CHEMISTRY

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1. Atomic Structure

Atomic Orbitals (Shapes and energies of s, p, d, f orbitals, nodal surfaces)

- Quantum Numbers:
  - Principal Quantum Number (n): Determines the main energy level and size of the orbital. Values: 1, 2, 3...
  - Azimuthal/Angular Momentum Quantum Number (l): Determines the shape of the orbital and subshell. Values: 0 to n-1.
    - l=0: s orbital (spherical)
    - l=1: p orbital (dumbbell)
    - l=2: d orbital (double dumbbell, except d_{z^2})
    - l=3: f orbital (complex)
  - Magnetic Quantum Number (m_l): Determines the orientation of the orbital in space. Values: -l to +l.
- Orbital Shapes:
  - s-orbitals: Spherically symmetric. Example: 1s, 2s.
  - p-orbitals: Dumbbell-shaped along axes. Three orientations: p_x, p_y, p_z.
  - d-orbitals: Five orientations. Four are double-dumbbell (d_{xy}, d_{yz}, d_{xz}, $d_{x^2-y^2}$) and one is dumbbell with a collar (d_{z^2}).
  - f-orbitals: Seven complex orientations.
- Orbital Energies (Multi-electron atoms):
  - Energy depends on both 'n' and 'l'.
  - For a given principal quantum number 'n', the order of increasing energy is: s < p < d < f (due to varying penetration and shielding effects).
  - General filling order (Aufbau principle, (n+l) rule): 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p.
- Nodal Surfaces (Nodes): Regions where the probability of finding an electron is zero.
  - Total number of nodes = n - 1.
  - Angular Nodes (Nodal Planes): Equal to 'l'. These define regions where the probability density is zero due to orbital shape/orientation.
    - For p-orbitals (l=1): 1 angular node.

        - For d-orbitals (l=2): 2 angular nodes.
        - d_{xy} orbital: Has 2 nodal planes (xz and yz planes).
    - Radial Nodes (Spherical Nodes): Equal to n - l - 1. These define spherical regions of zero probability.
    - For a 4d orbital (n=4, l=2):
        - Angular nodes = l = 2.
        - Radial nodes = n - l - 1 = 4 - 2 - 1 = 1.
- Stability of Electron Configurations: Atoms achieve extra stability with half-filled or completely-filled subshells.
    - Europium (Eu, Z=63): Electronic configuration is [Xe] $4f^7$ $6s^2$ (half-filled 4f orbital).
    - Gadolinium (Gd, Z=64): Electronic configuration is [Xe] $4f^7$ $5d^1$ $6s^2$ (half-filled 4f orbital, with an additional electron in 5d for stability).
- Pi (π) bond: Formed by the lateral (sideways) overlap of parallel p-orbitals.

Atomic Spectra (Origin, Types, Fine Structure - brief)

- Origin of Atomic Spectra: Arises from electron transitions between discrete energy levels within an atom. When an electron gains energy, it moves to a higher energy state; when it returns to a lower state, it emits electromagnetic radiation (photon) with energy `E = hν = E*final - Einitial*`.
- *Emission Spectra: Produced when excited atoms release energy by emitting photons as electrons fall to lower energy levels. Appears as bright lines on a dark background, unique for each element.*
- *Absorption Spectra: Produced when atoms absorb specific wavelengths of incident radiation, causing electrons to jump to higher energy levels. Appears as dark lines on a bright background at the same wavelengths as the emission lines.*
- *Line Spectra: Characteristic of individual atoms in the gaseous state, consisting of distinct, discrete lines due to quantized energy levels.*
- *Types of Transitions & Spectroscopic Regions:*
    - *Electronic Transitions: Involve electron jumps between different electronic energy levels. High energy. Occur primarily in the Ultraviolet (UV) and Visible regions of the electromagnetic spectrum. (Relevant to UV spectrometer)*
    - *Vibrational Transitions: Involve changes in molecular vibrational energy levels (not typically for isolated atoms, but relevant for comparison). Medium energy. Occur in the Infrared (IR) region.*
    - *Rotational Transitions: Involve changes in molecular rotational energy levels (not for isolated atoms). Low energy. Occur in the Microwave and Far-IR regions.*
- *Fine Structure: The splitting of main spectral lines into multiple, closely spaced lines. Primarily caused by spin-orbit coupling, which is the interaction between the electron's spin angular momentum and its orbital angular momentum. Relativistic corrections also contribute.*

Electronic Configuration Principles (Aufbau Principle, Pauli's Exclusion Principle, Hund's Rule)

- *Electronic Configuration Principles: Rules governing the distribution of electrons in atomic orbitals.*
- *1. Aufbau Principle:*
    - *Electrons occupy atomic orbitals in order of increasing energy.*
    - *Lowest energy orbitals are filled first.*
    - *General order of filling: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p.*
- *2. Pauli's Exclusion Principle:*
    - *No two electrons in an atom can have all four quantum numbers identical.*
    - *An atomic orbital can hold a maximum of two electrons, and these two electrons must have opposite spins (↑↓).*

- *3. Hund's Rule of Maximum Multiplicity:*
  - *For degenerate orbitals (orbitals of the same energy level, e.g., p, d, f subshells), electron pairing in an orbital does not occur until all orbitals of that subshell are singly occupied.*
  - *All singly occupied electrons in degenerate orbitals must have parallel spins.*
- *Electron Shell Capacities ($2n^2$ Rule): The maximum number of electrons a main energy shell (n) can hold is `$2n^2$`.*
  - *K-shell (n=1): Max 2 electrons (2^1).*
  - *L-shell (n=2): Max 8 electrons (2^2).*
  - *M-shell (n=3): Max 18 electrons (2^3).*
  - *N-shell (n=4): Max 32 electrons (2^4).*
- *Valence Electrons:*
  - *Electrons located in the outermost main energy shell of an atom.*
  - *Primarily determine an atom's chemical properties and reactivity.*
  - *For an element with Z=35 (Bromine), configuration is [Ar] $3d^{10}$ $4s^2$ $4p^5$. The outermost shell is n=4, thus valence electrons = 2 (from 4s) + 5 (from 4p) = 7.*

Electronic Configurations of Multi-electron Atoms and Ions

- *Electronic Configuration (EC) describes the distribution of electrons in an atom's orbitals.*
- *Aufbau Principle: Electrons fill atomic orbitals of the lowest energy first.*
- *Pauli Exclusion Principle: An orbital holds a maximum of two electrons with opposite spins.*
- *Hund's Rule of Maximum Multiplicity: For degenerate orbitals, electrons fill each orbital singly with parallel spins before pairing up.*
- *Shells and Subshells:*
  - *Shells (n=1, 2, 3...) or (K, L, M...). Maximum electrons in a shell (n) = $2n^2$. E.g., M shell (n=3) accommodates 18 electrons. (MCQ 11)*
  - *Subshells: s (max $2e^-$), p (max $6e^-$), d (max $10e^-$), f (max $14e^-$).*
- *Valence Electrons: Electrons in the outermost shell. They determine an element's chemical properties and valency. (MCQ 1, 8, 25, 59)*
- *Chemical Properties: Primarily determined by the number of valence electrons. (MCQ 1, 8, 25)*
- *Valency: Combining capacity, based on electrons gained, lost, or shared to achieve a stable octet/duet.*
  - *Silicon (Z=14): Valency 4.*
  - *Fluorine (Z=9): Valency 1.*
  - *Aluminium (Z=13): Valency 3.*
  - *Sulphur (Z=16): Valency 2.*
  - *Nitrogen (Z=7): Valency 3. (MCQ 33)*
  - *Magnesium (Z=12): Valency 2.*
  - *Sodium (Z=11): Valency 1.*
  - *Oxygen (Z=8): Valency 2. (Gains 2 electrons to form an octet; MCQ 30)*
  - *Neon (Z=10): Valency 0. (MCQ 9, 27, 28)*
  - *In CaO, Ca atom transfers its two valence electrons to the oxygen atom. (MCQ 23)*
- *Stability: Half-filled (e.g., d^5, p^3) and fully-filled (e.g., d^{10}, p^6, s^2) subshells confer extra stability.*
- *Exceptions to Aufbau Principle:*
  - *Chromium (Cr, Z=24): [Ar] $3d^5$ $4s^1$ (not [Ar] $3d^4$ $4s^2$) due to half-filled d-subshell stability. (MCQ 15)*
  - *Copper (Cu, Z=29): [Ar] $3d^{10}$ $4s^1$ (not [Ar] $3d^9$ $4s^2$) due to fully-filled d-subshell stability. (MCQ 14, 62)*
  - *Vanadium (V, Z=23): [Ar] $3d^3$ $4s^2$. (MCQ 49)*
- *Unpaired Electrons: Electrons occupying an orbital singly. For $1s^2 2s^2 2p^5$ (Fluorine, Z=9), there is 1 unpaired electron. (MCQ 34)*
- *Ions: Formed by gaining (anions) or losing (cations) electrons.*
  - *Cations: Smaller than parent atom. (MCQ 57)*
  - *Anions: Larger than parent atom.*
  - *Example: Cl (Z=17, $17e^-$). Cl^- has 18 electrons (7 valence + 1 gained = 8 valence electrons). (MCQ 10, 38)*
  - *Example: An ion with +3 charge, mass number 27, 14 neutrons: Protons = 27-14=13. Neutral atom has 13 elec-*

trons. Ion has 13-3=10 electrons. (MCQ 36)

- Isoelectronic Species: Atoms or ions with the same total number of electrons.
  - Cl^- ($18e^-$) is isoelectronic with K^+ ($18e^-$), S^{2-} ($18e^-$), P^{3-} ($18e^-$). (MCQ 6, 12, 26)
  - O^{2-} ($10e^-$) is isoelectronic with F^- ($10e^-$), Ne ($10e^-$), Na^+ ($10e^-$), Mg^{2+} ($10e^-$), Al^{3+} ($10e^-$). (MCQ 26, 35, 63)
  - N^- ($8e^-$) is NOT isoelectronic with F^- ($10e^-$). (MCQ 32)
  - S ($16e^-$) is NOT isoelectronic with Cl^- ($18e^-$). (MCQ 29)
- Atomic Number (Z): The most fundamental characteristic of an element. (MCQ 31)
- Atomic Radius (AR):
  - Decreases across a period (increasing effective nuclear charge).
  - Increases down a group (increasing number of shells). (MCQ 22)
  - Order of AR: Na > Li > Be > O. (MCQ 3)
  - Order of AR: C < B < Si < Al. (MCQ 13)
  - Correct comparisons: Mg > Be, I > Br, K > Na. Rb < Cs (Cs has larger radius than Rb). (MCQ 42)
  - Ionic Radii (Isoelectronic Series): For isoelectronic species, radius decreases with increasing positive nuclear charge (Z). E.g., $Mg^{2+} < Na^+ < F^- < O^{2-}$ (all $10e^-$). (MCQ 41)
  - Lanthanide Contraction: Gradual decrease in atomic/ionic radii across lanthanides due to poor shielding by 4f electrons. Order of increasing ionic size: $Yb^{3+} < Pm^{3+} < Ce^{3+} < La^{3+}$. (MCQ 50)
- Ionization Energy (IE): Energy required to remove an electron from a gaseous atom/ion.
  - Generally increases across a period and decreases down a group.
  - Exceptions: IE_1 of Be ([He] $2s^2$) > IE_1 of B ([He] $2s^22p^1$) due to stable $2s^2$; IE_1 of N ([He] $2s^22p^3$) > IE_1 of O ([He] $2s^22p^4$) due to stable half-filled $2p^3$.
  - Order for 2nd period elements (IE_1): Li < B < Be < C < O < N < F. (MCQ 48)
  - Successive IE: $IE_1 < IE_2 < IE_3$... (removing electrons from a positive ion requires more energy). Mg^{2+} has the highest IE among Mg, Mg^+, Mg^{2+}. (MCQ 24)
  - Highest IE: Configuration $1s^22s^22p^3$ (Nitrogen) has highest IE among options due to stable half-filled 2p subshell. (MCQ 39)
- Electron Gain Enthalpy (EGE): Energy change when an electron is added to a neutral gaseous atom.
  - Generally more negative (energy released) across a period and less negative down a group.
  - Exceptions: Noble gases have positive EGE (energy required).
  - High negative EGE: Halogens (e.g., F, Cl).
  - EGE Matching: Ne ($1s^22s^22p^6$): +48 kJ/mol; Na ($1s^22s^22p^63s^1$): -53 kJ/mol; F ($1s^22s^22p^5$): -328 kJ/mol; O ($1s^22s^22p^4$): -141 kJ/mol. (MCQ 47)
  - Assertion: Electron affinity (EGE) does NOT always increase from top to bottom in a group; it generally decreases (becomes less negative). (MCQ 37)
- Diamagnetism and Paramagnetism:
  - Paramagnetic: Contains one or more unpaired electrons (attracted to magnetic field).
  - Diamagnetic: All electrons are paired (weakly repelled by magnetic field).
  - Example (via MO theory): O_2^+, O_2^-, O_2 are paramagnetic. O_2^{2-} (18 electrons) is diamagnetic. (MCQ 56, 58)
- Bond Order (BO) from Molecular Orbital Theory (for diatomic molecules):
  - BO = ½ (Number of bonding electrons - Number of antibonding electrons).
  - Higher BO indicates greater bond strength.
  - Calculations: O_2 (BO=2.0), O_2^+ (BO=2.5), O_2^- (BO=1.5), O_2^{2-} (BO=1.0). (MCQ 53, 61)
  - Order: $O_2^+ > O_2 > O_2^- > O_2^{2-}$. (MCQ 44)
  - Transformations: N_2 (BO=3.0) -> N_2^+ (BO=2.5): BO decreases. O_2 (BO=2.0)

-> O_2^+ (BO=2.5): BO increases. (MCQ 46)
- C_2 molecule consists of only pi (π) bonds. (MCQ 51)

Historical Development of Atomic Models (Dalton, Thomson, Rutherford)

- Dalton's Atomic Theory (1808):
  - Matter consists of indivisible particles called atoms.
  - Atoms of a given element are identical in mass and properties.
  - Atoms are the smallest particles that participate in chemical reactions.
  - Proposed unique symbols for elements; e.g., Phosphorus: A circle with three lines radiating from the center to the circumference.
- Thomson's Atomic Model (1904) - "Plum Pudding Model":
  - Proposed after discovery of electron (1897).
  - Atom is a uniform sphere of positive charge with negatively charged electrons embedded in it, like plums in a pudding.
  - Overall electrically neutral.
- Rutherford's Atomic Model (1911) - "Nuclear Model":
  - Gold Foil Experiment (Alpha-scattering): Most alpha particles passed through, some deflected at small angles, very few bounced back.
  - Conclusions:
    - Most of the atom's space is empty.
    - A tiny, dense, positively charged region at the center called the nucleus contains nearly all the atom's mass.
    - Electrons revolve around the nucleus in distinct orbits.
    - An atom is electrically neutral: The positive charge on the nucleus is exactly balanced by the total negative charge of the orbiting electrons.
  - Limitations: Failed to explain the stability of the atom and the discrete atomic spectra.

Isotopes, Isobars, and Isotones

1. Atomic Number (Z):
   - Number of protons in an atom's nucleus.
   - Defines the identity of an element (its most characteristic property).
   - In a neutral atom, Z = number of electrons.
2. Mass Number (A):
   - Total number of protons and neutrons (collectively called nucleons) in the nucleus.
   - A = Z + N (where N = number of neutrons).
   - Number of neutrons (N) = A - Z.
3. Atomic Particles & Charge:
   - Protons (p^+): Positive charge, negligible mass relative to nucleus, located in nucleus.
   - Neutrons (n^0): No charge (neutral), located in nucleus.
   - Electrons (e^-): Negative charge, negligible mass, orbit the nucleus.
   - A neutral atom has an equal number of protons and electrons, making it electrically neutral.
   - Ions are formed when atoms gain or lose electrons, resulting in a net charge (e.g., Fe^{3+} has 3 fewer electrons than protons).
4. Isotopes:
   - Atoms of the same element.
   - Have the same atomic number (Z) (same number of protons).
   - Have different mass numbers (A) (different number of neutrons).
   - Exhibit similar chemical properties (due to same Z and electron configuration).
   - Exhibit different physical properties (due to different masses).

- Examples:
  - Hydrogen: Protium (^{1}H, 0n), Deuterium (^{2}H, 1n), Tritium (^{3}H, 2n).
  - Carbon: ^{12}C (6n), ^{14}C (8n).
  - Chlorine: ^{35}Cl (18n), ^{37}Cl (20n).
  - Uranium: ^{235}U (143n), ^{238}U (146n) – both have 92 protons.
  - Oxygen naturally occurs in 3 isotopic forms (e.g., ^{16}O, ^{17}O, ^{18}O).
- Applications of Radioactive Isotopes:
  - Uranium-235 (^{235}U): Nuclear fuel in reactors.
  - Cobalt-60 (^{60}Co): Treatment of cancer.
  - Iodine-131 (^{131}I): Treatment of goitre and thyroid disorders.
  - Carbon-14 (^{14}C): Carbon dating for archaeological and geological specimens.
  - Iron-59 (^{59}Fe): Diagnosis of anemia (as a tracer).
  - Phosphorus-32 (^{32}P): Treatment of leukemia.
  - Radioactive isotopes can pose environmental hazards (e.g., Chernobyl accident involved radioactive Uranium and Caesium).

5. Isobars:
   - Atoms of different elements.
   - Have the same mass number (A) (same total number of nucleons).
   - Have different atomic numbers (Z) (different number of protons).
   - Exhibit different chemical and physical properties.
   - Example: Argon-40 ($^{40}Ar_{18}$) and Calcium-40 ($^{40}Ca_{20}$).
     - $^{40}Ar_{18}$: 18 protons, 22 neutrons.
     - $^{40}Ca_{20}$: 20 protons, 20 neutrons.
6. Isotones:
   - Atoms of different elements.
   - Have the same number of neutrons (N).
   - Have different atomic numbers (Z) and different mass numbers (A).
   - Example: Tritium ($^{3}H_{1}$) and Helium-4 ($^{4}He_{2}$).
     - $^{3}H_{1}$: 1 proton, 2 neutrons.
     - $^{4}He_{2}$: 2 protons, 2 neutrons.
7. Average Atomic Mass:
   - Calculated as the weighted average of the masses of an element's naturally occurring isotopes, taking into account their relative abundances.
   - Example: Chlorine (^{35}Cl and ^{37}Cl) in a 3:1 ratio has an average atomic mass of approximately 35.5 u.
8. Valency: The combining capacity of an atom.

Quantum Numbers (Principal, Azimuthal, Magnetic, Spin and their significance)

- Principal Quantum Number (n):
  - Values: 1, 2, 3, ... (positive integers).
  - Significance: Defines main energy shell/level, size of the orbital, and largely determines electron energy.
  - Maximum electrons in a shell: $2n^2$.
  - Total number of orbitals in a shell: n^2.
- Azimuthal/Angular Momentum Quantum Number (l):
  - Values: 0, 1, 2, ..., (n-1).
  - Significance: Defines subshell within a shell, shape of the orbital.
  - l=0 corresponds to s subshell (spherical).
  - l=1 corresponds to p subshell (dumbbell).
  - l=2 corresponds to d subshell (complex).
  - l=3 corresponds to f subshell (more complex).
  - Number of orbitals in a subshell: (2l + 1).
  - Orbital notation combines 'n' and subshell letter (e.g., n=3, l=2 is 3d).
- Magnetic Quantum Number (m_l):
  - Values: -l, (-l+1), ..., 0, ..., (+l-1), +l.
  - Significance: Defines the orientation of the orbital in space. Each m_l value represents a unique orbital.
- Spin Quantum Number (m_s):
  - Values: +1/2 or -1/2.

- *Significance: Defines the intrinsic spin angular momentum of the electron (spin up or spin down).*
- *Each orbital can hold a maximum of two electrons with opposite spins (Pauli Exclusion Principle).*

- *Energy Order of Orbitals (Aufbau Principle / (n+l) Rule):*
  - *Orbitals are filled in increasing order of (n+l) values.*
  - *If two orbitals have the same (n+l) value, the orbital with the lower 'n' value has lower energy and is filled first.*
  - *Example order: 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p.*
- *Electronic Configuration Example: For an element with atomic number 16 (Sulfur): $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^4$. The principal quantum number (n) of its outermost electrons is 3.*

2. Physical and Chemical Changes, Solution

Concept of Solutions: Solute, Solvent, and Components

- *A solution is a homogeneous mixture of two or more components.*
- *Solute: Component present in a smaller amount.*
- *Solvent: Component present in a larger amount, determining the physical state of the solution.*
- *Tincture of iodine is a solution of iodine (solute) in alcohol (solvent).*
- *Mole fraction (X) of a component = Moles of component / Total moles in the solution.*
- *Number of moles = Given mass / Molar mass.*
- *Number of molecules is directly proportional to the number of moles.*
- *Solutions are classified by their ability to conduct electricity:*
  - *Electrolytes: Conduct electricity due to mobile ions (e.g., copper sulphate, sodium chloride, potassium chloride, sodium hydroxide, HCl, KOH).*
  - *Strong Electrolytes: Dissociate completely into ions (e.g., potassium chloride, NaCl, HCl, KOH).*
  - *Non-electrolytes: Do not conduct electricity as they do not form ions (e.g., sugar solution, glucose solution).*
- *Water is a good solvent for ionic compounds and polar substances (e.g., glucose, urea, acetone) due to its high dipole moment (polarity) and high dielectric constant.*
- *Sugar dissolves in water primarily due to the formation of hydrogen bonds.*
- *Reverse osmosis is a process used for transforming salty water into pure water.*
- *Colligative properties of dilute solutions depend only on the number of solute particles, not their chemical nature or size.*
- *Adding common salt (a non-volatile solute) to water increases its boiling point (boiling point elevation) and decreases its freezing point (freezing point depression).*
- *Glass is best described as a supercooled solution (an amorphous solid with high viscosity).*
- *On calibrated pipettes:*
  - *TD means "To Deliver" (the specified volume).*
  - *TC means "To Contain" (the specified volume).*
- *Pasture pipettes are commonly used for transferring small volumes of liquids (< 5 mL).*
- *Safety Data Sheets (SDS) contain 16 standardized sections.*
- *Flammable gases should be stored at a minimum distance of 20 feet from oxygen.*
- *The Van't Hoff factor (i) for an electrolyte is generally greater than one due to dissociation into multiple ions.*
- *The Ideal Gas Equation is PV = nRT.*
- *Air is always a mixture of various gases.*

Definition and Characteristics of Chemical Changes (Chemical Reactions)

- *Chemical Change (Chemical Reaction): A process where substances are transformed into new substances with different chemical compositions and properties.*

- *Characteristics of Chemical Changes:*
  - *Formation of one or more new substances.*
  - *Often irreversible or difficult to reverse.*
  - *Accompanied by significant energy changes (heat, light, sound produced or absorbed).*
  - *Examples: Burning (paper, magnesium ribbon, coal, glucose), rusting of iron, milk turning sour, cooking food, fermentation, mixing acid and base, rancidity.*
- *Physical Change: A change in the physical properties (e.g., state, shape, size) of a substance, without altering its chemical composition or forming new substances.*
- *Characteristics of Physical Changes:*
  - *No new substances formed.*
  - *Generally reversible.*
  - *Minimal energy changes.*
  - *Examples: Melting ice, boiling water, dissolving sugar/salt, magnetization, cutting hair/fruit/foil, evaporation, sublimation of camphor.*
- *Types of Chemical Reactions:*
  - *Combination Reaction: Two or more reactants combine to form a single product (e.g., $CaO + H_2O \rightarrow Ca(OH)_2$, $2Mg + O_2 \rightarrow 2MgO$).*
  - *Decomposition Reaction: A single compound breaks down into two or more simpler substances (e.g., $CaCO_3 \rightarrow CaO + CO_2$, $2HgO \rightarrow 2Hg + O_2$, Electrolysis of water, $2AgCl \rightarrow 2Ag + Cl_2$ in sunlight). Thermal decomposition can apply to ZnO, Ag_2O.*
  - *Displacement (Substitution) Reaction: A more reactive element displaces a less reactive element from its compound (e.g., $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$; Iron is more reactive than copper. $Pb + CuCl_2 \rightarrow PbCl_2 + Cu$).*
  - *Double Displacement Reaction: Exchange of ions between two compounds, often leading to precipitation (e.g., $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$).*
  - *Redox (Oxidation-Reduction) Reactions: Involve simultaneous oxidation and reduction.*
    - *Oxidation: Gain of oxygen, loss of hydrogen, loss of electrons, increase in oxidation state.*
    - *Reduction: Loss of oxygen, gain of hydrogen, gain of electrons, decrease in oxidation state.*
    - *Oxidizing Agent: Causes oxidation (gets reduced itself).*
    - *Reducing Agent: Causes reduction (gets oxidized itself).*
    - *Examples: Combustion (e.g., $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, where Carbon is oxidized), rusting, rancidity, $2Mg + O_2 \rightarrow 2MgO$, $ZnO + C \rightarrow Zn + CO$ (C is reducing agent, ZnO is oxidizing agent).*
- *Stoichiometry & Laws of Chemical Combination:*
  - *Law of Conservation of Mass: Matter cannot be created or destroyed in a chemical reaction; total mass of reactants equals total mass of products (e.g., 10g $CaCO_3$ decomposed to 5.6g CaO produces 4.4g CO_2).*
  - *Law of Constant (Definite) Proportion: A given compound always contains its component elements in fixed ratios by mass (e.g., Ammonia always has the same N:H proportion).*
  - *Avogadro's Hypothesis: Equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules.*
  - *Glucose Combustion: $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$; 1 mole of glucose requires 6 moles of oxygen.*
- *Reaction Rate & Catalysis:*
  - *Reaction Rate: Generally decreases with time as reactant concentrations decrease.*
  - *Factors Influencing Reaction Rate: Temperature (higher temperature increases rate due to more rapid molecular motion), concentration of reactants, presence of catalyst, surface area of reactants/catalyst.*

- *Catalyst: A substance that changes (usually increases) the rate of a chemical reaction without being consumed; it works by lowering the activation energy of the reaction.*
- *Catalytic Promoter: Increases the activity of a catalyst.*
- *Catalytic Inhibitor: Slows down or stops a reaction.*
- *Surface Area of Catalyst: Higher surface area (e.g., powdered form) leads to higher catalytic activity.*
- *Enzyme Catalysis: Biological catalysts, often act as homogeneous catalysts (in solution) and function via an 'induced fit' mechanism.*
- *Rate Constant (k): Depends on temperature (Arrhenius concept); defined as the reaction rate when each reactant's concentration is unity; unit for a first-order reaction is s^{-1}.*

- *Solutions & Hard Water:*
  - *Hard Water: Water containing significant concentrations of dissolved metallic ions.*
  - *Metallic Constituents of Hard Water: Primarily Calcium (Ca^{2+}) and Magnesium (Mg^{2+}) ions. Iron (Fe^{2+}/Fe^{3+}) also contributes.*
  - *Temporary Hardness: Caused by dissolved bicarbonates of Ca and Mg; can be removed by boiling.*
  - *Permanent Hardness: Caused by dissolved chlorides and sulfates of Ca and Mg; cannot be removed by boiling.*
  - *Disadvantages of Hard Water: Not suitable for drinking, washing clothes (forms scum with soap), or use in boilers (forms scale).*
- *Desiccants/Drying Agents: Substances that absorb moisture (exhibit hygroscopic nature). Example: Silica gel.*
- *Fire Extinguishers: Carbon dioxide (CO_2) and Bromochlorodifluoromethane (BCF) are common.*
- *Fermentation: A process where organic compounds are slowly decomposed into simpler substances, typically in the absence of oxygen (anaerobic).*
- *Corrosion (e.g., Rusting): A chemical reaction (oxidation) where a metal reacts with its environment. During rusting of iron, the weight of iron increases due to the incorporation of oxygen.*
- *Precipitation Reaction: A reaction where an insoluble solid (precipitate) forms when two solutions are mixed.*
- *Egg Boiling: A chemical change (protein denaturation) leading to a decrease in disorder and a negative ΔG (spontaneous process).*
- *Air: Is always a mixture.*
- *Glass Etching: Can be achieved using diamond or hydrofluoric acid (HF).*
- *Nanomaterials Hindrance: Prone to coarsening effect from Ostwald ripening or agglomeration on aging.*
- *Chemical Equilibrium: A dynamic state where the rates of forward and backward reactions are equal, resulting in no net change in reactant and product concentrations. Catalysts help reach equilibrium faster but do not shift its position.*
- *Adsorbate: A substance that is deposited or held on the surface of another substance (adsorbent).*

Definition and Characteristics of Physical Changes

- *Physical Change Definition: Alters the form or state (e.g., solid, liquid, gas) of a substance, but its chemical composition and identity remain the same; no new substance is formed.*
- *Physical Change Characteristics:*
  - *The chemical identity of the substance remains unchanged.*
  - *No new chemical bonds are formed or broken.*
  - *Often reversible (e.g., melting ice to water and freezing water back to ice).*
  - *Energy changes involved are typically smaller compared to chemical changes.*
- *Examples of Physical Changes:*
  - *Melting of ice, wax, or other solids (note: melting of wax is physical, but*

burning a candle involves both physical and chemical changes).
    - Boiling or evaporation of water.
    - Condensation of steam.
    - Freezing of water.
    - Dissolution of sugar in water.
    - Heating of a platinum crucible.
    - Sublimation of substances like camphor, naphthalene, ammonium chloride, or anthracene.
    - Deposition (change of gaseous state directly to solid state).
- States of Matter: Generally considered 5 states: Solid, Liquid, Gas, Plasma, and Bose-Einstein Condensate.
- Intermolecular Forces:
    - The force of attraction among molecules is highest in solids (e.g., sugar) and lowest in gases (e.g., helium).
    - The melting point of a solid is an indication of the strength of the intermolecular forces of attraction.
- Phase Changes (Physical Processes):
    - Melting/Fusion: Solid to liquid at a specific melting point.
    - Boiling: Liquid to gas (a bulk phenomenon, occurring at the boiling point throughout the liquid); bubbles in vigorously boiling pure water are composed primarily of water vapour.
    - Evaporation: Liquid to gas (a surface phenomenon, occurring below the boiling point).
    - Sublimation: Solid directly to gas without changing into the liquid state (e.g., naphthalene, camphor, ammonium chloride, anthracene).
    - Deposition: Gas directly to solid without changing into the liquid state.
    - Condensation: Gas to liquid.
    - The state of matter can be changed by altering its temperature and/or pressure.
- Separation Techniques (based on physical properties):
    - Sublimation: Separates a sublimable solid (e.g., naphthalene, camphor, anthracene) from non-sublimable substances (e.g., sand, salt).
    - Distillation: Separates liquids with different boiling points, or a non-volatile dissolved solid from a liquid.
    - Chromatography: Separates components based on their differential distribution between a stationary and a mobile phase (e.g., colors from a dye, pigments from natural colors, drugs from blood). It cannot separate radio-isotopes.
    - Filtration: Separates insoluble solids from liquids.
    - Evaporation: Separates a soluble non-volatile solid from its volatile solvent.
- Other Relevant Physical Concepts:
    - Diffusion: The intermixing of particles of two different types of matter on their own.
    - Crystalline Solid: Possesses a regular, ordered arrangement of particles (e.g., table salt).
    - Distilled Water: Considered pure and free from impurities.
    - Nanomaterials:
        - Their properties strongly depend on the surface atoms.
        - They have a large surface area to volume ratio.
        - Nanoparticles typically have sizes < 100 nm.
        - Particles of nanomaterials cannot be seen with the naked eye.
    - Boyle's Law: At a constant temperature, the volume (V) of a fixed mass of gas is inversely proportional to its pressure (P), mathematically represented as V $\propto$ 1/P, V = K/P, or VP = K.
    - Compressed Air Technique: Used for cleaning and drying documents.

Distinction between Physical and Chemical Changes

- Physical Change: A change where no new substance is formed, and the chemical composition remains the same.
- Physical changes are generally reversible.