# GENERAL CHEMISTRY 101

**Course Outline**

- Atoms, molecules, elements, and compounds and chemical reactions
- Modern electronic theory of atoms
- Electronic configuration, periodicity, and building up of structures of solids
- Chemical equations and stoichiometry
- Hybridization and shapes of simple molecules
- Valence forces
- Structure of solids
- Chemical bonding and intermolecular forces
- Kinetic theory of matter
- Elementary Thermochemistry
- Rate of reactions, equilibrium, and thermodynamics
- Acids, bases, and salts
- Properties of gases
- Redox reaction
- Introduction to electrochemistry
- Radioactivity

## ELECTRONIC CONFIGURATION

Electronic Configuration shows the arrangement or distribution of electrons in the orbitals present in an atom.

An orbital is a space or region in an atom where the probability of binding an electron is high.

### Energy of Orbitals

When writing electron configuration, orbitals of lower energies are filled to the maximum before those of higher energies. The types of orbitals include

1. S-orbital
2. P-Orbital
3. D-Orbital
4. F Orbital
    - The s orbital has the lower energies followed by subsequent orbitals (s<p<d<f).
    - There are three types of p Orbitals Px,Py,Pz. For the p orbitals since they are 3 types, 6 electrons are required to maximally fill the p orbital. The three types of P have the same energies. Therefore, the orbitals are said t0be “Degenerated”.
    - !0electrons are required to completely fill the d orbital

Aufbau filling Rule

The Aufbau filling rule places electrons in atoms according to their energy levels(shell)

The number of orbitals is very important. The number of orbital within in a particular shell or energy level correspond to the shell number and subsequent shell number increase their numbers consecutively.

1s

2s 2p

3s 3p 3d

4s 4p 4d 4f

5s 5p 5d 5f 5g

*__Electronic Configuration__*

$13=1s^2 2s^2 2p^6 3s^2 3p^1$

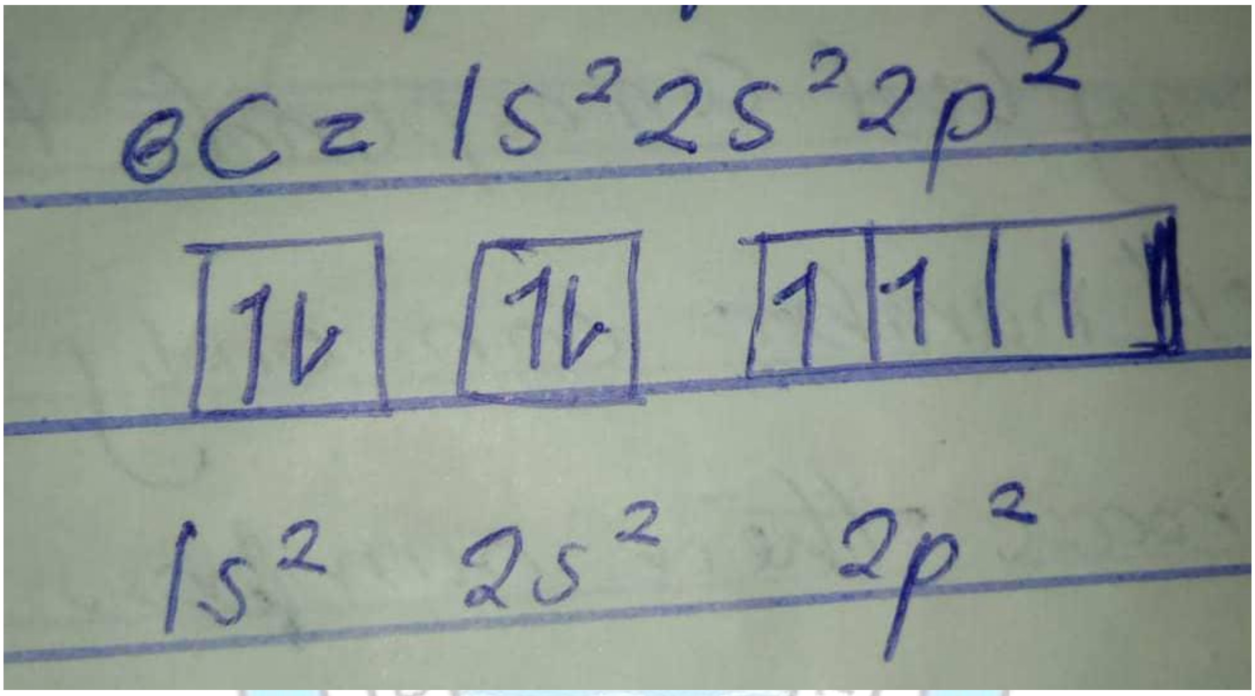
$27=1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$

$49=1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^1$

## Hund's Multiplicity Rule

It applies to regenerate orbitals. It allows us to distribute electrons in subshells of the same energy (regenerate orbitals).

If orbitals have the same energy, electrons are distributed singly first before pairing them.




## The Periodic Table

The periodic table shows the arrangement of elements according to their atomic numbers.

The periodic table is divided into two categories:

1. Group
2. Period

### Group

The group in a periodic table shows the arrangement of elements according to their valence electrons (electrons in the outer shell).

There are eight groups in the periodic table.

**GROUP 1:** Groups that have 1 electron in the valence orbital (1s^1). The valence shell is the shell with the highest principal quantum number.

**GROUP 2:** These are elements that have two electrons in the valence s-orbital e.g Beryllium.

** Groups 1 & 2 elements are known as s-block elements. Electrons are entering their valence orbital. **

**GROUP 3:** These are elements that have 3 valence electrons e.g Boron.

**GROUP 4:** These are elements that have 4 valence electrons e.g Carbon (1s^2 2s^2 2p^2) Silicon (1s^2 2s^2 2p^6 3s^2 3p^2).

**GROUP 5:** Nitrogen, N = 1s^2 2s^2 2p^3.

**GROUP 6:** Oxygen, Sulphur.

**GROUP 7:** Fluorine, Chlorine.

**Groups 3 to 7 elements are known as P block. Because electrons enter their orbit.**

**GROUPS 8** elements are elements with filled shells. They are said to have octet electron configuration. They include neon, Helium, Argon, Krypton.

**<u>The Period</u>**

The period in the periodic table shows the arrangement of elements according to their number of shells or principal quantum numbers. The highest principal quantum number in the atom or element gives the period.

Example: $_{47}$X= Group 3    Group 5

| | I | | | | | | | VIII |
|---|---|---|---|---|---|---|---|---|
| 1 | $_1$H | II | III | IV | V | VI | VII | $_2$He |
| 2 | $_3$Li | $_4$Be | $_5$B | $_6$C | $_7$N | $_8$O | $_9$F | $_{10}$Ne |
| 3 | $_{11}$Na | $_{12}$Mg | $_{13}$AL | $_{14}$Si | $_{15}$P | $_{16}$S | $_{17}$Cl | $_{18}$Ar |
| 4 | $_{19}$K | $_{20}$Ca | | | | | | |

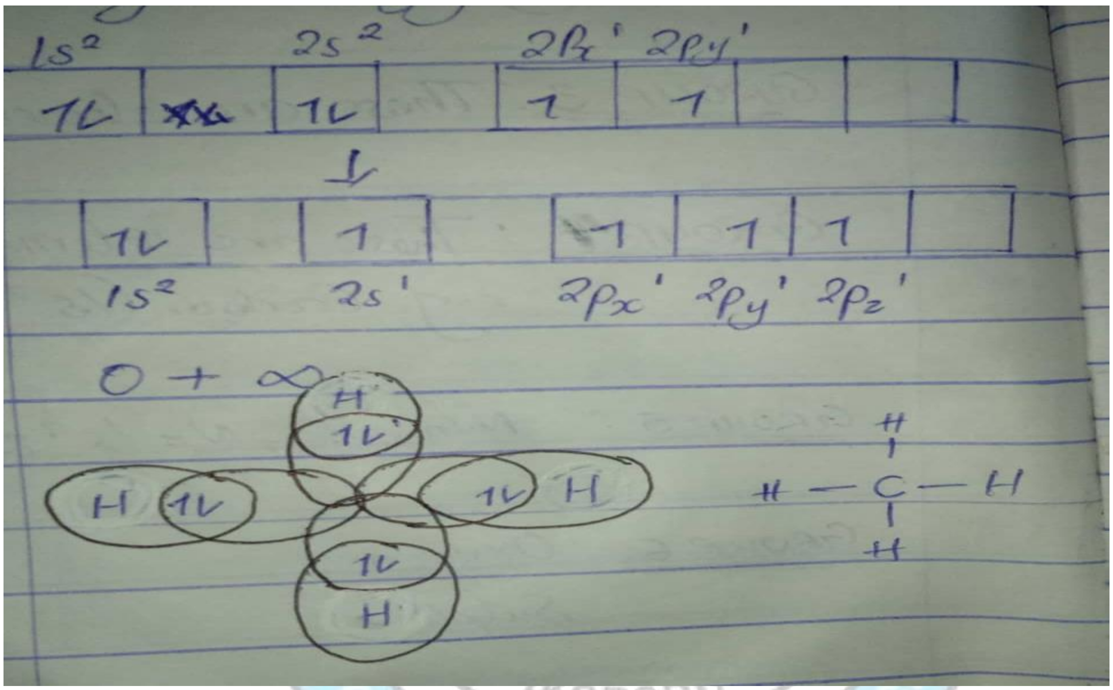
## HYBRIDIZATION

Hybridization is the mixing of two orbitals to obtain a new orbital of different properties.

**<u>Hybridization of Carbon</u>**:

Carbon - $1s^2\ 2s^2\ 2p^2$ (Ground state electron configuration)

Moving electron from a lower energy to higher energy




Carbon atom has 6 electrons and the electronic configuration is $1s^2 2s^2 2p^6$ from this electron configuration carbon appears to have a valency of 2, but we know that Carbon has valency of 4

**Question:** How to give Carbon a valency of 4 based on the electronic configuration above

**Answer:** One of the electrons in the 2s orbitals must be promoted to the empty $2P_2$ sub orbital

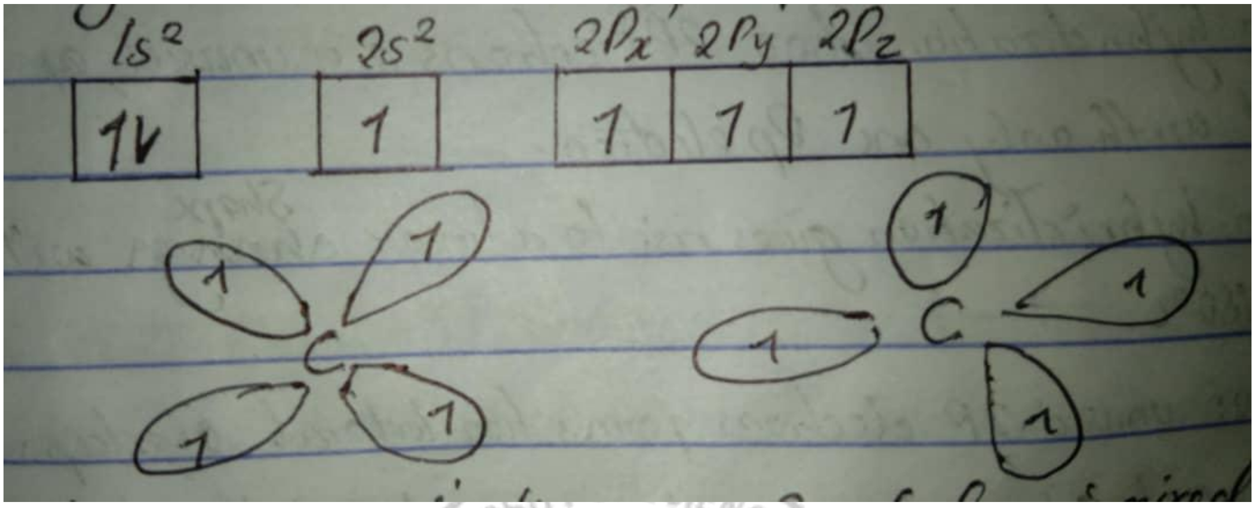
The 2s electron is mixed with 3p electrons to give a new set of elements known as the $Sp^3$ Orbital. This process is called HYBRIDIZATION

Then carbon now possess four (4) single electrons, thus acquiring a valency of 4. This is also known as tetrahedral hybridization because the orbital formed has a tetrahedral shape

($sp^3$ bond angle 109.5°)

($sp^2$ hybridization bond angle 120°)

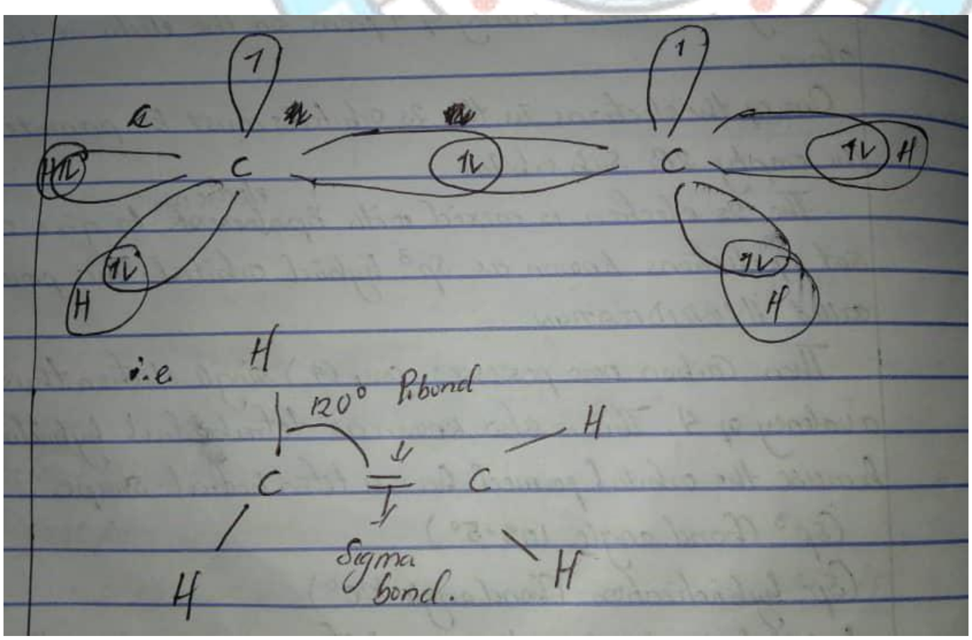
In $sp^2$ hybridization, the $2p_2$ electron is unused for the formation of a sigma bond.




For $sp^2$ hybridization, one 2s electron is mixed with the two 2p electrons to give $sp^2$ hybrid orbitals. The shape of the $sp^2$ orbital is trigonal. For this reason, it is called trigonal hybridization.

The unused 2Pz electron forms a lateral overlap with an adjacent carbon atom to produce a Pi-bond.

$\pi$ bond.

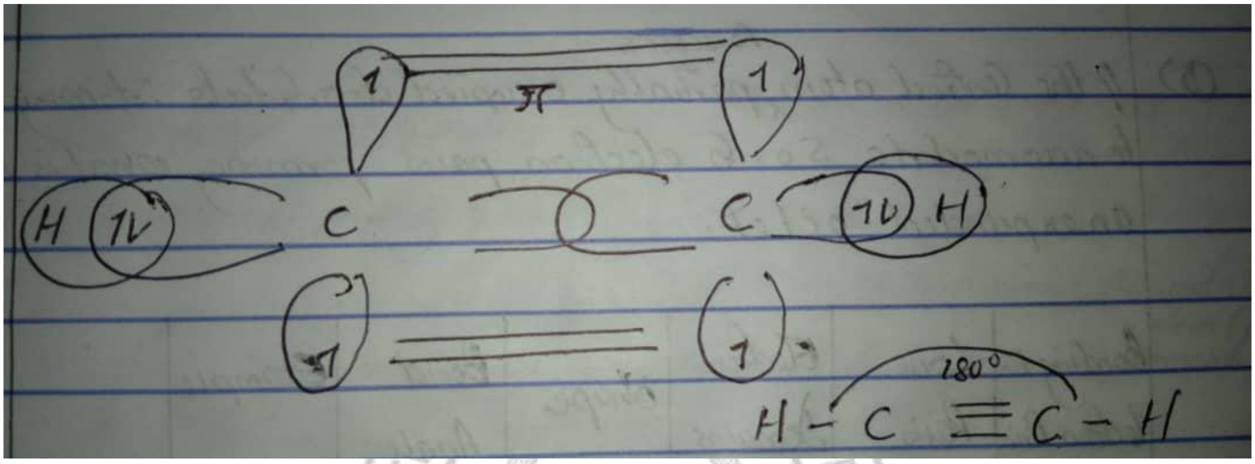


S.P. Hybridization (Diagonal Hybridization)

In Sp hybridization, two 2P electrons are unused, 2S electron is mixed with only one 2P electron.

This Sp hybridization gives rise to a linear shape with a bond angle of 180°.

The two unused 2P electrons form two lateral overlaps between two adjacent carbon atoms, one above and the other beneath the plane of the other and those lateral overlaps are Pi bonds.




VSEPR model is used to determine the shapes of molecules based on electron pair repulsion.

**VSEPR (Valence Shell Electron Pair Repulsion)**

It focuses on the bonding and non-bonding electrons present in the valence shell of an atom that connect with two or more other atoms.

H refers to the following:

(a) Negative Electron Charge: Negative electron charges will repel each other causing them as well as the chemical bonds they form to stay as far apart as possible.

Example: The two electrons of the fraction $AX_2$ will extend out in the opposite direction. It therefore will extend regularly at 180°c, so that the bonds will be as far apart as possible in opposite directions to give a linear shape.

(b) If the central atom A contains one or two pairs of non-bonding electrons. The orbitals containing the bonding and non-bonding pairs in the valence shell will extend out from the central atom in the directions that minimizes their mutual repulsion.

(c) If the central atom partially occupied d-orbitals, it may be able to accommodate 5 or 6 electron pairs, forming what is known as an expanded octet.

| Bonding E Pairs | Lone Pairs | Electron Domains | Shape | Bond Angle | Example |
|---|---|---|---|---|---|
| 2 | 0 | 2 | Linear | 180° | $CO_2$ |
| 3 | 0 | 3 | Trigonal Planar | 120° | $BF_3$/$PF_3$ |
| 2 | 1 | 3 | Bent | 119° | $SO_2$ |
| 4 | 0 | 4 | Tetrahedral | 109.5° | $CH_4$ |
| 3 | 1 | 4 | Trigonal Pyramidal | 107.5° | $NH_3$ |
| 2 | 2 | 4 | Angular | 104.5° | $H_2O$ |
| 5 | 0 | 5 | Trigonal Bipyramidal | 90° | $PCl_5$ |
| 4 | 1 | 5 | Seesaw | 90° | $SF_4$ |
| 3 | 2 | 5 | T-shaped | 101.6° | $ClF_3$ |

## Atoms, molecules, elements, and compounds and chemical reactions

Atoms are the smallest unit of matter that retain the chemical property of an element.

(Diagram of an atom with labels "Electrons" and "Nucleus")

**Structure of an atom:**

- Electron moves constantly around the nucleus through a channel called the orbital. It is negatively charged.
- Neutron is the atomic character found inside the nucleus with a neutral charge.
- Proton is found in the nucleus and is positively charged.

**Examples:**

- Hydrogen is the first element with 1 electron and 1 proton.
- Carbon has 6 electrons and protons, having a tetrahedral structure that forms complex compounds leading to organic chemistry.

**MODERN ATOMIC THEORY**

**Dalton:** Dalton talked about the indivisibility of and identity of an atom (an atom represents an element).

**Rutherford:** Rutherford shed more light on the presence of a nucleus inside an atom. Highlighting its dense and positively charged properties as a result of protons +ve charge.

**Bohr's Model:** Bohr's model gave the idea of electron orbit round the nucleus at a fixed energy level which are not the same in all orbitals.

**Molecules**

Molecules are a group of two or more different atoms that are chemically bonded together.

**Difference Between Molecule & Compound**

- Molecules are a combination of two or more atoms.
- Compounds are a combination of two or more elements.
- The properties of a compound are different from the properties of the element that make up the compound.

**Chemical Reactions**

- The process that converts reactants to the products could be more than one. $R_1 + R_2 \rightarrow \text{Product}$

**Features of Chemical Reactions**

1. Enables breaking of bonds in the reactants and forming new bonds in the products. $\mathrm{NaOH + HCl \rightarrow NaCl + H_2O}$
2. They obey the law of combination according to mass.

**Examples of Chemical Reactions**

1. Combination Reaction $2H_2 + O_2 \rightarrow 2H_2O$
2. Decomposition Reaction: It generally means breakdown of reactants involved $2H_2O_2 \rightarrow 2H_2O + O_2$
3. Displacement reaction $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
4. Redox reaction $2H_2O_2 + 2KI + H_2SO_4 \rightarrow I_2 + 2H_2O + K_2SO_4$

# KINETIC STUDY OF MATTER

**Matter is anything that has mass and occupies space, divided into three states:**

**- Solid**

**- Gas**

**- Liquid**

**The Kinetic Theory of Matter explains the state of matter in solid, liquid, and gaseous states based on the following assumption:**

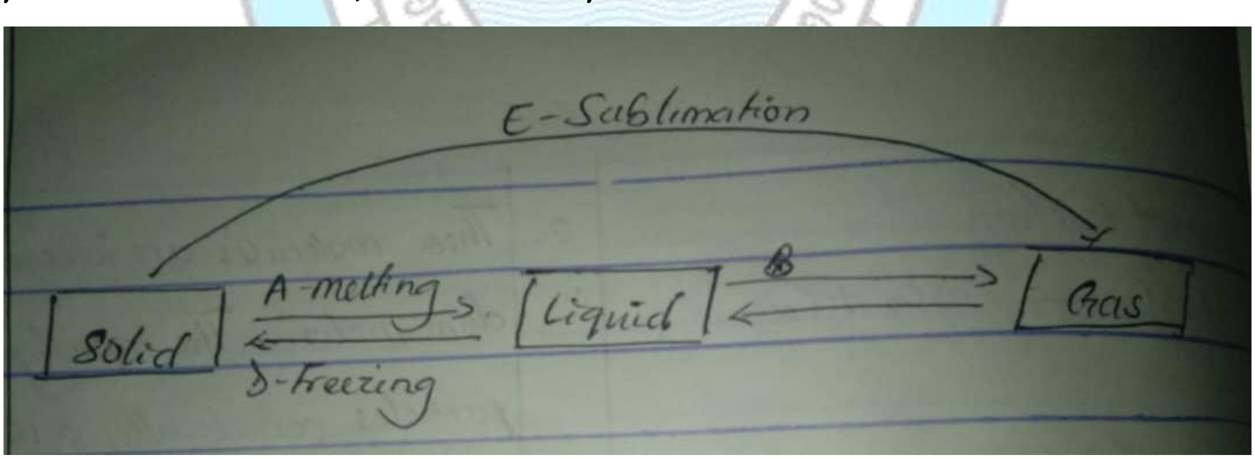
**1. Matter consists of very small particles, each of which has a mass.**

**2. These molecules are in constant random motion. The rapidly moving particles constantly collide with each other and with the walls of the container.**

**3. There are forces of attraction between particles of matter, these attraction forces will increase as the distance between the particles become closer.**

**4. The average kinetic energy of the gas particles depends only on the temperature of the system, that is, the higher the temperature, the higher the average kinetic energy; the lower the temperature, the lower the average kinetic energy.**

***When a substance is heated, temperature increases and the average kinetic energy also increases. If a substance is cooled, the kinetic energy of the particles will decrease, and room temperature will decrease. ***




***Solids that sublimes: Iodine, Camphor***

**The difference between the three different states of matter is the degree of movement of the particles**

**<u>Solids:</u> The molecules, atoms and ions in solids are very closely packed and held firmly together by forces of attraction.**

**Characteristics of solid state:**

(i) Solids have fixed mass
(ii) They have definite shape
(iii) They have definite volume
(iv) Solids have high density
(v) Get deformed when compressed
(vi) Solids can be crystalline or amorphous
(vii) Solids can only vibrate and rotate about their fixed positions

**Liquid state**

Liquids have more kinetic energy and are no longer held in fixed positions, They are inelastic and brittle

**Characteristics of Liquid State:**

(i) Liquid has fixed mass
(ii) Liquid has no definite shape, but takes the shape of the container
(iii) Liquid has definite volume
(iv) Liquids are less dense than solids
(v) Liquid cannot be compressed
(vi) Liquid can be volatile or non-volatile

**Gaseous state:**

The particles of gases have much more kinetic energy than liquids. The cohesive force in a gas is negligible and the particles of gas move about in all directions at great speeds.

**Characteristics of Gases:**

(i) Gases have fixed mass
(ii) No definite shape
(iii) No definite volume
(iv) Gases are less dense than water
(v) Gases are compressible
(vi) Gases mix with other gases in all proportions

# THERMOCHEMISTRY

Thermo means heat.

It is the study of the chemistry of heat change that accompanies a chemical reaction.

It is the study of the energy or enthalpy changes associated with chemical reactions as various bonds are being broken or formed.

Every chemical substance has a unique enthalpy (H) which is also known as its intrinsic energy or heat content. The value of the enthalpy cannot be measured, experimentally but its change during a reaction can be measured. (ΔH)

Every element in its free state or uncombined state has an enthalpy value of zero but different compounds have different enthalpies at the same temperature and pressure. The change in enthalpy of a reaction is the difference between the total enthalpy of the product and the total enthalpy of the reactant.

$$\Delta H = \Sigma H_p - \Sigma H_r$$

If the molar enthalpy of A, B, & C are -20, 72, 19 kJ/mol respectively, then the enthalpy changes of the reaction A+2B→2Cis given by:

$$\Delta H = \sum H_p - \sum H_r$$

$$\Delta H = [2(19) - (1(-20) + 2(72))]$$

$$= [38 - (-20 + 144)]$$

$$= 38 - 124$$

$$\Delta H = -116\text{kJ}$$

A thermochemical equation always indicates the nature and magnitude of the enthalpy change written after the equation for a reaction:

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)\ \Delta H = -286\text{kJ}$$

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)\ \Delta H = -572\text{kJ}$$

**The magnitude of ΔH\Delta H varies with the number of moles of substance in the reaction as shown in the reaction process. It is only true for the physical state of the substance in the equation.**

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)\ \Delta H = -244\text{kJ}$$

**Exothermic and Endothermic Reactions**

- **Exothermic Reaction: One in which the evolved loses heat to the surroundings in order to form the product. ΔH is negative for exothermic reaction because ΔHp < ΔHr**

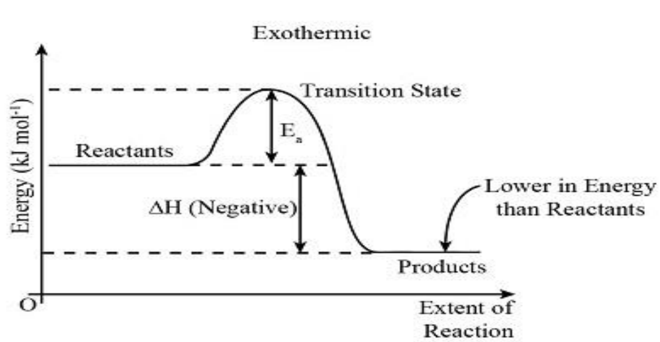
**A hypothetical exothermic reaction given by the thermodynamic equation:**

$$2W + X \rightarrow Q$$

$$\Delta H = -98\text{ kJ}$$

$$2W + X \rightarrow Q + 98\text{ kJ}$$

**An exothermic reaction may be represented on an energy level diagram to show the relative enthalpy, activation energy, and the enthalpy change of the reaction.**



Examples of Exothermic Reactions:

- Dissolution of NaOH (Sodium Hydroxide) in water
- Combustion reactions

**ENDOTHERMIC REACTION**

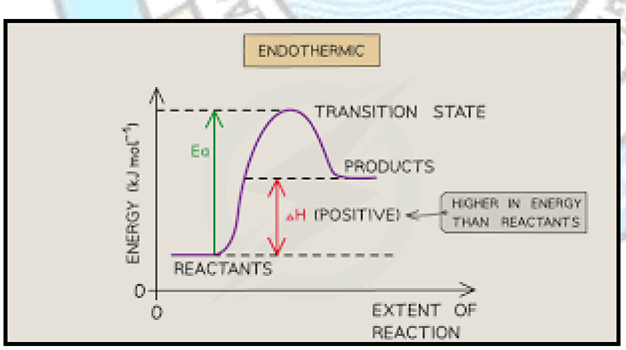
An endothermic reaction is one in which the reactant absorbs heat from the surroundings in order to form the product. ΔH is positive for an endothermic reaction because the sum of the enthalpy of the products is greater than the sum of the enthalpy of the reactants ($\Sigma H_p > \Sigma H_r$). An endothermic reaction is given by the reaction:

$$A + B \rightarrow C \quad \Delta H = \ + 150\text{kJ}$$

$$A + B + 150\text{kJ} \rightarrow C$$

$$A + B \rightarrow C - 150\text{kJ}$$

An endothermic reaction can be represented on an energy level diagram:




**Examples of endothermic reactions:**

- The thermal decomposition of calcium bicarbonate.
- Dissolution of ammonium chloride in water.

**Important Enthalpy Change:**

- Standard enthalpy changes for a reaction: $\Delta H^0$

- **Standard enthalpy of formation $\Delta H°f$**
- **Standard enthalpy of combustion $\Delta H°c$**
- **Standard enthalpy of neutralization $\Delta H°n$**
- **Standard enthalpy of hydrogenation $\Delta H°hy$**

## Wave Particle Duality

- This phenomenon makes us know that wave exhibit specific properties. As waves house has a particles Electron movement around the nucleus is not in a straight line, and it moves as a wave and appears as a particle.
- Heisenberg theory of uncertainty It is impossible to determine the position and momentum of an electron simultaneously and you cannot pinpoint the exact location of the electron as it moves around fast and

Types of quantum numbers

Azimuthal

Angular Momentum Quantum Number: The shape of the orbital where the electrons are to be found.

- S, P, d, f.

Magnetic Quantum Number: Orientation of the orbital in space.

Spin Quantum Number

Bohr's Contribution: Bohr's atomic theory does not do justice to the basic atomic structure.

Limitations of Bohr's Theory:

1. Bohr's theory postulated that an atom has only one electron but it has many more based on the element in question (e.g., Helium has 2 electrons).
2. The theory did not indicate the wave particle duality nature of an atom.
3. It did not explain the fine structure of spectral lines.
4. It did explain the splitting of spectral lines in a magnetic field (Zeeman's effect).
5. Hydrogen atoms are also found in spectral lines.

Spatial Science Subgroup

Lyman’s Series → $n_1 = 1$ (UV)

Balmer’s Series → $n_2 = 2$ ( )

Paschen’s Series → $n_3 = 3$ (IR)

**S orbital** has only one energy level, with a minimum of 2 electrons. It is the most common orbital

**P orbital** has a maximum of 6 electrons (dumbbell-shaped).

**d** (maximum of 8 electrons; Soluble and shows complex shape).

**f** (maximum of 14 electrons).

The movement of electrons are

Absorption Spectrum Electron gains energy as it moves from a lower energy level to a higher energy level.

Emission Spectrum Electron loses energy as it moves from a higher energy level to a lower energy level.

# Chemical Equilibrium

Chemical equilibrium reactions are reversible reactions carried out in closed containers, which cause the product formed recombines to form the reactants at the same rate at which reactants form the products.

Reactant and products in a chemical equilibrium are separated by the sign (⇌) called the reversible arrow.

**Examples of reversible reactions:**

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

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

$$3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$$

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

**Irreversible and Reversible Reactions**

Chemical reactions which only go in one direction are known as irreversible reactions, because the product formed cannot be reconverted to give the reactants under any condition.

Examples of Irreversible reactions

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

$$MgO + 2HCl \rightarrow MgCl_2(aq) + H_2$$

A reversible reaction is one which can go either forward or backward depending on the condition of the reaction. The sign (⇌) is used and the equation is reversible.

**Homogeneous and Heterogeneous Equilibrium**

A homogeneous equilibrium reaction is one in which the reactants and products are in the same physical state of matter (liquid, gas).

A heterogeneous equilibrium reaction is one in which the reactants and products are not in the same physical state.

Example: $H_2O(l) \rightleftharpoons H_2O(g)$

## Characteristics of a System in Chemical Equilibrium

A chemical equilibrium reaction is a special reversible reaction with the following characteristics:

- The rate of the forward reaction is exactly equal to that of the backward reaction.
- The free energy change of an equilibrium is zero $\Delta G=0$.
- It has a dynamic nature, because the forward and backward reactions never cease.
- The concentration of the reactants and products at equilibrium are constant at constant temperature regardless of whether the equilibrium point is approached from the direction of pure reactants or pure products or a mixture of reactant and product.
- Every chemical equilibrium system has a unique equilibrium constant (Keq) at constant temperature.

## Principles That Guide Chemical Equilibrium: Le Chatelier's Principle

It states that "If a system in chemical equilibrium is disturbed by changing one of the factors maintaining the equilibrium, the system will shift its equilibrium point in a direction that opposes or minimizes the imposed disturbance."

The effect of changing conditions on any equilibrium can be predicted by Le Chatelier's principle. In all cases, a disturbed equilibrium system always reduces the effect of an imposed disturbance by using one of its two reversible reactions to develop the exact opposite of the effect created by the imposed disturbance.

The position of equilibrium is an indication of the proportion of products to the reactants in an equilibrium mixture.

**Note:** If a large amount of the reactant is changed to products, the position of equilibrium will favor the products or lie to the right.

$$N2+3H2 \rightleftharpoons 2NH3(g)$$

**Note:** If the conversion of reactants to products is small, the position of equilibrium favors the reactants or lies to the left.

**Factors that affect a system in Chemical Equilibrium:**

1. **Concentration**
    - Increase in temperature of the reactants will yield more product
    - Increase in concentration of product forms more reactant.
    - Decrease in concentration of reactant will yield more product.
    - Decrease in concentration of product forms more product.
    - Addition of external reactant will yield more products.
    - Addition of external product will yield more products
2. **Pressure:**

**Gay-Lussac's Law of Volume:**

- When pressure is added, volume is reduced.
- Decrease in pressure increases volume.
- Example Reaction: $N_2(g)+3H_2(g) \leftrightarrow 2NH_3(g)$

3. **Temperature:**

- **Endothermic Reactions:**

Increase in temperature reduces volume and favors the forward reaction.

Decrease in temperature reduces volume and favors the backward reaction.

- **Exothermic Reactions:**

Increase in temperature favors the backward reaction.

Decrease in temperature favors the forward reaction.

**Equilibrium Constant** (Keq):

**Equilibrium constant is a quantitative measure of how far to the right the equilibrium reaction goes. If the value of equilibrium constant is very large, a large proportion of the reactant is converted to product. But if equilibrium constant is low, no significant reaction takes place.**

**Equilibrium constant for concentration (Kc)**

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

**Equilibrium constant for partial pressure (Kp)**

$$K_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

***Assuming both reactants and products are gases***

**Equilibrium constant in terms of concentration expressions (Kc) may be written from the balanced equation of the chemical equilibrium reactions shown above. If the equation of a chemical equilibrium is reversed, the new equilibrium constant expression is the reciprocal of the power.**

**If a heterogeneous equilibrium reaction involves pure solids or pure liquids, the concentration of such pure solids or pure liquids are constant and so, do not appear in equilibrium constant expressions.**

***[Concentration of pure solids or liquids are constant]***

$$Fe_{(s)} + 4H_2O_{(g)} \rightleftharpoons Fe_3O_{4(s)} + 4H_{2(g)}$$

$$k_c = \frac{[H_2]^4}{[H_2O]^4} \qquad k_p = \frac{P^4H_2}{P^4H_2O}$$

$$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$$

$$k_c = [CO_2] \qquad k_p = PCO_2.$$

**Note:** **Pressure and concentration cannot alter the value of Kc and Kp only temperature can.**

**OXIDATION-REDUCTION REACTION**

Every oxidation-reduction reaction always involves the interaction between a reducing agent and an oxidizing agent which brings about two opposite but complementary reactions or processes known as Reduction and Oxidation.

Oxidation-Reduction can be defined in terms of:

1. Hydrogen and Oxygen
2. Electron Transfer
3. Oxidation Number

**Hydrogen and Oxygen**

**Oxidation** can be defined as the addition of oxygen to a substance or the removal of hydrogen from a substance. While an oxidizing agent is one which brings oxygen to another substance or removes hydrogen from that substance.

$$Cu_2O\,(s) + H_2 \rightarrow Cu + H_2O$$

From the reaction above, the process of removing hydrogen to water is oxidation. Thus copper (II) oxide which releases the oxygen to hydrogen is the oxidizing agent.

Another example is the oxidation of ethanol to ethanol.

$CH_3CH_2OH \rightarrow K_2Cr_2O_7CH_3CHO$

$K_2Cr_2O_7$: which removes hydrogen atoms from ethanol is the oxidizing agent.

**Reduction** is the addition of hydrogen to a substance or the removal of oxygen from a substance. A reducing agent is a substance which brings hydrogen to another substance or removes oxygen from that substance.

Example:

$$3Fe + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$$

Conversion of steam to hydrogen is reduction because it involves the removal of oxygen from steam. In this reaction, iron is the reducing agent.

**Electron Transfer** Oxidation simply means loss of electrons while reduction means gain of electrons.

Oxidation=loss of $e^-$

Reduction=gain of $e^-$

An oxidizing agent is an electron acceptor, while a reducing agent is an electron donor.

Example:

$$Na(s)+\frac{1}{2}Cl2(g)\rightarrow Na^{+}cl^{-}(s)$$

Another good example is the electrovalent formation of NaCl from sodium and chlorine atoms. In this reaction, sodium, like all metals, is the reducing agent because it donates its outermost electron to the outermost shell of the chlorine atom which accept the electron is the oxidizing agent (electron acceptor). In all redox processes, the total number of electrons donated is equal to the total number of electrons accepted.

**Oxidation Number** Oxidation is increase in oxidation number Reduction is decrease in oxidation number Oxidizing agent is one whose oxidation number decreases Reducing agent is one whose oxidation number increases

$$Na(s)+\frac{1}{2}Cl2(g)\rightarrow Na^{+}cl^{-}(s)$$

Na(S) (Sodium) is the reducing agent because its oxidation number increases from 0 to +1.

Cl2(g) (Chlorine) is the oxidizing agent because its oxidation number decreases f rom 0 to -1.

The change from sodium to Sodium ion (Na -> $Na^{+}$) is an oxidation process because oxidation number increases and Chlorine's change (Cl2 -> $2Cl^{-}$) is a reduction process because oxidation number decreases.

ASSIGNING RULES FOR OXIDATION NUMBER

1. The oxidation number of an element in its free or uncombined state is equal to zero.
2. The oxidation number of hydrogen in its compounds is always +1, except in metallic hydrides where it is -1.
3. The oxidation number of oxygen in its compounds is -2 except in peroxides (-1) and superoxides (-1/2).
4. The oxidation number of any group 2 element in its compound is -2.
5. The oxidation number of an element which exists as a simple ion is equal to the charge of the ion.
6. The oxidation number of fluorine in all its compounds is -1
7. The sum of the oxidation number of all the atoms in a neutral compound is equal to zero.
8. The sum of the oxidation numbers of all the atoms in a charged number or radical is equal to the overall charge on the charged compound.

Examples:

1. Calculate the oxidation number of chromium in $Cr_2O_7^{2-}$

$2Cr + 7(-2) = -2$

$2Cr + (-14) = -2$

$2Cr = 12$

$Cr = +6$

2. Calculate the oxidation of Manganese in $KMnO_4$

$1 + Mn + 4(-2) = 0$

$Mn - 8 = 0$

$Mn = +7$

**Examples of Oxidizing Agent or common Oxidizing agents**

1. Oxygen Molecules
2. Potassium tetraoxomanganate (V) $KMnO_4$
3. Potassium heptachlorodimanganate (V)
4. Concentrated Tetraoxosulphate (VI) acid
5. Chlorine Gas
6. Bromine Gas
7. Aqueous Iodine
8. Hydrogen Peroxide
9. Manganese (IV) oxide
10. Iron (III) Salts (Iron (III) chloride)

**Common Reducing Agents**

1. Metals
2. Carbon
3. Hydrogen Sulphide
4. Sulphur (IV) oxide
5. Tin (II) Chloride ($SnCl_2$)
6. Sodium thiosulphate ($Na_2S_2O_3$) (Sodium hexadiosulphate 2)
7. Ethanoic acid or acetic acid
8. Carbon (IV) oxide
9. Acidified Potassium iodide solution (KI)
10. ron (III) salt ($FeCl_3$)

**How to Test For Oxidizing Agent**

1. Oxidizing agent will react with hydrochloric acid solution to liberate chlorine gas.
2. Oxidizing agent with hydrogen sulphide to form a precipitate of sulphur.
3. Oxidizing agent react with Potassium iodine solution to liberate iodine.

### How to test for Reducing Agents

1. Reducing agent will react with acidified potassium hexa solution and change its color from orange to green.
2. Reducing agent will react with acidified potassium and remove purple color up to colorless.

### VALENCE FORCES

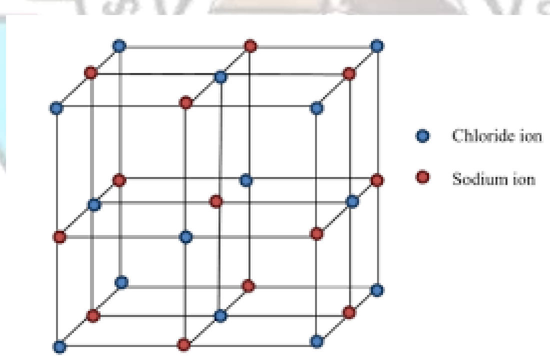
Valence forces are forces that hold atoms in a molecule to form chemical bonds. They are caused by the attraction of valence electrons.

Valence forces determine how atoms bond together to form solid solids, which can be crystalline or amorphous.

### CRYSTALLINE SOLIDS

In Crystalline Solids, the atoms, ions, or molecules are arranged in a definite repeating pattern as shown in the diagram of NaCl.

The strong force of attraction holding the atoms together causes the atoms in the solid to be ordered. Examples of crystalline solids are Metals, Sodium chloride, etc.




### AMORPHOUS SOLIDS

The particles of such solids lack an ordered internal structure and are randomly arranged as glass. Examples of amorphous solids include candle wax, glass, and most organic compounds.

*Classification of Crystalline Solids*

Crystalline solids are generally classified according to the nature of the forces that hold the particles together.

This includes:

1. Metallic Solids
2. Covalent Network Solids
3. Ionic Solids
4. Molecular Solids

**Metallic Solids** These are crystals formed by metals, e.g. copper, aluminum, iron, etc. Crystals of metals are described as uniform distribution of atomic nuclei within a sea of delocalized electrons. The atoms are held together by metallic bonds.

Features

They exhibit properties such as:

a. Electrical conductivity

b. Metallic lustre

c. Malleability

d. Hardness

e. Ductility

f. High melting points

**Covalent Network Solids** Atoms in these solids are held together by a network of covalent bonds. Examples include:

- Diamond
- Silicon (e.g., $SiO_2$)
- Silicon carbide (SiC)

PROPERTIES

1. Hardness: These solids are extremely hard, e.g. Diamond
2. High melting points

**Ionic Solids**

They are composed of positive and negative ions that are held together by electrostatic attractions, e.g., Sodium Chloride, Magnesium Oxide, Calcium Sulphate, etc.

**Properties:**

1. High melting points

2. They are hard but brittle
3. They are non-conductors
4. They are non-ductile
5. They possess ions and they exhibit electrical conductivity when in molten state
6. They form ions

**Molecular Solids**

They exist from weak structure forces (Van der Waals) between molecules in solid. Examples are:

- Ice, Iodine, Sucrose (Cane sugar) and other examples are hydrogen, nitrogen, oxygen in form molecules called strong low melting point (less than zero)

  **Properties:**

- Low melting points
- Non-Solid

**CHEMICAL BONDS**

Chemical bond is the attractive force which holds the constituent particles (atoms/ions or molecules) together in a compound.

**Types of Chemical Bonding**

1. Ionic/Electrovalent bonding
2. Covalent bonding
3. Coordinate bonding

**Ionic/Electrovalent bonding** It occurs when there is a transfer of electrons from one atom to another. The atom that donates electrons possesses a positive charge while the atom that accepts electrons acquires a negative charge. A positively charged atom is called a cation while a negatively charged atom is called an anion. Ionic bonding is usually between electro-positive and electro-negative atoms i.e. atoms with wide differences in electronegativity.

Ionic Bond = Electro-positive atom + Electro-negative atom

Ionic bonding is common between elements of groups 1, 2, and 6, 7. The atoms undergo bonding to attain a stable electronic configuration making them resemble inert gas.

**Properties of Ionic compounds:**

1. High melting point & boiling points
2. They conduct electricity in solution or molten form because of the presence of ions.
3. Most of them are crystalline.
4. They are generally soluble in water (polar solvent).

**Covalent Bonding** Covalent bonding results when two atoms share two electrons thereby making the participating atoms.

**Properties of Covalent Compounds**

1. Low melting point.
2. They are insulators (No electrical conductivity both in aqueous and solid state).
3. Their solids are generally amorphous.
4. They are generally soluble in non-polar solvent (some are also soluble in polar solvent).

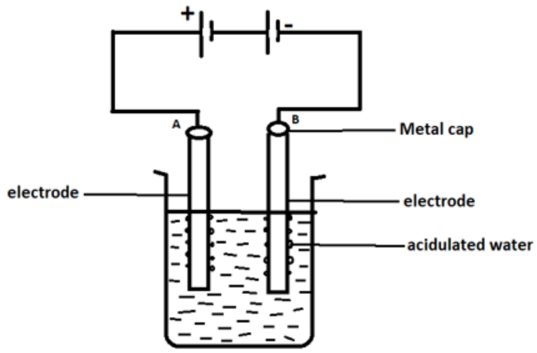
**Coordinate Covalent Bonding** This type of bonding involves the sharing of two electrons donated by only one of the participating atoms.

**Examples of compounds where covalent bonding occurs:**

1. Ammonium chloride
2. Complexes of transition metals [Cr(NH3)6]3+

**Electrolysis**

Electrolysis is the process of using a direct current to cause the decomposition of an electrolyte by means of redox reactions which occur at the anode and cathode of an electrolytic process.



Electrolysis of water

**Terms Used in Electrolysis**

**Conductors** are solid substances that allow the flow of electric current. Metals and graphite are examples.

**Insulators** (non-conductors) are substances that do not allow the flow of electric current. For example, protein (macromolecules).

**SEMICONDUCTORS** are substances that conduct electric current partially and are mainly used as transistors, e.g., Germanium, Silicon.

**ELECTROLYTES** are substances that conduct electricity in molten state or when dissolved in water. Examples include salts, bases, and acids.

**STRONG ELECTROLYTES** are substances that ionize almost completely in water. Examples include H2SO4

**WEAK ELECTROLYTES** are substances that ionize partially in water. Example: CH3COOH→CH3COO−+H+

**NON-ELECTROLYTES** are substances that do not ionize in water. Examples include macromolecules and covalent compounds.

**ELECTRODES** are metals or poles of carbon (graphite) at which electric current enters or leaves the electrolyte.

**ANODE** is the positively charged electrode in an electrolytic cell through which conventional current enters the electrolyte, or the point where electrons leave the electrolyte.

**CATHODE** is the negatively charged electrode in an electrolytic cell through which conventional current leaves the electrolyte, or the point where electrons enter the electrolyte.

**ACTIVE ELECTRODES** is an electrode that undergoes a chemical change during electrolysis.

**INERT ELECTRODES** is an electrode which do not undergo any chemical change during electrolysis.

**IONIC THEORY.**

**The Ionic Theory explains the behavior of ions in solution and their crucial role in conducting electricity. When substances like salts, acids, and bases dissolve in water, they dissociate into positive and negative ions. These ions move freely, allowing the solution to conduct electric current. The ease with which ions are discharged during electrolysis depends on their position in the electrochemical series, concentration, and the nature of the electrode. Cations are positively charged ions, while anions are negatively charged ions. Factors such as the ion's position in the series and the nature of the electrode influence the discharge of ions during electrolysis, which is a key concept in the Ionic Theory.**

**FACTORS THAT AFFECT THE DISCHARGE OF ION DURING ELECTROLYSIS**

1. Position of the ion in the Electrochemical series
2. Concentration of the ion in the electrolyte
3. Nature of the electrode

**Position of the ion in the Electrochemical series**

**Cations:** $K^+$, $Na^+$, $Ca^{2+}$, $Mg^{2+}$, $Al^{3+}$, $Zn^{2+}$, $Fe^{2+}$, $Sn^{2+}$, $Pb^{2+}$, $H^+$, $Cu^{2+}$, $Hg^{2+}$, $Ag^+$, $Au^{3+}$

**Anions:** $F^-$, $SO_4^{2-}$, $NO_3^-$, $Cl^-$, $Br^-$, $I^-$, $OH^-$

**Anode**

$$\mathrm{NaCl} \rightarrow \mathrm{Cl^-(aq)} + \mathrm{Na^+(aq)}$$

$$\mathrm{H_2O} \rightarrow \mathrm{OH^-(aq)} + \mathrm{H^+(aq)}$$

At the anode, oxidation goes on.

The selectively discharged ion is the one that is lower in the electrochemical series.

$$Cl^- \rightarrow Cl_2 + 2e^-$$

$$OH^- - 2e^- \rightarrow H_2O + O_2$$

$$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$$

**At the anode, the selectively discharged ion is Cl–**

**Cathode**

**At the cathode, hydrogen will be selectively discharged.**

$$2H^+ + 2e^- \rightarrow 2H$$

$$4H^+ + 4e^- \rightarrow 2H_2$$

$$OH^- + 2H^+ + 2e^- \rightarrow 2H_2O + O_2 + 4e^- + 2H_2$$

$$4OH^- + 4H^+ \rightarrow 2H_2O + O_2 + 2H_2$$

$$4H_2O \rightarrow 2H_2$$

***BRINE (concentrated NaCl)***

- **$Cl^-$ will be more selectively discharged at the anode in preference to $OH^-$ in water.**
- **2Cl–→Cl2+2e–$2Cl^-$**

**Anode**

- **$H^+$ will be selectively discharged.**
- **2H++2e–→H22$H^+ + 2e^-$**

**Cathode**

- **Overall reaction**
- **2Cl–+2H++2e–→2Cl2+H22$Cl^- + 2H^+ + 2e^-$**
- **2Cl–+2H+→Cl2+H22$Cl^- + 2H^+$**

**Nature of the Electrode**

*Electrolysis of Copper (II) tetraoxosulphate (VI) using copper electrodes.*

This electrolyte process is used to purify copper.

- $CuSO_4$
- $H_2O$
- **Cathode**
    - $Cu^{2+}$ $H^+$ $OH^-$
    - $Cu^{2+}$ will be selectively discharged.
    - Cathode's reaction:
    - Cu2++2e−→Cu$Cu^{2+} + 2e^-$
- **Anode**
    - The potential of hydroxyl ion is high to copper ion, so copper ion will be selectively discharged.
    - Cu will be selectively discharged.

**Anode reaction:**

- $Cu \rightarrow Cu^{2+} + 2e^-$

INTERMOLECULAR FORCES

Intermolecular forces are forces that hold molecules together. They are electrostatic in nature and arise from the interaction between positively charged and negatively charged species. They are most important for solids and liquids because the molecules are close together. However, in gases, the forces are significant at very high pressures (compression) resulting in liquefaction and hence, deviation from the ideal gas law.

Intermolecular forces determine the following properties:

1. Melting point of solids and boiling point of liquids (Higher → Higher)
2. Mechanical properties of solids (Brittle, Malleability, Ductility)
3. Vaporization of liquids (Vapor Pressure of liquids) (Evaporation)

Types of Intermolecular Force or Interaction

1. Dipole-dipole Attraction (Interaction)

This occurs when two bonded atoms in a covalent bond generate a dipole (charges that are equal but opposite).

δ**+** δ−

These dipoles can be attractive (repulsive) depending on the degree of alignment. When the positive end of one dipole (δ**+**) is near the negative end of another (δ−) it produces attraction; when equally charged ends are near, it produces repulsion.

Examples of molecules where dipole dipole attraction occurs are

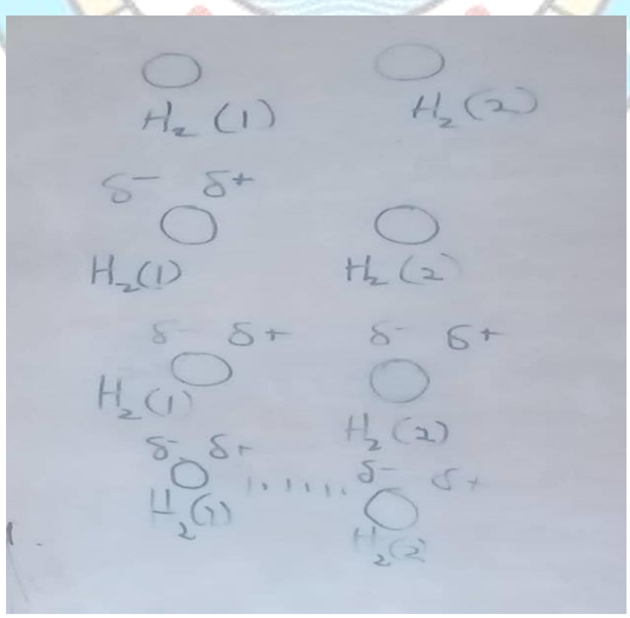
- Hcl
- Hf
- $H_2O$

**London Dispersion Forces (Van der Waals)**

Both atoms have similar polarity (electronegativity) and so

C - C

There are instantaneous dipole-induced dipole interactions between non-polar molecules. Fluctuations in the electron distributions called POLARISABILITY within and non-polar molecules could result to the formation of instantaneous dipole moments which produce attractive forces called "LONDON DISPERSION FORCES" between non-polar substances.



London Dispersion Forces are responsible for the general trend towards higher boiling points with increases in molecular mass. Eg.: Alkane homologous series

In summary, intermolecular forces increase with molecular mass."

**Hydrogen Bonds**

- It occurs when a hydrogen atom forms a bridge between two highly electronegative atoms e.g. O, N, F. It is a special type of dipole-dipole interaction and it is especially strong because of large attraction occurring between the molecules.

**Note:**

- The higher the electronegativity difference, the stronger the dipole moment.
- Hydrogen bonding occurs in substances such as water, alcohol, phenol, and ammonia.

**Properties of Hydrogen Bonded Compounds**

- High melting and boiling point

- Toughness of solids

- Hydrogen bonds increase the toughness of solids. E.g. Polyesters, Nylon 6,6

- Increase in Viscosity: The higher the hydrogen bonds, the higher the viscosity of the liquid. (It turns from a liquid to a syrup e.g. glycerine)

- Vapour Pressure: The more the hydrogen bonds, the lower the vapour pressure of the liquid. (The liquid will not evaporate easily or at all)

| **Molecule (liq)** | **Boiling Point** |
|---|---|
| $C_2H_5OH$ (liq) | 78°C |

**Stoichiometry**

Stoichiometry is the study of the quantitative relationship or ratio between reactants and products in a chemical reaction.

**Example reaction:** $A + B \rightarrow C + D$ (Reactants Products)

**Key Aspects of Stoichiometry**

1. Balancing of reaction equations
2. Converting grams to moles
3. Calculating molar mass

4. Calculating mole ratio

**Illustration:**

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

$$Fe^{2+} : MnO_4^{-} = 5:1$$

**Mass-Mass Stoichiometry Problems:** To solve a mass-mass problem using stoichiometry, follow these steps:

1. Correctly identify the problem as a mass-mass problem:
   - For example, in a chemical equation like A+2B→C, you may be given the mass of A and asked to determine the mass of B that will completely react with A.
   - Determine how many grams of CC will be produced.
2. Balance the chemical equation:
   - Ensure that the same number of atoms of both reactants and products are the same.
   - Obey the law of conservation of mass.
3. Convert any mass value in the problem into moles:
   - Use the molar mass to do this.

$$\text{moles} = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{mass} = \text{moles} \times \text{molar mass}$$

4. Use molar proportion to determine unknown quantities of moles:

$$\frac{n_a}{a} = \frac{n_b}{b}$$

5. Convert the mole value (you just found) into mass using the molar mass of that substance:

$$m = n \times \text{molar mass}$$

**Example:**

1. **Calculate the amount of $CO_2$ that will be liberated when 5g of $Na_2CO_3$ reacts with excess dilute HCl. Also, determine the amount of $CO_2$ formed in moles.**

**Equation:**

$$2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$$

**Moles:**

**1 mole $Na_2CO_3$ = 106g 1 mole $CO_2$= 44g**

**Amount in moles liberated:**

$$\text{mole} = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{mole} = \frac{5g}{106g}$$

$$\text{mole} = 0.047\text{mol of } \mathrm{CO_2} \text{ (liberated)}$$

**Amount in grams liberated:**

$$106g\ Na_2CO_3 \rightarrow 44g\ CO_2$$

$$5g\ Na_2CO_3 \rightarrow x$$

$$x = \frac{44g \times 5g}{106g}$$

$$x = 2.08g\ CO_2 \text{ (liberated)}$$

2. **How many moles of oxygen $O_2$are required for the combustion of methane $CH_4$to produce 10g of $CO_2$?**

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

**Moles:**

$$x\,g\,O_2 \rightarrow 10g\,CO_2$$

$$x = \frac{64g \times 10g}{44g}$$

$$x = 14.55g\,O_2\ (\text{required for combustion})$$

**Moles of $O_2$:**

$$\text{moles} = \frac{14.55}{32}$$

$$= 0.45\text{moles of } O_2$$