P.V = n.R.T

CHEMISTRY REVISION
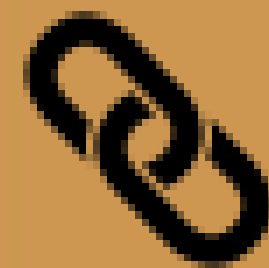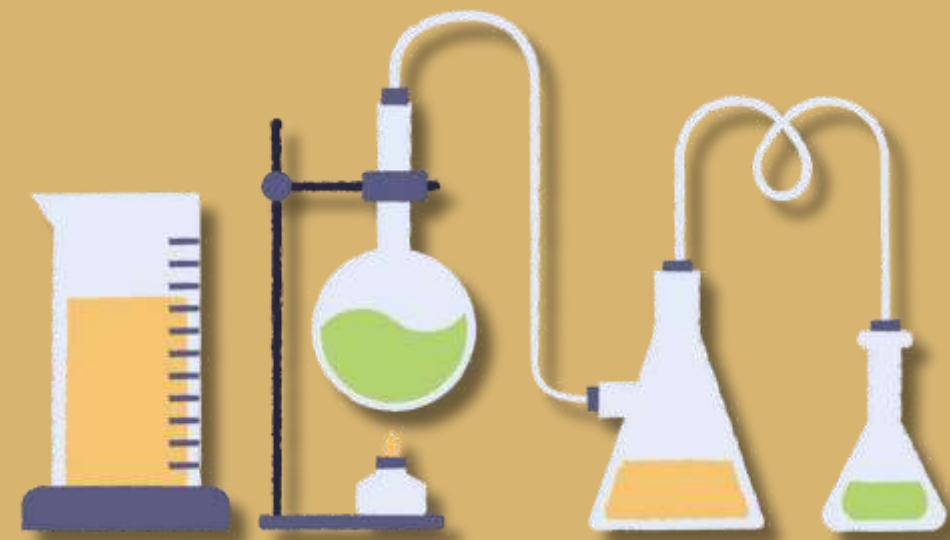
Classification of matter:

MATTER

PURE SUBSTANCES

MIXTURES

elemental substances

Gold (Au)

Helium (He)

Oxygen gas (O_2)

Ozone (O_3)

Diamond (carbon atoms arranged in a repeating three-dimensional structure)

compounds

Water (H_2O)

Carbon dioxide (CO_2)

Ammonia (NH_3)

Table Sugar ($C_{12}H_{22}O_{11}$)

Salt (NaCl)

heterogeneous mixtures

Dirt (mixture of minerals, organic matter, and air)

Mud (mixture of dirt and water)

Fog (water droplets suspended in air)

Granite (a kind of rock made of grains of various mineral compounds)

Lemonade with pulp

Fruit smoothie

homogeneous mixtures

Vinegar (water with acetic acid dissolved in it)

Carbonated water (carbon dioxide dissolved in water)

Metal alloys (like bronze, brass, and steel)

Air (nitrogen, oxygen, argon, carbon dioxide and other gases mixed evenly together, usually with water vapor as well)

Pink lemonade from powdered mix

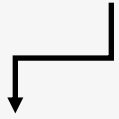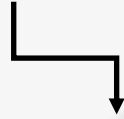## Pure substances:

- Substances that are made from one kind of particles and has a fixed or a constant composition.

Pure Substance is classified as:

Elements such as: Gold   Compounds such as: Water (H_2O)

How to separate pure substance:

- Pure substances cannot be separated physically it must do a chemical reaction to separate a compound into elements.

➢ Compounds : broke into elements by chemical processes
        such as : electrolysis of water

➢ Elements : chemical or physical processes cannot break
        the elements into simpler substances.

Mixtures

- **Some substances not have chemically combined but they are mixed with each other.**
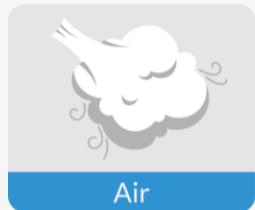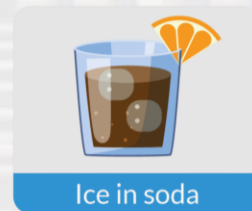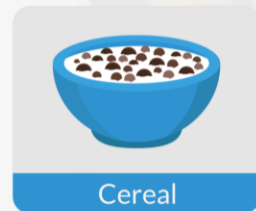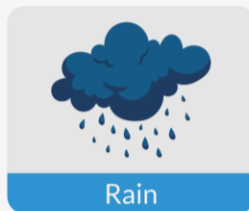
Mixtures are classified as:

Homogeneous mixture　　　　**Heterogeneous mixture**

Example of mixtures

Heterogeneous　　　　　Homogeneous

Steal　　　Rain　　　Cereal　　　Ice in soda

Air　　　Wine　　　Soil　　　Blood

Comparison between heterogeneous mixture and homogeneous mixture:

HOMOGENEOUS MIXTURES
VERSUS
HETEROGENEOUS MIXTURES

| Homogeneous mixtures	Heterogeneous mixtures
Homogenous mixtures have a uniform composition throughout the mixture	Heterogenous mixtures have a mixed composition which may vary from point to point
Components are not visible to the naked eye	Components can be seen easily
The whole mixture is in the same phase	Substances can be of two phases and layers may separate
Particle size is often at atomic or molecular level	Heterogenous mixtures have large particle sizes
Components cannot be separated easily	Components can be separated easily

Visit www.pediaa.com

Separation of mixtures:

❑ **Distillation: used to separate 2 liquids depends on volatility "how readily substances become gas" or boiling point.**
The most volatile component vaporizes at the lows temperature and the vapor passes through a cooled tube (a condenser) where it condenses back into liquid state.

❑ **Filtration: used to separate a mixture of insoluble solid and liquid using filter paper and filter funnel**

❑ **Evaporation: used to separate a mixture of soluble solid and liquid.**
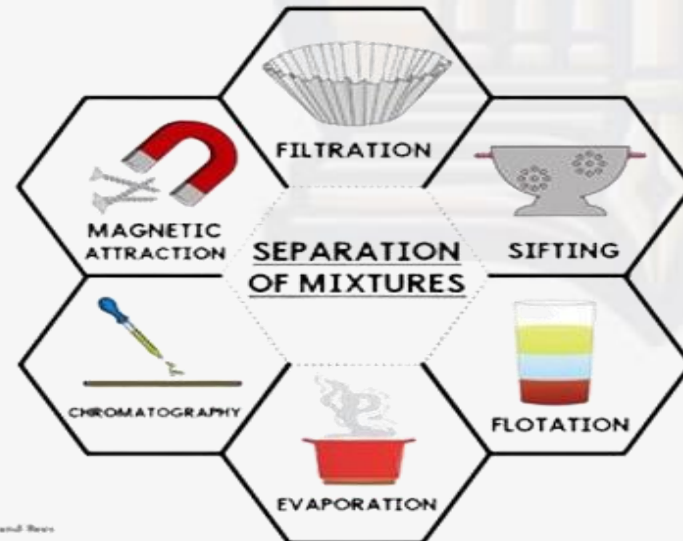
Solutions:

- **A homogenous mixture of two or more substances in relative amounts that can be varied continuously up to what is called the limit of solubility**

❑ **Solute: A substance dissolved in another substance, known as a solvent**

❑ **Solvent: A substance that dissolves a solute, resulting in a solution.**

❑ **Solubility of a solute : the amount of solute is dissolved in an amount of solvent.**

Types of solutions according to concentrations:

Supersaturated solution	Unsaturated solution	Saturated solution
Contains a large amount of solute at a temperature where it is going to be reduced, causing the extra solute to crystallize quickly.	A solution capable of taking up more of the solute at a given temperature	A solution where the solvent is incapable of dissolving any more solute at a given temperature.

Types of solutions according to the phase:

Types of binary solutions	Solute	Solvent	Examples
Solid solution	Solid	Solid	Copper dissolved in gold (alloys)
	Liquid	Solid	Mercury with sodium (amalgam)
Liquid solution	Solid	Liquid	Sodium chloride dissolved in water
	Liquid	Liquid	Ethyl alcohol dissolved in water
	Gas	Liquid	Carbon dioxide dissolved in water (soda water)
Gaseous solution	Liquid	Gas	Water vapour in air (cloud)
	Gas	Gas	Mixture of Helium-Oxygen gases

Colloids:

- Is a heterogeneous mixture of undissolved particles , is an intermediate state between a solution and suspension.

- Do not settle down, Intermediate particles(1-1000nm)

Examples: milk and dust in the air

Classifying Colloids

- A common method of classifying colloids is based on the phase of the dispersed substance and what phase it is dispersed in. The types of colloids includes sol, emulsion, foam, and aerosol.

- Sol is a colloidal suspension with solid particles in a liquid.

- Emulsion is between two liquids.

- Foam is formed when many gas particles are trapped in a liquid or solid.

- Aerosol contains small particles of liquid or solid dispersed in a gas.

Types of Colloids:

Dispersion medium	Dispersed phase	Name of the colloid	Examples
Gas	Liquid	Liquid Aerosol	Fog Aerosol spray
Gas	Solid	Solid Aerosol	Smoke, Air pollutants like fumes, dust.
Liquid	Gas	Foam	Whipped cream, Shaving cream, Soda water, Froth.
Liquid	Liquid	Emulsion	Milk, Cream, Mayonnaise
Liquid	Solid	Sol	Inks, Paints, colloidal gold.
Solid	Gas	Solid foam	Pumice stone, Foam rubber bread.
Solid	Liquid	Gel	Butter, cheese
Solid	Solid	Solid sol	Pearls, opals coloured glass alloys colloidal dispersed eutetics.

Tyndall Effect

- Is the effect of light scattering in colloids and is not observed through solution and suspension.

- In suspension may scatter light when the particle are large, and if the particle are small, it doesn't scatter the light and can be opaque.

- This effect is used to know the mixture is colloids or true solution.

Flashlight

Solution   Colloid   Suspension
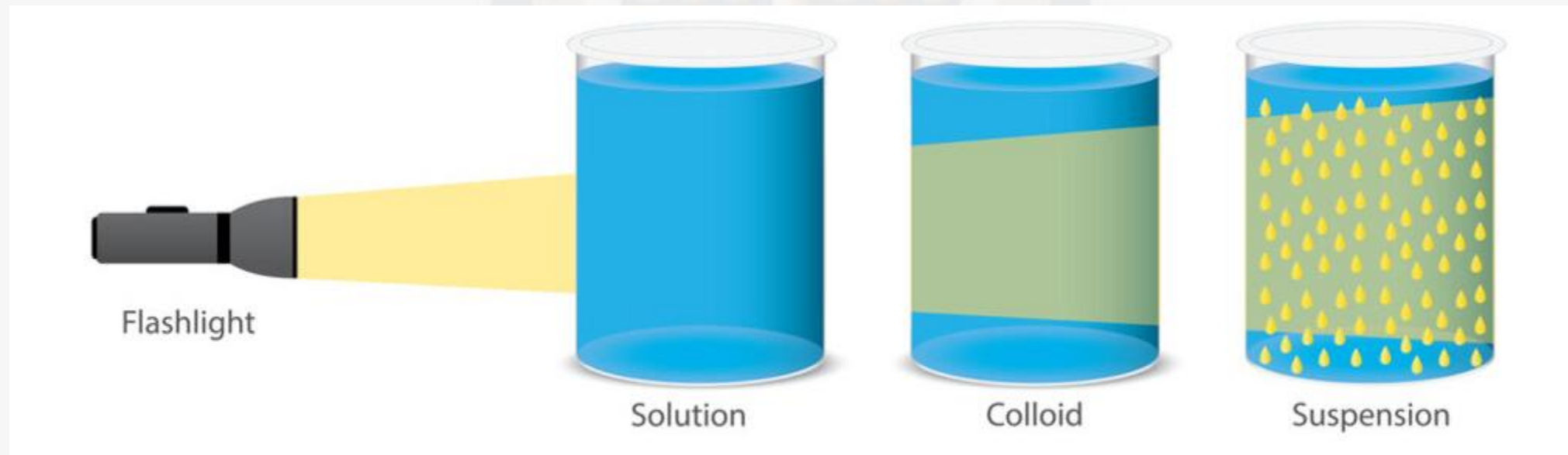
Suspensions:

- Is a heterogeneous mixture in which the solid particles are spread throughout the liquid without dissolving in it.

Properties of Suspension

- The size of solute particles in a suspension is quite large.
- The particles of a suspension can be seen easily.
- The particles of a suspension do not pass through a filter paper. So, a suspension can be separated by filtration.
- The suspension is unstable. The particles of a suspension settle down after some time.
- A suspension scatters a beam of light passing through it because of its large particle size.

Examples of Suspension

- Muddy water
- Milk of magnesia
- Sand particles suspended in water
- Flour in water
- Slaked lime for whitewashing
- Paints in which dyes are suspended in turpentine oil.

Physical property:
• It is the change of matter from state to another (solid-liquid-gas and making a mixture).
No change in the composition of matter, no new substance.

Chemical property:
• Changing in the composition of matter , New substance formed.

Alloy:
• Metallic substance composed of two or more elements, as either a compound or a solution.

Physical properties of matter:

❑ Texture: It refers to the physical appearance or feel of an object, the physical feel of something smooth rough, fuzzy, slimy, and lots of textures something in between.

❑ Uniformity; It determines if the matter is regular "uniform distribution" (homogenous mixture))
or irregular "non-uniform distribution" (heterogeneous mixture)

❑ Malleability: The property of a material to be able to be hammered and reshaped into various shapes without breaking.

❑ Bounce: The tendency to return to its original shape

❑ **Luster:** Metals are lustrous, which means they have a shiny surface

❑ **Elasticity:** The property of solid material that it gains its original shape and size after the removal of applied force.

❑ **Plasticity:** The property of solid material that it does not gain its original shape and size after the removal of applied force.

❑ **Flexibility:** The ability of a material to deform elastically and return to its original shape when the applied stress is removed.

❑ **Ductility:** A crucial property that describes the ability of a material to be stretched, pulled, or drawn into a thin wire or thread without breaking.

❑ **Strength:** Refers to a material's capacity to endure applied forces without breaking or deforming permanently.

❑ **Hardness:** A measure of how well a material resists deformation.

Composites:

- It is a solid heterogeneous mixture of two of more substances that makes use of the properties of each component. This means that the produced composite has extraordinary properties such as high strength, low density, corrosion resistance, cracking resistance,…etc unlike its individual components.

- For example: carbon-composites are fiber- reinforced, chemically inert materials which possess high strength, rigidity, thermal stability, high resistance to thermal shock and retain their mechanical properties at high temperature. Such properties have led to their use in external body parts of the space shuttle

- Also, composites are seen in building block, car body, planes body

The Mole:

- The amount of a chemical substance that contains Avogadro's number of particles (atoms, molecules, ions, electrons).

Avogadro's Number = 6.02 x 10^23
- 1 Mole = 6.02 x 10^23
- Mole = mass /molar mass
- Mole = number of units / 6.02 x 10^23

Polyatomic elements (diatomic) like:

- Oxygen
- Sulfur
- Nitrogen
- Chlorine
- Phosphors
- Florine
- Hydrogen
- Bromine
- Iodine

No. of particles	
Moles	6.02x10x^23

Mass	
Moles	Molar mass
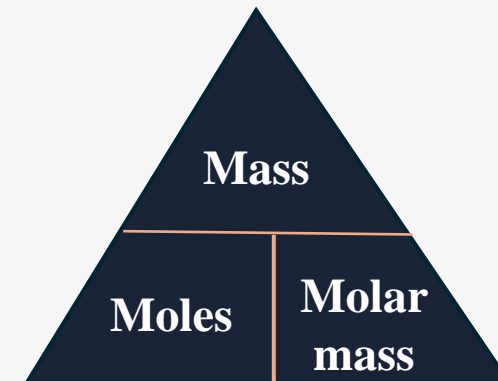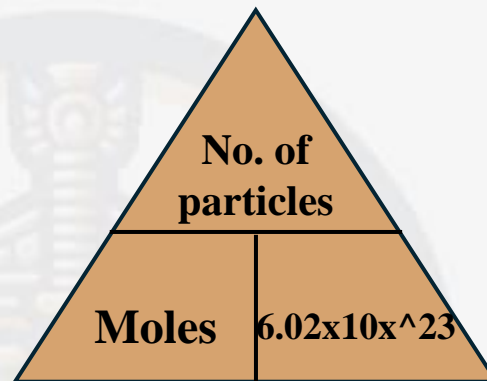
Calculating molar volume:

- The molar volume is the volume occupied by one mole of a gas.
- The units used for the molar volume are l mol -1 (liters per mole).

Limiting reagent

- This method is useful when there are only two reactance.
- This limiting reagent can be derived by comparing the amount of produces that can be formed from each reactant .

α/β = Mole / ratio = (limiting reagent)

Theoretical and actual yield :

- Actual yield is the amount of produce actually obtained from a reaction.
- Theoretical yield is the amount of produce that would be resulted if the limiting reagent reacted.
- % yield = actual / theoretical x 100

How to Calculate Percent Yield Theoretical Yield?

- The amount of product formed when the limiting reactant is completely consumed.
- percent yield = (Actual yield)÷(Theoretical yield) x 100

Chemical Equations and Balancing Chemical Equations

- **Chemical Reaction a process in which one or more substances, the reactants, are converted to one or more different substances, the products. Substances are either chemical elements or compounds.**

- **A chemical reaction rearranges the constituent atoms of the reactants to create different substances as products. Reactants are on the left side of the equation, and products are on the right.**

Balancing Chemical Equations:

- **Making sure that the amounts and charges of substances are the same on both sides of a chemical reaction.**

Example: **Balance the equation:**
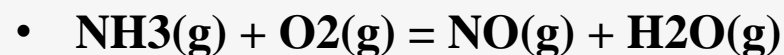
- $NH_3(g) + O_2(g) = NO(g) + H_2O(g)$
Using lowest whole number coefficients

Answer: **$4NH_3(g) + 5O_2(g) = 4NO(g) + 6H_2O(g)$**

Molar Mass:

- The mass in grams of one mole of an element or compound

Example: Calculate the molar mass of Methane (CH4)

- (1 atom of Carbon) x (12.0 g/mol Carbon) + (4 atoms of Hydrogen) x (1.00 g/mol Hydrogen)
= 16.0 g/mol Methane.

Percent Composition:

- How abundant one element in a compound is compared to the other(s) in the same compound.

- % Composition = (M.M. of element / M.M. of compound) ×100

Chemical reactions :

- **Beginning with reactants and end with products.**
- **Synthesis reactions bond reactants together, a process that requires energy.**
- **Decomposition reactions break the bonds within a reactant and thereby release energy.**
- **In exchange reactions, bonds are both broken and formed, and energy is exchanged.**
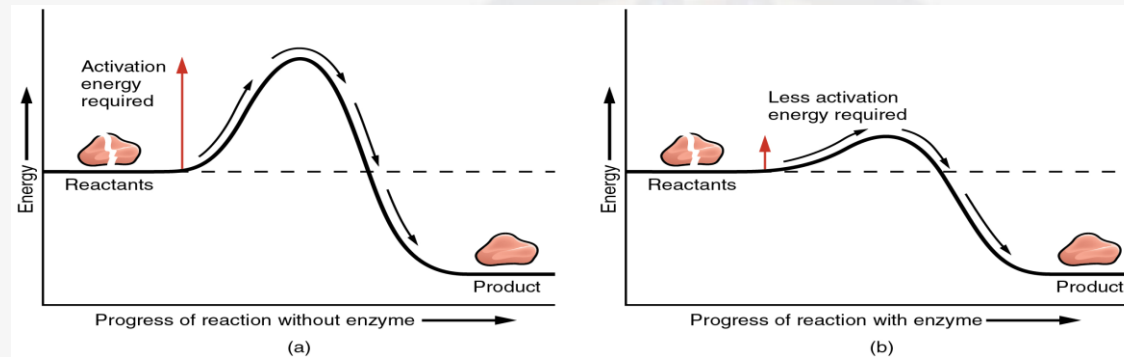
Reactions:

- **A chemical process in which substances act mutually on each other and are changed into different substances, or one substance changes into other substances.**
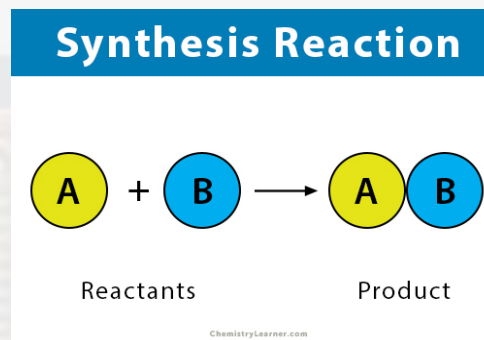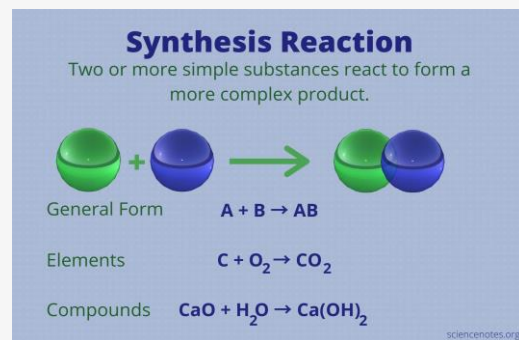
Types of Reactions:

- **Synthesis reaction**
- **Decomposition reaction**
- **Single-replacement reaction**
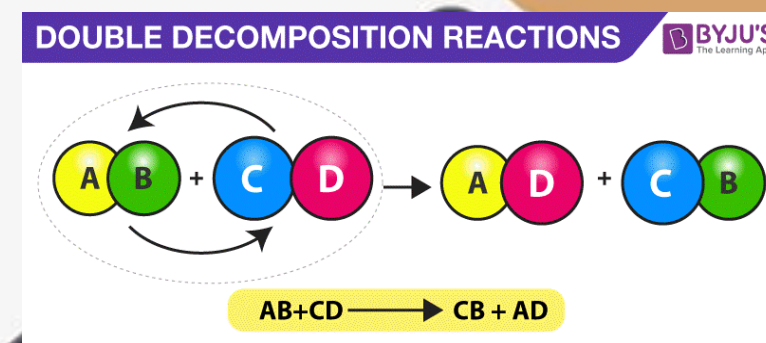- **Double-replacement reaction**

Synthesis reactions:

- It is s a reaction in which two or more substances react to produce a single produce.
- When two element react, the reaction is always a synthesis.


Decomposition reaction:

- It is a reaction can be defined as a chemical reaction in which one reactant breaks down into two or more products.

- This type of reaction is represented by the general equation: A + B → AB. An example of a synthesis reaction is the combination of sodium (Na) and chlorine (Cl) to produce sodium chloride (NaCl).

- This reaction is represented by the chemical equation: $2Na + Cl_2 → 2NaCl$.

Single replacement reaction:

- A single-replacement reaction is a chemical reaction in which one element replaces another in a compound.

- **Example:** A specific example of a single-replacement reaction is $Zn + 2HCl \longrightarrow ZnCl2 + H2$
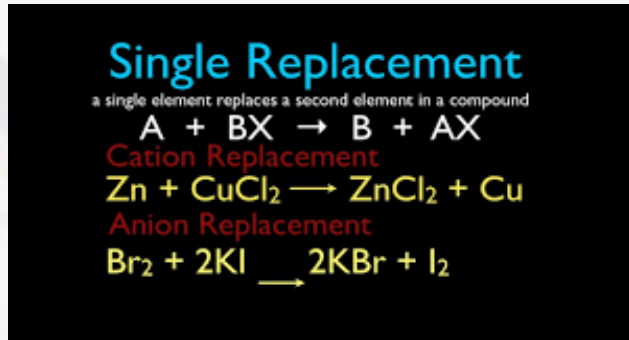
Double replacement reaction:

- Double replacement reactions are defined as those in which two ionic compounds exchange their ions. Ions of the same kind and two new products are formed. An ionic compound is neutral but has cations that carry positive charges and anions that carry negative charges

- **Example 1:** when solutions of sodium sulfate and silver nitrate are mixed, there is an immediate formation of a precipitate .The precipitate is silver sulfate

- **Example 2:** Potassium hydroxide solution reacts with dilute nitric acid, undergoing a double replacement reaction
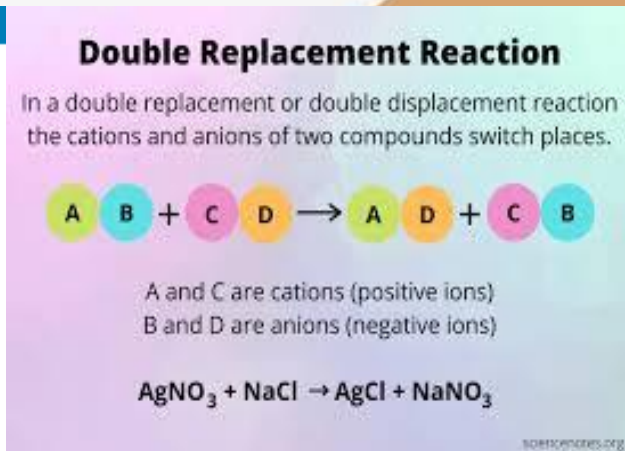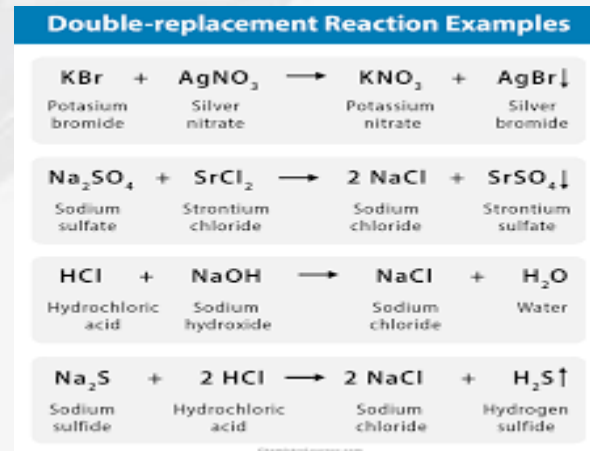
Intramolecular forces:

- Are the forces that hold atoms together within a molecule

- These forces are responsible for the structural stability of molecules and influence their physical and chemical properties

- The three main types of intramolecular forces are:

Ionic Bonds:

- Ionic bonds are formed when electrons are transferred from one atom to another, resulting in the formation of positively charged ions (cations) and negatively charged ions (anions).

- These ions are held together by the electrostatic attraction force between opposite charges.

- Ionic bonds are typically found in compounds composed of metal and nonmetal atoms.

Ionic Bond

Transfer of electrons

Atom 1 (metal)   Atom 2 (nonmetal)   Ionic molecule

ChemistryLearner.com
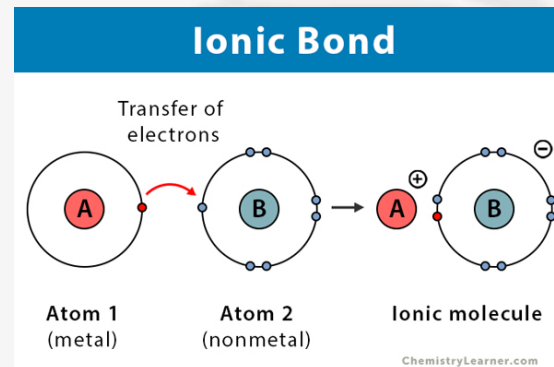
Covalent Bonds:

- Covalent bonds are formed when atoms share pairs of electrons to achieve a stable electron configuration to be like the closest Nobel gas.

- This sharing of electrons creates a strong bond between the atoms, holding them together in a molecule.

- Covalent bonds can be further classified as polar covalent bonds and nonpolar

- Covalent bonds, depending on the electronegativity difference between the bonded atoms.

Polar Covalent Bonds:

- Polar covalent bonds are occurred when there is an unequal sharing of electrons between atoms due to differences in electronegativity.

- This results in the formation of partial positive and partial negative charges within the molecule.
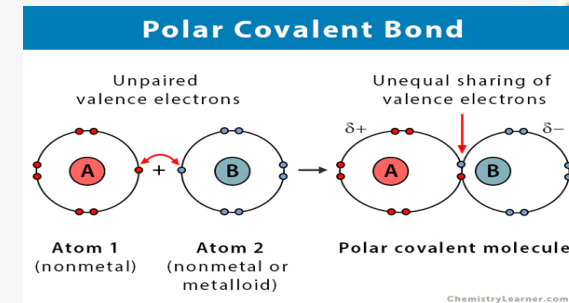
Nonpolar Covalent Bonds:

- Nonpolar covalent bonds are occurred when atoms with similar electronegativities share electrons equally, resulting in a balanced distribution of charge within the molecule.
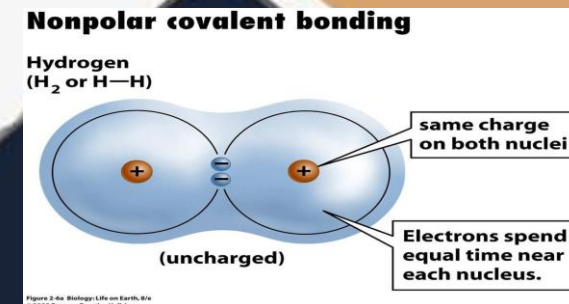
Metallic Bonds:

- Metallic bonds are formed between metal atoms in a metallic structure.

- In metallic bonding, valence electrons are delocalized and free to move throughout the entire metal lattice.

- This electron delocalization results in a strong electrostatic attraction between the positively charged metal ions and the negatively charged electrons, giving metals their characteristic properties such as conductivity, malleability, and ductility.

Intermolecular forces :

- They are the forces of attraction or repulsion that occur between molecules.

- These forces are responsible for determining the physical properties of Substances such as melting and boiling points, viscosity, surface tension and solubility.

- Unlike intramolecular forces, which occur within molecules, intermolecular forces act between molecules.

- The main types of intermolecular forces include:
    London Dispersion Forces
    Dipole-Dipole Interactions
    Hydrogen Bonding
    Ion-Dipole Interactions
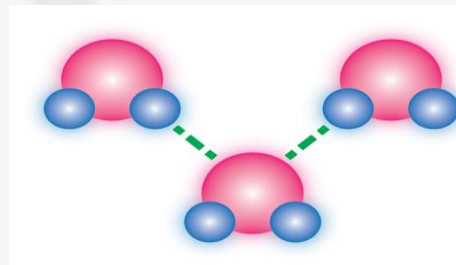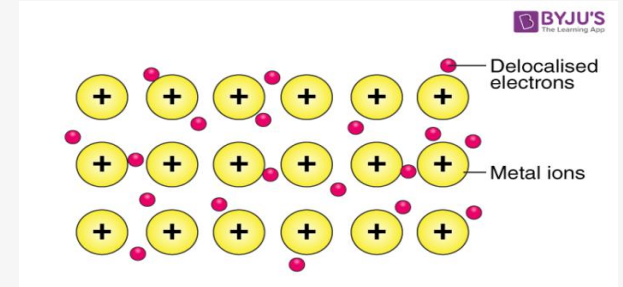
Van der Waals Forces:-

❑ **London Dispersion Forces:**
• These forces result from temporary fluctuations in electron distribution within molecules, creating temporary dipoles.
• These temporary dipoles induce similar temporary dipoles in neighboring molecules, leading to an attractive force between them.
• London dispersion forces are present in all molecules, regardless of polarity, but are generally weaker in nonpolar molecules compared to polar molecules.
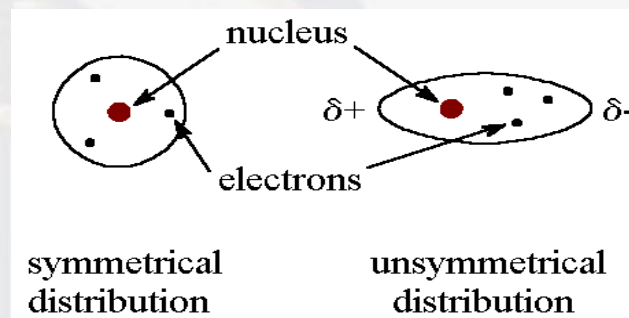
❑ **Dipole-Dipole Interactions:**
• Dipole-dipole interactions occur between polar molecules that possess permanent dipole moments due to differences in electronegativity between atoms. The positive end of one
• polar molecule is attracted to the negative end of another polar molecule, resulting in an attractive force between them.
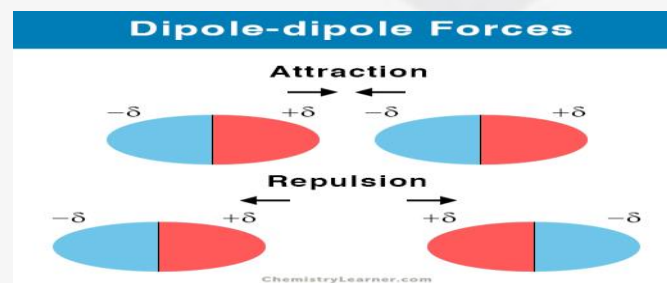
❑ Hydrogen Bonding:

- Hydrogen bonding is a special type of dipole-dipole interaction that occurs when a hydrogen atom covalently bonded to an electronegative atom (such as nitrogen, oxygen, or fluorine) interacts with another electronegative atom in a neighboring molecule.

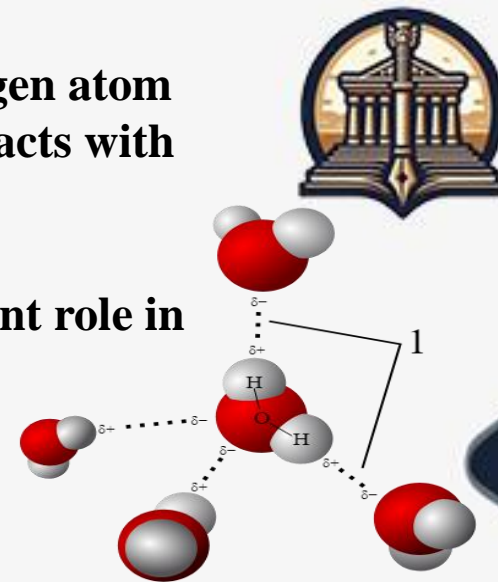- Hydrogen bonding is stronger than typical dipole-dipole interactions and plays a significant role in determining the properties of substances like water, ammonia, and hydrogen fluoride.

❑ Ion-Dipole Interactions:

- Ion-dipole interactions occur between ions (charged particles) and polar molecules.

- The positive or negative end of a polar molecule is attracted to oppositely charged ions, leading to an attractive force between them. Ion-dipole interactions are particularly important in solutions where ions are solvated by polar solvent molecules.

- Understanding intermolecular forces is crucial for explaining the behavior of substances in various physical and chemical processes, including phase transitions, solution formation and the structure of condensed phases (liquids and solids).
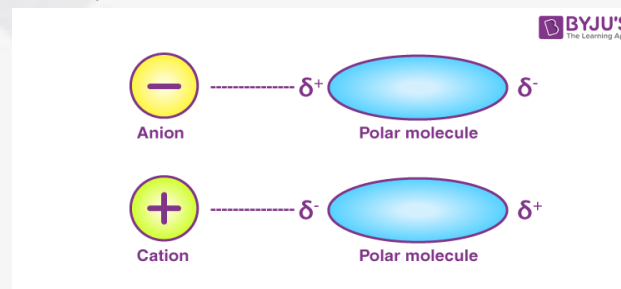
The strength of this interaction depends on:-

- The magnitude of the dipole moment.
- Size of the polar molecule.
- The distance between ion and molecular dipole.

Intermolecular Interactions	kJ/mol
Ionic	700-4000
Ion-Dipole	40-600
Hydrogen Bonding	10-40
Dipole-Dipole	5-25
Ion-Induced Dipole	3-15
Dipole-Induced Dipole	2-10
Dispersion (London)	0.05-40

Type of IMFA	Involves	Occurs Between...	Strength of Attraction	Effect on Boiling/Melting Points	Effect on Freezing Points
London/Dispersion	Temporary Dipoles	Nonpolar molecules	Low	Low elevation (↑BP)	(↓↓↓↓FP)
Dipole-Dipole	Permanent dipoles	Polar molecules	Medium	Medium elevation (↑↑BP)	(↓↓↓FP)
Hydrogen Bonds	Permanent dipoles between H & F,O, N	Polar molecules	Medium-High	Medium-High elevation (↑↑↑BP)	(↓↓FP)
Ion-Dipole	Full ion and dipole	Polar molecules	High	High elevation (↑↑↑↑BP)	(↓FP)

Electronegativity:

- It is the ability of an atom in a covalent molecule to attract the electrons of the bond toward itself.

- The difference in Electronegativity of the covalent compounds measures the polarity of these molecules.

Electronegativity Difference

Nonpolar Covalent Bond	0.4	Polar Covalent Bond	1.8	Ionic Bond

Examples:

C-H	C-O	NaCl
2.5-2.1 = 0.4	3.5 − 2.5 = 1.0	3.0 − 0.9 = 2.1
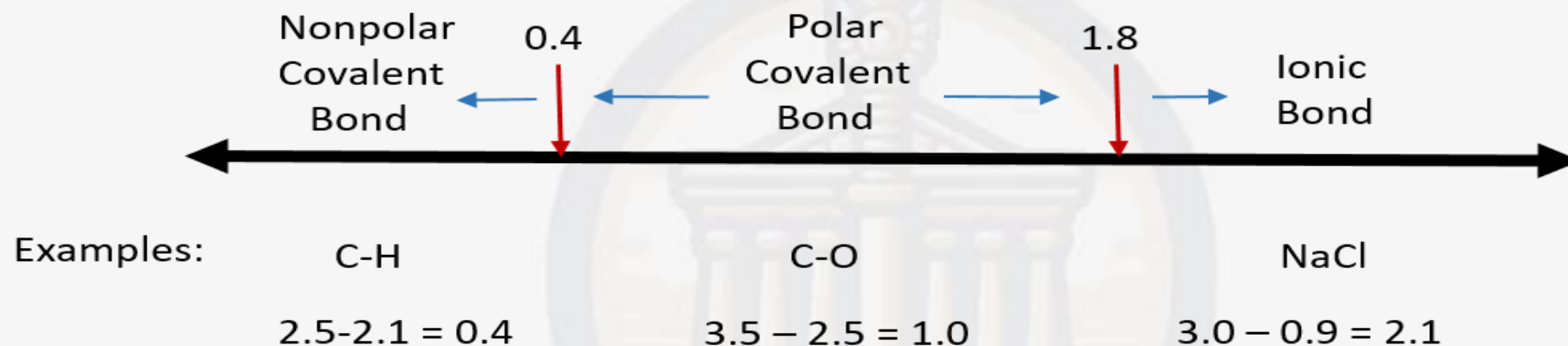
Notes

- Inert gases don't have electronegativity. That's because most of them don't enter chemical reactions so they can't form compound.

- By increasing in difference in electronegativity, polarity increases then the melting point, boiling point and conductivity also increase.

Lewis Structure

- Is a representation for the valence shell electrons of a molecule

- It is used to show how the electrons in a molecule are arranged around individual atoms.

- Electrons are shown as "dots" or, in the case of a bond, as a line connecting the two atoms.

Valence Shell Electron Pair Repulsion
(VSEPR Theory)

l.p.	b.a.	Parent Shape	Molecular Shape	Picture	Angles
0	2	Linear	Linear	X—E—X	180°
0	3	Trigonal Planar	Trigonal Planar		120°
1	2		Bent		< 120°
0	4	Tetrahedral	Tetrahedral		109.5°
1	3		Trigonal Pyramidal		107°
2	2		Bent		104°

What is the difference between metals and non-metals?

Metals	Non-metals
Metals are sonorous, which means produces sound when hit by an object	Non-metals do not produce any sound
Metals are good reducing agent because it loses electrons easily from their valance shell	Non-metals are good oxidizing agent because it accepts electrons in their outermost shell
React with oxygen to form basic oxides	Usually react with oxygen to form acidic oxides
These have low ionization enthalpy	These have high ionization enthalpy
Loses electrons to form electropositive ions	Accepts electrons to form electronegative ions
Metals contain 1-3 electrons in their outermost shell	Non-metals contain 4-8 electrons in their outermost shell
Metals are lustrous, i.e., having a shining appearance	Non-metals are dull in appearance
Metals are malleable	Non-metals are not malleable

Periodic Table of the Chemical Elements

Chemical activity series

- It is the arrangement of metals in the descending order of their reactivities.

- The activity series is a list of elements in decreasing order of their reactivity.

- The element can replace the element that follows it in the chemical activity serious

- Elements above hydrogen are reducing agents

- Elements below hydrogen are oxidizing agents

What is an oxidation-reduction redox reaction?

- It is a reaction thus consists of two half-reactions oxidation half reaction, reduction half reaction

In oxidation

- Oxygen atoms increase
- Hydrogen atoms decrease
- Oxidation number increase
- Loses electron

In reduction

- Oxygen atoms decrease
- Hydrogen atoms increase
- Oxidation number decreases
- Gains electron

Increase in O.N = 3

$$\overset{+2}{Cu}O + \overset{-3}{N}H_3 \longrightarrow \overset{0}{Cu} + \overset{0}{N_2} + H_2O$$

Decrease in O.N = 2
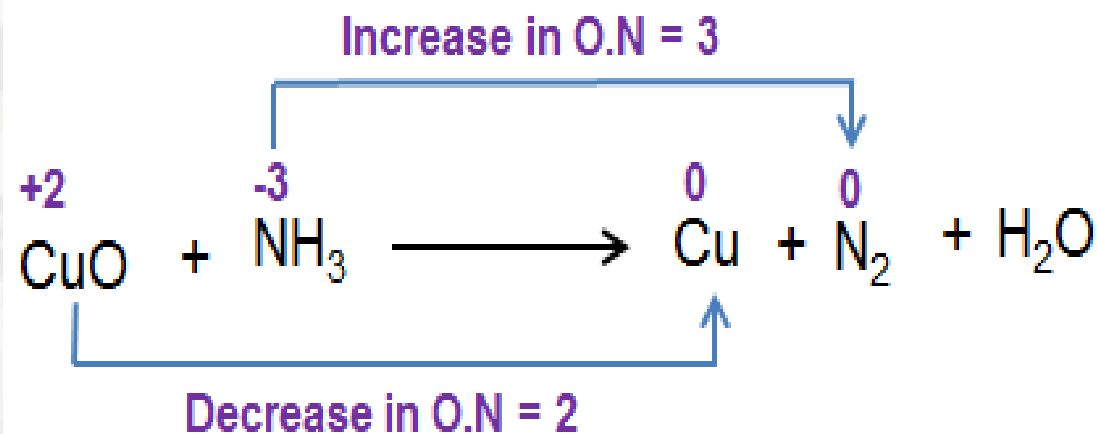
Oxidation number

In elements:
- If they have no charge, then the oxidation number = zero

O_2, H_2, Cl_2, N_2, C, Na, K

- If it is an ion, then the oxidation number = the charge

Na^+, Cl^-, Mg^{2+}, Al^{3+}

In compounds
- The charge = the sum of the oxidation number of the two elements
- In H_2O O.N of H_2 + O.N OF O = zero
- In $NH_4 +$ O.N of N + O.N of H_4 =1
- In PO_4 3- O.N of P + O.N of O_4 = -3

In some cases, the oxidation number is constant
- group 1A has O.N =1
- group 2A has O.N =2
- group 3A has O.N =3
- Hydrogen has O.N = -1 when reacting with metals, and O.N = 1 when reacting with non-metals
- Oxygen normally has O.N = -2, in peroxide = -1, in super oxide = -1/2
- Florine has O.N = -1, other halogens vary

We have to calculate the oxidation number to know what is oxidized and what is reduced.

Fe 2+ + Cr2O7 2- →Fe 3+ +Cr 3+
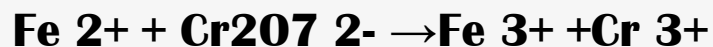
- First, we start by identifying the Oxidation States of Initial Compounds:

Fe 2+ has O.N = 2
- In Cr2O7 2- : here are seven oxygen atoms, each with an oxidation state of -2, so the total for oxygen is $7x-2=-14$
- Since the overall charge of the ion is -2, we can find the total O.N for the two chromium atoms:

- The sum of all oxidation states must be equal to the total charge, so $2x+(-14)=-2$. Solving for x, we get $2x=12$, so $x=6$. Thus, each Cr has an oxidation state of +6.

- Second, we Identify the Oxidation States of Products:
Fe3+ has an oxidation state of +3.
Cr3+ has an oxidation state of +3.

- Third, we Determine Changes in Oxidation States
Fe2+ to Fe3+: The oxidation state increased by one unit (from +2 to +3). This is an oxidation process.
Cr6+ to Cr3+: The oxidation state decreased by three units (from +6 to +3), so This is
a reduction process.

Balancing redox reaction

1. determine what is oxidized and what is reduced
2. split it into 2 half-reactions
3. balancing the atoms(except oxygen and hydrogen)
4. balance oxygen with water (H2O)
5. balance hydrogen with H+
6. balancing the electrons
7. the overall rection

$$Fe^{2+} + Cr_2O_7^{2-} \rightarrow Fe^{3+} + Cr^{3+}$$
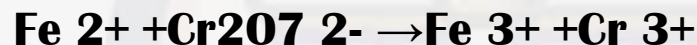
Fe is oxidized, Cr is reduced

Half oxidation: $Fe^{2+} \rightarrow Fe^{3-} + e^-$

Half reduction: $6e^- + Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

In $Cr_2O_7^{2-}$ there are 7 Oxygen atoms, they can be balanced by 7 H2O

And the 14 hydrogen (added previously in H2O) can be balanced by 14 H+

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7 H_2O$$

Then we have to equalize the electron in the two half-reactions

Oxidation half-reaction : $(Fe^{2+} \rightarrow Fe^{3-} + e^-) \times 6 = 6Fe^{2+} \rightarrow 6 Fe^{3+} + 6e^-$

Reduction half-reaction : $(6e^- + Cr_2O_7^{2-} \rightarrow 2Cr^{3+}) \times 1 = 6e^- + Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

Then we add the two half-reactions together to get the overall reaction

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$
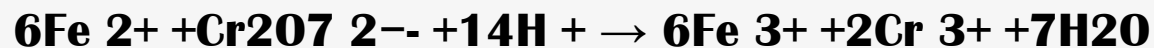
Changing the medium from acidic to base

First, we have to determine whether the solution is acidic or basic

An acidic solution contains **H+**

A Basic solution contains **OH-**

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$
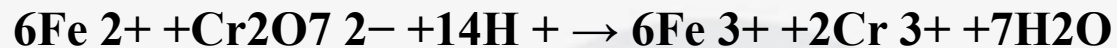
This an acidic solution as it contains H+ so we add OH- to both sides with the number of H+ (which is 14)

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ + 14OH^- \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O + 14OH^-$$
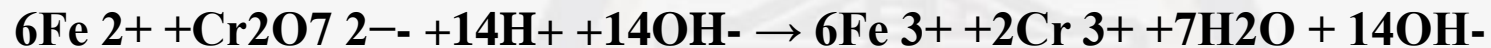
The H+ reacts with OH- forming H2O, at this point we have water in both sides, so we subtract the amount of water in the fewer side from both sides

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H_2O - 7H_2O \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O + 14\ OH^- - 7H_2O$$
$$6Fe^{2+} + Cr_2O_7^{2-} + 7H_2O \rightarrow 6Fe^{3+} + 2Cr^{3+} + 14OH^-$$
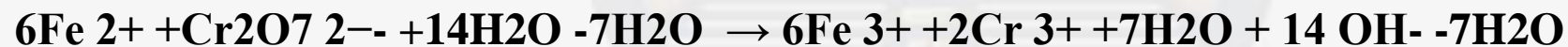
Electrochemical cell

- It is a device that converts the chemical energy into electric energy.
- It consists of two half-reactions; oxidation occurs at the anode, and reduction occurs at the cathode
- The electrode (which is a metal rod) is put in the electrolyte (which is the solution).
- As in the galvanic cell, the Zn rod is put in $ZnSO_4$ solution (Anode), and the Cu rod is put in $CuSO_4$ solution(cathode)
- Then they are connected, and the circuit is closed by using a salt bridge.
- An oxidation reaction occurs at the anode, and the electrons transfer from the anode to the cathode, and the reduction reaction occurs.
- The Zn mass decreases, Zn 2+ concentration increases, Cu mass increases, and Cu 2- decreases.

Martin S. Silberberg, *Chemistry: The Molecular Nature of Matter and Change*, 2nd Edition. Copyright © The McGraw-Hill Companies, Inc. All rights reserved.

A

Voltmeter

Anode Zn (−)   Salt bridge   Cathode (+) Cu

Na SO_4^{2-}

$2e^-$

Zn Zn^{2+}

Zn^{2+}

$2e^-$

Cu^{2+}

Cu^{2+}

Oxidation half-reaction
$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

Reduction half-reaction
$2e^- + Cu^{2+}(aq) \longrightarrow Cu(s)$

Overall (cell) reaction
$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

© McGraw-Hill Higher Education/Stephen Frisch, photographer

B

Zinc-Copper Reaction Voltaic Cell

The E.M.F can be calculated by these laws:

E.M.F = SRP of the cathode + SOP of the anode
E.M.F = SRP of the cathode – SRP of the anode
E.M.F = SOP of the anode – SOP of the cathode

For example:
E.M.F of Zn + Cu 2- → Zn 2+ + Cu
=SRP of the Cu + SOP of Zn
=0.16 + 0.34 = 1.1 Volts

N.P. Any rule can be used depending on the givens

Anodic - exhibits greater tendency to lose electrons

Reduction Reaction	E^0(V)	Oxidation Reaction	E^0 (V)
$Li^+ + e^- \rightarrow Li$	-3.04	$Li \rightarrow Li^+ + e^-$	3.04
$K^+ + e^- \rightarrow K$	-2.92	$K \rightarrow K^+ + e^-$	2.92
$Ba^{2+} + 2e^- \rightarrow Ba$	-2.90	$Ba \rightarrow Ba^{2+} + 2e^-$	2.90
$Ca^{2+} + 2e^- \rightarrow Ca$	-2.87	$Ca \rightarrow Ca^{2+} + 2e^-$	2.87
$Na^+ + e^- \rightarrow Na$	-2.71	$Na \rightarrow Na^+ + e^-$	2.71
$Mg^{2+} + 2e^- \rightarrow Mg$	-2.37	$Mg \rightarrow Mg^{2+} + 2e^-$	2.37
$Al^{3+} + 3e^- \rightarrow Al$	-1.66	$Al \rightarrow Al^{3+} + 3e^-$	1.66
$Mn^{2+} + 2e^- \rightarrow Mn$	-1.18	$Mn \rightarrow Mn^{2+} + 2e^-$	1.18
$2H_2O + 2e^- \rightarrow H_2 + 2 OH^-$	-0.83	$H_2 + 2 OH^- \rightarrow 2H_2O + 2e^-$	0.83
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76	$Zn \rightarrow Zn^{2+} + 2e^-$	0.76
$Cr^{2+} + 2e^- \rightarrow Cr$	-0.74	$Cr \rightarrow Cr^{2+} + 2e^-$	0.74
$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44	$Fe \rightarrow Fe^{2+} + 2e^-$	0.44
$Cr^{3+} + 3e^- \rightarrow Cr$	-0.41	$Cr \rightarrow Cr^{3+} + 3e^-$	0.41
$Cd^{2+} + 2e^- \rightarrow Cd$	-0.40	$Cd \rightarrow Cd^{2+} + 2e^-$	0.40
$Co^{2+} + 2e^- \rightarrow Co$	-0.28	$Co \rightarrow Co^{2+} + 2e^-$	0.28
$Ni^{2+} + 2e^- \rightarrow Ni$	-0.25	$Ni \rightarrow Ni^{2+} + 2e^-$	0.25
$Sn^{2+} + 2e^- \rightarrow Sn$	-0.14	$Sn \rightarrow Sn^{2+} + 2e^-$	0.14
$Pb^{2+} + 2e^- \rightarrow Pb$	-0.13	$Pb \rightarrow Pb^{2+} + 2e^-$	0.13
$Fe^{3+} + 3e^- \rightarrow Fe$	-0.04	$Fe \rightarrow Fe^{3+} + 3e^-$	0.04

Arbitrary Neutral : H_2

Reduction Reaction	E^0(V)	Oxidation Reaction	E^0 (V)
$2H^+ + 2e^- \rightarrow H_2$	0.00	$H_2 \rightarrow 2H^+ + 2e^-$	0.00

Cathodic - exhibits greater tendency to gain electrons

Reduction Reaction	E^0(V)	Oxidation Reaction	E^0 (V)
$S + 2H^+ + 2e^- \rightarrow H_2S$	0.14	$H_2S \rightarrow S + 2H^+ + 2e^-$	-0.14
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	0.15	$Sn^{2+} \rightarrow Sn^{4+} + 2e^-$	-0.15
$Cu^{2+} + e^- \rightarrow Cu^+$	0.16	$Cu^+ \rightarrow Cu^{2+} + e^-$	-0.16
$SO_4^{2+} + 4H^+ + 2e^- \rightarrow SO_2 + 2H_2O$	0.17	$SO_2 + 2H_2O \rightarrow SO_4^{2+} + 4H^+ + 2e^-$	-0.17
$AgCl + e^- \rightarrow Ag + Cl^-$	0.22	$Ag + Cl^- \rightarrow AgCl + e^-$	-0.22
$Cu^{2+} + 2e^- \rightarrow Cu$	0.34	$Cu \rightarrow Cu^{2+} + 2e^-$	-0.34
$ClO_3^- + H_2O + 2e^- \rightarrow ClO_2^- + 2OH^-$	0.35	$ClO_2^- + 2OH^- \rightarrow ClO_3^- + H_2O + 2e^-$	-0.35
$2H_2O + O_2 + 4e^- \rightarrow 4OH^-$	0.40	$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$	-0.40
$Cu^+ + e^- \rightarrow Cu$	0.52	$Cu \rightarrow Cu^+ + e^-$	-0.52
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$2I^- \rightarrow I_2 + 2e^-$	-0.54
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$	-0.68
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Fe^{2+} \rightarrow Fe^{3+} + e^-$	-0.77
$NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$	0.78	$NO_2 + H_2O \rightarrow NO_3^- + 2H^+ + e^-$	-0.78
$Hg^{2+} + 2e^- \rightarrow Hg$	0.78	$Hg \rightarrow Hg^{2+} + 2e^-$	-0.78
$Ag^+ + e^- \rightarrow Ag$	0.80	$Ag \rightarrow Ag^+ + e^-$	-0.80
$NO_3^- + 4H^+ +3 e^- \rightarrow NO + 2H_2O$	0.96	$NO + 2H_2O \rightarrow NO_3^- + 4H^+ +3 e^-$	-0.96
$Br_2 + 2e^- \rightarrow 2Br^-$	1.06	$2Br^- \rightarrow Br_2 + 2e^-$	-1.06
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$	-1.23
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.28	$Mn^{2+} + 2H_2O \rightarrow MnO_2 + 4H^+ + 2e^-$	-1.28
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$2Cr^{3+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+ + 6e^-$	-1.33
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$2Cl^- \rightarrow Cl_2 + 2e^-$	-1.36
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.44	$Ce^{3+} \rightarrow Ce^{4+} + e^-$	-1.44
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Au \rightarrow Au^{3+} + 3e^-$	-1.50
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.52	$Mn^{2+} + 4H_2O \rightarrow MnO_4^- + 8H^+ + 5e^-$	-1.52
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78	$2H_2O \rightarrow H_2O_2 + 2H^+ + 2e^-$	-1.78
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.82	$Co^{2+} \rightarrow Co^{3+} + e^-$	-1.82
$S_2O_8^{2-} + 2e^- \rightarrow 2SO_4^{2-}$	2.01	$2SO_4^{2-} \rightarrow S_2O_8^{2-} + 2e^-$	-2.01
$O_3 + 2H^+ + 2e^- \rightarrow O_2 + H_2O$	2.07	$O_2 + H_2O \rightarrow O_3 + 2H^+ + 2e^-$	-2.07
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$2F^- \rightarrow F_2 + 2e^-$	-2.87

What is the thermodynamics ?

- the study of the interrelation of heat and work with chemical reactions or with physical changes of state within the confines of the laws of thermodynamics.

Different Branches of Thermodynamics :

- ❑ Classical Thermodynamics
- ❑ Statistical Thermodynamics
- ❑ Chemical Thermodynamics
- ❑ Equilibrium Thermodynamics

- **Classical Thermodynamics**

In classical thermodynamics, the behavior of matter is analyzed with a macroscopic approach. Units such as temperature and pressure are taken into consideration, which helps the individuals calculate other properties and predict the characteristics of the matter undergoing the process.

- **Statistical Thermodynamics**

In statistical thermodynamics, every molecule is under the spotlight, i.e., the properties of every molecule and how they interact are taken into consideration to characterize the behavior of a group of molecules.

- **Chemical Thermodynamics**

Chemical thermodynamics is the study of how work and heat relate to each other in chemical reactions and in changes of states.

- **Equilibrium Thermodynamics**

Equilibrium thermodynamics is the study of transformations of energy and matter as they approach the state of equilibrium.

What is the Thermodynamic Systems?

- A thermodynamic system is a specific portion of matter with a definite boundary on which our attention is focused. The system boundary may be real or imaginary, fixed or deformable.

What is the boundary ?

- A conceptual line that divides the system that you want to study from everything else
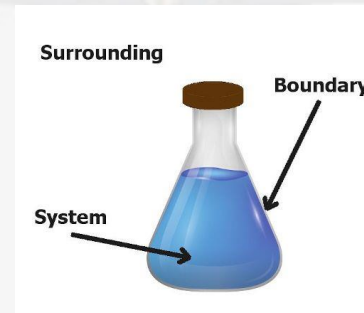
Types of systems:

- Isolated System – An isolated system cannot exchange energy and mass with its surroundings. The universe is considered an isolated system.

- Closed System – Across the boundary of the closed system, the transfer of energy takes place, but the transfer of mass doesn't take place. Refrigerator, compression of gas in the piston-cylinder assembly are examples of closed systems.

- Open System – In an open system, the mass and energy both may be transferred between the system and surroundings. A steam turbine is an example of an open system.

Open system   Closed system   Isolated system
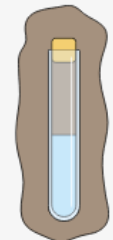
Thermodynamic Process:

- Adiabatic Process – A process where no heat transfer into or out of the system occurs.

- Isochoric Process – A process where no change in volume occurs and the system does no work.

- Isobaric Process – A process in which no change in pressure occurs.

- Isothermal Process – A process in which no change in temperature occurs.

Thermodynamic Properties:

• **Intensive properties** are properties that do not depend on the quantity of matter.
Ex. Pressure and temperature

• In the case of **extensive properties**, their values depends on the mass of the system.
Ex. Volume, energy, and enthalpy

What is Enthalpy?

• Enthalpy is the measurement of energy in a thermodynamic system. The quantity of enthalpy equals the total heat content of a system, equivalent to the system's internal energy plus the product of volume and pressure.

• $H = E + PV$

The enthalpy (H) equals the sum of the internal energy (E) and the product of the pressure (P) and volume (V) of the system.

First Law of Thermodynamics:

- First law of thermodynamics, also known as the law of conservation of energy, states that energy can neither be created nor destroyed, but it can be changed from one form to another.

Ex: Switching on light may seem to produce energy, but it is electrical energy that is converted.

Second Law of Thermodynamics:

- Second law of thermodynamics states that the entropy in an isolated system always increases. Any isolated system spontaneously evolves towards thermal equilibrium—the state of maximum entropy of the system.

Third Law of Thermodynamics:

- Third law of thermodynamics states that the entropy of a system approaches a constant value as the temperature approaches absolute zero.

Ex: The molecules within it move freely and have high entropy.

Thermodynamics Summary and Overview

- In simple terms, thermodynamics deals with the transfer of energy from one form to another. The laws of thermodynamics are:

- First law of thermodynamics: Energy can neither be created nor be destroyed, it can only be transferred from one form to another.

- Second law of thermodynamics: The entropy of any isolated system always increases.

- Third law of thermodynamics: The entropy of a system approaches a constant value as the temperature approaches absolute zero.

- Zeroth law of thermodynamics: If two thermodynamic systems are in thermal equilibrium with a third system separately, then they are in thermal equilibrium with each other.

- Entropy is the measure of the number of possible arrangements the atoms in a system can have.

- Enthalpy is the measurement of energy in a thermodynamic system.

Endothermic reaction: (Heat is absorbed)

- Chemical reactions in which the reactants absorb heat energy from the surroundings to form products.

➢ Photosynthesis: Plants absorb heat energy from sunlight to convert carbon dioxide and water into glucose and oxygen.

➢ Cooking an egg: Heat energy is absorbed from the pan to cook the egg.
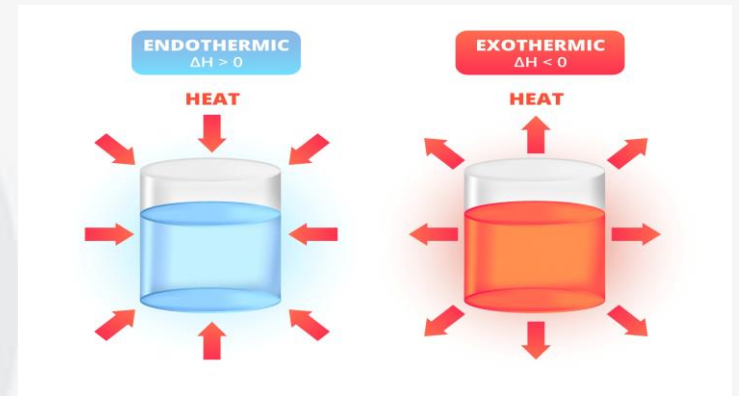
Exothermic reaction: (Heat is released)

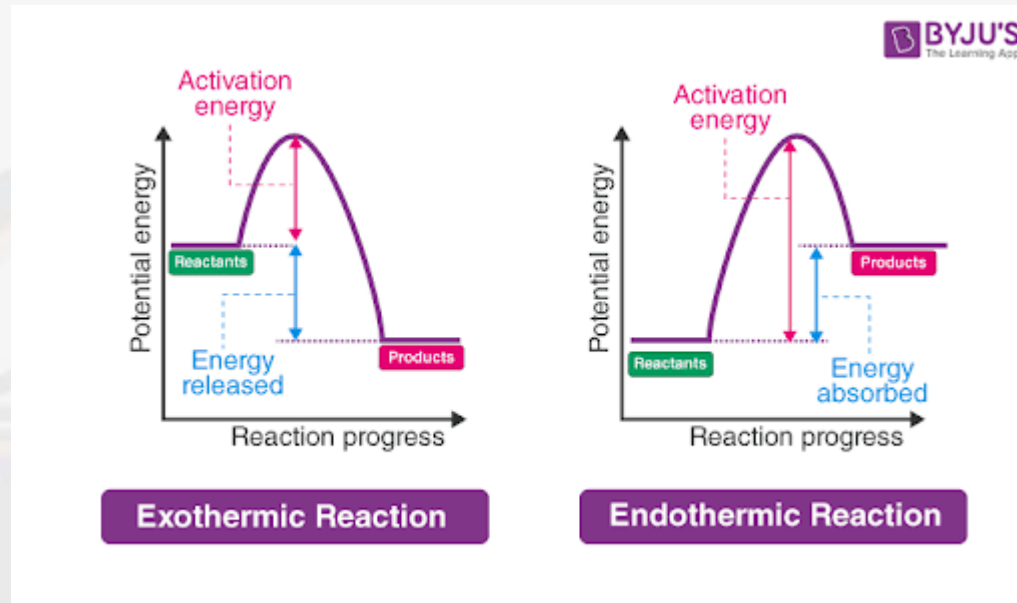- Chemical reaction that releases energy by light or heat.

➢ Combustion: The burning of carbon-containing compounds uses oxygen, from air, and produces carbon dioxide, water, and lots of heat.

➢ Rain: Condensation of water vapor into rain releasing energy in the form of heat

Endothermic Vs Exothermic

Comparison	Endothermic reactions	Exothermic reactions
Main theme	endo = inside	exo = outside
Definition	Reactions that absorb heat	Reactions that evolve heat
Surrounding Temperature	Decreases	Increases
$\Delta E = E_p - E_r$	Positive	Negative
Stability of products	Less	More
ΔH (Enthalpy change)	Positive	Negative
ΔS (Entropy change)	Negative	Positive
Spontaneity	Non-spontaneous	Spontaneous
Endergonic / Exergonic reactions	Endergonic	Exergonic
Examples	Melting of ice Photosynthesis Cooking an egg Thermal cracking	Freezing of water Combustion Rusting of iron Fission and fusion

Difference between Heat and Temperature

Quantity Used to Compare	Heat	Temperature
Definition	The heat of an object is the total energy of all the molecular motion inside that object.	Temperature is the measure of the thermal energy or average heat of the molecules in a substance.
SI Unit	Joule	Kelvin
Unit	Joules, Calories	Fahrenheit, Celsius, Kelvin
Symbol	Q	T
Particles	It's a measure of the number of atoms multiplied by the energy possessed by each atom.	It is like a level which determines the direction in which the heat will flow. It is related to how fast the molecules within a substance are moving.
Ability to do Work	It has the ability to do work	It can be used to measure the degree of heat.

- The energy changes that occur during a chemical reaction can be shown in a diagram called a potential energy diagram, or sometimes called a reaction progress curve.

- A potential energy diagram shows the change in potential energy of a system as reactants are converted into products.

- The figure below shows basic potential energy diagrams for an endothermic (A) and an exothermic (B) reaction.
- Recall that the enthalpy change (ΔH) is positive for an endothermic reaction and negative for an exothermic reaction. This can be seen in the potential energy diagrams.

- The total potential energy of the system increases for the endothermic reaction as the system absorbs energy from the surroundings. The total potential energy of the system decreases for the exothermic reaction as the system releases energy to the surroundings.
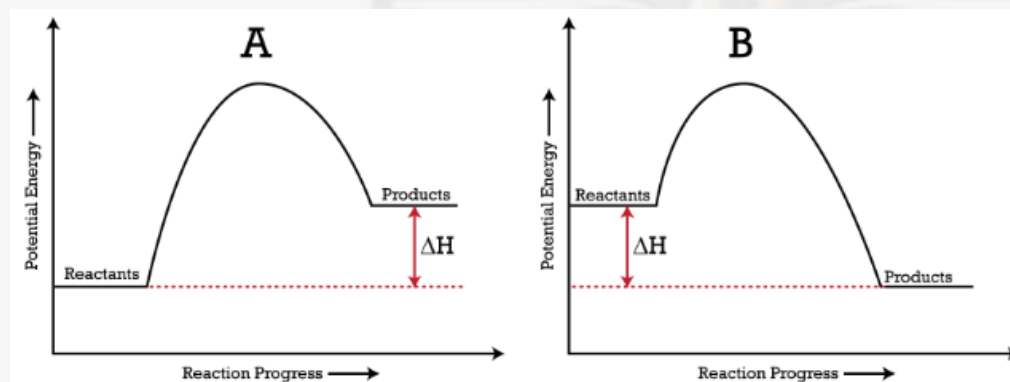
- The activation energy for a reaction is illustrated in the potential energy diagram by the height of the hill between the reactants and the products. For this reason, the activation energy of a reaction is sometimes referred to as the activation energy barrier. Reacting particles must have enough energy so that when they collide, they can overcome that barrier (see figure below).
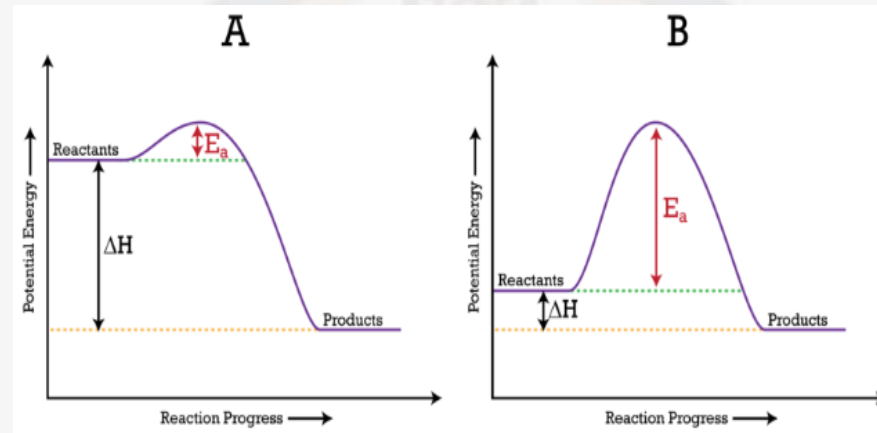
Summary
- A potential energy diagram shows the change in potential energy of a system as reactants are converted into products.

- **Enthalpy of a reaction** is the change in heat energy (ΔH) that occurs when reactants transform into products.
- So, ΔH is positive if heat is absorbed during the reaction and ΔH is negative if heat is emitted during the reaction.
- ΔH value negative : energy released –> exothermic reaction
- ΔH value positive : energy absorbed –> endothermic reaction

$$\Delta H = \sum \Delta H \text{ (bonds broken in reactants) } - \sum \Delta H \text{ (bonds made in products)}$$

Activation energy:
- Is the minimum amount of extra energy required by a reacting molecule to get converted into product.

- It can also be described as the minimum amount of energy needed to activate or energize molecules or atoms so that they can undergo a chemical reaction or transformation.

Its measuring unit is in:
➢ joules (J) or kilojoules per mole (kJ/mol)

Activated complex:

- It is a transitional structure that results from an effective collision and that persists while old bonds are breaking, and new bonds are forming

And it also describing,

- a molecule in an unstable state intermediate to the reactants and the products in the chemical reaction.

- When reactant molecules collide with each other at their highest energy point, an intermediate is formed which remains in equilibrium with the main reactant. If this intermediate complex has an energy equal to or greater than the Threshold Energy, then it will be converted into a product

Entropy(S):

- Entropy is a measure of the disorder in a system and is a thermodynamic property

It is also defined as,

- A process is more likely to occur if it is accompanied by an increase in entropy; that is, ΔS is positive.

- Extensive property which means that it scales with the size or extent of a system.

- Entropy order: gas > liquid > solids
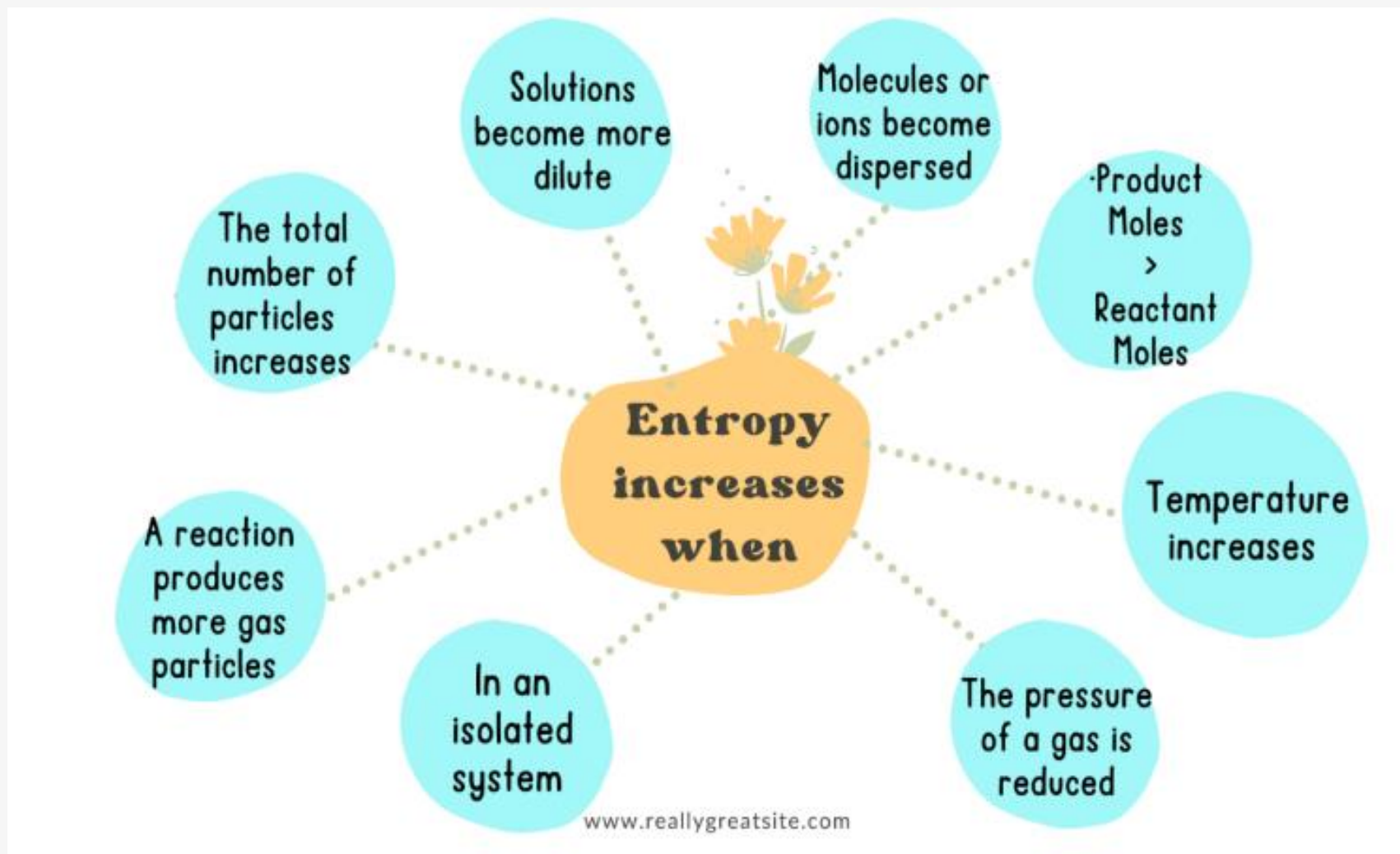
- **From these examples, it is clear that entropy increases with a decrease in regularity.**

Gibbs Free Energy Equation

Gibbs free energy is equal to the enthalpy of the system minus the product of the temperature and entropy. The equation is given as:

$$G = H - TS$$

G = Gibbs free energy

H = enthalpy

T = temperature

S = entropy

Or more completely as:

$$G = U + PV - TS$$

U = internal energy (SI unit: joule)

P = pressure (SI unit: pascal)

V = volume (SI unit: m^3)

T = temperature (SI unit: kelvin)

S = entropy (SI unit: joule/kelvin)

Catalyst:

- Any substance that increases the rate of a reaction without itself being consumed. Enzymes are naturally occurring catalysts responsible for many essential biochemical reactions.

- Most solid catalysts are metals or the oxides, sulfides, and halides of metallic elements and of the semi metallic elements boron, aluminum, and silicon.

- Gaseous and liquid catalysts are commonly used in their pure form or in combination with suitable carriers or solvents; solid catalysts are commonly dispersed in other substances known as catalyst supports.

- **Positive Catalysts**

Catalysts that increase the rate of a chemical reaction are positive catalysts. It increases the rate of reaction by lowering the activation energy barriers such that a large number of reaction molecules are converted into products, and thereby the percentage of yield of products increases.
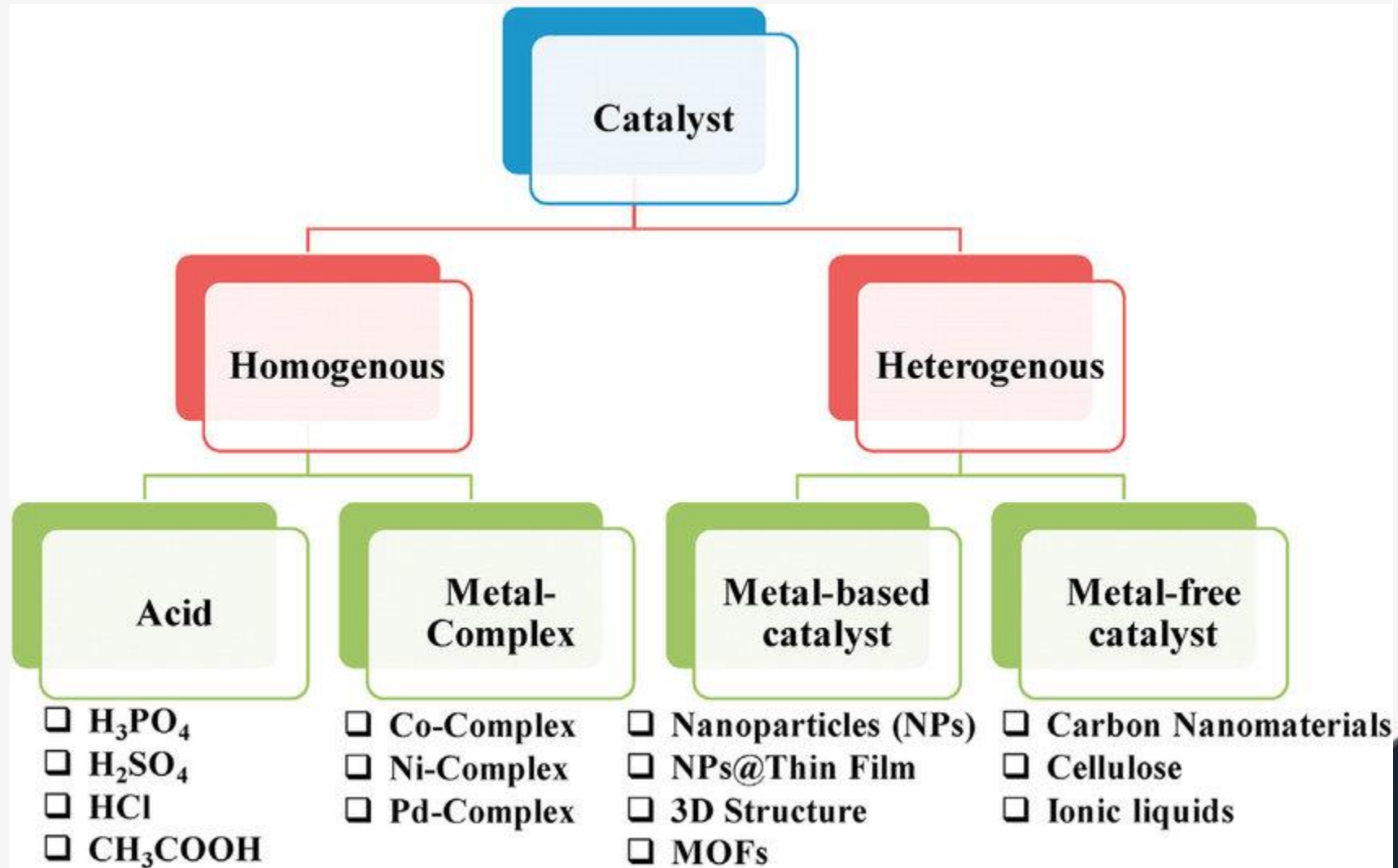
- **Negative Catalysts**

Catalysts that decrease the rate of reaction are negative catalysts. It decreases the rate of reaction by increasing the activation energy barrier, which decreases the number of reactant molecules to transform into products, and hence the rate of reaction decreases.

Types of Catalyst:

Gas:

- Gas is a substance that is in a gaseous or vaporous state of matter. The term gas is also used in reference to the state itself when applied to matter that has the properties of a gaseous substance.
- Gas is one of four natural states of matter, along with liquid, solid and plasma. A gas has no fixed shape or volume.
- A gas uniformly fills any container, is easily compressed, and mixes completely with any other gas. One of the most obvious properties of a gas is that it exerts pressure on its surroundings.
- The force which the substance exerts on another substance per unit area is known as pressure.

Pressure

- Pressure is usually represented by the formula $p = F/A$.
- Here, 'F' is the applied force, and 'A' is the area of the region.

- The CGS unit of pressure is barye (Ba). It is equal to 0.1 Pa

Common units

- 1 torr = 1 mm Hg
- 1 atm =760 torr
- SI unit: pascal 1 atm =101,325 Pa

Barometer

- A barometer is a scientific instrument used to measure atmospheric pressure, also called barometric pressure.
- The atmosphere is the layers of air wrapped around Earth. That air has a weight and presses against everything it touches as gravity pulls it to Earth. Barometers measure this pressure

Gas laws

- The gas laws were developed towards the end of the 18th century by numerous scientists (after whom the individual laws are named).

➢ The five gas laws are listed below:
- **Boyle's Law:** It provides a relationship between the pressure and the volume of a gas.
- **Charles's Law:** It provides a relationship between the volume occupied by a gas and the absolute temperature.
- **Gay-Lussac's Law:** It provides a relationship between the pressure exerted by a gas on the walls of its container and the absolute temperature associated with the gas.
- **Avogadro's Law:** It provides a relationship between the volume occupied by a gas and the amount of gaseous substance.
- **The Combined Gas Law** (or the **Ideal Gas Law**): It can be obtained by combining the four laws listed above.

Boyle's Law

- Boyle's law gives the relationship between the **pressure** of a gas and the **volume** of the gas at a constant temperature. Basically, the volume of a gas is inversely proportional to the pressure of a gas at a constant temperature

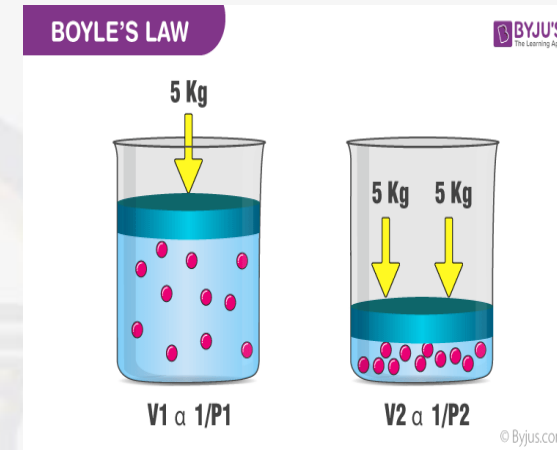Boyle's law equation is written as:

- ❏ $V \propto 1/P$

Or

- ❏ $P \propto 1/V$

Or

- ❏ $PV = k1$

- Where V is the volume of the gas, P is the pressure of the gas, and K1 is the constant.

- It can be used to determine the current pressure or volume of gas and can also be represented as:

$$P1V1 = P2V2$$

Rectangular Hyperbola:

- It is the nature of Boyle's Law's pressure vs volume (P vs V) graph.

Charle's Law:

- Charle's law states that at constant pressure, the **volume** of a gas is **directly** proportional to **the temperature** (in Kelvin) in a closed system.

- Basically, this law describes the relationship between the temperature and volume of the gas.

$V \propto T$

- Where, V = volume of gas, T = temperature of the gas in Kelvin.

- Another form of this equation can be written as,
- V1 / T1 = V2 / T2

Charles's Law

-65°C   250°C
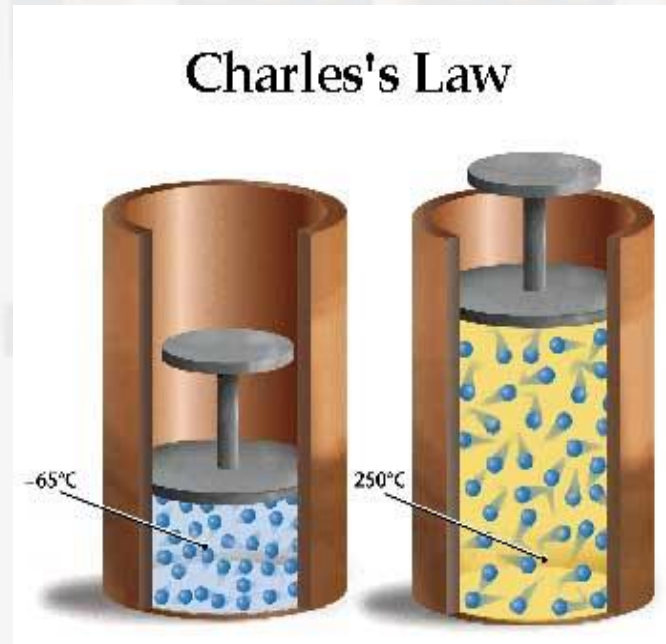
Gay-Lussac Law

- Gay-Lussac law gives the relationship between temperature and pressure at constant volume. The law states that at a constant volume, the **pressure** of the gas is directly proportional to the **temperature** of a given gas
- If you heat up a gas, the molecules will be given more energy; they move faster. If you cool down the molecules, they slow down, and the pressure decreases. The change in temperature and pressure can be calculated using the Gay-Lussac law, and it is mathematically represented as:
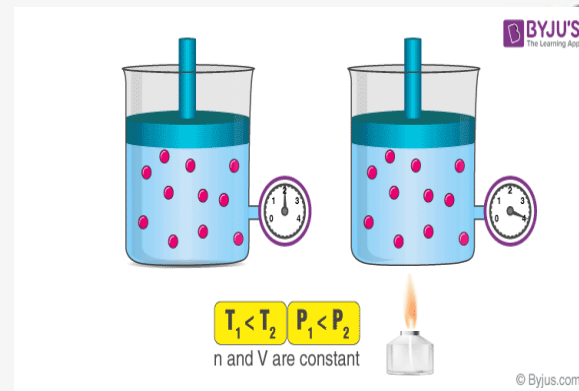 - ❏ $P \propto T$

Or

  - ❏ $P / T = k1$

or

  - ❏ $P1 / T1 = P2 / T2$

Avogadro's law

- Avogadro's law states that if the gas is an ideal gas, the same number of molecules exists in the system. The law also states that equal **volumes** of gases at the same temperature and pressure contain equal **numbers of molecules**. This statement can be mathematically expressed as,
  - ❏ $V / n = $ constant

Or

  - ❏ $V1 / n1 = V2 / n2$

Where V is the volume of an ideal gas and n represents the number of gas molecules.

Ideal Gas

- Ideal gases are also known as perfect gas. It establishes a relationship among the four different gas variables such as pressure (P), Volume (V), Temperature (T) and amount of gas (n).

➢ **Ideal Gas Properties and Characteristics**

- The motion of ideal gas in a straight line is constant and random.
- The gas occupies a very small space because the particle in the gas is minimal.
- There is no force present between the particle of the gas. Particles only collide elastically with the walls of the container and with each other.
- The average kinetic energy of the gas particle is directly proportional to the absolute temperature.
- The gases are made up of many of the same particles (atoms or molecules), which are perfectly hard spheres and very small.
- The actual volume of the gas molecule is considered negligible as compared to the space between them, and because of this reason, they are considered as the point masses.

Absolute zero

- At a temperature of absolute zero there is no motion and no heat. Absolute zero occurs at a temperature of 0 kelvin, or -273.15 degrees Celsius, or at -460 degrees Fahrenheit

Summary

Gas Law	Formula	Description
Charle's Law	$V_1/T_1 = V_2/T_2$	At constant P, as the volume increases, the temperature also increases.
Boyle's Law	$P_1V_1 = P_2V_2$	At constant T, if pressure increases, then volume decreases.
Gay-Lussac Law	$P_1/T_1 = P_2/T_2$	At constant V, as pressure increases, the temperature also increases.
Avogadro's Law	$V / n = $ constant	When the amount of gas increases, the volume of the gas also increases.
Ideal Gas Law	$PV = nRT$	

Kinetic Theory of Gases Postulates

- The kinetic theory of gas postulates is useful in understanding the macroscopic properties from the microscopic properties.

- Gases consist of a large number of tiny particles (atoms and molecules). These particles are extremely small compared to the distance between the particles. The size of the individual particle is considered negligible, and most of the volume occupied by the gas is empty space.

- These molecules are in constant random motion, which results in colliding with each other and with the walls of the container. As the gas molecules collide with the walls of a container, the molecules impart some momentum to the walls. Basically, this results in the production of a force that can be measured. So, if we divide this force by the area, it is defined to be the pressure.

- The collisions between the molecules and the walls are perfectly elastic, which means when the molecules collide, they do not lose kinetic energy. Molecules never slow down and will stay at the same speed.

- The average kinetic energy of the gas particles changes with temperature; i.e., the higher temperature, the higher the average kinetic energy of the gas, resulting in faster molecular motion.

- The molecules do not exert any force of attraction or repulsion on one another except during collisions.

Effusion

- Effusion is the movement of a gas through a tiny hole into a vacuum.
- We want to know the rate of effusion, which is how much gas moves through the hole per unit time.
- We assume that the gas particles don't bump into each other while they move through the hole (this means it's a narrow hole in a thin wall).
- So, the rate of effusion just depends on how often the particles bump the hole. This depends on their density and speed.
- Because at a given temperature, all gases have the same kinetic energy, their speed is inversely proportional to m1/2, the square root of the mass.
- Thus, the relative effusion rates for different gases at the same temperature is

$$\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$
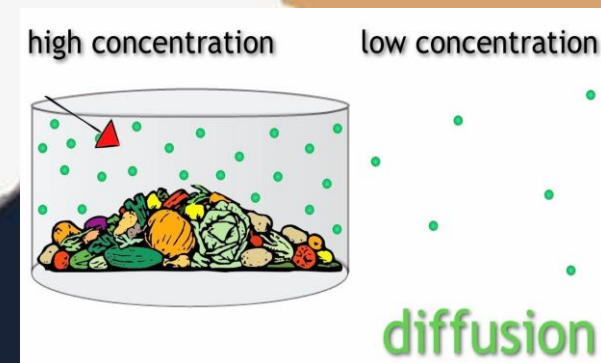
Diffusion

- Diffusion is a more complicated process. is the term used to describe the mixing of gases.
- In diffusion the gases move from the region of high concentration to the region of low concentration. It means the movement of gases through each other or the spreading of one gas through another.
- Because there are many collisions, the gases move much slower than we might expect from the average speeds near 400-700 m/s.
- (This is why it will take a moment to smell perfume when someone walks into a room.)

Technically, many processes that sound like this are not exactly simple diffusion.
- We have to be careful about whether there are pressure differences or flows of gases (like wind).

high concentration   low concentration

diffusion

Graham's law

- Graham's law states that the rate of diffusion or of effusion of a gas is **inversely proportional** to the square root of its molecular weight.

- According to Graham's Law, at constant pressure and temperature, molecules or atoms with lower molecular mass will effuse faster than the higher molecular mass molecules or atoms.
- Thomas even found out the rate at which they escape through diffusion.
- In other words, it states that the rate of effusion of a gas is inversely proportional to the square root of its molecular mass.
- This formula is generally used while comparing the rates of two different gases at equal pressures and temperatures. The formula can be written as

- M1 is the molar mass of gas 1
- M2 is the molar mass of gas 2
- Rate1 is the rate of effusion of the first gas
- Rate2 is the rate of effusion for the second gas

$$\frac{Rate_1}{Rate_2} = \sqrt{\frac{M_2}{M_1}}$$

GRAHAM'S LAW: DIFFUSION AND EFFUSION

BBYJU'S
The Learning App