxide (MgO):**
    - Mg (atomic number 12): 2,8,2 → loses 2 electrons → Mg^{2+}.
    - O (atomic number 8): 2,6 → gains 2 electrons → O^{2-}.
    - Result: MgO is formed with strong ionic bonds.
3. **Calcium Fluoride (CaF_2):**
    - Ca (2,8,8,2) → loses 2 electrons → Ca^{2+}.
    - Each F atom (2,7) → gains 1 electron → $2F^-$.
    - Result: CaF_2 crystal structure.

Properties of Ionic Compounds

- **State:** Crystalline solids at room temperature due to strong lattice.
- **Melting and boiling points:** Very high because of strong ionic bonds.
- **Solubility:** Generally soluble in polar solvents like water but insoluble in non-polar solvents (benzene, petrol).
- **Electrical conductivity:**
    - Do not conduct electricity in solid state (ions fixed in lattice).
    - Conduct electricity in molten or aqueous state (ions are mobile).
- **Hardness:** Hard but brittle. Ionic crystals shatter when stress is applied because like charges are forced close together, causing repulsion.

COVALENT BONDING

Definition

Covalent bonding is the type of chemical bonding formed when **two atoms share one or more pairs of electrons** in order to attain stability (noble gas configuration).

Conditions for Covalent Bond Formation

- Occurs mainly between **non-metal atoms**.
- Difference in electronegativity is small (less than 1.7).
- Atoms need to complete their octets by **sharing electrons** instead of losing or gaining them.

Types of Covalent Bonds

1. **Single Covalent Bond:** Formed when one pair of electrons is shared.
    - Example: Hydrogen molecule (H_2), H–H.
    - Each hydrogen contributes 1 electron to the shared pair.
2. **Double Covalent Bond:** Formed when two pairs of electrons are shared.
    - Example: Oxygen molecule (O_2), O=O.
    - Each oxygen atom shares 2 electrons.
3. **Triple Covalent Bond:** Formed when three pairs of electrons are shared.
    - Example: Nitrogen molecule (N_2), N≡N.
    - Each nitrogen atom shares 3 electrons.
4. **Coordinate (Dative) Covalent Bond:** A special type where one atom donates both electrons for the shared pair.
    - Example: Formation of ammonium ion (NH_4^+) from ammonia and hydrogen ion.

Examples of Covalent Compounds

- **Water (H_2O):** O (2,6) shares 2 electrons with 2 H atoms.
- **Methane (CH_4):** C (2,4) shares 4 electrons with 4 H atoms.
- **Carbon dioxide (CO_2):** C (2,4) shares 2 electrons with each O atom, forming two double bonds.

Properties of Covalent Compounds

- **State:** Usually gases (O_2, CO_2), liquids (H_2O), or soft solids (sulphur, iodine).
- **Melting and boiling points:** Generally low because molecules are held together by weak intermolecular forces. Exception: diamond and graphite (giant covalent structures).
- **Solubility:** Mostly insoluble in water but soluble in organic solvents.
- **Electrical conductivity:** Poor conductors as they lack free ions or electrons. Exception: graphite (due to delocalized electrons).
- **Bond directionality:** Bonds are directional (exist between specific atoms), unlike ionic bonds which are non-directional.

COMPARISON BETWEEN IONIC AND COVALENT BONDING

Feature	**Ionic Bonding**	**Covalent Bonding**
Mode of formation	Transfer of electrons (metal to non-metal)	Sharing of electrons (non-metal with non-metal)
Particles formed	Ions (cations and anions)	Molecules
Bond strength	Strong (electrostatic attraction)	Strong within molecules but weak intermolecular forces
Melting/Boiling points	Very high	Generally low (except giant covalent structures)
Electrical	Conduct in molten/aqueous	Do not conduct (except graphite)

conductivity	state	
Solubility	Soluble in water, insoluble in organics	Insoluble in water, soluble in organic solvents
Structure	Giant ionic lattice	Simple molecules or giant covalent structures

IMPORTANCE OF UNDERSTANDING CHEMICAL BONDING

- **Explains stability of matter:** Why atoms combine and how they achieve noble gas configuration.
- **Predicts properties of compounds:** Whether a compound will be soluble, hard, volatile, or conductive.
- **Industrial relevance:** Used in material science (semiconductors, ceramics, polymers).
- **Biological relevance:** Covalent bonds are fundamental in biomolecules (proteins, carbohydrates, DNA).
- **Everyday life:** Explains why salt dissolves in water, why water is liquid, and why metals conduct electricity.

EVALUATION

1. Define chemical bonding. Why do atoms form bonds?
2. Describe, with the aid of electron configurations, how magnesium oxide (MgO) is formed.
3. Differentiate between ionic bonding and covalent bonding using five points.
4. Explain the difference between a coordinate bond and an ordinary covalent bond. Give one example.
5. Why do ionic compounds have higher melting points than covalent compounds?

6. Explain why graphite conducts electricity while diamond does not, although both are covalently bonded forms of carbon.

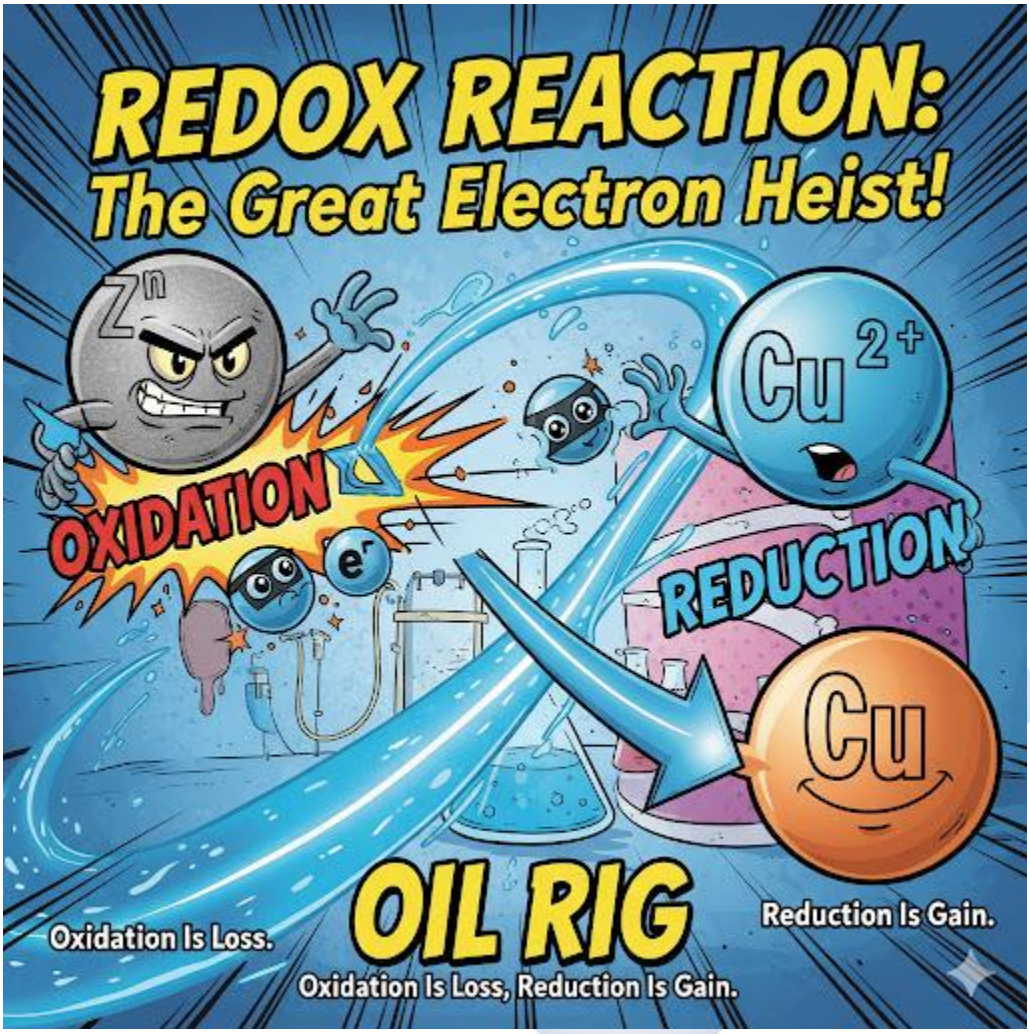
ASSIGNMENT

1. Draw and explain the formation of the following:
   (a) Calcium chloride ($CaCl_2$)
   (b) Carbon dioxide (CO_2)
   (c) Ammonium ion (NH_4^+)
2. Construct a detailed table of **seven differences** between ionic and covalent compounds, with two examples each.
3. Write a detailed note explaining the properties of ionic and covalent compounds and relate these properties to their structures.

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WEEK 9: OXIDATION AND REDUCTION

Definitions, Examples, and Redox Reactions



Introduction

Every chemical reaction involves the rearrangement of atoms. In many reactions, electrons are transferred from one species to another. Such reactions are called **oxidation-reduction (redox) reactions**.

- Redox reactions play a vital role in **biological systems** (respiration, photosynthesis), **industrial processes** (extraction of metals, production of energy in batteries), and **environmental processes** (corrosion, water purification).
- To understand redox processes, we use **three main approaches**:

- **Classical (oxygen/hydrogen transfer)**
- **Electron transfer**
- **Oxidation number method**

Definitions of Oxidation and Reduction

A. Classical Definitions

- **Oxidation**: Addition of oxygen or removal of hydrogen.
- **Reduction**: Removal of oxygen or addition of hydrogen.

Examples:

1. $2Mg + O_2 \rightarrow 2MgO_2$
   - Magnesium oxidized (oxygen added).
   - Oxygen reduced (electrons accepted).
2. $CuO + H_2 \rightarrow Cu + H_2O$
   - CuO reduced (oxygen removed).
   - H_2 oxidized (oxygen added).

B. Electron Transfer Definitions

- **Oxidation**: Loss of electrons. (OIL – Oxidation Is Loss).
- **Reduction**: Gain of electrons. (RIG – Reduction Is Gain).

Examples:

- $Na \rightarrow Na^+ + e^- \rightarrow$ Na is oxidized.
- $Cl_2 + 2e^- \rightarrow 2Cl^- \rightarrow Cl_2$ is reduced.

C. Oxidation Number Definitions

- **Oxidation**: Increase in oxidation number.
- **Reduction**: Decrease in oxidation number.

Examples:

1. $Fe^{2+} \rightarrow Fe^{3+} + e^-$
    - Fe^{2+} is oxidized (oxidation number increases from +2 to +3).
2. $Cl_2 + 2e^- \rightarrow 2Cl^-$
    - Cl_2 is reduced (oxidation number decreases from 0 to –1).

Redox Reactions

A **redox reaction** occurs when oxidation and reduction take place **simultaneously**.

- The substance **oxidized** is the **reducing agent** (it donates electrons).
- The substance **reduced** is the **oxidizing agent** (it accepts electrons).

Example 1:

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

- Zn oxidized ($Zn \rightarrow Zn^{2+} + 2e^-$).
- Cu^{2+} reduced ($Cu^{2+} + 2e^- \rightarrow Cu$).
- Zn = reducing agent, Cu^{2+} = oxidizing agent.

Example 2:

$2Na + Cl_2 \rightarrow 2NaCl$

- Na oxidized ($Na \rightarrow Na^+ + e^-$).
- Cl_2 reduced ($Cl_2 + 2e^- \rightarrow 2Cl^-$).

Oxidizing and Reducing Agents

- **Oxidizing Agent (Oxidant):** Accepts electrons, gets reduced.
    - Examples: O_2, Cl_2, $KMnO_4$, $K_2Cr_2O_7$, HNO_3.
- **Reducing Agent (Reductant):** Donates electrons, gets oxidized.
    - Examples: H_2, CO, Na, Zn, Fe.

Examples of Redox Reactions

1. **Combination Reactions:**
   $4Na + O_2 \rightarrow 2Na_2O$
2. **Displacement Reactions:**
   $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
3. **Combustion Reactions:**
   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
    - Carbon oxidized from –4 to +4.
    - Oxygen reduced from 0 to –2.
4. **Corrosion (Rusting):**
   $4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3$

Balancing Redox Reactions

There are two main methods:

(i) Oxidation Number Method

Steps:

1. Assign oxidation numbers to all elements.
2. Identify oxidation and reduction changes.

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3. Balance the number of electrons lost and gained.
4. Balance atoms and charges.

Example: Balance $Fe^{2+} + MnO_4^- \rightarrow Fe^{3+} + Mn^{2+}$ in acidic solution.

- Oxidation: $Fe^{2+} \rightarrow Fe^{3+} + e^-$.
- Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$.
- Multiply Fe reaction by 5: $5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$.
- Combine:

$$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$$

(ii) Ion–Electron Method

- Write oxidation and reduction half-reactions separately.
- Balance atoms (except H and O).
- Balance O using H_2O.
- Balance H using H^+ (or OH^- in alkaline medium).
- Balance charges using electrons.
- Add half-reactions, canceling electrons.

Applications of Redox Reactions

1. **Biological Systems**
    - **Respiration:** Glucose oxidized to CO_2 and H_2O, releasing energy.
    - **Photosynthesis:** CO_2 reduced to glucose using sunlight.
2. **Industrial Applications**
    - Extraction of metals (e.g., Fe from Fe_2O_3 by reduction with CO).
    - Electroplating and electrolysis.

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- Production of bleach and disinfectants.

3. **Environmental Applications**
   - Water treatment (Cl_2 as oxidizing agent).
   - Removal of pollutants.
4. **Everyday Life**
   - Rusting of iron (undesirable redox).
   - Batteries (dry cell, lead–acid accumulator, lithium-ion cell).

Common Misconceptions

- Oxidation is not only "addition of oxygen"; it may involve **electron loss** or **increase in oxidation number**.
- Reduction is not only "removal of oxygen"; it may involve **electron gain** or **decrease in oxidation number**.
- Oxidation and reduction **always occur together**; one cannot happen without the other.
- The **oxidizing agent is reduced** and the **reducing agent is oxidized**.

Evaluation

1. Define oxidation and reduction according to:
   - (a) Classical definition
   - (b) Electron transfer definition
   - (c) Oxidation number definition
2. Identify the oxidizing and reducing agents in:
   - (a) $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
   - (b) $2Mg + O_2 \rightarrow 2MgO$
3. State two differences between an oxidizing agent and a reducing agent.
4. Balance the following redox reactions:
   - (a) $Cr_2O_7^{2-} + SO_2 \rightarrow Cr^{3+} + SO_4^{2-}$ (acidic solution).
   - (b) $NO_3^- + Cu \rightarrow Cu^{2+} + NO$.
5. List four applications of redox reactions in everyday life.

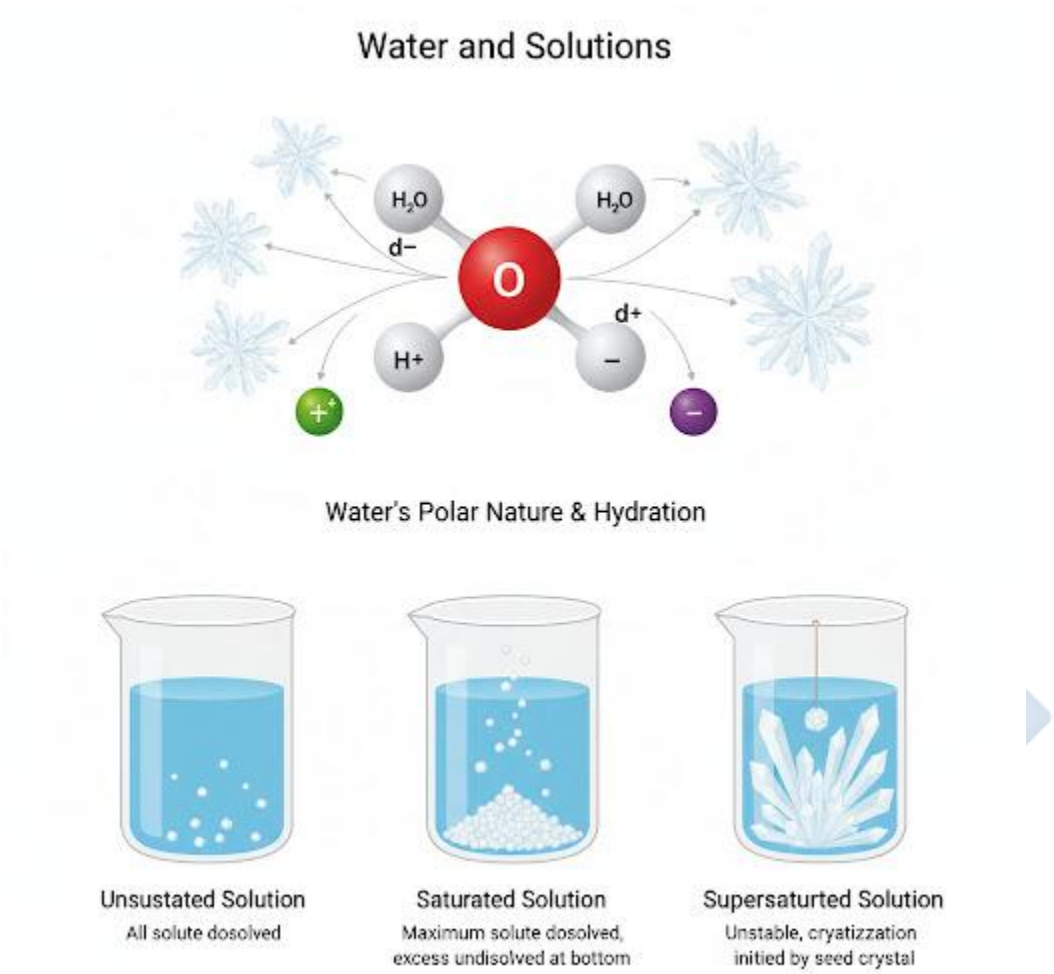
Assignment

1. Explain why the reaction below is classified as a redox reaction:

$$2Na + Cl_2 \rightarrow 2NaCl$$

2. Using the oxidation number method, balance:
   - (a) $H_2S + HNO_3 \rightarrow S + NO + H_2O$
   - (b) $MnO_2 + HCl \rightarrow MnCl_2 + Cl_2 + H_2O$.
3. Research and prepare a **case study** on one application of redox reactions in:
   - (i) Corrosion and its prevention.
   - (ii) Batteries and energy storage.

WEEK 10: WATER AND SOLUTION – PROPERTIES, TYPES OF SOLUTIONS, SOLUBILITY



WATER

Water is the most abundant compound on Earth and is often described as the **universal solvent** because of its ability to dissolve a wide variety of substances. It exists naturally in three states: solid (ice), liquid (water), and gas (water vapour).

(a) Importance of Water

- **Biological importance:**
    - Acts as a transport medium for nutrients and waste products in plants and animals.
    - Maintains body temperature through sweating and evaporation.

- Participates in photosynthesis:
    - $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

- **Industrial importance:**
    - Used as a raw material in the manufacture of chemicals (e.g., H_2, O_2, H_2SO_4).
    - Acts as a coolant in power plants and vehicles.
    - Used as a cleaning and dissolving agent.
- **Domestic importance:**
    - Drinking, cooking, washing, sanitation.

(b) Physical Properties of Water

1. Pure water is a **colourless, odourless, and tasteless** liquid.
2. **Boiling point:** 100°C at 1 atm pressure.
3. **Freezing point:** 0°C at 1 atm pressure.
4. **Density:** 1 g/cm³ at 4°C.
5. Water **expands upon freezing**, making ice less dense (ice floats on water).
6. Water has a **high specific heat capacity**, meaning it can absorb or release large amounts of heat without a large change in temperature. This moderates climate.
7. Water has **high surface tension** due to hydrogen bonding, allowing small insects to walk on its surface.
8. **Electrical conductivity:** Pure water is a poor conductor of electricity, but natural water conducts electricity due to dissolved ions.

(c) Chemical Properties of Water

1. **Reaction with metals:**
    - Active metals such as sodium, potassium, and calcium react with cold water to form hydroxides and release hydrogen gas.
        - $2Na + 2H_2O \rightarrow 2NaOH + H_2\uparrow$
    - Less reactive metals (e.g., Mg, Zn) react only with steam.
        - $Mg + H_2O$ (steam) $\rightarrow MgO + H_2\uparrow$
2. **Reaction with non-metals:**
    - Chlorine dissolves in water to form a mixture of hydrochloric acid and hypochlorous acid:
        - $Cl_2 + H_2O \rightarrow HCl + HOCl$
3. **Hydrolysis:** Some salts undergo hydrolysis in water.
    - Example: $NH_4Cl + H_2O \rightarrow NH_4OH + HCl$
4. **Electrolysis:**
    - Water decomposes when an electric current is passed through acidified water:
        - $2H_2O \rightarrow 2H_2\uparrow + O_2\uparrow$

SOLUTION

A **solution** is a homogeneous mixture of two or more substances in which one substance (solute) is uniformly dispersed within another (solvent).

- **Solute:** Substance being dissolved (lesser amount).
- **Solvent:** Substance that dissolves the solute (greater amount).

Examples:

- Salt solution: NaCl (solute) + Water (solvent).

- Sugar solution: Sugar (solute) + Water (solvent).

(a) Types of Solutions

1. **According to Physical State of Solvent:**
    - Solid solutions: Brass (Cu + Zn), Bronze (Cu + Sn).
    - Liquid solutions: Vinegar (CH_3COOH in H_2O), sugar solution.
    - Gaseous solutions: Air (mixture of gases).
2. **According to Solubility Level:**
    - **Unsaturated solution:** More solute can dissolve at a given temperature.
    - **Saturated solution:** Contains the maximum amount of solute at a given temperature; any additional solute remains undissolved.
    - **Supersaturated solution:** Contains more solute than normal; prepared by heating a saturated solution and cooling gently without disturbance.
3. **According to Electrical Conductivity:**
    - **Electrolytic solutions:** Conduct electricity (e.g., NaCl solution, HCl solution).
    - **Non-electrolytic solutions:** Do not conduct electricity (e.g., sugar in water, ethanol in water).

(b) Properties of Solutions

1. They are **homogeneous** mixtures with uniform composition.
2. Particles are **molecular-sized** and cannot be seen with the naked eye.
3. They are **transparent** and do not scatter light (no Tyndall effect).
4. Solute cannot be separated by filtration but by **evaporation or distillation**.

5. Concentration varies — solutions may be dilute or concentrated.

SOLUBILITY

Definition:

The solubility of a substance is the maximum amount of solute that can dissolve in **100 g of solvent at a given temperature** to form a saturated solution.

Unit: g solute / 100 g solvent (at a specified temperature).

(a) Factors Affecting Solubility

1. **Nature of solute and solvent:**
    - Polar solutes dissolve in polar solvents (e.g., NaCl in H_2O).
    - Non-polar solutes dissolve in non-polar solvents (e.g., iodine in Cl_4).
2. **Temperature:**
    - For most solids, solubility increases with temperature (e.g., KNO_3).
    - For gases, solubility decreases with increase in temperature.
3. **Pressure:**
    - Has little effect on solids and liquids.
    - For gases, solubility increases with pressure (Henry's Law). Example: CO_2 in soft drinks.
4. **Particle size:**
    - Smaller particles dissolve faster due to larger surface area.
5. **Agitation:**
    - Stirring increases the rate of dissolution.

(b) Solubility Curves

- A solubility curve shows how the solubility of a substance varies with temperature.
- It can be used to:
    1. Determine the solubility of a salt at a given temperature.
    2. Predict the temperature at which a solution becomes saturated.
    3. Calculate the amount of solute that crystallizes when a hot saturated solution is cooled.

Worked Example 1:

The solubility of KNO_3 at 40°C is 64 g per 100 g of water. How much KNO_3 will crystallize if a saturated solution containing 64 g at 40°C is cooled to 20°C where solubility is 32 g per 100 g of water?

Solution:

Amount crystallized = 64 g – 32 g = **32 g**

(c) Applications of Solubility

1. **Crystallization:** Used in purification of substances such as sugar and salts.
2. **Drug formulation:** Determines the maximum dose of drugs soluble in blood plasma.
3. **Food industry:** Explains dissolving of salt and sugar in water-based foods and drinks.
4. **Environmental science:** Solubility of oxygen in water is vital for aquatic life.
5. **Industrial processes:** Fertilizer production, beverage carbonation, soap and detergent manufacturing.

EVALUATION

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1. List five physical properties of water and explain their significance.
2. Write balanced equations for the reaction of sodium and magnesium with water.
3. Define the following terms: solute, solvent, saturated solution, supersaturated solution.
4. State and explain three factors affecting solubility.
5. Differentiate between electrolytic and non-electrolytic solutions with examples.
6. Using a solubility curve, explain how you can obtain crystals from a solution.

ASSIGNMENT

1. At 60°C, the solubility of KCl is 45 g per 100 g of water, while at 20°C, it is 30 g per 100 g of water. Calculate the mass of KCl that will crystallize out if 100 g of a saturated solution at 60°C is cooled to 20°C.
2. Explain why sugar readily dissolves in water but does not dissolve in kerosene.
3. Describe three industrial applications of solubility principles.

WEEK 11: SOLUBILITY AND CRYSTALLIZATION

(Saturated Solutions, Solubility Curves, and Crystallization)

Meaning of Key Terms

Term	Full Explanation	Example
Solute	A substance that is dissolved in another substance (solvent) to form a solution. Solutes can be solid, liquid, or gas.	Sugar in water; CO_2 in soft drinks.
Solvent	The substance that dissolves the solute. In chemistry, water is called the **universal solvent** because it dissolves many substances.	Water dissolving salt.

Solution	A **homogeneous mixture** of a solute and solvent. Particles are uniformly distributed, cannot be filtered out, and do not settle.	Sugar solution.
Solubility	The maximum mass of solute (in grams) that dissolves in **100 g of solvent at a fixed temperature** to form a saturated solution.	36 g NaCl/100 g H_2O at 25°C.
Saturated Solution	A solution containing the **maximum possible amount of solute** at a given temperature. No more solute can dissolve under these conditions.	Fully dissolved salt in water at room temperature.
Unsaturated Solution	A solution that contains **less solute** than it can hold at a given temperature. More solute can dissolve.	Lightly sugared tea.
Supersaturated Solution	Contains **more solute than the solubility limit** at a given temperature. This is unstable and can deposit crystals if disturbed.	Cooled sugar solution in candy making.
Crystallization	The process of forming pure, solid crystals from a solution by cooling or evaporating the solvent. Used for purification.	Copper sulfate crystals.

Factors Affecting Solubility

a) Nature of Solute and Solvent

- Solubility depends on **chemical polarity**:
    - **Polar solvents (e.g., water)** dissolve **polar/ionic solutes** like sugar, NaCl.
    - **Non-polar solvents (e.g., benzene)** dissolve **non-polar solutes** like iodine, fats, oils.
- Principle: **"Like dissolves like."**

b) Temperature

- For **most solid solutes**, solubility **increases** with temperature (endothermic dissolution).
- For **gases**, solubility **decreases** as temperature increases.
    - Example: Soda loses fizz when warmed because CO_2 escapes.

c) Pressure (Important for gases)

- Increasing pressure increases gas solubility (Henry's Law).
    - Example: Soft drinks are bottled under high CO_2 pressure.

d) Particle Size

- Smaller particles dissolve faster because of larger surface area.

e) Stirring/Agitation

- Stirring speeds up dissolution but doesn't change solubility.

f) Impurities

- Impurities in a solution may alter solubility.

Experimental Determination of Solubility (Step-by-Step)

Aim:

To determine the solubility of a solid (e.g., KNO_3) in water at 40°C.

Apparatus:

Beaker, thermometer, balance, stirrer, heat source, evaporating dish.

Procedure:

1. Weigh **100 g of distilled water** into a beaker.
2. Heat water to 40°C.
3. Add potassium nitrate gradually while stirring until no more dissolves (solution becomes saturated).

4. Record the **mass of solute added**.
5. Compute solubility:

$$\text{Solubility} = \frac{\text{Mass of solute}}{\text{Mass of solvent}} \times 100$$

Example Calculation:

If **50 g of KNO_3** dissolves in 100 g of water at 40°C:

$$\text{Solubility} = \frac{50}{100} \times 100 = 50\,\text{g}/100\,\text{g water}$$

Solubility Curves (Highly Elaborated)

A **solubility curve** is a graph of **temperature vs solubility**.

- **X-axis:** Temperature (°C)
- **Y-axis:** Solubility (g solute/100 g water)
- The curve shows how much solute dissolves at different temperatures.

Key Observations:

- For most salts (e.g., KNO_3), the curve **slopes upward** – solubility increases with temperature.
- For gases, the curve **slopes downward** – solubility decreases as temperature rises.

Applications of Solubility Curves:

1. Predict the amount of solute that will crystallize when a solution cools.
2. Determine the best temperature for dissolving substances.
3. Compare solubility of different solutes.

5. Crystallization Process (Industrial and Lab Detail)

Laboratory Method (Copper (II) Sulfate Crystals):

1. Dissolve excess $CuSO_4$ in hot water to form a saturated solution.
2. Filter to remove impurities.
3. Cool slowly – crystals form as solubility decreases.
4. Filter crystals, wash with cold water, dry between filter papers.

Industrial Applications:

- Sugar refining, salt purification, alum and drug production, fertilizer manufacture.

Worked Examples

Example 1:

At 80°C, **150 g of KNO_3** dissolves in **100 g water**. If cooled to 30°C, solubility is **45 g/100 g water**.

$$\text{Mass crystallized} = 150 - 45 = 105\,\text{g}$$

Example 2:

The solubility of NaCl at 25°C is **36 g/100 g water**. Calculate the mass of NaCl needed to prepare a **saturated solution in 250 g water**.

$$\text{Mass of solute} = \frac{36}{100} \times 250 = 90\,\text{g}$$

Solubility Concepts

Concept	Meaning	Example
Solubility	g solute/100 g solvent at fixed temp	36 g NaCl in 100 g water at 25°C

Saturated Solution	Max solute at given temp	Saturated sugar water
Supersaturated Solution	Contains excess solute, unstable	Candy making solution
Crystallization	Pure crystal formation from solution	Alum, $CuSO_4$, sugar

Evaluation

1. Define **solubility** and give its SI unit.
2. Explain with a diagram what happens when a saturated solution is cooled.
3. A solution contains 90 g of KNO_3 in 100 g water at 60°C. If the solubility is **110 g/100 g water**, classify the solution.
4. Give two factors that affect gas solubility but not solid solubility.
5. Differentiate between **saturated** and **supersaturated** solutions.

Assignment

1. Design a **step-by-step lab procedure** for obtaining crystals of potassium trioxalate.
2. Explain why cooling a saturated solution leads to crystal formation.
3. Draw and explain a solubility curve for KNO_3.
4. Describe two industrial applications of crystallization in drug and fertilizer production.
5. Why is water considered a **universal solvent**? Give examples.

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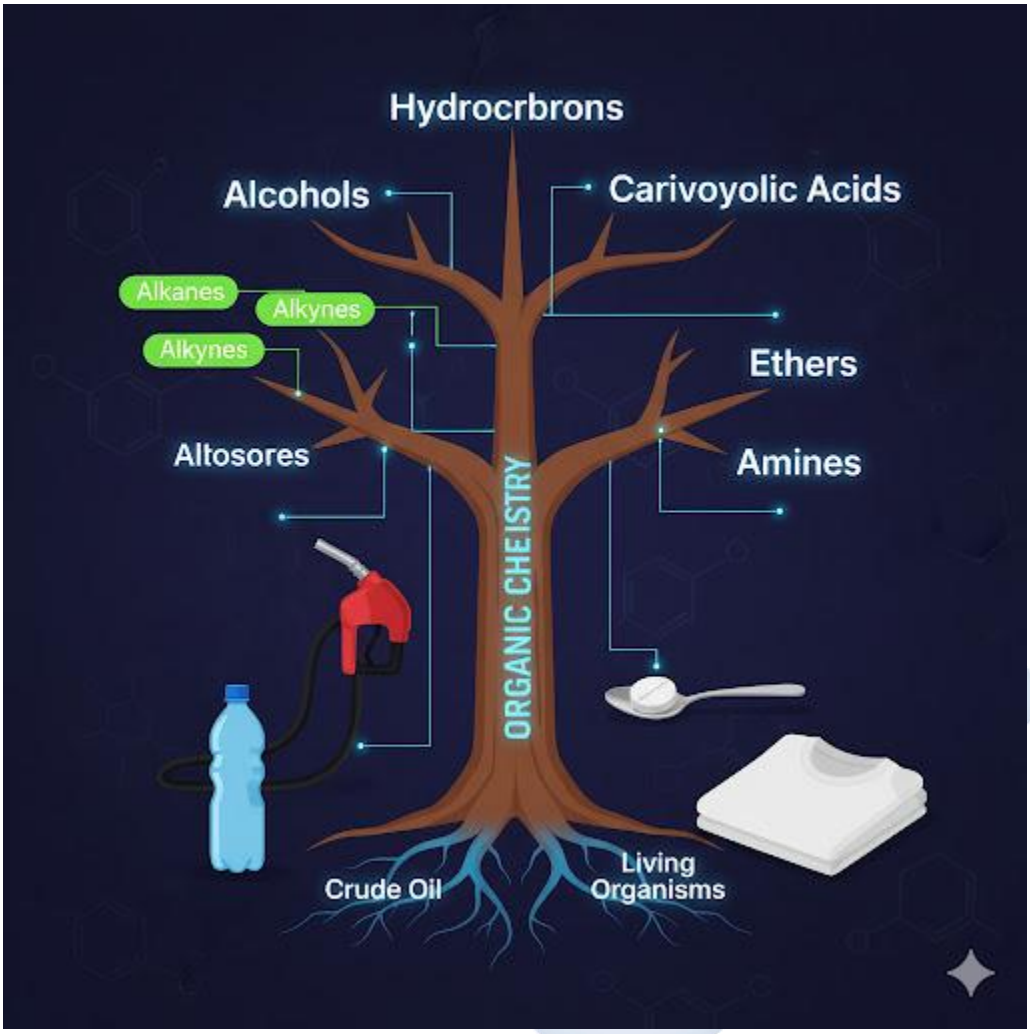
SS1 CHEMISTRY LESSON NOTES

THIRD TERM

1. Introduction to Organic Chemistry – Characteristics of organic compounds.
2. Alkanes, Alkenes, and Alkynes – Structure, properties, examples.
3. Isomerism – Structural and geometric.
4. Petroleum and Petrochemicals – Refining, fractional distillation.
5. **Midterm Examination**
6. Chemical Reactions and Types – Combination, decomposition, displacement.
7. **Midterm break**
8. Energy Changes in Reactions – Endothermic and exothermic.
9. Rates of Chemical Reactions – Factors affecting rate.
10. Reversible Reactions and Chemical Equilibrium – Le Chatelier's principle.
11. Metals and Non-Metals – Properties and uses.
12. Examination
13. Closing

WEEK 1: INTRODUCTION TO ORGANIC CHEMISTRY

Characteristics of Organic Compounds



MEANING OF ORGANIC CHEMISTRY

Organic Chemistry is the branch of Chemistry that deals with the **study of carbon compounds**, particularly those containing **carbon–hydrogen bonds (hydrocarbons)** and their derivatives.

The term *organic* originally came from the belief that compounds from living organisms could only be formed by a "vital force." This was disproved in 1828 when **Friedrich Wöhler** accidentally synthesized **urea** (an organic compound found in urine) from **ammonium cyanate**

(an inorganic salt). This historic event marked the **birth of modern Organic Chemistry**, proving that organic compounds can be synthesized artificially.

Today, Organic Chemistry is not limited to compounds from living organisms; it covers **all carbon compounds**, except a few simple ones like carbon dioxide, carbonates, bicarbonates, and carbides (which are studied under Inorganic Chemistry).

IMPORTANCE OF ORGANIC CHEMISTRY

Organic Chemistry is crucial because it influences every aspect of human life. Its importance can be grouped as follows:

(a) In Everyday Life

- The food we eat (carbohydrates, proteins, lipids) is organic.
- The medicines we take (antibiotics, painkillers, vitamins, vaccines) are organic.
- Clothing materials such as cotton, wool, nylon, polyester, and rayon are organic.
- Fuels (petrol, diesel, kerosene, cooking gas, coal) are organic.

(b) In Industry

- Petrochemical industries rely on hydrocarbons for fuels and raw materials.
- Pharmaceutical industries depend on organic synthesis for drug production.
- Plastic, paint, rubber, and cosmetic industries are based on organic compounds.

(c) In Agriculture

- Fertilizers, pesticides, herbicides, and growth hormones contain organic compounds.

(d) In Nature

- Life itself is built around **carbon chemistry**. DNA, RNA, enzymes, and hormones are organic molecules that control heredity and metabolic processes.

GENERAL CHARACTERISTICS OF ORGANIC COMPOUNDS

Organic compounds have unique properties due to the **nature of carbon atoms** and the **types of bonds** they form. The main characteristics include:

(a) Presence of Carbon and Hydrogen

- Organic compounds always contain carbon, usually bonded with hydrogen.
- Carbon's ability to **catenate** (form long chains, branched chains, and rings with itself) makes the variety of organic compounds almost limitless.
- For example:
    - Methane (CH_4) → simplest hydrocarbon.
    - Hexane (C_6H_{14}) → long chain.
    - Benzene (C_6H_6) → ring compound.

(b) Tetravalency of Carbon

- Carbon has an atomic number of 6, with an electronic configuration of 2, 4.
- It has 4 valence electrons and needs 4 more to complete its octet.
- Therefore, it always forms **four covalent bonds** with other atoms, allowing for stable molecules.
    - Example: Methane, CH_4.

(c) Covalent Bonding

- The atoms in organic compounds are linked mainly by covalent bonds (shared electron pairs).
- This makes organic compounds generally less soluble in water and have lower melting/boiling points compared to ionic salts.

(d) Combustibility

- Most organic compounds are combustible.

- They burn in oxygen to produce carbon dioxide and water, often with much heat energy.
    - Example: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + heat.
- Hydrocarbons with higher carbon content burn with a **sooty flame** due to incomplete combustion.

(e) Solubility

- Organic compounds are generally insoluble in water (polar solvent) but soluble in organic solvents such as benzene, chloroform, ethanol, and ether.
- Reason: Organic compounds are mostly non-polar or weakly polar, and solubility depends on similarity of polarity ("like dissolves like").

(f) Thermal Instability

- Many organic compounds decompose easily when heated strongly.
- For instance, sugar ($C_{12}H_{22}O_{11}$) chars on strong heating to give carbon and water vapour.

(g) Slow Reaction Rate

- Organic reactions are usually slow and often require special conditions such as heat, pressure, or catalysts (e.g., platinum, nickel).
- Example: Hydrogenation of ethene to ethane requires nickel catalyst.

(h) Formation of Isomers

- Organic compounds exhibit **isomerism** – the existence of two or more compounds with the same molecular formula but different structural or spatial arrangements.
    - Example: C_4H_{10} → Butane and 2-methylpropane (isobutane).
- This greatly increases the number of possible organic compounds.

(i) Large Number of Compounds

- Over 90% of known chemical compounds are organic. This is due to:
    1. **Catenation** – ability of carbon to bond with itself extensively.

2. **Tetravalency** – carbon forms up to 4 bonds.
3. **Multiple bonding** – carbon forms single, double, and triple bonds.
4. **Isomerism** – different arrangements of the same atoms.

(j) Low Melting and Boiling Points

- Most organic compounds are liquids or gases at room temperature due to weak intermolecular forces (Van der Waals forces).
- Solids like glucose and urea decompose before melting.

(k) Distinctive Odours

- Many organic compounds have characteristic odours.
    - Esters smell fruity.
    - Amines have a fishy smell.
    - Aldehydes and ketones have sharp odours.

COMPARISON BETWEEN ORGANIC AND INORGANIC COMPOUNDS

Feature	**Organic Compounds**	**Inorganic Compounds**
Main elements	Carbon, hydrogen (plus O, N, S, halogens, P)	Metals, non-metals, salts, oxides
Bonding	Mostly covalent	Mostly ionic
Melting/boiling points	Generally low (except network solids like diamond)	Generally high
Solubility	Insoluble in water; soluble in organic solvents	Soluble in water; insoluble in organic solvents
Combustibility	Mostly combustible	Few combustible (e.g., carbon,

		sulphur)
Stability to heat	Less stable; decompose on heating	More stable
Reaction rate	Slow; require catalysts/conditions	Fast
Number of compounds	Very large (millions)	Relatively fewer
Isomerism	Very common	Rare

REASONS FOR THE STUDY OF ORGANIC CHEMISTRY

1. To understand the **chemical basis of life**.
2. To explain the composition, properties, and uses of hydrocarbons and their derivatives.
3. To aid in the synthesis of **new materials** such as drugs, polymers, and fuels.
4. To support industries such as **petrochemical, textile, pharmaceutical, plastic, and cosmetic industries**.
5. To provide solutions to problems in **energy, environment, and health**.

EVALUATION

1. Define Organic Chemistry.
2. Briefly explain the historical development of Organic Chemistry.
3. List six industries that depend on Organic Chemistry for survival.
4. State and explain **seven general characteristics** of organic compounds.
5. Differentiate between organic and inorganic compounds under four headings.
6. Why are organic compounds so numerous compared to inorganic compounds?

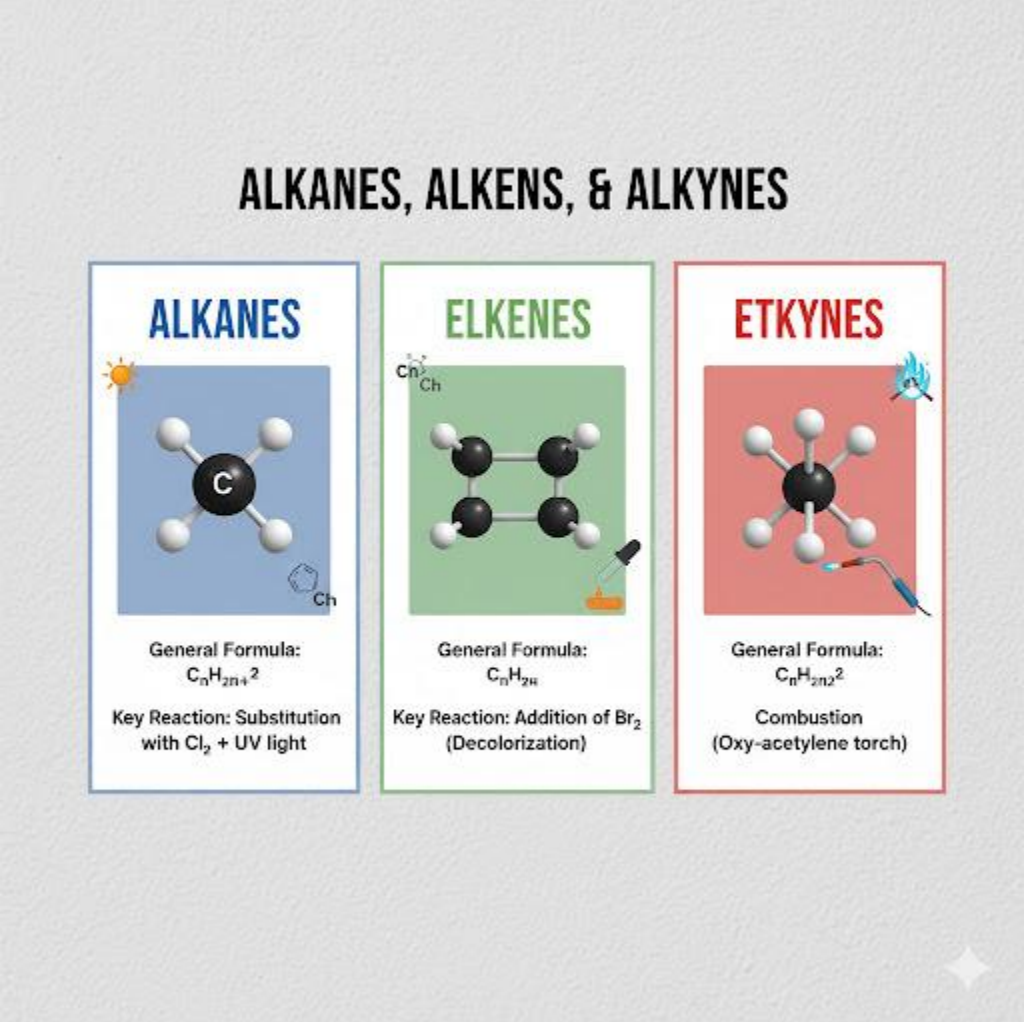
ASSIGNMENT

1. Discuss **five reasons** why carbon forms a very large number of compounds.
2. Organic compounds are mostly covalent in nature. Explain with two examples.
3. Draw the structures of the two isomers of C_4H_{10} and name them.

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WEEK 2: ALKANES, ALKENES, AND ALKYNES

Structure, Properties, and Examples



INTRODUCTION TO HYDROCARBONS

- **Definition:** Hydrocarbons are organic compounds made up of only **carbon (C)** and **hydrogen (H)** atoms.
- **Classification:**
    1. **Alkanes (Saturated hydrocarbons):** All bonds are **single covalent bonds (σ-bonds)**.

2. **Alkenes (Unsaturated hydrocarbons):** Contain at least **one double bond (C=C)** consisting of one sigma (σ) and one pi (π) bond.
3. **Alkynes (Unsaturated hydrocarbons):** Contain at least **one triple bond (C≡C)** consisting of one sigma (σ) and two pi (π) bonds.

- **Homologous Series:**
  A family of organic compounds with:
  - The same **general formula**.
  - A difference of a **$–CH_2–$ group** between successive members.
  - Similar chemical properties.
  - Gradual change in physical properties.

Alkanes (Saturated Hydrocarbons)

General Formula: CnH2n+2C_nH_{2n+2}

Bonding: Only single bonds (σ-bonds).

Examples:

- Methane (CH_4)
- Ethane (C_2H_6)
- Propane (C_3H_8)
- Butane (C_4H_{10})

Isomerism:

- From **butane (C_4H_{10})** upwards, alkanes show **structural isomerism**.
  - **n-butane:** $CH_3–CH_2–CH_2–CH_3$
  - **iso-butane:** $CH_3–CH(CH_3)–CH_3$

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Physical Properties:

- **State:** C_1–C_4 are gases, C_5–C_{17} are liquids, C_{18} and above are waxy solids.
- **Melting/boiling points:** Increase with molecular size due to stronger van der Waals forces.
- **Polarity:** Non-polar, insoluble in water but soluble in organic solvents.

Chemical Properties:

Alkanes are chemically **inert** due to strong C–C and C–H bonds.

- **Combustion:**
  - Complete: $CH_4+2O_2 \rightarrow CO_2+2H_2OCH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
  - Incomplete: Produces CO or C (soot).
- **Substitution reaction:**
  - In the presence of **UV light or heat**, halogens replace hydrogen: $CH_4+Cl_2 \rightarrow CH_3Cl+HClCH_4 + Cl_2 \rightarrow CH_3Cl + HCl$
  - Further substitution yields CH_2Cl_2, $CHCl_3$, and CCl_4.

ALKENES (UNSATURATED HYDROCARBONS – DOUBLE BONDS)

General Formula: CnH2nC_nH_{2n}

Bonding: At least one double bond (σ + π).

Examples:

- Ethene (C_2H_4): $H_2C=CH_2$
- Propene (C_3H_6): $CH_3–CH=CH_2$
- But-1-ene (C_4H_8): $CH_2=CH–CH_2–CH_3$
- But-2-ene (C_4H_8): $CH_3–CH=CH–CH_3$

Isomerism:

- **Structural isomerism:** Different positions of the double bond (but-1-ene vs but-2-ene).
- **Geometrical isomerism (cis/trans):**
  - cis-but-2-ene (CH_3 groups on same side).
  - trans-but-2-ene (CH_3 groups on opposite sides).

Physical Properties:

- Similar to alkanes but slightly more polarizable due to π-electrons.
- Lower members are gases; middle members are liquids; higher ones are waxy solids.

Chemical Properties:

Highly reactive due to **π-bond** (weak, easily attacked by electrophiles).

1. **Combustion:**
   $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$
2. **Addition reactions:**
   - Hydrogenation: $C_2H_4 + H_2 \xrightarrow{Ni} C_2H_6$
   - Halogenation: $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$ (bromine water test).
   - Hydrohalogenation: $C_2H_4 + HCl \rightarrow C_2H_5Cl$.
   - Hydration: $C_2H_4 + H_2O \xrightarrow{H_2SO_4} C_2H_5OH$.
3. **Polymerization:**
   - Ethene polymerizes under high pressure to give **polyethene**.

ALKYNES (UNSATURATED HYDROCARBONS – TRIPLE BONDS)

General Formula: CnH2n−2C_nH_{2n-2}

Bonding: At least one triple bond (σ + 2π).

Examples:

- Ethyne (acetylene) C_2H_2: H–C≡C–H

- Propyne (C_3H_4): $CH{\equiv}C–CH_3$
- Butyne (C_4H_6): $CH{\equiv}C–CH_2–CH_3$ (1-butyne), $CH_3–C{\equiv}C–CH_3$ (2-butyne).

Isomerism:

- From C_4H_6 (butyne) upward, alkynes exhibit **structural isomerism** (position of triple bond).

Physical Properties:

- Similar to alkanes and alkenes (non-polar, insoluble in water).
- Ethyne is a colourless gas with faint odour, highly flammable.

Chemical Properties:

1. **Combustion:**
   $2C2H2+5O2 \rightarrow 4CO2+2H2O2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$.
2. **Addition reactions:**
   - With hydrogen:
     $C2H2+2H2 \rightarrow NiC2H6C_2H_2 + 2H_2 \xrightarrow{Ni} C_2H_6$.
   - With halogens:
     $C2H2+2Br2 \rightarrow C2H2Br4C_2H_2 + 2Br_2 \rightarrow C_2H_2Br_4$.
3. **Substitution reaction (terminal alkynes):**
   - Ethyne reacts with sodium:
     $HC{\equiv}CH+2Na \rightarrow 2NaC{\equiv}CH+H2HC{\equiv}CH + 2Na \rightarrow 2NaC{\equiv}CH + H_2$.
4. **Industrial importance:**
   - Ethyne burns with an oxy-acetylene flame (very hot) used in **welding and cutting metals**.

Comparison of Alkanes, Alkenes, and Alkynes

Property	Alkanes	Alkenes	Alkynes
General Formula	CnH2n+2C_nH_{2n+2}	CnH2nC_nH_{2n}	CnH2n−2C_nH_{2n-2}
Bond Type	Single bonds only (σ)	Double bond (σ + π)	Triple bond (σ + 2π)
Saturation	Saturated	Unsaturated	Unsaturated
Isomerism	Structural (from C_4H_{10})	Structural + Geometrical	Structural (from C_4H_6)
Reactivity	Least reactive	Reactive	Very reactive
Major Reaction	Substitution	Addition	Addition + substitution
Industrial Use	Fuels (petrol, diesel, kerosene)	Plastics, alcohols, detergents	Welding, chemical synthesis

Applications and Industrial Importance

- **Alkanes:** Used as **fuels** (methane – natural gas, propane – cooking gas, petrol, kerosene, diesel).
- **Alkenes:** Raw materials for **plastics (polyethene, PVC)**, synthetic alcohols, detergents.
- **Alkynes:** Ethyne for welding (oxy-acetylene flame reaches ~3000°C), starting material in synthesis of ethanol, ethanoic acid, and plastics.

Evaluation

1. Differentiate between alkanes, alkenes, and alkynes with respect to:
    - General formula
    - Type of bond

- Reactivity.

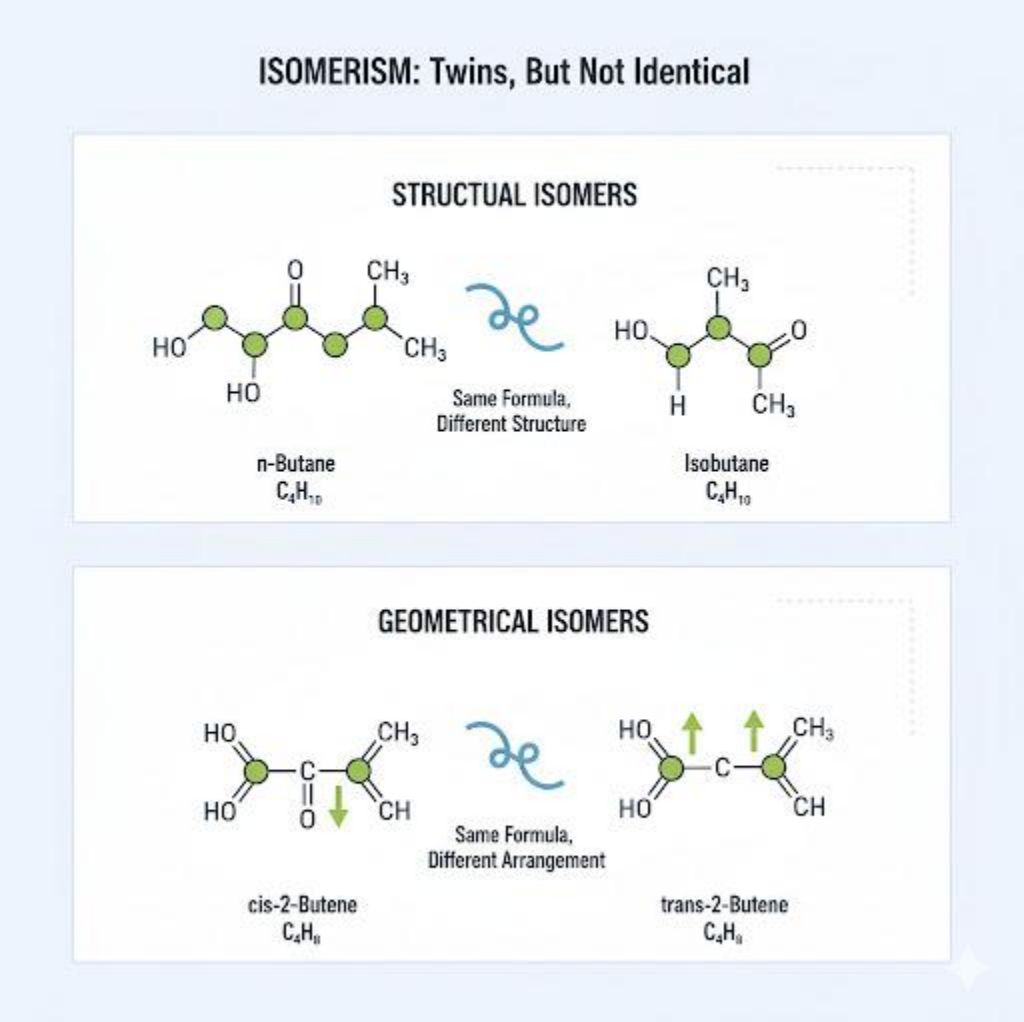
2. Draw and name all possible isomers of C_4H_{10}, C_4H_8, and C_4H_6.
3. Write equations for the complete and incomplete combustion of:
    - Methane (CH_4)
    - Ethene (C_2H_4)
    - Ethyne (C_2H_2).
4. Explain why alkenes readily undergo addition reactions while alkanes do not.
5. Ethene reacts with hydrogen bromide (HBr). Write the mechanism of this reaction.

Assignment

1. Give three reasons why hydrocarbons are classified into alkanes, alkenes, and alkynes.
2. Write the **structural and condensed formulae** of the first four members of:
    - (a) Alkanes
    - (b) Alkenes
    - (c) Alkynes.
3. Explain **cis-trans isomerism** in alkenes with a diagram.
4. Research: Why is **ethyne flame mixed with oxygen** used in welding instead of an ordinary ethyne flame?
5. State **four industrial applications** each of alkanes, alkenes, and alkynes.

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WEEK 3: ISOMERISM – STRUCTURAL AND GEOMETRIC



MEANING OF ISOMERISM

Isomerism is a very important concept in organic chemistry. It is the phenomenon where two or more compounds have the **same molecular formula** (same number and type of atoms) but are arranged differently, leading to **different structural formulas** and often different physical and chemical properties.

- Compounds that show this relationship are called **isomers**.
- Isomerism explains why organic chemistry has such a vast number of compounds, even though they are made mainly of carbon, hydrogen, oxygen, nitrogen, and halogens.

Example:

- Butane (C_4H_{10}) has two isomers:
    1. n-Butane → $CH_3–CH_2–CH_2–CH_3$
    2. 2-Methylpropane → $CH_3–CH(CH_3)–CH_3$

Both have the same molecular formula (C_4H_{10}) but different structures and properties.

TYPES OF ISOMERISM

Isomerism is broadly divided into:

1. **Structural Isomerism** – when isomers differ in how atoms are bonded together.
2. **Stereoisomerism** – when isomers differ in the spatial arrangement of atoms (same connectivity but different 3D orientation).

In SS1, we focus on **structural isomerism** and one type of stereoisomerism: **geometrical isomerism**.

STRUCTURAL ISOMERISM

Structural isomerism occurs when compounds have the same molecular formula but different bonding arrangements of atoms.

Types of Structural Isomerism

1. **Chain Isomerism (Skeletal Isomerism):**
    - Occurs when compounds differ in the arrangement of the carbon skeleton (straight chain vs branched chain).
    - **Example:**
        - n-Butane: $CH_3–CH_2–CH_2–CH_3$
        - 2-Methylpropane: $CH_3–CH(CH_3)–CH_3$

→ Both are C_4H_{10} but have different structures and boiling points.

2. **Position Isomerism:**
    - Occurs when a functional group or substituent is attached at different positions on the carbon chain.
    - **Example:**
        - Butan-1-ol: $CH_3–CH_2–CH_2–CH_2OH$ (OH at position 1)
        - Butan-2-ol: $CH_3–CH_2–CHOH–CH_3$ (OH at position 2).

→ Both are alcohols with formula $C_4H_{10}O$ but have different properties.

3. **Functional Group Isomerism:**
    - Occurs when compounds have the same molecular formula but belong to different homologous series (different functional groups).
    - **Example:**
        - Ethanol (CH_3CH_2OH) – an alcohol.
        - Dimethyl ether (CH_3OCH_3) – an ether.

→ Both are C_2H_6O but very different in chemistry.

4. **Tautomeric (Dynamic) Isomerism:**
    - A special type where two isomers exist in equilibrium and interconvert easily.
    - Most common is **keto-enol tautomerism**.
    - **Example:**
        - Keto form: $CH_3–CO–CH_2–CH_3$ (Butan-2-one).

- Enol form: $CH_2=CH–CH(OH)–CH_3$ (But-2-en-ol).

5. **Metamerism:**
    - Occurs in compounds with a divalent functional group (like –O–, –S–, or –CO–NH–) where alkyl groups on either side are different.
    - **Example:**
        - Diethyl ether ($CH_3CH_2–O–CH_2CH_3$).
        - Methyl propyl ether ($CH_3–O–CH_2CH_2CH_3$).

GEOMETRICAL ISOMERISM (Cis-Trans Isomerism)

Geometrical isomerism is a form of **stereoisomerism**. It occurs when there is **restricted rotation** around a double bond or within a ring structure, leading to different spatial arrangements of atoms.

Features of Geometrical Isomerism

- Found mainly in **alkenes** and **cyclic compounds**.
- Due to the rigidity of the C=C double bond, groups attached cannot rotate freely.
- Isomers are named as **cis** (same side) and **trans** (opposite side).

Examples

1. **But-2-ene (C_4H_8):**
    - **Cis-but-2-ene:** CH_3 groups on the same side of the double bond.
    - **Trans-but-2-ene:** CH_3 groups on opposite sides of the double bond.
2. **1,2-Dichloroethene ($C_2H_2Cl_2$):**
    - **Cis form:** Both Cl atoms on the same side.

- **Trans form:** Cl atoms on opposite sides.

PROPERTIES OF ISOMERS

- **Structural isomers:**
  - May differ in **boiling and melting points**.
  - Different solubilities and densities.
  - Different chemical reactivities, depending on the functional groups.
- **Geometrical isomers:**
  - Cis isomers are usually **more polar** (higher boiling points, lower melting points).
  - Trans isomers often **pack better** in crystals (higher melting points).
  - Their reactivities in addition or substitution reactions may differ.

SIGNIFICANCE OF ISOMERISM

1. Explains the **diversity of organic compounds**.
2. Crucial in **drug design** – e.g., one isomer of a drug may be beneficial while another is harmful.
3. Important in **materials science**, polymers, and plastics.
4. Explains why two substances with the same formula behave completely differently.
5. Provides insight into **structure–property relationships** in chemistry.

SOLVED EXAMPLES

Example 1: Write and name all structural isomers of **$C_4H_{10}O$**.

Solution:

$C_4H_{10}O$ could be an alcohol or ether.

1. Butan-1-ol → $CH_3–CH_2–CH_2–CH_2OH$
2. Butan-2-ol → $CH_3–CH_2–CHOH–CH_3$
3. 2-Methylpropan-1-ol → $(CH_3)_2CH–CH_2OH$
4. 2-Methylpropan-2-ol → $(CH_3)_3COH$
5. Diethyl ether → $CH_3CH_2–O–CH_2CH_3$
6. Methyl propyl ether → $CH_3–O–CH_2CH_2CH_3$

Thus, there are **6 isomers**.

Example 2: Draw the cis and trans isomers of But-2-ene.

- **Cis-but-2-ene:**

  H CH3

  \ /

  C=C

  / \

  CH3 H

- **Trans-but-2-ene:**

  H CH3

  \ /

  C=C

  / \

  H CH3

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EVALUATION

1. Define isomerism.
2. Distinguish between chain and position isomerism with one example each.
3. Write the structural formulas of two functional group isomers of C_2H_6O.
4. Why is geometrical isomerism not possible in alkanes?
5. Draw the cis and trans forms of 1,2-dichloroethene.

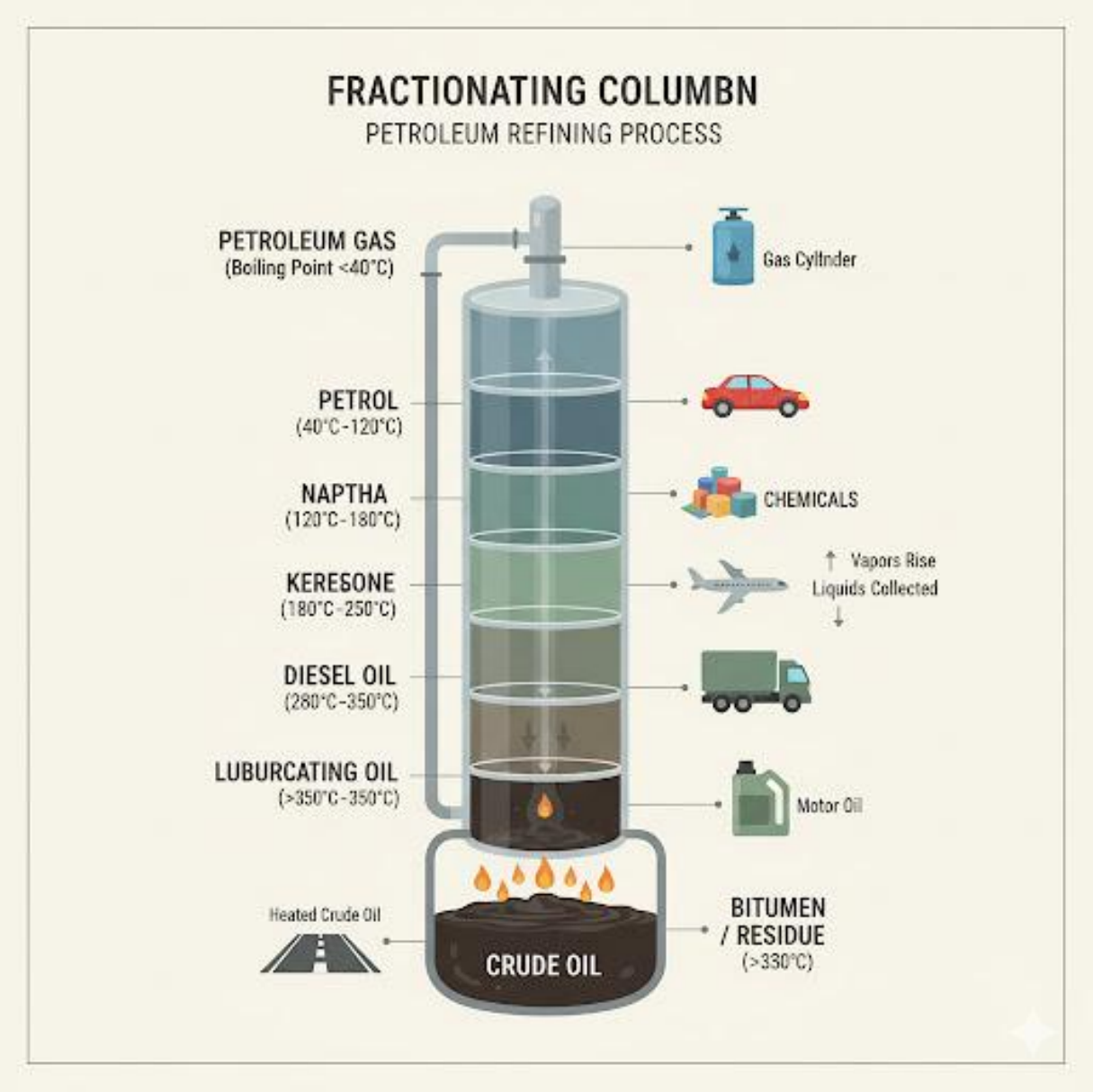
ASSIGNMENT

1. Give three differences between structural and geometrical isomerism.
2. Write and name **all possible chain isomers of C_5H_{12}**.
3. With diagrams, show the cis and trans isomers of But-2-ene and explain how they differ in properties.

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WEEK 4: PETROLEUM AND PETROCHEMICALS

(Refining and Fractional Distillation)



Meaning and Origin of Petroleum

- **Petroleum** is a **naturally occurring, viscous, dark brown or black, flammable liquid** found in underground geological formations, often trapped in **porous rocks** beneath impermeable rock layers.

- It is a **mixture of hydrocarbons** (molecules containing hydrogen and carbon) along with impurities like sulphur, nitrogen, oxygen, and trace metals.
- **Origin:**
    - Petroleum is a **fossil fuel** formed from the remains of microscopic plants and marine animals buried under sedimentary rocks **50–500 million years ago**.
    - Heat and pressure over millions of years converted these remains into crude oil and natural gas.
    - Petroleum often occurs with **natural gas** in reservoirs capped by non-porous rock layers.

Chemical Composition of Petroleum

Petroleum contains:

- **Alkanes (Paraffins):** 30–60% – Straight or branched saturated hydrocarbons. Example: Methane, Ethane.
- **Cycloalkanes (Naphthenes):** 30–40% – Saturated cyclic hydrocarbons. Example: Cyclohexane.
- **Aromatic Hydrocarbons:** 3–15% – Compounds with benzene rings. Example: Benzene, Toluene.
- **Impurities:**
    - Sulphur compounds (H_2S), Nitrogen compounds, Oxygen compounds.
    - Trace metals like Vanadium (V), Nickel (Ni).

Importance of Petroleum

Petroleum is the **lifeblood of the modern economy**:

1. **Energy Source:** Fuels like petrol, diesel, kerosene power cars, ships, aircraft, homes, and industries.
2. **Industrial Raw Material:** Foundation for **petrochemical industries** producing plastics, synthetic fibers, fertilizers, paints, detergents, pharmaceuticals, and more.
3. **Lubricants:** Petroleum lubricating oils reduce friction in machines.
4. **Economic Backbone:** Petroleum is Nigeria's **largest export and revenue source**.
5. **Asphalt and Bitumen:** Road surfacing and roofing materials.
6. **Electricity Generation:** Heavy fuel oils power turbines in thermal plants.

Petrochemicals and Their Applications

Petrochemicals are chemicals derived from petroleum or natural gas. They are **building blocks** for numerous synthetic materials:

Petrochemical	**Source**	**End Products**
Ethylene (C_2H_4)	Cracking of naphtha	Polyethylene (bags, bottles), PVC, ethanol
Propylene (C_3H_6)	Cracking of gas oil	Polypropylene (plastic chairs, ropes), acrylonitrile
Benzene	Catalytic reforming	Nylon, detergents, drugs, explosives
Toluene	Catalytic reforming	Paints, adhesives, explosives (TNT)
Methanol (CH_3OH)	Synthesis gas	Formaldehyde, resins, solvents
Ammonia (NH_3)	Natural gas (via Haber process)	Fertilizers, nitric acid

Stages of Petroleum Refining

Petroleum refining is a **multi-step industrial process** that converts crude oil into usable products.

A. Separation Stage (Fractional Distillation):

- Principle: **Different hydrocarbons have different boiling points.** Heating crude oil causes components to vaporize and separate at different temperatures.
- Crude oil is first **heated to ~400°C** and fed into a **fractionating tower**.
- The column is **hottest at the bottom** and cooler at the top, causing vapors to condense at different heights.

B. Conversion Processes:

To improve quality and quantity of useful products:

1. **Cracking:** Breaking large molecules into smaller, high-demand ones (e.g., petrol, LPG).
    - *Thermal cracking*: High temperature, high pressure.
    - *Catalytic cracking*: Uses zeolite catalysts.
2. **Reforming:** Rearranging hydrocarbons to form branched and aromatic hydrocarbons, increasing octane number.
3. **Polymerization:** Small olefins (e.g., ethene, propene) combined to make large molecules.
4. **Alkylation:** Light hydrocarbons combined to form high-octane fuels.

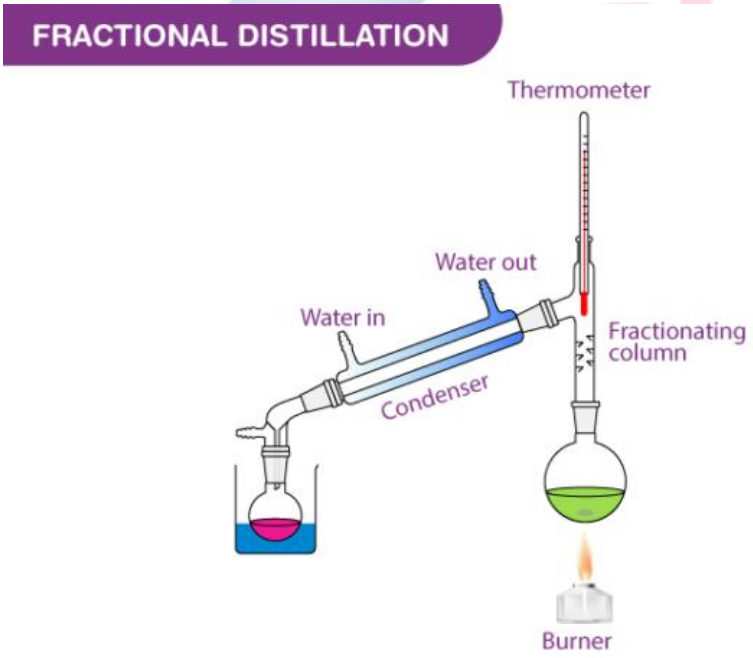
C. Treatment and Blending:

- Desulphurization (removal of sulphur compounds).
- Blending different fractions to meet desired fuel specifications.

Fractional Distillation in Detail

Fraction	Boiling Range	Condensation Point in	Uses

	(°C)	Tower	
Petroleum Gas (LPG)	<40°C	Top of tower	Cooking gas, heating, chemical feedstock
Petrol (Gasoline)	40–120°C	Upper section	Car fuel, solvents
Naphtha	120–180°C	Upper-middle	Petrochemicals, plastics
Kerosene	180–250°C	Middle	Jet fuel, lamps
Diesel Oil	250–320°C	Middle-lower	Lorries, generators
Lubricating Oil	320–400°C	Lower	Lubricants, wax, polishes
Fuel Oil	400°C+	Bottom	Ships, factories
Bitumen/Residue	Solid	Bottom residue	Road tarring, roofing



Environmental and Safety Concerns in Petroleum Industry

1. **Air Pollution:** Combustion releases greenhouse gases (CO_2, CO, NO_x, SO_2).
2. **Oil Spills:** Devastate aquatic ecosystems (e.g., Niger Delta).
3. **Resource Depletion:** Petroleum is **non-renewable**; reserves are finite.
4. **Fire Hazards:** Petroleum products are highly flammable.
5. **Industrial Accidents:** Explosions, toxic gas leaks.

Control Measures:

- Use of renewable energy (solar, wind).
- Improved pipeline maintenance to prevent spills.
- Efficient refinery operations to minimize emissions.
- Environmental laws and safety training for workers.

Nigeria's Petroleum Industry

- **Refineries:** Port Harcourt, Kaduna, Warri, and Dangote Refinery (Lagos).
- **Major Products:** Petrol, diesel, aviation fuel, lubricants, asphalt.
- **Importance:** Provides **90% of Nigeria's foreign exchange earnings.**

Comparison of Petroleum, Natural Gas, and Coal:

Feature	**Petroleum**	**Natural Gas**	**Coal**
Main form	Liquid	Gas	Solid
Main component	Hydrocarbons	Methane (CH_4)	Carbon (C)
Energy content	High	Very high	Moderate
Ease of transport	Pipelines, tankers	Pipelines	Difficult

Cleanliness	Less clean than gas	Cleanest fossil fuel	Dirtiest (ash, soot)

Evaluation

1. Define petroleum and describe its origin.
2. Explain fractional distillation and why it is effective for separating crude oil components.
3. List **four conversion processes** in petroleum refining and their purposes.
4. State the boiling point range, position in column, and uses of:
    - Naphtha
    - Kerosene
    - Bitumen
5. Explain **three environmental hazards** of petroleum refining and suggest solutions.
6. Differentiate between **petroleum products** and **petrochemicals** with examples.
7. Describe how Nigeria benefits economically from petroleum.

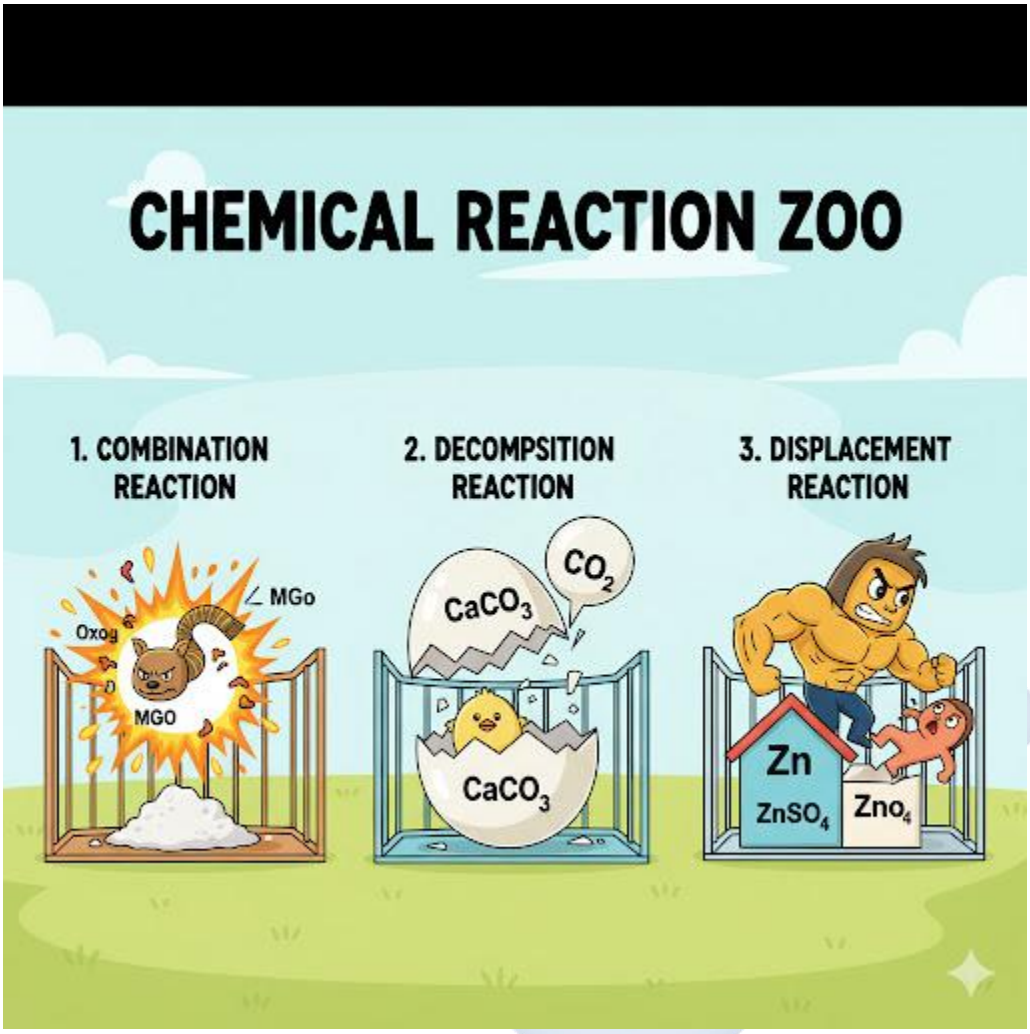
Assignment:

1. Draw a fully labeled **fractionating column** of petroleum, showing all major collection points.
2. Write a comprehensive note on **cracking** and its industrial importance.
3. List **five petrochemicals** and describe two uses for each.
4. Suggest three ways Nigeria can reduce dependence on petroleum.
5. Compare petroleum, coal, and natural gas as sources of energy in terms of availability, efficiency, and environmental impact.

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WEEK 6: CHEMICAL REACTIONS AND TYPES

Combination Reactions, Decomposition Reactions, Displacement Reactions



Meaning of Chemical Reactions

- A **chemical reaction** is a process in which substances (reactants) interact to form new substances (products) with different chemical compositions and properties.
- In a chemical reaction:
    - **Bonds in reactants are broken**.
    - **New bonds are formed in products**.

- Example:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Hydrogen and oxygen (reactants) form water (product).

Characteristics of Chemical Reactions:

1. **Energy change** → reactions may release or absorb energy (exothermic or endothermic).
2. **Observable changes** → gas evolution, colour change, heat/light, precipitate formation.
3. **Mass is conserved** → Law of Conservation of Mass states: *"Matter is neither created nor destroyed in a chemical reaction."*
4. **Irreversibility** → many reactions cannot be easily reversed.

Everyday Examples of Chemical Reactions:

- Cooking of food (chemical change in starch and proteins).
- Burning of fuel (combustion reaction).
- Rusting of iron (slow chemical reaction).
- Photosynthesis in plants (sunlight-driven chemical process).

Types of Chemical Reactions

We shall study three common types: **Combination, Decomposition, and Displacement**.

(A) Combination (Synthesis) Reactions

- **Definition:** Two or more simple substances react chemically to form a single, more complex product.
- **General Form:**

$$A + B \rightarrow AB$$

Examples:

1. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ → burning of magnesium ribbon in air forms magnesium oxide (a white ash).
2. $C(s) + O_2(g) \rightarrow CO_2(g)$ → burning of carbon in oxygen forms carbon dioxide.
3. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ → synthesis of ammonia (Haber process).
4. $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$ → formation of slaked lime.

Features:

- Exothermic (releases heat/light).
- Only one product is formed.
- Common in combustion and industrial synthesis.

Everyday Applications:

- Combustion of fuels for energy.
- Photosynthesis in plants ($6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + O_2$).
- Cement industry (lime reacting with water).

(B) Decomposition Reactions

- **Definition:** A single compound breaks down into two or more simpler substances.
- **General Form:**

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$$AB \rightarrow A + B$$

Types of Decomposition:

1. **Thermal Decomposition** → breakdown caused by heat.
   - $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$
   - $2KClO_3(s) \xrightarrow{MnO_2, \Delta} 2KCl(s) + 3O_2(g)$
2. **Electrolytic Decomposition** → breakdown using electricity.
   - Electrolysis of water:

$$2H_2O(l) \xrightarrow{electricity} 2H_2(g) + O_2(g)$$

3. **Photolytic Decomposition** → breakdown by light.
   - $2AgCl(s) \xrightarrow{sunlight} 2Ag(s) + Cl_2(g)$
   - This explains why **silver salts are stored in dark bottles**.

Features:

- Requires energy input (endothermic).
- Produces two or more products.
- Used in laboratory preparation of gases (e.g., O_2 from $KClO_3$).

Everyday Applications:

- Photosynthesis (photolytic decomposition of water in chloroplasts).
- Rechargeable batteries (electrolysis processes).
- Decomposition of food → spoilage.

(C) Displacement Reactions

- **Definition:** A more reactive element displaces a less reactive element from its compound.

- **General Form:**

$$A + BC \rightarrow AC + B$$

Examples:

1. $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$
2. $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$
3. $Cl_2(g) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$

Explanation:

- In (1), zinc is more reactive than copper, so zinc displaces copper.
- In (3), chlorine is more reactive than bromine, so it displaces bromine.

Features:

- Guided by the **reactivity series of metals and halogens**.
- More reactive element replaces a less reactive one.
- Used in extraction of metals (e.g., Fe from Fe_2O_3 using C).

Everyday Applications:

- Galvanization (zinc protects iron by reacting first).
- Recovery of metals from solutions.
- Halogen displacement reactions in water purification.

Summary Table

Reaction Type	**General Form**	**Example**	**Nature & Applications**
Combination	A + B → AB	2Mg+O2→2MgO2Mg + O_2 → 2MgO	Exothermic; combustion, industrial synthesis
Decomposition	AB → A + B	CaCO3→CaO+CO2CaCO_3 → CaO + CO_2	Endothermic; lab prep of gases, photography,

			electrolysis
Displacement	A + BC → AC + B	Zn+CuSO4→ZnSO4+CuZn + $CuSO_4 \rightarrow ZnSO_4 + Cu$	Depends on reactivity; metal extraction, galvanization

Evaluation

1. Define a chemical reaction and list **four observable signs** that indicate a chemical reaction has taken place.
2. Write balanced equations for:
   a. Burning of magnesium in oxygen.
   b. Thermal decomposition of calcium carbonate.
   c. Displacement of copper from copper (II) sulphate by iron.
3. Distinguish between:
   a. Combination and decomposition reactions (with examples).
   b. Thermal and photolytic decomposition.
4. Why is the reaction between zinc and copper (II) sulphate possible, but the reverse is not?
5. Mention one industrial process that involves:
   a. Combination reaction
   b. Decomposition reaction
   c. Displacement reaction

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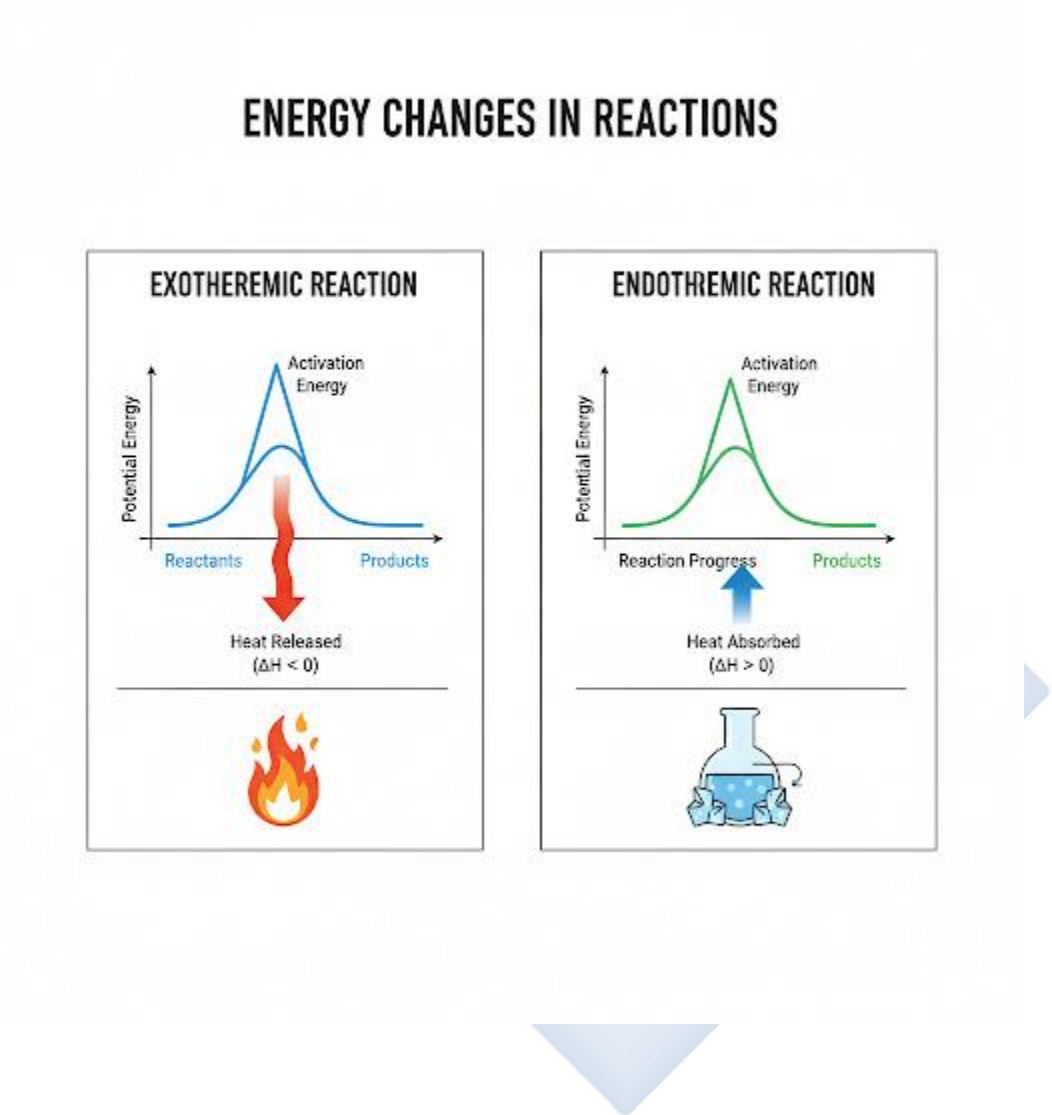
Assignment

1. Classify the following reactions and give reasons for your classification:
   a. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
   b. $2H_2O(l) \xrightarrow{electricity} 2H_2(g) + O_2(g)$
   c. $Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$
   d. $2KClO_3(s) \xrightarrow{MnO_2,\Delta} 2KCl(s) + 3O_2(g)$
   e. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
2. Explain why decomposition reactions are generally **endothermic** while combination reactions are usually **exothermic**.
3. State two real-life applications each of:
   - Combination reactions
   - Decomposition reactions
   - Displacement reactions
4. Research Task: Draw and explain the **reactivity series of metals**. Use it to predict which of the following reactions will occur spontaneously:
   a. $Fe + CuSO_4$
   b. $Cu + ZnSO_4$
   c. $Mg + HCl$

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WEEK 8: ENERGY CHANGES IN REACTIONS

Endothermic and exothermic



INTRODUCTION

All chemical reactions involve **changes in energy**. Energy can exist in various forms such as:

- **Heat (thermal energy)**
- **Light**
- **Sound**
- **Electrical energy**

During a reaction, energy can either be **released to the surroundings** or **absorbed from the surroundings**. The direction of energy flow determines whether a reaction is **exothermic** or **endothermic**.

Understanding energy changes is crucial because:

- It helps predict whether a reaction will **occur spontaneously**.
- It explains **temperature changes** in the surroundings.
- It underpins the use of **chemical energy** in industries, biology, and everyday life.

DEFINITIONS

- **Exothermic Reaction:** A chemical reaction that **releases heat energy** (or other forms of energy) to the surroundings. The temperature of the surroundings **increases**.
    - **Example:** Combustion of fuels – $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + heat.
- **Endothermic Reaction:** A chemical reaction that **absorbs heat energy** (or other forms of energy) from the surroundings. The temperature of the surroundings **decreases**.
    - **Example:** Thermal decomposition of calcium carbonate – $CaCO_3 \rightarrow CaO + CO_2$ (heat absorbed).

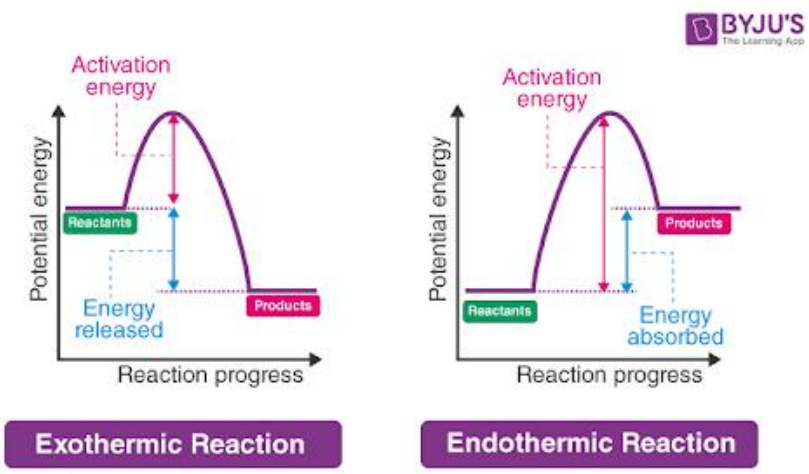
PRINCIPLES OF ENERGY CHANGES IN CHEMICAL REACTIONS

1. **Law of Conservation of Energy:** Energy cannot be created or destroyed; it can only be **transferred** from one form to another.
    - In chemical reactions, energy is transferred **between the reacting system and its surroundings**.
2. **Enthalpy (ΔH):**
    - Enthalpy is the **total heat content** of a system at constant pressure.
    - **$\Delta H < 0$:** Exothermic reaction (energy released).

- **ΔH > 0:** Endothermic reaction (energy absorbed).

3. **Activation Energy (Ea):**
    - Even exothermic reactions require a **minimum energy** to initiate the reaction.
    - This is called the **activation energy**, necessary to break bonds in reactants before new bonds can form in products.

ENERGY PROFILE



(a) Exothermic Reaction

Description:

- Reactants have **higher energy** than products.
- Energy is released as the reaction progresses.
- **ΔH is negative**.

Example: Combustion of methane – $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + heat.

(b) Endothermic Reaction

Description:

- Reactants have **lower energy** than products.

- Energy is absorbed from surroundings.
- **ΔH is positive**.

Example: Thermal decomposition of calcium carbonate – $CaCO_3 \rightarrow CaO + CO_2$.

CHARACTERISTICS OF EXOTHERMIC REACTIONS

1. Heat energy is **released to the surroundings**.
2. Surroundings experience a **temperature increase**.
3. Enthalpy change (ΔH) is **negative**.
4. Products are **more stable** than reactants.
5. Occurs spontaneously if sufficient activation energy is provided.
6. **Examples:**
    - Combustion of fuels (CH_4, C_3H_8, petrol)
    - Respiration ($C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ + energy)
    - Neutralization reactions ($HCl + NaOH \rightarrow NaCl + H_2O$ + heat)
    - Freezing of water (exothermic phase change)

CHARACTERISTICS OF ENDOTHERMIC REACTIONS

1. Heat energy is **absorbed from the surroundings**.
2. Surroundings experience a **temperature decrease**.
3. Enthalpy change (ΔH) is **positive**.
4. Products are **less stable** than reactants.
5. Often **non-spontaneous** without continuous energy supply.
6. **Examples:**

- Photosynthesis: $6CO_2 + 6H_2O + \text{sunlight} \rightarrow C_6H_{12}O_6 + 6O_2$
- Thermal decomposition of calcium carbonate: $CaCO_3 \rightarrow CaO + CO_2$
- Dissolution of ammonium chloride in water
- Electrolysis of water: $2H_2O + \text{energy} \rightarrow 2H_2 + O_2$

PRACTICAL DEMONSTRATIONS

Reaction	**Observation**	**Type**
Dissolving NaOH in water	Beaker becomes hot	Exothermic
Dissolving NH_4Cl in water	Beaker becomes cold	Endothermic
Combustion of candle	Flame, heat released	Exothermic
Thermal decomposition of $KClO_3$	Oxygen released, requires heating	Endothermic

REAL-LIFE APPLICATIONS

(a) Industrial

- **Exothermic:** Combustion for energy, neutralization in manufacturing processes, Haber process (ammonia production).
- **Endothermic:** Electrolysis for metal extraction, thermal cracking of hydrocarbons.

(b) Biological

- **Exothermic:** Respiration supplies energy for life processes.
- **Endothermic:** Photosynthesis stores solar energy in chemical bonds.

(c) Daily Life

- **Exothermic:** Burning fuels for cooking and heating; fireworks.
- **Endothermic:** Cold packs for injuries, melting ice.

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COMPARISON OF EXOTHERMIC AND ENDOTHERMIC REACTIONS

Feature	Exothermic	Endothermic
Heat energy flow	Released to surroundings	Absorbed from surroundings
Effect on surroundings	Temperature rises	Temperature falls
ΔH (enthalpy change)	Negative	Positive
Stability of products	More stable than reactants	Less stable than reactants
Occurrence	Often spontaneous	Often requires continuous energy input
Activation energy	Usually lower than energy released	Often higher than energy absorbed
Examples	Combustion, neutralization, respiration	Photosynthesis, decomposition, electrolysis

EVALUATION

1. Define exothermic and endothermic reactions with examples.
2. State five characteristics each of exothermic and endothermic reactions.
3. Draw and label energy profile diagrams for one exothermic and one endothermic reaction.
4. Explain the difference between ΔH in exothermic and endothermic reactions.
5. Describe two real-life examples of exothermic reactions and two of endothermic reactions.
6. Why is photosynthesis considered an endothermic reaction?

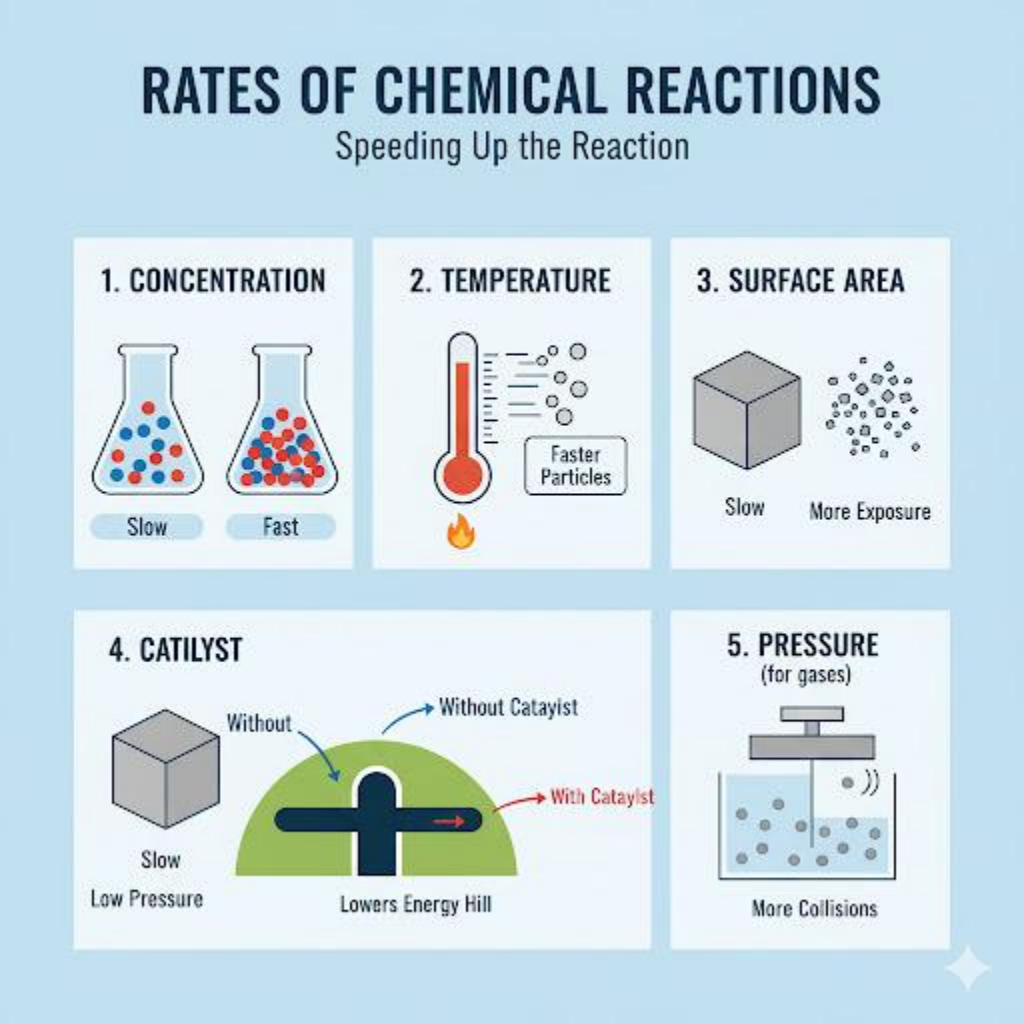
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ASSIGNMENT

1. Explain, with equations, why dissolving NaOH in water is exothermic while dissolving NH_4Cl in water is endothermic.
2. List six examples each of exothermic and endothermic reactions in daily life and industry.
3. A student performed the reaction $CaCO_3 \rightarrow CaO + CO_2$ using a bunsen burner.
    - (a) Identify the type of reaction.
    - (b) Explain the energy change involved.
    - (c) Sketch the energy profile diagram for the reaction.

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WEEK 9: RATES OF CHEMICAL REACTIONS – FACTORS AFFECTING RATE



Concept of Rate of Reaction

- **Definition:** The rate of a chemical reaction is the measure of how fast reactants are converted into products in a chemical process.
- It tells us **how much reactant is used up** or **how much product is formed** in a given time.
- Mathematically:

$$\text{Rate} = \frac{\Delta[\text{Reactant}]}{\Delta t} = \frac{\Delta[\text{Product}]}{\Delta t}$$

where Δ = change, [] = concentration, t = time.

- **Units of rate:**
  - mol dm^{-3} s^{-1} (for concentration changes)
  - cm^3 s^{-1} (for volume of gas evolved)
  - g s^{-1} (for mass change)

Ways of Measuring Rate of Reaction

- **Gas evolution:** Using gas syringe or inverted burette (e.g., Mg + HCl).
- **Mass change:** Place reaction vessel on a balance and record loss in mass (e.g., $CaCO_3$ + HCl → CO_2).
- **Colour change:** Observe using colorimeter or naked eye (e.g., iodine clock reaction).
- **Turbidity/Precipitate formation:** Measure time for a mark to disappear through cloudy solution ($Na_2S_2O_3$ + HCl → S).

Factors Affecting Rate of Reaction

(i) Nature of Reactants

- Bond type and bond strength affect rate.
- Ionic reactions (in solution) are faster than covalent reactions (involving bond breaking).
- Example:
  - Neutralization: HCl + NaOH → NaCl + H_2O (very fast).
  - Reaction between methane and chlorine (CH_4 + Cl_2) requires UV light (slow without).

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(ii) Concentration of Reactants

- Higher concentration = more particles per unit volume = greater collision frequency = faster reaction.
- Example:
    - 2M HCl reacts faster with Mg ribbon than 0.5M HCl.
- **Graphical Representation:** Steeper slope of volume vs time curve for higher concentration.

(iii) Temperature

- Higher temperature increases kinetic energy of particles.
- More particles have energy ≥ **activation energy (E_a)**.
- Increases both **frequency and effectiveness** of collisions.
- Rule: Rate approximately doubles for every 10°C rise in temperature.
- Example: Reaction between $Na_2S_2O_3$ and HCl is much faster at 60°C than 25°C.

(iv) Pressure (for gaseous reactions)

- Increasing pressure reduces gas volume → more molecules per unit volume → more collisions.
- Example: Haber process:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

works best at **high pressure**.

(v) Surface Area of Reactants

- Reaction occurs at surfaces. Finely divided solids expose more area for collision.

- Example:
    - Powdered $CaCO_3$ reacts faster with HCl than marble chips.
    - Sawdust burns faster than logs.

(vi) Presence of a Catalyst

- **Definition:** A catalyst is a substance that increases the rate of reaction without being consumed or permanently changed.
- **Mechanism:** Provides an **alternative pathway** with lower activation energy.
- **Examples:**
    - MnO_2 catalyzes decomposition of $H_2O_2 \rightarrow O_2$.
    - Fe catalyzes Haber process (NH_3 synthesis).
    - Ni catalyzes hydrogenation of vegetable oils.

Types of Catalysis:

1. **Homogeneous Catalysis:** Catalyst in the same phase as reactants.
    - Example: $2SO_2 + O_2 \xrightarrow{NO} 2SO_3$.
2. **Heterogeneous Catalysis:** Catalyst in a different phase.
    - Example: Fe in Haber process.
3. **Autocatalysis:** One of the products acts as catalyst.
    - Example: Hydrolysis of esters catalyzed by H^+ ions produced.

(vii) Light (Photochemical Reactions)

- Certain reactions are activated by light energy.
- Examples:
    - Photosynthesis: $6CO_2 + 6H_2O \xrightarrow{light/chlorophyll} C_6H_{12}O_6 + 6O_2$.
    - Decomposition of silver halides in photography: $2AgCl \xrightarrow{light} 2Ag + Cl_2$.

Collision Theory of Chemical Reactions

- For reaction to occur:
    1. Particles must collide.
    2. Collisions must be **effective** (with correct orientation and sufficient energy).
    3. The energy required to initiate reaction is called **activation energy (E_a)**.

Energy Profile Diagram:

- Without catalyst: High activation energy barrier.
- With catalyst: Lower activation energy, faster reaction.

Experimental Demonstrations

1. **Effect of Concentration:** Magnesium + HCl with varying molarity; measure H_2 evolution.
2. **Effect of Temperature:** $Na_2S_2O_3$ + HCl at different temperatures; time disappearance of cross mark.
3. **Effect of Catalyst:** Decomposition of H_2O_2 with MnO_2; rapid oxygen evolution.
4. **Effect of Surface Area:** $CaCO_3$ powder vs marble chips in HCl.

Industrial Relevance of Rate of Reaction

- **Haber Process:** Ammonia production (Fe catalyst, 200 atm, 450°C).
- **Contact Process:** Sulphuric acid production (V_2O_5 catalyst, 450°C).
- **Hydrogenation of oils:** Manufacture of margarine (Ni catalyst).
- **Polymerization:** Catalysts control rate and yield of plastics.

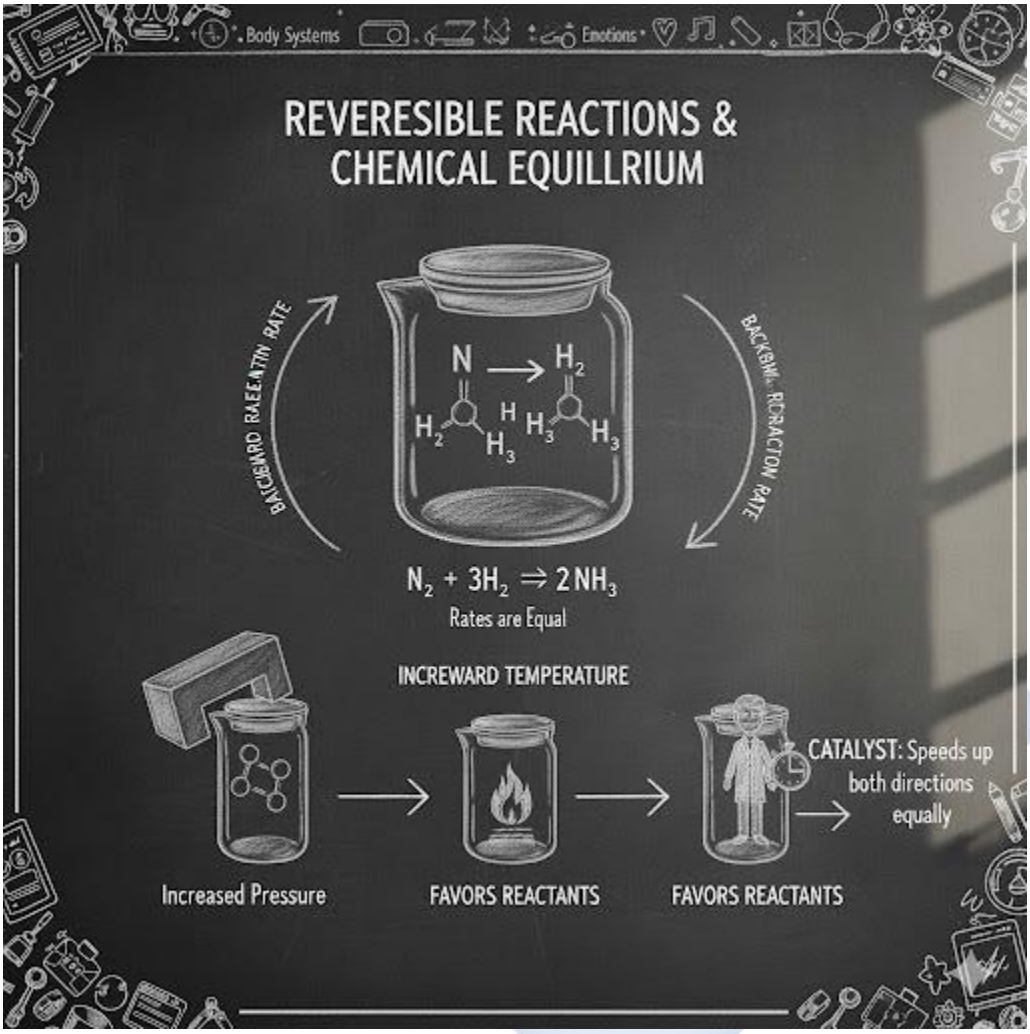
Evaluation

1. Define rate of reaction.
2. State five factors affecting rate of chemical reaction.
3. Explain collision theory and its relation to activation energy.
4. Why does powdered zinc react faster with dilute HCl than zinc granules?
5. Draw and explain an energy profile diagram for a catalyzed and uncatalyzed reaction.
6. Give two examples of photochemical reactions.
7. Why is high pressure used in the Haber process?

Assignment

1. Differentiate between homogeneous and heterogeneous catalysis with examples.
2. Explain why reactions between ionic compounds in aqueous solution are generally faster than reactions involving covalent compounds.
3. Write a detailed report of an experiment to show how concentration affects rate of reaction.
4. Research: Sketch an energy profile diagram showing activation energy with and without a catalyst.
5. Discuss four ways rate of reaction is applied in industry.

WEEK 10: REVERSIBLE REACTIONS AND CHEMICAL EQUILIBRIUM – LE CHATELIER'S PRINCIPLE



REVERSIBLE AND IRREVERSIBLE REACTIONS

(a) Irreversible Reactions

- Irreversible reactions proceed **only in one direction**, from reactants to products, and cannot revert under normal conditions.
- Once completed, the reactants are fully converted to products.

Examples:

1. Combustion of methane:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

2. Neutralization of acids and bases:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

(b) Reversible Reactions

- Reversible reactions can proceed in **both forward and backward directions** under the same conditions.
- Represented with a **double arrow (⇌)**.

Examples:

1. Formation of ammonia (Haber process):

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

2. Thermal decomposition of calcium carbonate:

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

- **Observation:** In a closed system, the products can recombine to form reactants.

CHEMICAL EQUILIBRIUM

- When a reversible reaction is carried out in a **closed system**, the reaction proceeds in both directions until a **dynamic equilibrium** is reached.
- At equilibrium:
    1. **Rate of forward reaction = Rate of backward reaction**
    2. **Concentrations of reactants and products remain constant**
    3. The system is **dynamic**: molecules continue to react, but overall composition does not change

Example:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Initially, formation of NH_3 is fast, but as NH_3 accumulates, the backward decomposition becomes significant. Eventually, the rates equalize.

FACTORS AFFECTING CHEMICAL EQUILIBRIUM

Le Chatelier's Principle:

If a system at equilibrium is disturbed by a change in **concentration, temperature, or pressure**, the system shifts to counteract the change and restore equilibrium.

(a) Effect of Concentration

- Increasing the concentration of **reactants** shifts equilibrium toward **products**.
- Increasing the concentration of **products** shifts equilibrium toward **reactants**.

Example:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

- Adding more H_2 shifts equilibrium → **right** (more NH_3).
- Adding NH_3 shifts equilibrium → **left** (more $N_2 + H_2$).

(b) Effect of Pressure (gaseous reactions)

- Increasing pressure favors the side with **fewer moles of gas**.
- Decreasing pressure favors the side with **more moles of gas**.

Example:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

- Left: 4 moles of gas; Right: 2 moles
- Increasing pressure → equilibrium shifts **right** (more NH_3).
- Decreasing pressure → equilibrium shifts **left** (more $N_2 + H_2$).

(c) Effect of Temperature

- For **exothermic reactions** (release heat):
  - Increasing temperature → equilibrium shifts **backward** (toward reactants).
  - Decreasing temperature → equilibrium shifts **forward** (toward products).
- For **endothermic reactions** (absorb heat):
  - Increasing temperature → equilibrium shifts **forward** (toward products).
  - Decreasing temperature → equilibrium shifts **backward** (toward reactants).

Example:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{Heat}$$

- Forward reaction exothermic.
- Increase temperature → less NH_3 formed.
- Decrease temperature → more NH_3 formed.

(d) Effect of Catalyst

- A catalyst **speeds up both forward and backward reactions equally**.
- It allows the system to reach equilibrium faster but **does not change the position of equilibrium**.

EQUILIBRIUM CONSTANT (Kc)

- At equilibrium, the ratio of **product concentrations** to **reactant concentrations**, raised to the power of their coefficients, is constant.

$$K_c = \frac{[Products]^p}{[Reactants]^r}$$

Example:

For the reaction:

$$aA + bB \rightleftharpoons cC + dD$$

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

- A **large Kc** → products favored

  A **small Kc** → reactants favored

Example Calculation:

If at equilibrium, $[NH_3]$ = 0.2 M, $[N_2]$ = 0.1 M, $[H_2]$ = 0.3 M in:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{0.2^2}{0.1 \times 0.3^3} = \frac{0.04}{0.1 \times 0.027} = \frac{0.04}{0.0027} \approx 14.81$$

INDUSTRIAL APPLICATIONS

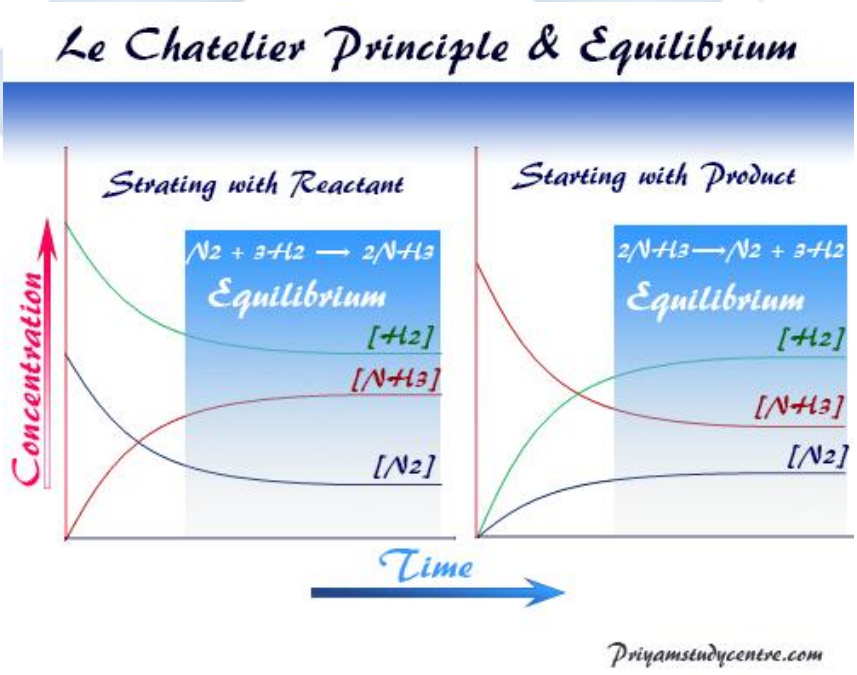
1. **Haber Process (NH_3 production):**
   - $N_2 + 3H_2 \rightleftharpoons 2NH_3$ + Heat
   - High pressure (200 atm) favors fewer gas molecules → more NH_3
   - Moderate temperature (450 °C) compromises yield & rate

- Iron catalyst with promoters increases rate

2. **Contact Process (SO_3 production):**
   - $2SO_2 + O_2 \rightleftharpoons 2SO_3$ $\Delta H = -197$ kJ/mol
   - High pressure & low temperature favors forward reaction
   - Vanadium(V) oxide catalyst speeds up reaction
3. **Food & Biological Systems:**
   - Dissolved gases in blood (O_2, CO_2) obey equilibrium principles.
   - Le Chatelier's principle explains gas exchange in lungs.

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1. **Graph of Concentration vs Time:**



   - Forward reaction: steep rise in products, then plateau at equilibrium
   - Backward reaction: products decreasing initially, then plateau at equilibrium
2. **Equilibrium Shifts:**

- Arrows showing movement toward products (right) or reactants (left) depending on concentration, pressure, or temperature changes.

EVALUATION

1. Differentiate between reversible and irreversible reactions with examples.
2. Define chemical equilibrium and state four characteristics.
3. Explain Le Chatelier's principle using $N_2 + 3H_2 \rightleftharpoons 2NH_3$.
4. A reaction is exothermic in the forward direction. How will increasing temperature affect equilibrium? Explain.
5. Why does a catalyst not change the position of equilibrium?

ASSIGNMENT

1. A system at equilibrium has the reaction:

$$CO + 2H_2 \rightleftharpoons CH_3OH$$

Explain how the yield of methanol will change if:

- a) Pressure is increased
- b) Temperature is decreased
- c) A catalyst is added

2. Calculate the equilibrium constant (Kc) for the reaction:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

if at equilibrium, $[NH_3] = 0.2$ M, $[N_2] = 0.1$ M, $[H_2] = 0.3$ M.

3. Explain how Le Chatelier's principle is applied in the **Contact Process** to optimize SO_3 production.

WEEK 11: METALS AND NON-METALS – PROPERTIES AND USES



Introduction

- **Metals and non-metals** are two broad categories of elements found on the **Periodic Table**.
- They are classified based on their **physical and chemical properties**.
- **Metals** are generally found on the **left and center** of the Periodic Table, while **non-metals** are found on the **right side**.
- Understanding their properties is crucial because they form the building blocks of materials used in **industry, agriculture, construction, medicine, and technology**.

Metals

Metals are elements that readily **lose electrons** to form **positive ions (cations)** and exhibit **metallic bonding**.

General Physical Properties of Metals:

1. **Malleability:** Can be hammered or rolled into sheets (e.g., aluminium foil).
2. **Ductility:** Can be drawn into wires (e.g., copper wires).
3. **Good Conductors of Heat and Electricity:** Due to free movement of delocalized electrons.
4. **High Melting and Boiling Points:** Most metals, except mercury, are solid at room temperature.
5. **High Density:** Metals like iron, copper, and gold are heavy.
6. **Metallic Lustre:** Shiny appearance when polished.
7. **Sonority:** Produce ringing sounds when struck.
8. **Strength:** Many metals (iron, steel) are strong and durable.
9. **Tensile Strength:** Ability to resist breaking under tension.
10. **Alloy Formation:** Metals can be mixed to form alloys (e.g., brass, bronze, steel).

General Chemical Properties of Metals:

1. **Electropositivity:** Metals lose electrons easily to form positive ions.
2. **Reaction with Oxygen:** Metals form basic oxides, e.g., $4Na + O_2 \rightarrow 2Na_2O$.
3. **Reaction with Water:** Metals like sodium react vigorously; some, like calcium, react slowly; metals like gold are unreactive.

4. **Reaction with Acids:** Most metals (except Cu, Hg, Ag, Au, Pt) react with dilute acids to form salts and hydrogen gas.
5. **Reaction with Chlorine:** Metals readily form chlorides, e.g., $2Na + Cl_2 \rightarrow 2NaCl$.
6. **Form Ionic Compounds:** Metals form ionic bonds by donating electrons.

Non-Metals

Non-metals are elements that tend to **gain or share electrons** to form **negative ions (anions)** or covalent compounds.

General Physical Properties of Non-Metals:

1. **Poor Conductors:** Generally insulators (except graphite).
2. **Low Melting and Boiling Points:** Many are gases or brittle solids at room temperature.
3. **Non-Malleable and Non-Ductile:** They shatter when hammered.
4. **Low Density:** Non-metals like sulfur and phosphorus are light.
5. **No Metallic Lustre:** Usually dull in appearance.
6. **Brittle:** Break easily when struck.
7. **Existence in Different States:** Some non-metals are gases (oxygen), liquids (bromine), or solids (sulfur).
8. **Allotropes:** Some non-metals exist in different physical forms, e.g., carbon (diamond, graphite), oxygen (O_2, O_3).

General Chemical Properties of Non-Metals:

1. **Electronegativity:** Tend to gain or share electrons.
2. **Acidic Oxides:** Non-metals form acidic oxides, e.g., $CO_2 + H_2O \rightarrow H_2CO_3$.
3. **Reaction with Hydrogen:** Forms covalent compounds (HCl, CH_4).
4. **Reaction with Metals:** Form ionic compounds (e.g., NaCl).

5. **Reactivity Trend:** Increases across a period (left to right on the Periodic Table).

Metalloids

- Metalloids are elements with **properties of both metals and non-metals**.
- Examples: Boron, Silicon, Arsenic, Antimony, Tellurium.
- They are often **semiconductors** (used in electronics).

Differences Between Metals and Non-Metals

Property	**Metals**	**Non-Metals**
Position on Periodic Table	Left and center	Right side
Electroconductivity	Good	Poor (except graphite)
Appearance	Lustrous, shiny	Dull
State at Room Temp	Solid (except mercury)	Solid, liquid, or gas
Malleability/Ductility	Malleable and ductile	Brittle
Type of Oxides	Basic or amphoteric	Acidic
Bonding	Metallic/Ionic	Covalent
Electronegativity	Low	High

Extraction of Metals

- Metals are obtained from **ores** (naturally occurring rocks containing metals).
- Methods include:
    1. **Electrolysis** (e.g., aluminium extraction from bauxite).

2. **Reduction with Carbon** (e.g., iron from haematite).
3. **Roasting and Smelting** for sulfide ores.

Uses of Metals

1. **Iron:** Construction, tools, machinery.
2. **Copper:** Electrical wiring, plumbing.
3. **Aluminium:** Aircraft, packaging, cooking utensils.
4. **Gold and Silver:** Jewelry, electronics.
5. **Zinc:** Galvanizing, batteries.
6. **Lead:** Car batteries, shielding from radiation.
7. **Titanium:** Aerospace, medical implants.

Uses of Non-Metals

1. **Oxygen:** Breathing, welding, steel production.
2. **Nitrogen:** Fertilizers, refrigeration, explosives.
3. **Sulfur:** Sulfuric acid, vulcanization of rubber.
4. **Chlorine:** Disinfectant, PVC production.
5. **Carbon:** Fuel, steelmaking, lubricants (graphite), gemstones (diamond).
6. **Phosphorus:** Matches, fertilizers, detergents.

Environmental and Industrial Importance

- Metals and non-metals play vital roles in **technological development**, **transportation**, **medicine**, **communication**, **energy generation**, and **construction**.

- Recycling metals reduces **environmental pollution** and conserves resources.
- Proper disposal of non-metallic wastes (plastics, toxic gases) is crucial for sustainable development.

Evaluation

1. Define metals and non-metals, giving two examples of each.
2. List five **physical properties of metals** and five of non-metals.
3. Explain why metals are good conductors of electricity.
4. Differentiate between **acidic oxides** and **basic oxides**.
5. State three uses each of:
    - (a) Copper
    - (b) Sulfur
6. What is a metalloid? Give two examples and uses.
7. Why are some metals extracted using electrolysis?

Assignment:

1. Draw a **table** showing at least **six differences between metals and non-metals**.
2. Outline the **steps in the extraction of aluminium** from bauxite.
3. Identify five industries that rely heavily on metals.
4. Discuss the **role of non-metals** in agriculture and medicine.
5. Prepare a chart summarizing **10 metals, their sources, properties, and uses**.

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