Chemistry Notes

Unit 1: Matter Matters

- **Matter** is anything that takes up space and has mass.
  - Mass is denoted by "m" and its unit is "g" or grams.
- Everything is made up of **particles**. Every element, compound, or mixture contains **atoms**.

Atoms

- An **atom** is the smallest particle in the universe.
- **Elements** are substances made from one kind of atom (e.g., Oxygen).
- **Compounds** are made from two or more elements that are chemically combined.

Subatomic Particles

- **Nucleus**: Positively charged central region of an atom, composed of protons and neutrons.
- **Protons**: Positively charged particles located in the nucleus.
  - Increasing the number of protons increases the atomic number.
- **Neutrons**: Located in the nucleus with no charge.
  - Impact the stability of the element.
  - More neutrons increase mass number, making the atom heavier.
  - To calculate the number of neutrons: $MassNumber - AtomicNumber(A - Z)$
- **Electrons**: Negatively charged particles located on the shells of an atom.
  - The first shell holds 2 electrons, and subsequent shells can hold 8 electrons.
  - **Valence electrons** are electrons on the outermost shell of an atom.

States of Matter

- **Solid**: Fixed volume and shape. Particles do not move much.
- **Liquid**: Fixed volume but not a fixed shape. Particles move somewhat freely.
  - Takes the shape of its container.
- **Gas**: No fixed volume or shape. Particles move very randomly.
  - Expands to take the shape and volume of the container.
  - Lighter than a solid and liquid of the same volume.

Phase Changes

- Freezing: Liquid → Solid
- Melting: Solid → Liquid
- Evaporation: Liquid → Gas
- Condensation: Gas → Liquid
- Deposition: Gas → Solid
- Sublimation: Solid → Gas

When a substance is heated, its molecules move faster with greater energy, increasing collisions and causing the substance to become less dense.

Kinetic Theory of Particles

- A substance can be a solid, liquid, or gas, and can change from state to state.
- Each state has different characteristics.
- Each state differs in how the particles are arranged and move.
- The temperature of a substance increases when the average kinetic energy of the particles increases.

> More particles moving = more collisions between particles = higher temperature.

Diffusion

> **Diffusion** is when particles flow from an area of higher concentration to an area of lower concentration until all particles are evenly distributed.

Factors Affecting Diffusion

- **Temperature**: Increasing temperature increases the rate of diffusion by increasing the energy of the particles, enabling them to move faster.
- **Concentration Difference**: A higher concentration difference results in a faster rate of diffusion.
- **Diffusion Distance**: The shorter the distance particles have to move, the faster they diffuse.
- **Mass of the Molecule**: The more mass a molecule has, the slower the rate of diffusion, as greater mass requires more energy to move it.

Pure and Impure Substances

- **Pure Substance**: A substance made up of only one type of molecule (e.g., oxygen or water).
- **Impure Substance**: A substance made up of different types of molecules (e.g., air or saltwater).
- **Mixture**: A material made up of two or more different substances that are physically combined and can be physically separated (never chemically combined).

Types of Mixtures

- **Homogenous Mixture**: Has a uniform composition.
- **Heterogenous Mixture**: Does not have a uniform composition.

It is possible to separate components in a heterogenous mixture but not in a homogenous mixture.

- **Alloy**: A substance formed from the combination of two or more metals.
- **Emulsion**: A mixture of two substances that originally don’t mix but can bind together with the aid of a chemical agent (emulsifier).
- **Colloid**: A state in which smaller particles are dispersed throughout a fluid.

Separation Techniques

- Solution: Mixture of two or more substances.
- Solute: Substance that is dissolved in a solution.
- Solvent: A substance that has the ability to dissolve a solute (normally a liquid).
- Soluble: Able to be dissolved.
- Filtrate: Product of filtration.
- Residue: The substance that remains in the filter paper.

Types of Separation Techniques

Technique	Description	Example
Filtration	Used to separate an undissolved solid from a mixture of the solid and a liquid solution. Enables the separation of insoluble solids from mixtures.	Sand and water, filtering coffee
Crystallization	Used to separate a dissolved solid from a solution when the solid is more soluble in a hot solvent than in cold.	Gentle heating of a saturated solution
Simple Distillation	Used to separate a liquid and a soluble solid from a solution.	
Fractional Distillation	Used to separate two or more liquids that are miscible with one another.	
Chromatography	Used to separate substances that have different solubilities in a given solvent (e.g., different colored inks mixed to make black ink).	Separating colored inks

Unit 2: Small Matters?

Atomic Structure (Including Isotopes)

- Atom: Smallest particle in the universe.

Subatomic Particles

- **Electron**: Negative charge, located in outer shells.
- **Neutron**: No charge, located in the nucleus.
- **Proton**: Positive charge, located in the nucleus.
- **Atomic Number**: Number of protons / Number of electrons.
- **Mass Number**: Number of Neutrons.
- To calculate the number of neutrons: $MassNumber(A) - AtomicNumber(Z)$

Isotopes

> **Isotopes** are atoms of the same element with the same number of protons (same atomic number) but different numbers of neutrons (different atomic mass).

- Example: Chlorine-35 and Chlorine-37.

Electron Configuration and Valency

Electronic Configuration

- The arrangement of electrons in the shells of an element.
- Shells are denoted by K, L, M, and so on.
- Example: Oxygen (atomic number 8) has an electronic configuration of "2,6".
- The first shell always has 2 electrons.

Valency

- Determined by the group (column) in the periodic table.
- Magnesium (Group 2) has a valency of 2+, meaning it gives away electrons to become stable.
- Chlorine (Group 7) has a valency of 1-, meaning it takes 1 electron to become stable.
- Elements in Group 4 can either give away or take electrons to become stable.

The Periodic Table

Organization

- Elements organized by atomic number.
- As you go down the rows, the number of outer shells increases.
- As you go through the columns (left to right), the atomic number increases.
- **Vertical Columns**: Groups, symbolize the number of valence electrons. Elements in the same group have the same valency.
- **Horizontal Rows**: Periods, represent the number of shells an element has.
- Atomic number = Number of protons OR Number of electrons.

Properties Based on Group

Group	Properties
Group 1 (Alkali Metals)	Good conductors of electricity, malleable.
Group 7 (Halogens)	Highly reactive with metals, different states at room temperature.
Group 8 (Noble Gases)	No reaction, gas at room temperature.

Trends

Group	Trends
Group 1 (Alkali Metals)	Melting point and boiling point go down as you go down the group, more reactive down the group.
Group 7 (Halogens)	Melting point and boiling point go up as you go down the group, less reactive down the group.
Group 8 (Noble Gases)	Melting point and boiling point go up as you go down the group, not reactive.

Properties Based on Atomic Mass

Group(s)	Ion Formation
1, 2, 3	Form +1, +2, +3 ions (lose electrons to become stable).
5, 6, 7	Form -3, -2, -1 ions (gain electrons to become stable).
4	Form either +4 or -4 ions (lose or gain electrons to become stable).

Transition Metals

- Charges vary (e.g., +2 or +3).
- Generally stronger with a high melting point.

Polyatomic Ions

- Ions made of two or more ions.

Metals and Non-metals

- Metals lose electrons to form cations with a positive charge.
- Non-metals gain electrons to form anions with a negative charge.
- Reactivity increases down the group in groups 1 and 2.

Group 1 (Alkalis)

- All alkalis are bases.
- Alkalis are basic substances that can be dissolved in water.
- Soft metals with 1 electron in their valence shell (+1 charge).
- Highly reactive and very shiny when cut.
- Form oxides when they react with O_2 or H_2O.
- Density increases, boiling point and melting point decreases.

Group 7 (Halogens)

- 7 electrons in their valence shell.
- Melting and boiling points increase down the group.
- Fluorine (pale yellow, gas), Chlorine (yellow-green, gas), Bromine (red-brown, liquid), Iodine (shiny purple, solid), Astatine (black, solid).
- Ready to gain 1 electron (-1 charge).
- Toxic as elements, but less toxic when they form ions.

Group 8 (Noble Gases)

- Full valence shells, completely stable.
- Do not react.
- He, Ne, Ar, Kr, Xe, Rn.
- Colorless.
- Exist as atoms in nature (monoatomic).
- Boiling point and density increase down the group.

Transition Metals

- High melting points.
- Low reactivity.
- Hard and have more strength.
- Form ions with different charges and a range of colors.
- Similar to each other.
- Some have magnetic properties.
- Malleable and ductile.

Bonding (Structure and Bonding, Properties)

Ionic Bonding

> **Ions** are atoms that are positively or negatively charged. Occurs when atoms exchange electrons with each other to fulfil this. Both atoms have a full valence shell.

- **Monatomic Ion**: A single ion (Helium He).
- **Polyatomic Ion**: Formed from groups of ions (Oxygen O_2).
- Electrostatic attraction between two oppositely charged ions.
- Occurs between metals and non-metals; metal loses an electron, non-metal gains an electron.
- Regular repeating arrangement called an ionic lattice.

When naming an ionic compound, the metal comes first.

- Atom loses electron = positively charged (vice versa).

Properties

- High melting points and boiling points.
- Ions with higher charge have higher melting points (stronger electrostatic forces).
- Can conduct electricity in liquid or aqueous solution.
- Dissolves in water.
- Does not dissolve in organic compounds.

Covalent Bonding

> **Covalent bonding** occurs between two non-metals. Share electrons with another atom instead of gaining or losing.

- No conductivity.
- **Single Bonds**: Single pair of electrons shared (2 electrons) (weak forces, easy to break).
- **Double Bonds**: Double pair of electrons shared (4 electrons).
- **Triple Bonds**: Triple pair of electrons shared (6 electrons) (very strong, hard to break).

Properties

- Low melting points and boiling points.
- Weak intermolecular forces.
- Do not conduct electricity.
- Bonds are strong.
- Need lots of energy to break them.
- Exist in nature as gases.

Unit 3: Energy and Development

Energy Changes in Reactions

Endothermic and Exothermic Reactions

- **Exothermic Reaction**: Releases heat energy as the reaction happens. The solution becomes warmer.
- **Endothermic Reaction**: Absorbs heat energy as the reaction happens. The solution becomes colder.

Examples

- **Exothermic**: Combustion, Respiration, Neutralization, Dissolving acids, Dissolving alkalis, Rusting, Oxidation of metals, Nuclear reaction.
- **Endothermic**: Thermal Decomposition, Dissolving some ionic salts in water (e.g., ammonium chloride, potassium nitrate, copper (II) sulphate), Photosynthesis, Action of light on a silver bromide.

Enthalpy Change (ΔH)

- Amount of heat evolved or absorbed in a reaction.
  - Negative value: Energy released (exothermic).
  - Positive value: Energy absorbed (endothermic).
- **Activation Energy**: The required amount of energy needed for the reaction to occur.
  - A catalyst may work by lowering the activation energy for a reaction.

Combustion of Fuels

- Combustion (burning) involves the reaction of a hydrocarbon and oxygen to produce carbon dioxide and water.
  - **Complete Combustion**: Sufficient oxygen produces carbon dioxide.
    - $C_xH_y + O_2 \rightarrow CO_2 + H_2O$
  - **Incomplete Combustion**: Insufficient oxygen produces carbon monoxide.
    - $C_xH_y + O_2 \rightarrow CO + H_2O$

Unit 4: Let’s Count

Chemical Formulas

- **Diatomic Elements**: Elements that always exist as two atoms together:
  - $H_2, O_2, Br_2, I_2, N_2, Cl_2$ (HOBrINCl)

Chemical Reactions and the Conservation of Mass

- **Relative Atomic Mass**: Calculated using the equation:
  - $A_r = \frac{(}{100}$
  - Example: Chlorine (75% Chlorine-35, 25% Chlorine-37)
    - $A_r = \frac{(75 \cdot 35) + (25 \cdot 37)}{100} = 35.5$

Balancing Equations

- **Atom Economy**: Measure of the amount of starting materials that end up as useful products.
- **Percentage Yield**: Shows how much product is obtained compared to the maximum possible mass.
- Atom economy of a reaction gives the percentage of atoms in reactants that form a desired product.

The Mole Concept and Chemical Calculations

- **Concentration**: Amount of solute / Volume of solute (mass / volume)
  - $Concentration = \frac{\text{moles}}{\text{volume}}$

Conversions

- $1000cm^3 = 1dm^3 = 1 \text{ liter}$
- $22.4L \text{ of gas} = 1 \text{ mole}$

Empirical Formula

> **Empirical Formula**: Shows the relative numbers of atoms of each element present, using the smallest whole numbers of the atoms.

- Example: Hydrogen peroxide's empirical formula is HO, with a 1:1 ratio.

Molecular Formula

Molecular Formula: The ACTUAL number of atoms of each element in a molecule.

- Example: Hydrogen peroxide's actual molecular formula is H_2O_2, with the ratio still being 1:1.
- $\text{Molecular formula} = \frac{\text{Molar Mass}}{\text{Empirical Mass}}$
- $\text{Moles} = \frac{\text{grams}}{\text{atomic mass}}$
- $1 \text{ mole/Avogadro's number} = 6.022 \times 10^{23} \text{ atoms or molecules}$
- Atoms: To convert from moles to atoms, multiply the molar amount by Avogadro's number. To convert from atoms to moles, divide the atom amount by Avogadro's number.
- Molecules: $\text{moles} \times \text{Avogadro's number}$
- $\text{Mass} = \text{number of moles} \times \text{molar mass}$
- $\text{Molar mass} = \frac{\text{mass}}{n}$
- $n = \frac{\text{mass}}{\text{molar mass}}$

Unit 5: How Fast? Unit 6: How Far?

Rates of Reaction

- Different reactions take place at different rates.
- Several indicators show whether a chemical reaction has taken place:
  - Color change
  - Effervescence
  - Precipitation
  - Energy change (temperature)
- Reactions with oxygen can be fast (burning) or slow (rusting).

Collision Theory

Collision theory is used to predict rates of chemical reaction, particularly for gases.

- The more particles in a container, the more the chances of collision between those particles.
- This theory helps us understand what's happening at the atomic level when reacting particles collide.
- For a reaction to occur:
  - The reacting particles must collide with each other.
  - The colliding particles must have the correct orientation at the time of collision.
  - The particles must have the minimum kinetic energy required to initiate a reaction (activation energy).
- Particles will collide more frequently if:
  - There is a higher concentration.
  - Particles are reacting faster.

Factors Affecting Rate of Reaction

- **Surface Area**: When a solid is powdered, the particles are more exposed, increasing the chance of a collision.
- **Concentration**: More particles result in a higher chance of collision.
- **Temperature**: When temperature increases, the particles gain kinetic energy, so more collisions take place.

Reversible Reactions

Dynamic Equilibrium

> **Equilibrium**: A state in which opposing forces or influences are balanced.

> **Dynamic Equilibrium**: A system in a steady state since forward reaction and backward reaction occur at the same time. No overall change in the amount of products and reactants, even though the reactions are ongoing.

- Takes place in a closed system; otherwise, the products would escape.
- Reversible reactions that happen in a closed system eventually reach dynamic equilibrium.
- The position of equilibrium is said to shift to the right when the forward reaction is favored.
  - This means that there is an increase in the amount of products formed.
- The position of equilibrium is said to shift to the left when the reverse reaction is favored.
  - So, there is an increase in the amount of reactants formed.

Le Chatelier's Principle

> **Le Chatelier's Principle** can be used to predict the effect of changes in temperature on systems in equilibrium.

- If the temperature of the reaction increases:
  - The equilibrium will shift in the direction of the endothermic reaction.
- If the temperature of a reaction decreases:
  - The equilibrium will shift in the direction of the exothermic reaction.
- An increase in pressure will favor the reaction that produces the least number of molecules.
- A decrease in pressure will favor the reaction that produces the greatest number of molecules.

The Haber Process

- The industrial process for the manufacture of ammonia from hydrogen and nitrogen.
- Hydrogen is obtained from the reaction of methane and steam, producing carbon monoxide as a byproduct.
- The hydrogen produced from this reaction also reacts with oxygen from the air, producing water and leaving nitrogen behind.
- The air is 77% nitrogen.
- These gases are then compressed and delivered to the reactor where ammonia is produced.
- These gases are then cooled off, and ammonia is liquified, ready to be tapped off.
- The unused hydrogen and nitrogen are recycled back to the reactor.

Unit 7: pHun Meeting @ Endpoint

Acids and Bases

Acids

Acid: A solution that has an excess of hydrogen (H^+ ions). The more H^+ ions, the more acidic the solution.

Properties

- Tastes sour.
- Conducts electricity.
- Corrosive.
- Some acids react strongly with metals, releasing H_2 (g).
- Reactions with carbonates release CO_2 (g).
- Turns blue litmus paper red.

Uses

- Acetic acid = vinegar.
- Citric acid = lemons, limes, and oranges.
- Ascorbic acid = vitamin C.
- Car batteries.

Bases

Base: A solution that has an excess of OH^- ions or it accepts H^+ ions. Forms aqueous solutions; also called Alkali.

- $H^+ + OH^- \rightarrow H_2O$

Properties

- Feels slippery.
- Tastes bitter.
- Corrosive.
- Destroy body tissue/ dissolve fatty (lipid) material.
- Can conduct electricity.
- Do not react with metals.
- Neutralize solutions containing hydrogen ions (H^+).
- Turns red litmus paper blue.

Uses

- Bases give soaps, ammonia, and many other cleaning products.
- The OH^- ions interact strongly with certain substances, such as dirt and grease.
- Chalk and oven cleaner are examples of familiar products that contain bases.
- Your blood is a basic solution (around pH 7.4).

The pH Scale

> The pH Scale is a scale running from 0 to 14 for expressing the acidity and the alkalinity of a solution.

- Measures the concentration of H^+ ions.
- $pH < 7 = \text{acid}$
- $pH > 7 = \text{Base}$
- $pH = 7 = \text{Neutral}$

Reactions of Acids and Bases

Oxides

- Metals form oxides that are basic, but non-metals form oxides that are acidic.
- Sulfur and Carbon are both non-metals that react with oxygen to form sulfur dioxide and carbon dioxide. These are gases which dissolve in rainwater making it acidic.

Metal and Non-Metal Oxides

- **Metal Oxides**: Calcium Oxide (CaO), Zinc Oxide (ZnO), etc.
- **Non-Metal Oxides**: CO_2, SO_2, NO_2, etc.
- Metal oxides react with water to form a hydroxide:
  - $CaO + H_2O \rightarrow Ca(OH)_2$
  - The pH level is greater than 7 making it alkaline ($pH > 7$)
- Non-metal oxides react with water to form a hydrogen:
  - $CO_2 + H_2O \rightarrow H_2CO_3$ which makes it a carbonic acid
  - The pH level is lower than 7 making it acidic ($pH < 7$)

Concentration and Strength

- **Concentration**: Deals with the amount of hydronium ions in the solution compared to the amount of water.
  - More acid or base and less water = more concentrated
- **Strength**: More ions and less molecules = stronger
- $Concentration = \frac{\text{how much acid}}{\text{certain volume}}$

Strong vs Weak Acids

- **Strong acids** ionize completely (e.g., H_2SO_4, HCl, HNO_3).
- **Weak acids** don’t fully ionize (less than 5%) and not all of the acid particles dissociate completely, it is a reversible reaction. (e.g., $H_3PO_4, HC_2H_3O_2$,organic acids)

Acid pH Indicators

- Litmus paper (blue to red)
- Phenolphthalein (colorless)
- Methyl orange (orange to red)
- Bromothymol Blue (Changes to yellow)

Salt Preparation Methods

Neutralization Reaction

- Acid + Metal Oxide → Salt + water
    - $HCl + Na_2O \rightarrow NaCl + H_2O$
- Acid + Metal Hydroxide → Salt + Water
    - $H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$
    - Combine the negative and positive ion on the reactants side to form the salt (Cl + Na)
- Acid + Metal Carbonate → Salt + Water + Carbon Dioxide
    - $HNO_3 + CaCO_3 \rightarrow Ca(NO_3)_2 + H_2O + CO_2$

Neutralization Reaction Apparatus

1. Take a dilute acid (HCl) in a beaker.
2. Heat with a Bunsen burner.
3. Put insoluble base like copper oxide a little bit at a time and it will disappear.
4. Once you can see some of the insoluble copper oxide in the beaker that means it is not disappearing, and this means we haveneutralized all of the acid.

Unit 8: ReDox

Oxidation Reduction

Redox Reactions

Redox reactions are of two types: Reduction & Oxidation

1. Oxidation: gain of oxygen, loss of electron, loss of hydrogen. (GOLEGH) The atom undergoing oxidation is the reducing agent
    - $2MgO + H_2O \rightarrow 2MgO$ (metal is oxidized)
    - $4Al + 3O_2 \rightarrow 2Al_2O_3$ (gain of oxygen to metal)
2. Reduction: loss of oxygen, gain of electron, loss of hydrogen (LOGELH) The atom undergoing reduction is the oxidizing agent.
    - $2MgO \rightarrow 2Mg + O_2$ (oxygen is reduced)
    - $2Al_2O_3 \rightarrow 4Al + 3O_2$ (loss of oxygen in metal)

Oxidation

- Most metals exist as oxides (magnesium oxide) in nature due to an abundance of oxygen. Only unreactive metals exist as pure (gold)
- Loss of Electrons With loss and gain of electrons, if they lose electrons we can say it has been oxidized, if it gains electrons, we can say it has been reduced.

Reduction

- Is Normally oxidation and reduction don't happen by themselves, both take place at the same time. When it happens at the same time, its called Redox reaction.
- Gain of Electrons

Example

- $\text{Oxidized (loses electrons)}: Mg + 2H^+ \rightarrow Mg^{2+} + H_2$
- $\text{Reduced (gains electrons)}$

Displacement Reactions

- A more reactive metal displacing a less reactive one.
    - $Ca + Fe^{2+}SO_4^{2-} \rightarrow Ca^{2+}SO_4^{2-} + Fe$
    - Calcium would displace the iron to form Calcium Sulphate and the Iron will precipitate out as a solid, CALCIUM SULPHATE ACT AS SPECTATOR IONS
    - $Ca + Fe^{2+} \rightarrow Ca^{2+} + Fe$
    - Ionic Equation, Spectator Ions got eliminated
    - The second step is known as an ionic equation, which only shows the particles that take part in the reaction and changes in some way

Individual Half Equations

- Shows us the gain and loss of electrons
    - $Ca \rightarrow Ca^{2+} + 2e^- \text{ (loses 2 electrons, Oxidation)}$
    - $Fe + 2e^- \rightarrow Fe^{2+} \text{ (gains 2 electrons, Reduction)}$

Electrochemistry

Electrolysis

- It is used to separate elements in insoluble ionic compounds like Lead Bromide.
- Electrolysis literally means splitting up with electricity, to pass electric current through the electrolyte, like if lead bromide ($PbBr_2$) was the electrolyte then the negative bromide ions present in the beaker would be attracted to the anode and it will be discharged. It will convert from a negatively charged ion to a neutral atom, these type of atoms normally form a gas and flow up. The positive lead ions will get attracted to the cathode and get discharged and become pure lead, this will cause it to fall down and become a solid, it will form a layer of molten lead.
- The ions are being oxidized and reduced at the electrodes.
  - At the anode it will be oxidized: $2Br^- \rightarrow Br_2 + 2e^-$
  - At the cathode it will be reduced: $Pb \rightarrow Pb^{2+} + 2e^-$

Equipment Used

- Cathode (-)
- Anode (+)
- Beaker
- Electrolyte

Electroplating

- An electric current is passed through a solution that conducts electricity called an electrolyte, to create this current, two electrodes are dipped into the electrolyte solution and connected to a battery or a power supply.

Copper Plating

- We have a copper anode and a brass cathode, with a copper sulphate solution ($CuSO_4$). The electrolyte will have positively charged Copper Ions (Cu^{2+}) and Negatively charged Sulphate ions (SO_4^{2-}). The Copper Ions will be attracted to the negative electrode (the Cathode), the copper ions will deposit onto to the brass cathode and form a thin copper plate.
- The positively charged copper ions will GAIN electrons which is known as reduction: $Cu^{2+} + 2e^- \rightarrow Cu(s)$
- The negatively charged sulphate ions will get attracted to the positively charged copper anode. The current supplied to the anode causes the copper atoms to oxidize (lose electrons) and then dissolve into the electrolyte solution.
- Electrons move from anode to cathode (positive to negative) along the wire.
- Copper anode bar gradually dissolves to replenish the copper ions in the electrolyte solution. The solution will stay at the same concentration.
- Loss of electrons – anode
- gain of electrons – cathode

Factors to Consider

- **Stronger Current**: Increases speed at which ions and electrons move through the circuit. One way to increase current is to increase concentration of solution.

Oxidation State

- In an ionic or covalent compound, the oxidation state is always 0.

Unit 9: Organic Chemistry

Alkanes and Alkenes

Organic Chemistry: particularly focused in Crit A&D.

Hydrocarbon: A compound made up of ONLY hydrogen (H) and carbon (C).

Crude Oil: Unfiltered oil. It is a mixture of hundreds of hydrocarbons that need to be separated out in an oil refinery

Environmental Impact of Oil Extraction

- One of the main impacts would be Pollution, Oil and gas operations / drilling could release many tons of harmful pollutants in the air and therefore discharged angerous chemicals into the water. Which will degrade the clean air of water.

Process of Crude Oil Extraction

- Mining/ Drilling
- Fractional Distillation
- Refining
- Smallest hydrocarbon chain length get collected at the top of the fractionating column
- Highest hydrocarbon chain length get collected at the bottom of the fractionating column.
- Carbon chain length is directly proportional to temperature.

Combustion and Substitution Reactions

> Fuel: All fuels are extracted from crude oil, it is a material that is used to produce heat or power by burning/ combustion

Types of Fuels

Liquified Petroleum Gas, Propane, Methane, CNG, Ethanol, Gasoline, etc.

Method of Extraction of Crude Oil

Fractional Distillation. Refer to separation techniques

- Crude oil is heated, and it gets transferred to a separating column and the different components have different boiling points, the ones with lower boiling points get elevated up (in the column) and they form more gaseous states, the different components with

Crude Oil Fractionation

Higher boiling points result in components settling at the bottom of the **fractionating column** in more solid states. This elevation and demotion occur due to **density**, **viscosity**, and the **carbon chain length**, which requires more energy to overcome attractive forces.

- Viscosity & Chain length increase down the column.
- Flammability decreases down the column.

Alkanes

Alkanes are characterized by **single covalent bonds**. Examples include:

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

Prefixes for naming alkanes:

- Pent - 5
- Hex - 6
- Hept - 7
- Oct - 8
- Non - 9
- Dec - 10

Physical Properties

- **Boiling Point**: Increases with carbon chain length (direct proportionality).
- **Melting Point**: Generally increases with molar mass and carbon chain length (direct proportionality).
- **Solubility**: Non-polar, immiscible with water, and soluble in organic solvents.

Sources

Alkanes are extracted from **non-renewable fossil fuels** like crude oil, natural gas, coal, and oil.

Combustion

Alkanes react with oxygen, though they are relatively unreactive in normal cases. The state of matter depends on **Van Der Waal's forces**. Liquids come from crude oil, and solids come from coal.

- **Ignition**: Alkanes with higher chain lengths are harder to ignite.
- **Flame Color**: Incomplete combustion produces a yellow flame; complete combustion produces a blue flame.

Saturation

Alkanes are **saturated**, meaning they only have carbon-carbon single bonds.

Alkenes

Alkenes have **one C=C double bond** in their structural formula and contain only carbon and hydrogen.

Unsaturation

Alkenes are **unsaturated**, allowing atoms to be added to the formula.

Physical Properties

Alkenes generally share similar physical properties with alkanes.

Isomers

> **Isomers**: Compounds with the same molecular formula but different structural formulas.

Example: But-1-ene and But-2-ene.

Reactivity

- Alkanes are saturated and less reactive.
- Alkenes have double bonds that can be broken, making them more reactive.

Reactions

- Alkanes undergo **substitution reactions** due to being saturated.
- Alkenes undergo **addition reactions**, forming alkanes.

During addition, the double bond becomes a single bond, allowing another element to attach.

Combustion of Propane

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

(Incomplete combustion yields carbon monoxide.)

Ethane and Bromine Reaction

$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$

The double bond between carbon atoms breaks, forming a product with all single bonds.

Cracking

> **Cracking**: The process of dividing hydrocarbons into shorter chains using heat or pressure.

It involves breaking long-chain hydrocarbon molecules into smaller, more useful products, such as smaller alkanes, alkenes, and hydrogen. Naphtha from crude oil is the main source.

Uses of Products

- Smaller alkanes are used as fuels.
- Smaller alkenes are used as polymers in the plastic industry.
- **Catalysts** lower activation energy and alter the reaction rate without being consumed.

Importance

Cracking produces:

- Smaller alkanes: better fuel
- Alkenes: starting material for plastics
- Hydrogen: used in ammonia production and as a fuel

General Equation

Higher Alkane → Smaller alkanes + alkenes + hydrogen

Conservation

The number of C and H atoms are conserved.

Reaction

There is no single unique reaction; hydrocarbons break randomly, forming a mixture of shorter hydrocarbons and H_2.

Products

Any combination of alkane, alkene, and hydrogen can be made, provided the numbers are balanced.

Testing for Alkenes

Prefixes:

- Meth – 1
- Eth – 2
- Prop – 3
- But – 4
- Pent – 5
- Hex – 6

With addition reactions for alkenes (e.g., C_3H_6), the carbon double bond opens up, allowing a new molecule to be added.

Types of Addition Reactions

1. **Water (H_2O)**
2. **Hydrogen (H_2)**
3. **Halogens**

Hydrogenation

Propene (alkene) reacts with hydrogen to form propane (alkane) by breaking the double bond.

Hydration

Ethene reacts with water to produce ethanol, a type of alcohol ($C_2H_4 + H_2O \rightarrow C_2H_5OH$). Ethanol is commonly used in industries and alcoholic beverages. Fractional distillation separates ethanol from unwanted ethene and water.

Halogenation

Add bromine water (orange solution) to two tubes, one with alkane gas and the other with alkene gas. The alkene tube will become colorless due to an addition reaction.

$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$ (dibromoethane)

This bromine test distinguishes between alkenes and alkanes.

Reactivity Comparison

- Alkenes: Double bonds - more reactive
- Alkanes: Single bonds - less reactive

Alcohols

Alcohols are another homologous series where one hydrogen in an alkane is replaced by a hydroxide group (OH).

Nomenclature

Use the same prefixes as alkanes, but replace the final "e" with "ol" (e.g., methanol, ethanol).

Properties

Properties change as molecules get bigger:

- Flammable
- Soluble
- Can be oxidized to form carboxylic acids

Combustion

Undergo complete combustion in air to form carbon dioxide and water.

Solubility

Dissolve in water to form neutral solutions.

Isomerism

> Isomers: Molecules with the same molecular formula but different structural formulas.

Example: C_4H_{10} can be butane or 2-methylpropane.

Naming Isomers

- Identify the longest carbon chain.
- Number the carbon atoms to give the substituent the lowest possible number.
- Name the substituents and indicate their position.

Example: 2-methylbutane. If a methyl group were connected to the third carbon, numbering would start from the other end.

If there are two methyl groups on the same carbon, it becomes 2,2-dimethylpropane (using the prefix "di" for two).

Isomer Properties

- **Long-chain hydrocarbons**: More surface contact, higher boiling and melting points, more reactive.
- **Branched structures**: Less surface contact, lower boiling and melting points, less reactive.

Positional Isomerism in Alkenes

In alkenes, isomerism is based on different positions for the double bonds. A double bond closer to the end of the chain is more reactive.

Important Rule: If the extension of the hydrocarbon is connected to the first carbon in the long chain, it remains the same name. Each carbon should have four bonds.

Alcohol Isomers

Example: Propan-1-ol (C_3H_7OH OR $CH_3CH_2CH_2OH$)

Alcohols can be oxidized to form carboxylic acids.

Carboxylic Acids

Carboxylic acids are a homologous series with the functional group -COOH.

- One oxygen is double-bonded to carbon, and the other is single-bonded to oxygen, with hydrogen bonded to the single-bonded oxygen.
- Names end in "anoic acid".

Properties

Weak acids that do not fully ionize.

$C_2H_5COOH \rightleftharpoons C_2H_5COO^- + H^+$

$C_2H_5COO^-$ is called a propanoate ion. Carboxylic acids react with metal carbonates to produce salt, water, and CO_2:

Carboxylic acid + metal carbonate → salt +water +CO_2

Example:

Ethanoic Acid + Potassium carbonate → Potassium Ethanoate + Water + CO_2

Uses

- Vinegar contains ethanoic acid.
- Oranges and lemons contain citric acid.
- Aspirin and Vitamin C are carboxylic acids.

Production

Carboxylic acids are made by oxidizing alcohols using an oxidizing agent.

Esters

Esters have the functional group -COO- (ester link or ester group) and often have pleasant smells. They are volatile and found in perfumes and food flavorings.

Formation

Carboxylic Acid + Alcohol → Ester + water

Example: Ethyl Ethanoate ($CH_3COOC_2H_5$) is produced from Ethanoic Acid (CH_3COOH) + Ethanol (C_2H_5OH) with an acid catalyst.

$$CH_3COOH + C_2H_5OH \rightarrow CH_3COOC_2H_5 + H_2O$$

The OH group from the carboxylic acid and the H from the OH group in the alcohol are lost, forming water.

> **Esterification**: Alcohols react with carboxylic acids; the reaction is reversible, and esters have characteristic smells.

Polymers

> **Polymers**: Multiple monomers combined together.

Unsaturated monomers form saturated polymers.

Example: Ethene + Ethene + ethene... → Poly(ethene)

This is called **addition polymerization**, which does not form any byproducts and uses only one type of unsaturated organic monomer.

Naming

Add "poly" before the monomer name and enclose the monomer's name in brackets.

Uses of Poly(ethene)

Plastic bags, plastic bottles, containers, cling film, plastic buckets, and hoses.