### **IB Chemistry Year 1 Final Study Document**

**General Advice:**

* **Keywords are crucial:** For definitions, focus on the bolded words – these are the core ideas.
* **Understand concepts:** Don't just memorize definitions; grasp the underlying principles to apply them.
* **Practice problems:** Use this guide to refresh your knowledge, then work through past questions.

#### I. Core Definitions & Concepts (Focus on Keywords)

**A. Fundamentals: Elements, Compounds, Mixtures & States of Matter**

1. **Element:** Pure substance, **cannot be broken down** into simpler substances by chemical means.
2. **Compound:** Two or more **elements chemically bonded** together in fixed ratio.
3. **Mixture:** Two or more substances **not chemically bonded**, retain individual properties.
   * **Homogeneous Mixture:** Uniform composition throughout (e.g., solution).
   * **Heterogeneous Mixture:** Non-uniform composition (e.g., sand and water).
4. **Separation Techniques:** Methods to separate components of a mixture based on **physical properties**.
   * *Examples:* Filtration (solids from liquids), Distillation (liquids with different boiling points), Chromatography (different affinities for phases).
5. **States of Matter:**
   * **Solid:** Fixed shape and volume, particles **tightly packed**, vibrate in fixed positions.
   * **Liquid:** Fixed volume, no fixed shape, particles **close but can move past** each other.
   * **Gas:** No fixed shape or volume, particles **far apart**, move randomly and rapidly.
6. **Kinetic Energy (KE):** Energy due to **motion** of particles.
7. **Temperature:** Measure of **average kinetic energy** of particles.

**B. Atomic Structure & Mass Spectrometry**

1. **Nuclear Atom:** Atom with a **small, dense, positively charged nucleus** and electrons orbiting it.
2. **Nucleon:** Particle in **nucleus** (proton or neutron).
3. **Atomic Number (Z):** Number of **protons**. Defines the element.
4. **Mass Number (A):** Number of **protons + neutrons**.
5. **Isotopes:** Atoms of **same element** (same Z), **different neutrons** (different A).
6. **Mass Spectrometry:** Analytical technique measuring **mass-to-charge ratio** of ions, used for identifying isotopes/molecular mass.
   * **Parent/Molecular Ion Peak (**M+**):** Line with **highest mass**, corresponds to molar mass of compound.
   * **Fragment Ions:** Smaller peaks, result from **breakdown of molecule**.
   * All fragments are assumed to be **1+** charged.
7. **Electron Configuration:** Distribution of **electrons in atomic orbitals**.
8. **Atomic Orbital:** Region of space where electron has **high probability**.
   * *Exceptions:* Cr and Cu (fill 4s then 3d for half/full stability: e.g., Cr: [Ar]3d54s1, Cu: [Ar]3d104s1).
9. **Emission Spectrum:** **Distinct lines** of specific frequencies/wavelengths, characteristic of an element, produced when excited electrons fall to lower energy levels.
   * **Types:** Line (atoms), Band (molecules), Continuous (blackbody radiation).
   * **Transitions:** Electrons move between specific energy levels.
10. **Electromagnetic (EM) Spectrum:** Range of all types of **electromagnetic radiation**.
11. **Hydrogen Emission Spectrum:** Series of lines for H, showing transitions.
    * **Convergence Limit:** Lines get **closer together** at higher energy (shorter wavelength), indicating ionization energy.
    * **Ionization Energy from Spectrum:** Energy of the convergence limit corresponds to the **first ionization energy**.

**C. Ionization Energy & Trends**

1. **First Ionization Energy:** Minimum energy to **remove one electron** from one mole of **gaseous atoms** to form one mole of **gaseous 1+ ions**.
   * *Formula:* X(g)→X+(g)+e−
2. **Successive Ionization Energy:** Energy required to remove **subsequent electrons**. Each successive IE is larger due to increased positive charge and smaller atomic/ionic radius.
3. **Ionization Energy Trends:**
   * **Across a Period (left to right):** Ionization energy **increases**. Due to increasing nuclear charge, decreasing atomic radius, similar shielding.
   * **Down a Group (top to bottom):** Ionization energy **decreases**. Due to increasing electron shells (larger radius), increased shielding, further distance from nucleus.

**D. Energetics/Thermochemistry**

1. **Exothermic Reaction:** **Releases heat** to surroundings, ΔH<0, products are **lower in energy** than reactants.
2. **Endothermic Reaction:** **Absorbs heat** from surroundings, ΔH>0, products are **higher in energy** than reactants.
3. **Enthalpy Change (**ΔH**):** **Heat change at constant pressure**.
4. **Lattice Enthalpy:** Energy change when **one mole of an ionic compound** is formed from its **gaseous ions** under standard conditions.
   * *General Equation for MX:* M+(g)+X−(g)→MX(s)
   * **Effect on Properties:**
     + **Volatility:** Higher lattice enthalpy = lower volatility (harder to break apart).
     + **Electrical Conductivity:** Ionic solids do not conduct (ions fixed), molten/aqueous conduct (ions mobile).
     + **Solubility:** Determined by balance between lattice enthalpy and hydration enthalpy.

**E. Stoichiometry & Gases**

1. **Reacting Solutions:** Mixing solutions to cause a reaction.
   * **Solution:** **Homogeneous mixture** where one substance (solute) is dissolved in another (solvent).
   * **Precipitate:** **Solid formed** from a solution during a chemical reaction.
2. **Titration:** Analytical technique to determine **unknown concentration** using a known concentration (standard solution).
   * **Equivalence Point:** Point where **reactants have completely reacted** stoichiometrically.
   * **Tips/Tricks:** Rinse apparatus correctly, use indicator for equivalence point, repeat for precision.
3. **Back Titration:** Titration where **excess of a known reactant** is added, then the unreacted excess is titrated. Useful for slow reactions or insoluble substances.
4. **Percentage Composition:** **Mass of each element** as a percentage of total compound mass.
5. **Atom Economy:** Measure of how many **atoms from reactants are incorporated into desired product**, not wasted.
   * Atom Economy=Molar mass of all reactantsMolar mass of desired product​×100
6. **Combustion Analysis:** Experimental method to determine **empirical formula** by burning a compound and measuring masses of products (CO2​, H2​O).
7. **Kinetic Theory of Gases:** Explains gas behavior: **random, continuous motion**; negligible particle volume; no intermolecular forces; collisions elastic; KE proportional to T.
8. **Molar Volume of a Gas:** Volume occupied by **one mole of any gas** at specific T and P (e.g., 22.7 dm3mol−1 at STP).

**F. Chemical Kinetics: Rates of Reaction**

1. **Reaction Rate:** **Change in concentration** of reactant/product per unit time.
   * **Units:** Can be mol dm−3 s−1 (or other concentration/time units like g s−1). The unit depends on what is being measured (e.g., mass change, volume change, concentration change). It always involves a change in quantity over time.
   * **Overall Rate vs. Independent Rate:** Overall rate is for the entire reaction. Independent rate refers to the rate of a specific reactant or product.
2. **Collision Theory:** For reaction to occur, particles must **collide with correct orientation** and **sufficient energy** (greater than or equal to Activation Energy).
3. **Factors Affecting Rate of Reaction:**
   * **Concentration (Reactants):** Higher concentration → more collisions (Collision Theory).
   * **Temperature:** Higher temperature → particles have higher kinetic energy, more frequent and more energetic collisions (more particles exceed Ea​) (Maxwell-Boltzmann distribution shifts).
   * **Surface Area (Solids):** Larger surface area → more particles exposed for collisions.
   * **Pressure (Gases):** Higher pressure → higher concentration → more frequent collisions.
   * **Catalyst:** Lowers activation energy by providing alternative pathway (more particles exceed Ea​).
4. **Activation Energy (**Ea​**):** **Minimum energy** required for reaction to occur (energy barrier).
5. **Catalyst:** Substance that **increases reaction rate** without being consumed, by **lowering** Ea​.
6. **Multi-step Reaction (Reaction Mechanism):** Series of **elementary steps** that lead to overall reaction.
   * **Elementary Step:** Individual step in a reaction mechanism.
   * **Molecularity:** Number of **reactant species** involved in an elementary step.
     + **Unimolecular:** One reactant species.
     + **Bimolecular:** Two reactant species.
     + **Termolecular (Trimolecular):** Three reactant species (rare).
   * **Rate Determining Step (RDS):** **Slowest step** in a multi-step reaction, determines overall reaction rate.
7. **Rate Equation (Rate Law):** Expresses relationship between **rate and reactant concentrations**.
   * For A+B→Products, Rate =k[A]m[B]n
   * **k:** Rate constant (temperature dependent).
   * **m, n:** Reaction orders with respect to A and B (experimentally determined, not stoichiometric coefficients).
   * **Overall Order:** m+n.
8. **Graphical Representations of Reaction Order:**
   * **Zero Order:** [A] vs time is linear (negative slope). Rate is constant.
   * **First Order:** ln[A] vs time is linear (negative slope). Rate is proportional to [A].
   * **Second Order:** 1/[A] vs time is linear (positive slope). Rate is proportional to [A]2.

**G. Chemical Bonding: Main Types & Structure**

1. **Bonding:** Force that **holds atoms together**.
2. **Ionic Bond:** **Electrostatic attraction** between **oppositely charged ions** (transfer of electrons).
   * **Lattice Structure (Ionic):** Giant repeating structure of **alternating positive and negative ions**.
3. **Covalent Bond:** **Sharing of electrons** between two atoms.
   * **Covalent Model:** Describes bonding by electron sharing.
4. **Metallic Bond:** **Electrostatic attraction** between **positive metal ions** and a **sea of delocalized electrons**. Occurs because metals tend to lose electrons.
   * **Lattice Structure (Metallic):** Regular arrangement of **positive metal ions** surrounded by **delocalized electrons**.
5. **Brief Oxidation vs Reduction:**
   * **Oxidation:** **Loss of electrons**.
   * **Reduction:** **Gain of electrons**.
6. **Transition Elements:** Elements with an **incomplete d subshell** in at least one of their common ions.
   * **vs d-block elements:** All transition elements are d-block, but not all d-block elements are transition elements (e.g., Zn, Sc).
7. **Electronegativity:** Ability of an atom to **attract electrons** in a covalent bond.
   * **Qualitative Analysis:** Compare electronegativity differences: large difference = more ionic, small difference = more covalent.
8. **Bonding Continuum:** Spectrum ranging from **pure covalent** (equal sharing) to **pure ionic** (complete transfer).
9. **Polyatomic Ion:** Ion consisting of **two** or more atoms **covalently bonded** together, with an overall charge.
10. **Bond Order:** Number of **electron pairs shared** between two atoms.
11. **Coordinate Bond (Dative Covalent Bond):** Covalent bond where **both shared electrons come from one atom**.
12. **Ionic vs Covalent Definition Difference:** Ionic involves **transfer of electrons**, forming ions. Covalent involves **sharing of electrons**, forming molecules.

**H. Covalent Bonding: Geometry, Polarity & Structures**

1. **VSEPR Theory (Valence Shell Electron Pair Repulsion):** Electron domains (bonding pairs and lone pairs) around a central atom **repel each other** and arrange to **minimize repulsion**.
   * **Electron Domain (ED):** Region of electron density (single, double, triple bonds count as one ED; lone pairs count as one ED).
   * **Electron Domain Geometry (EDG):** Arrangement of **all electron domains**.
   * **Molecular Geometry (MG):** Arrangement of **only atoms** (ignoring lone pairs for shape name).

| **# Electron Domains** | **# Bonding Domains** | **# Lone Pairs** | **Electron Domain Geometry** | **Molecular Geometry** | **Ideal Bond Angle(s)** |
| --- | --- | --- | --- | --- | --- |
| 2 | 2 | 0 | Linear | Linear | 180∘ |
| 3 | 3 | 0o | Trigonal Planar | Trigonal Planar | 120∘ |
| 3 | 2 | 1 | Trigonal Planar | Bent | <120∘ |
| 4 | 4 | 0 | Tetrahedral | Tetrahedral | 109.5∘ |
| 4 | 3 | 1 | Tetrahedral | Trigonal Pyramidal | <109.5∘ |
| 4 | 2 | 2 | Tetrahedral | Bent | 104.5∘ |
| 5 | 5 | 0 | Trigonal Bipyramidal | Trigonal Bipyramidal | 120∘,90∘ |
| 5 | 4 | 1 | Trigonal Bipyramidal | See-Saw | <120∘,<90∘ |
| 5 | 3 | 2 | Trigonal Bipyramidal | T-shaped | <90∘ |
| 5 | 2 | 3 | Trigonal Bipyramidal | Linear | 180∘ |
| 6 | 6 | 0 | Octahedral | Octahedral | 90∘ |
| 6 | 5 | 1 | Octahedral | Square Pyramidal | <90∘ |
| 6 | 4 | 2 | Octahedral | Square Planar | 90∘ |

1. **Bond Polarity:** Unequal sharing of electrons in a covalent bond due to **electronegativity difference**.
   * In polyatomic ions with coordinate bonds, the bond polarity still depends on the electronegativity of the atoms involved.
2. **Covalent Network Structures (Giant Covalent):** Atoms held by **strong covalent bonds** in a continuous **three-dimensional lattice**.
3. **Molecular Structures:** Discrete molecules held by **strong covalent bonds intramolecularly**, but weak **intermolecular forces** between molecules.
4. **Allotropes:** **Different structural forms** of the same element in the same physical state.

**I. Allotropes and Expanded Octets**

* **Diamond (Carbon):**
  + **Arrangement:** Each C atom bonded to **four other C atoms** in a **tetrahedral** arrangement.
  + **Conductivity:** Non-conductor (no delocalized electrons).
  + **IMFs:** Only strong covalent bonds (no IMFs as it's a network).
* **Graphite (Carbon):**
  + **Arrangement:** Each C atom bonded to **three other C atoms** in **hexagonal layers**. Layers held by weak LDFs.
  + **Conductivity:** Good conductor (delocalized electrons within layers).
  + **IMFs:** Weak London Dispersion Forces between layers.
* **Graphene (Carbon):**
  + **Arrangement:** Single layer of graphite, **hexagonal 2D network**.
  + **Conductivity:** Excellent conductor.
  + **IMFs:** N/A (single layer).
* **Fullerenes (Carbon, e.g.,** C60​ **Buckminsterfullerene):**
  + **Arrangement:** Spherical molecular cages of **hexagons and pentagons**.
  + **Conductivity:** Semiconductor/non-conductor.
  + **IMFs:** London Dispersion Forces between molecules.
* **Silicon (Si):**
  + **Arrangement:** Similar to diamond, **tetrahedral network**.
  + **Conductivity:** Semiconductor.
* **Silica (**SiO2​**):**
  + **Arrangement:** Each Si bonded to **four O atoms**, each O bonded to **two Si atoms** in a **tetrahedral network**.
  + **Conductivity:** Non-conductor.

1. **Expanded Octets:** When central atoms in period 3 or beyond (e.g., P, S, Cl, Br, I, Xe) can accommodate **more than 8 valence electrons** by utilizing their empty d-orbitals.
   * Can have **10 or 12** (even numbers) valence electrons.
   * *Example atoms:* Sulfur (S), Phosphorus (P), Chlorine (Cl), Iodine (I), Xenon (Xe).

**J. Intermolecular Forces (IMFs) & Physical Properties of Covalent Structures**

1. **Intermolecular Forces (IMFs):** **Attractive forces between molecules**. Weaker than intramolecular (covalent) bonds.
   * **London Dispersion Forces (LDFs):**
     + **Definition:** Weak, temporary attractive forces due to **instantaneous, induced dipoles** from momentary electron distribution.
     + **Existence:** Present in **all molecules** (polar and nonpolar).
     + **Strength:** Increases with **increasing number of electrons** (larger electron clouds are more polarizable), leading to stronger temporary dipoles.
   * **Dipole-Dipole Forces:**
     + **Definition:** Attractive forces between **permanent dipoles** of **polar molecules**.
     + **Existence:** Only in **polar molecules**.
   * **Hydrogen Bonding:**
     + **Definition:** Strongest type of IMF, special dipole-dipole attraction between a **hydrogen atom** (covalently bonded to a highly electronegative atom: **F, O, or N**) and a **lone pair** on another highly electronegative atom (F, O, or N) in a different molecule.
     + **Necessary Conditions:** H atom bonded to F, O, or N, AND a lone pair on F, O, or N in an adjacent molecule.
   * **Rank of Strengths (Weakest to Strongest):** LDFs < Dipole-Dipole < Hydrogen Bonding.
2. **Physical Properties of Covalent Structures:** Determined by the **type and strength of forces** present.
   * **Molecular Covalent:** Properties determined by **IMFs**.
     + **Volatility:** Generally **high** (weak IMFs are easily overcome).
     + **Solubility:** "Like dissolves like" (polar in polar, nonpolar in nonpolar).
     + **Electrical Conductivity:** **Non-conductors** (no free ions or delocalized electrons).
   * **Network Covalent:** Properties determined by **lattice features** (strong covalent bonds throughout).
     + **Volatility:** Generally **very low** (strong covalent bonds require high energy to break).
     + **Solubility:** **Insoluble** in common solvents (covalent bonds too strong to break).
     + **Electrical Conductivity:** Generally **non-conductors** (e.g., diamond, silica), except for those with delocalized electrons (e.g., graphite, graphene).

**K.** Chromatography & Analytical **Techniques**

1. **Chromatography:** Separation technique based on **differential partitioning** of components between a **stationary phase** and a **mobile phase**.
   * **Mobile Phase:** Solvent that **moves** through the stationary phase.
   * **Stationary Phase:** Material that **stays** fixed (e.g., paper, silica gel).
   * **Affinity:** Tendency of a substance to bind to or be attracted by another.
   * **Thin Layer Chromatography (TLC):** Uses a **silica gel plate** as stationary phase.
   * **Paper Chromatography:** Uses **chromatography paper** (cellulose) as stationary phase.
   * **Polarity & Movement:**

### **Polarity & Movement:**

* **If mobile phase is polar and stationary phase is non-polar (e.g., normal phase chromatography, but reversed phase is more common for this setup):**
  + More polar solutes travel **faster** (prefer mobile phase).
  + Less polar solutes travel **slower** (prefer stationary phase).
* **If mobile phase is non-polar and stationary phase is polar (e.g., normal phase chromatography):**
  + More polar solutes travel **slower** (prefer stationary phase).
  + Less polar solutes travel **faster** (prefer mobile phase).
  + **Retardation Factor (**Rf​**):**
    - Rf​=Distance travelled by solvent frontDistance travelled by solute​

**L.** Resonance, **Hybridization & Bonding Geometries**

1. **Resonance Structure:** When a single Lewis structure **cannot accurately represent** the true bonding, multiple resonance structures are drawn, representing electron delocalization.
   * **Conditions:** Multiple valid Lewis structures can be drawn by **moving only electrons** (atoms stay in place).
   * **Delocalization:** Electrons are **spread over multiple atoms** rather than localized between two.
   * **Bond Order:** Average of the bond orders in all resonance structures.
2. **Benzene (**C6​H6​**):** Aromatic compound.
   * **Structure:** Cyclic, planar, **hexagonal ring** of carbon atoms with delocalized π electrons. All C-C bond lengths are equal (intermediate between single and double).
   * **Physical Evidence:** **X-ray diffraction** shows all C-C bonds are of equal length (139 pm), intermediate between C-C single (154 pm) and C=C double (134 pm).
   * **Chemical Evidence:** Undergoes **substitution reactions** rather than addition reactions (like saturated compounds) and is **less reactive** than expected for an unsaturated compound (due to delocalization stability).
3. **Formal Charge:** Hypothetical charge on an atom in a molecule, assuming electrons in a chemical bond are shared equally between atoms.
   * **Formula 1 (Broad):** Formal Charge = (Valence electrons of neutral atom) - (Valence electrons in current structure)
   * **Formula 2 (Specific):** Formal Charge = (Valence electrons of neutral atom) - (Non-bonding electrons) - 21​(Bonding electrons)
4. **Sigma (**σ**) Bond:** Formed by **overlap** of atomic orbitals along the bond axis. **Must be present** in any bond (single, double, triple).
5. **Pi (**π**) Bond:** Formed by **sideways overlap of parallel p atomic orbitals** above and below the bond axis. Present in double (one π) and triple (two π) bonds, in addition to a σ bond.
6. **Hybridization:** The **mixing of atomic orbitals** (s, p, d) on a central atom to form new, degenerate **hybrid orbitals** that have different shapes and orientations than the original atomic orbitals, allowing for stronger bonds and predicting molecular geometry.
   * sp3 **Hybridization:**
     + **Formation:** One s orbital + three p orbitals → four sp3 hybrid orbitals.
     + **Geometry:** Tetrahedral.
     + **Example:** Methane (CH4​).
     + *Diagram:* Shows one electron from 2s promoted to 2p, then four sp3 hybrid orbitals form.
   * sp2 **Hybridization:**
     + **Formation:** One s orbital + two p orbitals → three sp2 hybrid orbitals, one unhybridized p orbital remains.
     + **Geometry:** Trigonal planar.
     + **Example:** Ethene (C2​H4​).
   * sp **Hybridization:**
     + **Formation:** One s orbital + one p orbital → two sp hybrid orbitals, two unhybridized p orbitals remain.
     + **Geometry:** Linear.
     + **Example:** Ethyne (C2​H2​).

**M. Metallic Bonding & Materials**

1. **Metallic Bonding:** **Electrostatic attraction** between **positive metal ions** (cations) and a **sea of delocalized valence electrons**.
   * **Why it happens:** Metal atoms have low ionization energies and tend to lose their valence electrons, which become delocalized.
2. **Factors Affecting Strength of Metallic Bonding:**
   * **Ionic Charge:** Higher positive charge on metal ion → stronger attraction → stronger bond.
   * **Ionic Radii:** Smaller ionic radius → closer to delocalized electrons → stronger attraction → stronger bond.
   * **Density of Sea of Electrons:** More delocalized electrons per unit volume → stronger bond.
3. **Properties of Metallic Substances:**
   * **Electrical Conductivity:** Excellent (delocalized electrons can move freely).
   * **Thermal Conductivity:** Excellent (delocalized electrons efficiently transfer heat).
   * **Malleability:** Can be hammered into sheets (ions can slide past each other without breaking bonds due to mobile electron sea).
   * **Ductility:** Can be drawn into wires (similar reason to malleability).
   * **Volatility:** Low (high melting/boiling points due to strong metallic bonds).
   * **Brittleness:** Generally not brittle.
   * **Corrodable:** Many metals are susceptible to corrosion (reaction with environment).
4. **Metalloids:** Elements with **properties intermediate** between metals and non-metals (e.g., Boron, Silicon, Germanium).
5. **Bonding Triangle:** A diagram that illustrates the **continuum of bonding types** (ionic, covalent, metallic) based on electronegativity difference and average electronegativity.
6. **Brittleness:** Tendency to **fracture or break** without significant deformation.
7. **Elasticity:** Ability to **return to original shape** after deformation.
8. **Plasticity:** Ability to undergo **permanent deformation** without fracturing.
9. **Corrosion:** Deterioration of a material due to **chemical reaction** with its environment.
10. **Exceptions of Bonding Triangle:** Some compounds don't fit perfectly, e.g., complex covalent compounds might have significant ionic character.
11. **Materials:**
    * **Alloys:** Mixtures of **two or more metals**, or a metal and a non-metal, designed to enhance properties. Generally **stronger and harder** than pure metals.
    * **Substitutional Alloy:** Atoms of similar size **replace** atoms in the original lattice.
    * **Interstitial Alloy:** Smaller atoms fit into the **gaps (interstices)** between larger atoms in the lattice.

**N. Summary of Properties for Different Bonding Types**

| **Property** | **Metallic** | **Ionic** | **Molecular Covalent** | **Network Covalent** |
| --- | --- | --- | --- | --- |
| **Melt/Boil Point** | High | High | Low | Very High |
| **Volatility** | Low | Low | High | Very Low |
| **Solubility** | Insoluble | Soluble in polar solvents | "Like dissolves like" | Insoluble |
| **Elec. Cond.** | Solid/Liquid: Good | Solid: Poor; Liquid/Aqueous: Good | Poor | Poor (Except graphite) |
| **Thermal Cond.** | Good | Poor | Poor | Poor (Except diamond) |
| **Brittleness** | Malleable/Ductile | Brittle | Brittle (if solid) | Very Hard/Brittle |
| **Corrodable** | Yes (often) | No | No | No |

#### II. Organic Chemistry

**A.** Hydrocarbons **& Polymers**

1. **Alkanes:** Saturated hydrocarbons, **single C-C bonds** (e.g., methane, ethane).
2. **Alkenes:** Unsaturated hydrocarbons, **at least one C=C double bond** (e.g., ethene).
3. **Alkynes:** Unsaturated hydrocarbons, **at least one C=C triple bond** (e.g., ethyne).
4. **Polymers:** **Macromolecules** formed by joining many small repeating units (monomers).
   * **Properties:** Strength (from chain length, entanglement, IMFs), cross-linking (increases rigidity), electrical/thermal conductivity (generally low), boiling/melting points (high).
5. **Addition Polymerization:** Monomers (alkenes/alkynes) **add to each other** without loss of atoms, breaking double/triple bonds.
   * **Naming:** (e.g., ethene → poly(ethene)).
6. **Biodegradable Polymers:** Polymers that **break down naturally** due to microorganisms (e.g., bacteria).
   * **Non-Biodegradable Polymers:** Do not break down easily (e.g., plastics with strong, unreactive C-C and C-H covalent bonds that bacteria struggle to cleave).
   * **Making more biodegradable:** Incorporating functional groups susceptible to hydrolysis (e.g., esters, amides) or using natural monomers.
7. **Condensation Polymers:** Monomers join by **losing a small molecule** (e.g., water, HCl) for each linkage.
   * **Esterification:** Reaction between a **carboxylic acid and an alcohol**, forming an **ester** and **water**.
     + You know it's esterification by seeing a carboxylic acid reacting with an alcohol, and the new compound formed contains an **-COO- (ester) linkage**.
   * **Naming Condensation Polymers:** Based on the functional groups formed (e.g., polyester, polyamide).
   * **Hydrolysis:** Reaction where **water breaks a bond**, usually in the presence of acid/base. Opposite of condensation.

**B. Organic Representations, Naming & Isomers**

1. **Full Structural Formula:** Shows **all atoms and all bonds**.
2. **Condensed Structural Formula:** Shows **atoms bonded to each carbon** but groups H atoms together (e.g., CH3​CH2​CH3​).
3. **Skeletal Formula:** Shows **carbon backbone as lines/zigzags**, with heteroatoms (non-C, non-H) and their H atoms shown. Carbons are at vertices/ends, H on carbons are implied.
4. **IUPAC** Naming Convention (Prefixes for **Carbon Chains):**
   * **Meth-**: 1 carbon
   * **Eth-**: 2 carbons
   * **Prop-**: 3 carbons
   * **But-**: 4 carbons
   * **Pent-**: 5 carbons, etc.
   * **Suffixes for bonds:** -ane (single), -ene (double), -yne (triple).
5. **Homologous Series:** Series of organic compounds with the **same general formula and functional group**, differing by a CH2​ unit, showing similar chemical properties and a gradual change in physical properties.
6. **Functional Groups:** Specific **atoms or groups of atoms** within a molecule that determine its characteristic chemical reactions.

### **I. Functional Groups & Priority Order**

Functional groups dictate a molecule's chemical behavior. Their priority influences IUPAC naming.

Priority Order (Highest to Lowest):

Carboxylic Acid > Ester > Amide > Aldehyde > Ketone > Alcohol > Amine > Alkyne (Triple Bond) > Alkene (Double Bond) > Ether (Alkoxy) > Halogenoalkane (Halogeno) > Alkane

| **Homologous Series** | **Formula** | **Suffix/Prefix** | **Naming Notes** |
| --- | --- | --- | --- |
| **Carboxylic Acid** | R-COOH | -oic acid | Carbonyl carbon is C1. |
| **Ester** | R-COO-R' | -oate | R' is the alkyl group (e.g., methyl), R is the alkanoate (e.g., ethanoate). |
| **Amide** | R-CO-NH2 | -amide | For N-substituted amides, use N-alkyl prefix (e.g., N-methylpropanamide). |
| **Aldehyde** | R-CHO | -al | Carbonyl carbon is C1. |
| **Ketone** | R-CO-R' | -one | Number chain to give carbonyl lowest number. |
| **Alcohol** (Hydroxyl) | R-OH | -ol / hydroxy- | Number chain to give -OH lowest number. |
| **Amine** | R-NH2 | -amine / amino- | Number chain to give -N lowest number. |
| **Halogeno** (Halogenoalkane) | R-X (X=F,Cl,Br,I) | halo- | Use prefix (fluoro-, chloro-, bromo-, iodo-). |
| **Alkoxy** (Ether) | R-O-R' | alkoxy- | Name as alkoxyalkane (e.g., methoxyethane). The smaller alkyl group forms the alkoxy part. |
| **Phenyl** | C6H5-  (one H is removed) | phenyl- | Used as a substituent on other chains or functional groups. |

### **II. Types of Molecules (Primary, Secondary, Tertiary)**

Classification is based on the number of **alkyl (R) groups** attached to a central atom (C for alcohols, N for amines/amides).

* **Alcohols:** The carbon bonded to the -OH group is classified by how many other alkyl groups it's bonded to (1, 2, or 3).
* **Amines:** The nitrogen atom is classified by how many alkyl groups it's bonded to (1, 2, or 3).
* **Amides:** Classified by the number of alkyl groups attached to the nitrogen atom of the amide group.

### **III. Naming Convention for Organic Compounds (IUPAC)**

1. **Identify the longest continuous carbon chain** containing the highest priority functional group.
2. **Number the parent chain** to give the highest priority functional group the lowest possible number.
3. **Identify and name substituents**.
4. Use **prefixes** (di-, tri-, tetra-) for multiple identical substituents.
5. **List substituents alphabetically** before the parent chain name.
6. Use **hyphens** to separate numbers and words (e.g., 2-methylpropane). Use **commas** to separate numbers (e.g., 2,2-dimethylpropane).

### **IV. Physical Trends in Homologous Series**

A homologous series is a group of compounds with the same general formula, differing by a -CH2- unit.

* **Boiling Point & Melting Point:** Generally **increase** with increasing chain length/molar mass due to stronger **London Dispersion Forces (LDFs)**.
* **Solubility in Water:** Generally **decreases** with increasing chain length as the nonpolar hydrocarbon portion becomes more dominant.
* **Viscosity & Density:** Generally **increase** with increasing chain length.

### **V. Different Types of Isomers**

Isomers have the same molecular formula but different arrangements of atoms.

* **Structural (Constitutional) Isomers:** Differ in the connectivity of atoms.
  + **Chain Isomers:** Different arrangement of the carbon skeleton (e.g., butane vs. 2-methylpropane).
  + **Positional Isomers:** Same carbon skeleton, but the position of a functional group or substituent differs (e.g., 1-propanol vs. 2-propanol).
  + **Functional Group Isomers:** Same molecular formula, but different functional groups.
    - **Alkoxy (Ether) and Hydroxyl (Alcohol):** e.g., C2H6O (ethanol and dimethyl ether)
    - **Aldehyde and Ketone:** e.g., C3H6O (propanal and propanone)
    - **Carboxyl (Carboxylic Acid) and Ester:** e.g., C2H4O2 (ethanoic acid and methyl formate)

### **VI. Stereoisomers**

Stereoisomers have the same connectivity but differ in the 3D spatial arrangement of atoms.

Stereoisomers  
├── Conformational Isomers (Conformers)  
│ └── Interconverted by rotation around single bonds (e.g., staggered vs. eclipsed ethane).  
└── Configurational Isomers  
 ├── Cis-Trans Isomers (Geometric Isomers)  
 │ └── Occur due to restricted rotation:  
 │ ├── \*\*Alkenes:\*\* Require different groups on each carbon of the C=C double bond.  
 │ │ ├── \*\*Cis:\*\* Same groups on the same side of the double bond.  
 │ │ └── \*\*Trans:\*\* Same groups on opposite sides of the double bond.  
 │ └── \*\*Cyclic Compounds:\*\* Substituents on different carbons of the ring.  
 │ ├── \*\*Cis:\*\* Substituents on the same side of the ring.  
 │ └── \*\*Trans:\*\* Substituents on opposite sides of the ring.  
 └── Optical Isomers (Enantiomers)  
 └── Require a \*\*chiral center\*\* (typically a carbon atom bonded to four \*different\* groups).  
 ├── \*\*Enantiomers:\*\* Non-superimposable mirror images of each other.  
 │ ├── Have identical physical properties except for their interaction with plane-polarized light (rotate it in opposite directions).  
 └── \*\*Diastereomers:\*\* Stereoisomers that are \*not\* mirror images of each other (e.g., compounds with multiple chiral centers).

### **VII. Analytical Techniques for Organic Chemistry**

* **Mass Spectrometry (MS):**
  + **Principle:** Ionizes molecules and separates them by mass-to-charge ratio (m/z).
  + **Information:**
    - **Molecular Ion Peak (M+):** The highest m/z peak, usually corresponding to the molecular mass.
    - **Fragment Ions:** Smaller peaks that provide structural clues; fragments are typically 1+.
* **Infrared (IR) Spectroscopy:**
  + **Principle:** Measures absorption of IR radiation, causing bonds to vibrate at specific frequencies.
  + **Information:** Identifies the presence or absence of specific functional groups based on characteristic absorption wavenumbers (e.g., O-H, C=O, N-H stretches).
* 1**H Nuclear Magnetic Resonance (NMR) Spectroscopy:**
  + **Principle:** Probes the magnetic environment of hydrogen nuclei (protons) in a molecule.
  + **Information:**
    - **Chemical Shift (**δ**):** Indicates the electronic environment of protons (shielding/deshielding).
    - **Integration:** The area under each peak is proportional to the number of equivalent protons.
    - **Splitting Pattern (Multiplicity):** Follows the **(n+1) rule**, where 'n' is the number of equivalent protons on adjacent carbons.
      * **Singlet (s):** n = 0 (no adjacent protons)
      * **Doublet (d):** n = 1
      * **Triplet (t):** n = 2
      * **Quartet (q):** n = 3
      * **Multiplet (m):** Complex splitting.
    - **Number of Signals:** Corresponds to the number of different types of chemically equivalent protons.
  + **TMS** is used since it’s chemically inert and has 12 chemically equivalent hydrogens whilst also producing a chemical shift way lower than most organic compounds