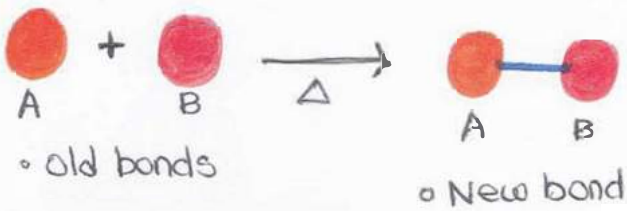
Chemical Reactions and Equ

1. What is chemical reactions?

- Chemical reactions are the processes in which new substance with new properties are form which involve chemical changes.
- It involves breaking of old chemical bonds which exist between the at of substance which act to form new chemical bonds.



- old bonds
- New bonds

- During a chemical reactions, atoms of one elements do not change in those of another elements.

Ex- Hydrogen react with oxygen to form water.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Hydrogen oxygen water

2. What is Reactants?

The substance which takes part in chemical reactions are called reactants

3. What is Products?

The new substance produced as a result of chemical reactions.

- EX- $C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)}$

Reactants Products

∴ Here reactants are transfer into products.

- The products thus formed have properties entirely differents from those of the reactants.

- EXPERIMENTS REGARDING CHEMICAL RNX -
- Magnesium ribbon experiment -

1. Before experiment please note that magnesium is silvery white metal It is available in the form of magnesium ribbion or wire

2. When magnesium ribbon is heated, it burns in air with dazzling white flame to form white powder called magnesium oxide.

3. Actually, on heating, magnesium combines with oxygen present in air to form magnesium oxide.

- Magnesium (As ribbon) + Oxygen (from air) $\xrightarrow{\text{Heat}}$ Magnesium oxide (white powder)

In these magnesium and oxygen are reactant, but only one product MgO_2 (magnesium oxide)

4. The magnesium ribbon which we use usually has a coating of magnesium oxide, on its surface which is formed by slow action of oxygen of air on it.

Que- Why the magnesium ribbon is cleaned by rubbing with sand paper before burning in air? (1 mark)

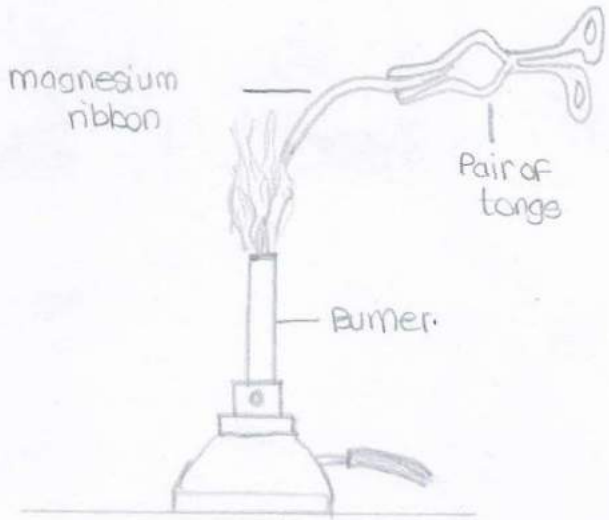
→ • This is done to remove the protective layer of magnesium oxide from the surface of magnesium ribbon before burning in air.

• During burning the dazzling white light given out during the burning of magnesium ribbon is harmful to eyes.

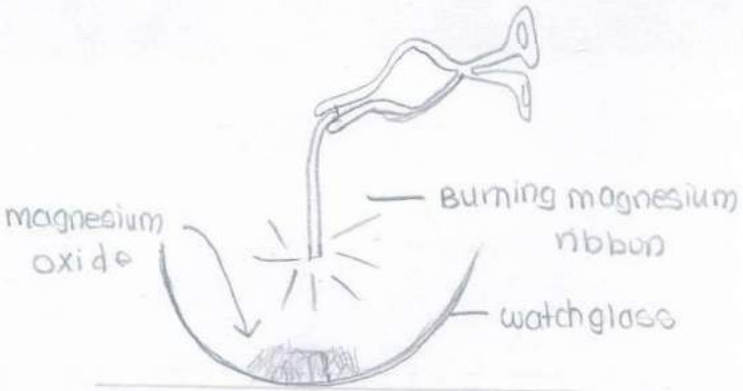
• So, the magnesium ribbon should be burned by keeping it far as possible from the eyes.

EXPERIMENT- (2M|3M)

1. Take about 2cm long magnesium ribbon and clean it by rubbing its surface with sand paper.
2. Hold mg ribbon with a pair of tongs at one end, and heats its others end over a burner.
3. The magnesium ribbon start burning with dazzling white flame.
4. Hold the burning magnesium ribbon over a watch glass so that the magnesium oxide powder calect in the watch glass.



a) magnesium ribbon being heated over a burner



b) magnesium ribbon burns in air to form magnesium oxide.

Example of chemical reaction occure in daily life-

1. Souring of milk
2. Formation of curd from milk.
3. cooking of food.
4. Digestion of food in our body
5. Process of Respiration
6. Rusting of iron
7. Burning of fuels
8. Burning of candle wax

- These all example involve chemical change.

•Characteristics of Chemical Reactions- (2M)

Important characteristics of chemical rx are-

1. Evolution of gas.
2. Formation of a precipitate
3. change in colour
4. change in temperature
5. change in state

1. Evolution of Gas-

Some chemical reactions are characteristics of some substance that react and evolve gas with some product.

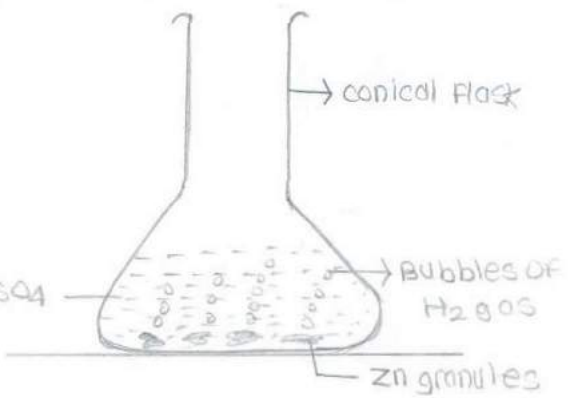
•Example- When zinc granules react with Dilute sulphuric acid, then the bubbles of hydrogen gas produce.

•EXPERIMENT-

1. Take some zinc granule in conical flask, and add zn over dilute sulphuric acid or dilute HCl.
2. we will see the bubbles of hydrogen gas formed around zn granules.
3. By touching flask, we can find it is somewhat hot, here change in temperature also occurs in chemical reaction.

- $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow$

zinc — dilute sulphuric acid — zinc sulphate — Hydrogen gas



•Example- 2 - when magnesium react with dilute acid (H_2SO_4 or HCl), then hydrogen gas evolved.

- $Mg + HCl \longrightarrow MgCl + H_2$ (evolve)

Example - 3 when sodium carbonate and dilute Hydrochloric acid is characteris by evolution of CO_2 gas.

- $Na_2CO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$

- sodium carbonate — sodium chloride — carbondioxide

- What is Precipitate?
- Precipitate is a 'solid state' which seperates out the from the solution during a chemical reaction.
- Precipitate usually seen under testtube at bottom part.
- It can form by passing a gas in aqueous solution of a substance (like passing carbon dioxide gas into lime water)
- Formation of insoluble solid precipitate is called Precipitation.

2. Formation of Precipitate–

Example - When potassium iodide solution is added to solution of lead nitrate, then yellow precipitate of lead iodide is formed.

Experiment – we can carry out this chemical reaction as follows -

1. Take some lead nitrate solution in a test tube
2. Add potassium iodide solution to it.
3. A yellow precipitate of lead iodide is formed at once.
4. A change in colour (colourless to yellow) also takes in this chemical reactions.

- $Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \longrightarrow PbI_{2(s)} + 2KNO_{3(aq)}$

lead nitrate | Potassium iodide | Lead iodide (Yellow) | Potassium nitrate

Example2 - When sulphuric acid and Barium chloride solution is formed by a white precipitate of Barium sulphate.

- $H_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2HCl$

sulphuric acid | Barium chloride | Barium sulphate (white) | Hydrochloric acid.

3. Change in Colour–

Example - When citric acid reacts with potassium permaganate solution, then the purple colour of potassium permanganate solution, disappears (it becomes colourless).

- Purple $\xrightarrow{\Delta}$ Colourless $3KMnO_4 + C_6H_8O_7 \xrightarrow{\Delta} 3HMnO_4 + C_6H_5K_3O_7$

(Purple) | citric acid | colourless

Experiment - we can carry out this reaction as follows–

1. Take some dilute potassium permanganate solution in test tube. It has purple colour.
2. Add some lemon juice with help of dropper and shake it.
3. The purple colour of potassium permanganate solution goes on fading and ultimately it becomes colourless.

Example 2 - when sulphur dioxide gas is passed through acidified potassium dichromoate solution by change in colour from orange to green.

$$SO_2 + \underset{\text{orange}}{K_2Cr_2O_7} + 3H_2SO_4 \longrightarrow K_2SO_4 + \underset{\text{green}}{Cr_2(SO_4)_3} + 3H_2O$$

- **Change in Temperature** - Exothermic reaction.

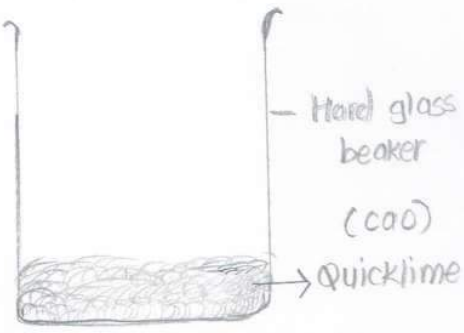
Example - When quicklime react with water, then slakedlime is formed and lot of heat energy is produced by a change in temperature. It is a exothermic reaction which means heat producing reaction.

Experiment -

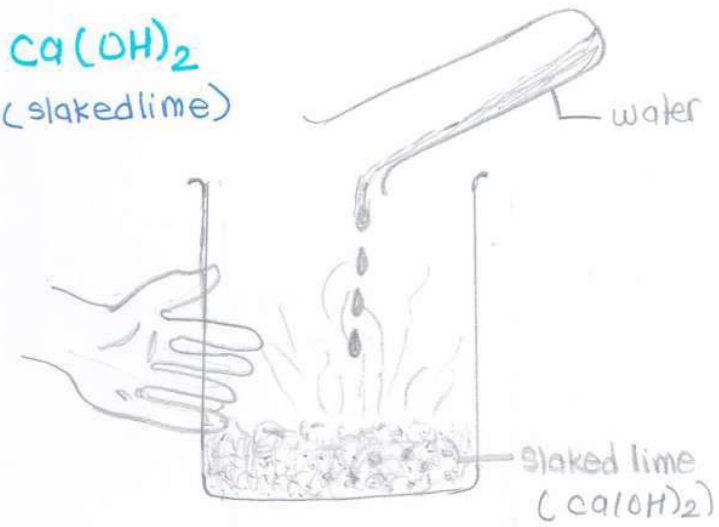
1. Take little of quicklime in hard glass beaker and add water to it slowly.
2. Touch the beaker carefully, the beaker feels to be quite hot its temp is high.

- Calcium oxide + water ⟶ calcium hydroxide
- $CaO + H_2O \longrightarrow Ca(OH)_2$

(quicklime) (slakedlime)



a. Quicklime reacts with water to form slaked lime release a lot of heat



b. The beaker becomes hot It temp. rises

Example 2 - The carbon burns in the air to form carbondioxide which also releases a large amount of heat.

$$\underset{\text{carbon}}{C} + \underset{\text{oxygen}}{O_2} \xrightarrow{\Delta} \underset{\text{carbondioxide}}{CO_2}$$

2. Endothermic reaction - The reaction which absorbs the heat are called as Endothermic reaction.

Example - The reaction between barium hydroxide and ammonium chloride to form Barium chloride, ammonia and water by change in temperature.

$$\underset{\text{Barium hydroxide}}{Ba(OH)_2} + \underset{\text{ammonium chloride}}{2NH_4Cl} \xrightarrow{\Delta} \underset{\text{Barium chloride}}{BaCl_2} + \underset{\text{ammonium hydroxide}}{2NH_4OH} \downarrow \text{(absorb) heat}$$

5. Change in state-

Example - When wax is burned, then water and carbondioxide are formed

- Now, wax is solid, water is a liquid whereas carbondioxide is a gas.
- This means that during the combustion reaction of wax, the physical state changes from solid to liquid and gas.
- The combustion of candle wax is characterised by change in state.

• CHEMICAL EQUATIONS-

- The method of representing a chemical reaction with help of symbol and formula of the substance involve in it known as chemical Equations.
- For example- zinc metal react with dilute sulphuric acid to form zinc sulphate and hydrogen gas. This rnx can be written as-

$$\text{Zinc} + \text{Sulphuric acid} \longrightarrow \text{Zinc sulphate} + \text{Hydrogen}$$

- This is known as the word equation.
- we can change the word equation into symbol and formula of various substance.
- Putting the symbol and formulae of all substance in above word equation

$$\underbrace{Zn + H_2SO_4}_{\text{Reactants}} \longrightarrow \underbrace{ZnSO_4 + H_2}_{\text{Products}}$$

- Here, zinc and sulphuric acid are reactant. and reactant are always written on the left hand side in an equation with plus(+) sign between them.
- Here, zinc sulphate and hydrogen are products, and it is always written on the right hand side in an equation with plus sign between them.
- The arrow sign (⟶) pointing towards the right hand side is put between the reactant and products.
- Arrow also indicate that the substance written on the left hand side are combining to give substances written on the right hand side in the equation.
- Chemical equation is a shorthand method of representing a chemical rnx.

Balanced and Unbalanced chemical Equations -

1. A balanced chemical equation has an equal number of atoms of different elements in the reactant and products.

Example - $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$

- In other words, a balanced equation has equal number of atoms of diff. elements on both the side must be equal.
- Now,
- Let us count the number of atoms of all the elements in the reactant and product sepretely.

	In reactant	In product
No. of Zn atoms -	1	1
No. of H atoms -	2	2
No. of S atoms -	1	1
No. of O atoms -	4	4

- Since, above there is an equal number of atoms of different elements in reactant and products is equal, so above equation is balanced equation.

2. A unbalanced chemical equation has unequal number of atoms of one or more elements in the reactant and product.

- Example - Hydrogen reacts with oxygen to form water, this point will written as -

- $H_2 + O_2 \longrightarrow H_2O$

- Let us count the number of hydrogen atoms and oxygen atoms in the reactant and product.

	In reactant	In Product
- No. of H atoms -	2	2
- No. of O atoms -	2	1

- In these equation thus, the no. of hydrogen atoms are equal, but the oxygen atoms are not equal on both side of reactant and product. So it is called unequal or unbalanced chemical equation.

- Reactant = Product (always)

- Note - The chemical equation should be balanced because it satisfy the law of conservation of mass in a reaction..

- Some rules about the balancing the equation -
- We should never change the formula of an element or a compound to balance the equation.

2. We can only multiply a symbol or formula by figures like 2, 3 and 4
3. It will be good to note here that here the elements which exist as diamatic molecules are oxygen O_2, hydrogen H_2, nitrogen N_2, Fluorine F_2, chlorine Cl_2, bromine Br_2 and iodine I_2.
4. All other element are usually considered monoatomic in equation writing and represented by their symbols.

Balancing of Chemical Equation-

- The process of making the number of different types of atoms equal on both the side of the equation called balancing of equation.
- The simple equations are balanced by hit and trial method.
- Example of the balancing the equations -
- Hydrogen burns in oxygen to form water. The reaction can be written in an equation.

$$H_2 + O_2 \longrightarrow H_2O$$

- count the number of hydrogen and oxygen atoms in reactant and product.

	In reactant	In Product
No. of H atoms -	2	2
No. of O atoms-	2	1

- The no. of H_2 atoms is equal on both side, but the number of oxygen atoms are unequal.
- There are 2 oxygen atoms on left side and 1 oxygen atoms on right side to have 2 oxygen atoms on the right side, we multiply H_2O by 2 and write $2H_2O$ so that:

$$H_2 + O_2 \longrightarrow 2H_2O$$

- Let us count the number of various atoms on both sides -

	In reactant	In product
No. of H atoms -	2	4
No. of O atoms-	2	2

- Though the number of oxygen atoms has become equal, but the no. of hydrogen atoms has become unequal.
- There are 2 atoms hydrogen atoms on the left side but 4 hydrogen atoms on the right side., we multiply it H_2 by 2 and write $2H_2$:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

- Let us count the no. of atoms on both side-

No. of H atoms -	4 (reactant)	4 (Product)
No. of O atoms -	2	2

- This chemical equations contains equal no. of atoms of hydrogen and oxygen on both the sides. so this is balanced equations.
- By these way we can balanced the equations.

• To make Equation More Informative -

- The chemical equations can be more informative - by three ways -

1. By indicating the "physical change or state" of the reactant and product.
2. By indicating the "heat change" taking place in the reaction.
3. By indicating the "conditions" under which the reaction takes place.

1. To indicate the physical change or state of the reactant and product -

The physical state for reactant and product of chemical reaction -

- solid • liquid • aqueous solution • gas
- solid state is indicated as - symbol (s).
- liquid state is indicated as - symbol (l)
- Aqueous state or solution is indicated as - symbol (aq).
- Gaseous state is indicated as - symbol (g).
- The physical state of the reactant and product are shown by putting the above "state symbol" just after their symbol or formulae in an eq. -

For Example -

zinc react with dilute sulphuric acid to form zinc sulphate solution and hydrogen gas.

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

Here,

- zinc metal is solid, so we write Zn(s)
- Dilute sulphuric acid is aqueous solution, so we can write $H_2SO_4(aq)$
- zinc sulphate is also an aqueous solution, so we can write $ZnSO_4(aq)$
- Hydrogen is gas, so we write hydrogen (g).
- The above equation written as -

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

- This equation is more informative because it tell us the physical state of the various substance involved in it.
- In same case 2 - Insoluble product called precipitate is formed by the mx. between solutions of reactants. and its is indicated by symbol "ppt".
- When calcium hydroxide solution (lime water) react with carbon-dioxide gas, a white precipitate of calcium carbonate is formed along with water.

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

Calcium hydroxide — carbondioxide — calcium carbonate (white ppt) — water

2. To indicate the Heat change in an Equation -

There are two types of reaction on the basis of heat changes involve-

1. Exothermic reaction
2. Endothermic reactions.

1. Exothermic Reactions-

Those reactions in which heat is evolved called as Exothermic reactions.

For Example-1. carbon burns in oxygen to form carbondioxide, a lot of heat is produced in this reaction.

$$C(s) + O_2(g) \longrightarrow CO_2(g) + \text{Heat}$$

carbon — oxygen — carbondioxide

The burning of carbon in oxygen is an exothermic reaction because heat is evolved in this reaction.

An exothermic reaction is indicated by writing + Heat or + Heat energy or just " + Energy on the products side.

Example 2 -

Que - why Burning of natural gas is an exothermic reaction? (2M)

- Natural gas is mainly methane (CH_4), when natural gas burns in the oxygen of air, it forms carbondioxide and water vapour. a large amount of heat is produced.
- Heat is mainly produce so called Exothermic reaction.
- All the combustion reaction are exothermic reaction.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) + \text{Heat energy}$$

Methane — Oxygen — carbondioxide — water

Example 3 -

Que - why respiration is an exothermic reaction? Explain -

- It is exothermic reaction because energy is produced during this process.
- During digestion food is broken down into simpler substances.
- The carbohydrate which we get from the foods like chappati, bread rice and potato we eat broke down into simple form called glucose.
- This glucose then undergoes slow combustion by combining with oxygen in cells of our body to produce energy in a process called Respiration. these energy maintain our body heat.

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l) + \text{Energy}$$

Glucose — Oxygen — carbondioxide — water

- It is also an combustion reaction.

2. Endothermic Reactions -

- Those reaction in which heat is absorbed are known as Endothermic reactions.
- Example - When nitrogen and oxygen are heated to a very high temperature (of about 3000°c) they combine to form nitrogen monoxide and lot of heat is absorbed. in this reaction.

$$N_2(g) + O_2(g) + \text{Heat} \longrightarrow 2NO(g)$$

Nitrogen oxygen Nitrogen monoxide

- An endothermic reaction is usually indicated by writing + Heat or + Heat energy on the reactant side of an equation.
- It is used inside the engine of motor vechiles.

- All the decomposition reaction are endothermic reaction because - it require energy (in the form of heat, light or electricity) to take place.
- For example - The decomposition of calcium carbonate is endothermic reaction.
- When calcium carbonate is heated, it decomposes to form calcium carbonate and carbondioxide.

$$CaCO_3(s) + \text{Heat} \longrightarrow CaO(s) + CO_2(g)$$

- calcium carbonate - Calcium oxide - carbondioxide

- Example -

Que - Why Photosynthesis is an endothermic reaction?

It is endothermic reaction because sunlight energy is absorbed during the process of photosynthesis by green plants.

- Example -

The electrolysis of water to form hydrogen and oxygen is an also endothermic reaction. this is because energy is absorbed during the reaction.

1. $6CO_2(g) + 12H_2O(l) \xrightarrow[\text{Plant enzyme}]{\text{light}} \underset{\text{glucose}}{C_6H_{12}O_6(aq)} + 6O_2(g) + 6H_2O(l)$

2. - Cathode - Reduction

$$2H^+_{(aq)} + 2e^- \longrightarrow H_2(g)$$

- Electrolysis of water

- Anode - Oxidation

$$2OH^-_{(aq)} \longrightarrow H_2O_{(l)} + \tfrac{1}{2}O_{2\,(g)} + 2e^-$$

3. To indicate the conditions under which the mx takes place-

- If heat is require for a mx. to take place, then the heat sign delta (Δ) is put over the arrow of the mx. of equations.
- If the reaction takes place in the presence of a catalyst, then the symbol or formula of catalyst is also written above or below the arrow sign in the equation.

For example –

- $2KClO_3(s) \xrightarrow[MnO_2]{\Delta} 2KCl(s) + 3O_2(g)$

Potassium chlorate → Potassium chloride + Oxygen

- Here, delta (Δ) stands for heat and MnO_2 is the catalyst.

2. The conditions of temperature and pressure at which reaction takes place can also be indicated on the above or below the arrow sign in the equation.

- For example – Methanoal or Methyl alcohol is manufactured from carbon monoxide and hydrogen.

- $CO(g) + 2H_2(g) \xrightarrow[ZnO + CrO_3]{300atm\ 300^\circ C} CH_3OH(l)$

Carbon monoxide + hydrogen → Methylalcohol

- Here 300 atm is pressure, $ZnO + CrO_3$ is catalyst, and 300°C is temp.

- **STEPS FOR WRITING EQUATIONS FOR CHEMICAL REACTION –**
- Mainly four steps are for writing equation in chemical reaction :-
- Step 1 - Write the chemical reaction in the form of word equation, keeping the reactant on left side and product on right side.
- Step 2 – Put the symbols and formula of all the reactant and product in the word equation.
- Step 3 – Balance the equation by multiplying the symbols and formulae by the smallest possible figure. (Do not change the formulae to balance the equation.
- Step 4 - If possible, make the equation more informative by indicating the physical states of reactant and products, by indicating the heat changes, if any, taking place in the reaction, and by indicating the conditions under which the reaction takes places. If however, you do not have sufficient information regarding the physical state, heat changes and conditions of the reaction, this step can be avoided.

Sample Problem - 1

- Write a balanced equation for the following reaction -
- Methane burns in oxygen to form carbon dioxide and water

→ Reaction in form of word equation -

- Methane + oxygen ⟶ carbondioxide + water

- Writing the formulae of all substance given above -

- $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

- Let us count the no. of various atom on reactant and product side-

	In reactant	In Products
1. No. of C atoms -	1	1
2. No. of H atoms -	4	2
3. No. of O atoms -	2	3

- The no. of carbon atoms are equal on both sides, but H and O atoms are unequal on both sides.
- To have 4 hydrogen atoms on right side, we multiply H_2O by 2 and write $2H_2O$, Thus

- $CH_4 + O_2 \longrightarrow CO_2 + 2H_2O$

- Counting the no. of atom on both sides again -

	In reactant	In products
1. No. of C atoms -	1	1
2. No. of H atoms -	4	4
3. No. of O atoms -	2	4

- only the number of oxygen atoms is unequal now. To have 4 oxygen atoms on the left side, we multiply O_2 by 2 and write $2O_2$.

- $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

- Let us count the number of various atoms on the both sides -

	In reactant	In Products
1. No. of C atoms -	1	1
2. No. of H atoms -	4	4
3. No. of O atoms -	4	4

- This chemical equation contains an equal numbers of various types of atoms in the reactant and product so this is a Balanced equation.

- This are the steps to balance the equations →
- It is not mandatory to write whole steps we can do as directed.

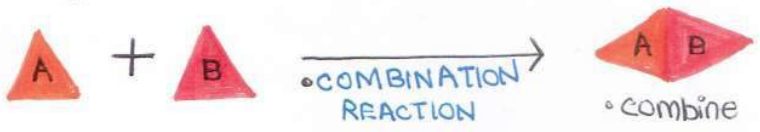
Oxides of Iron metal-

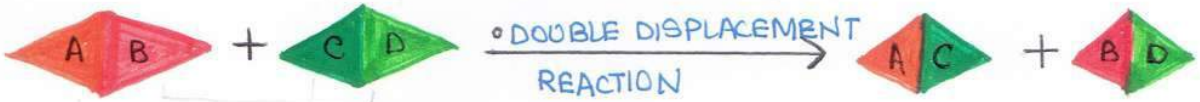
- Mainly two oxides of Iron -

1. Iron (II) oxide – FeO, it is the valency of iron in it is II (two). The common name is ferrous oxide (FeO).
2. Iron (III) oxide – Fe_2O_3, it is the valency of iron – III (three). The common name is ferric oxide.
3. Fe_3O_4 – It is the mixture of iron (II) oxide and iron (III) oxide. ($Fe_3O_4 = FeO + Fe_2O_3$). The common name is magnetic iron oxide.

• TYPES OF CHEMICAL REACTIONS-

- 4 main types of chemical reactions –





1. COMBINATION REACTION –

- Those reaction in which two or more substance combine to form a single substance.
- some examples of reaction –

1. Magnesium and oxygen combine, when heated to form magnesium oxide.

$$2Mg(s) + O_2(g) \xrightarrow{\text{combustion by combination}} 2MgO(s)$$

Magnesium oxygen Magnesium oxide

2. sodium metal burns in medium of chloride to form sodiumchloride.

$$2Na(s) + Cl_2(g) \xrightarrow{\text{combination}} 2NaCl(s)$$

sodium chlorine sodium chloride

3. Carbon (coal) burns in air to form carbon dioxide.

$$C(s) + O_2(g) \xrightarrow{\text{combination}} CO_2(g)$$

carbon air carbon dioxide

4. Ammonia reacts with hydrogen chloride to form ammonium chloride -

- $NH_3(g) + HCl(g) \xrightarrow{\text{combination}} NH_4Cl(s)$

Ammonia — hydrogen chloride — Ammonium chloride

5. Sulphur dioxide reacts with oxygen to produce sulphur trioxide -

- $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$

Sulphur dioxide — oxygen — sulphur trioxide

6. The calcium hydroxide solution, when applied to the walls, reacts slowly with the carbondioxide gas present in air to form a thin, shining layer of calcium carbonate on the walls of the house:

- $Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$

calcium hydroxide — carbondioxide — calcium carbonate — water

This process gives a white, shiny apperance to the walls of a house, it is called white-washing. The $CaCO_3$ is actually formed after two to three days of white-washing and gives a shiny finish to the walls.

2. DECOMPOSITION REACTIONS -

- Those reaction in which a compound splite up into two or more simpler substance are known as decomposition reaction.
- It is just the opposite of a combination reaction.
- Some examples of decomposition reactions -

1. When calcium carbonate is heated, it decomposes to give calcium oxide and carbondioxide.

- $CaCO_3(s) \xrightarrow[\text{decomp.}]{\text{Heat}} CaO(s) + CO_2(g)$

calcium carbonate (limestone) — calciumoxide (lime) — carbondioxide

- Calcium oxide (or lime) is used on a large scale in the manufacture of cement and glass.
- When a decomposition reaction is carried out by heating, it is called 'thermal decomposition'. (thermal means releasing of heat)

2. When potassium chlorate is heated in the presence of manganese dioxide catalyst, it decomposes to give potassium chloride and oxygen.

$$2KClO_3(s) \xrightarrow[\text{decomp.}]{\text{Heat}} 2KCl(s) + 3O_2(g)$$

Potassium chlorate — Potassium chloride — oxygen

- This decomposition process is used for preparing oxygen gas in laboratory.

Example 3- When ferrous sulphate is heated strongly, it decomposes to form ferric oxide, sulphur dioxide and sulphur trioxide.

- $2FeSO_4(s) \xrightarrow[\text{Decomposition}]{\text{Heat}} Fe_2O_3(s) + SO_2(g) + SO_3(g)$

Ferrous sulphate (Green colour) | Ferric oxide (Brown colour) | Sulphur dioxide | Sulphur trioxide

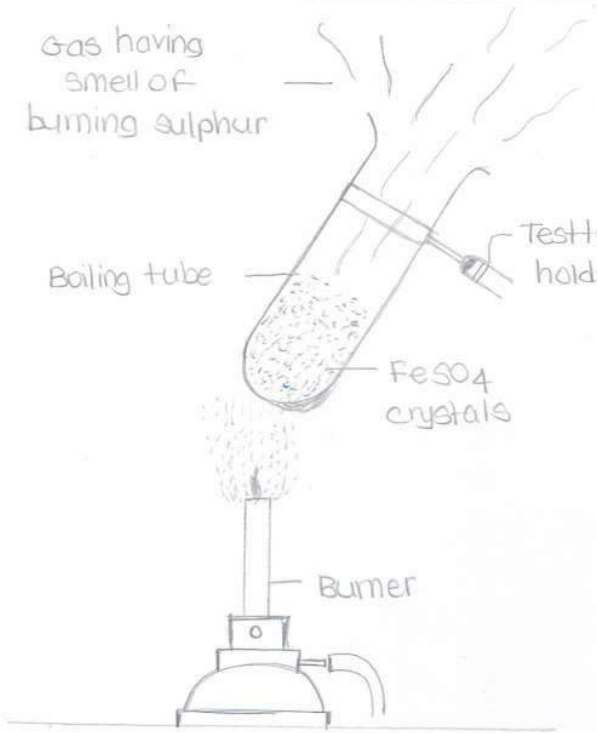
- Note - The Ferrous sulphate is also known as iron(II) sulphate and Ferric oxide is also known as iron(III) oxide.
- The ferrous sulphate crystals which are available in ferrous sulphate heptahydrate $FeSO_4.7H_2O$. They contain 7 molecules of water of crystallation These crystals are green in colour.
- When green colour ferrous sulphate heptahydrate crystals ($FeSO_4.7H_2O$) are heated, they lose 7 molecules of water of crystalliation to form anhydrous ferrous sulphate ($FeSO_4$) which is white in colour.
- Decomposition of Ferrous sulphate in the labroratory -
- Take 2 g of $2FeSO_4$ crystals in dry boiling tube. The ferrous sulphate crystals are green in colour.
- Heat the boiling tube over a burner
- The green colour of ferrous sulphate crystals first change into white and then brown solid is formed (which is ferric oxide).
- Gas having the smell of burning sulphur comes out of the boiling tube.

Example 4- When lead nitrate is heated strongly, it breaks down to form lead monoxide, nitrogen dioxide and oxygen -

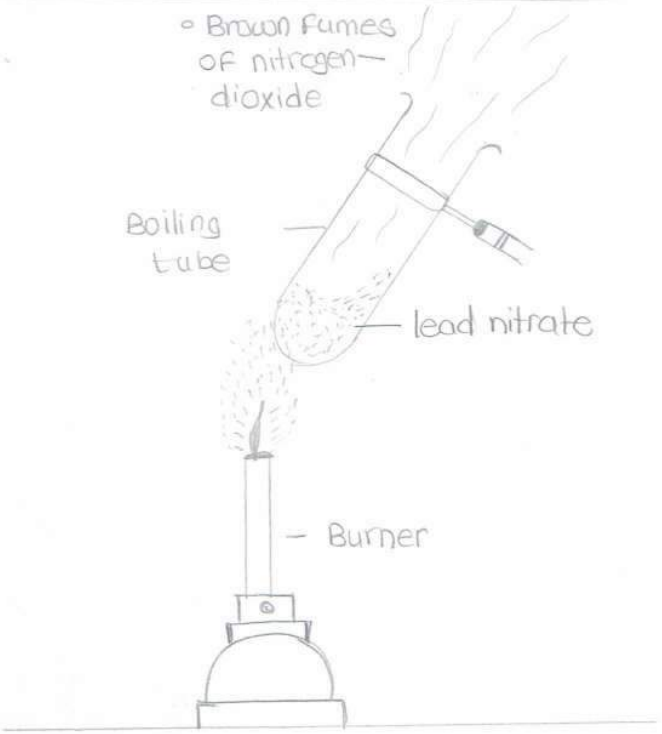
- $2Pb(NO_3)_2(s) \xrightarrow[\text{Decomposition}]{\text{Heat}} 2PbO(s) + 4NO_2(g) + O_2(g)$

- Lead nitrate (colourless) | Lead mono oxide (yellow) | Nitrogen dioxide (Brown fumes) | oxygen

- The decomposition of lead nitrate is brought about by heat, it is actually an example of thermal decomposition.
- Decomposition of lead nitrate in the labroratory -
- Take about 2g of lead nitrate powder in boiling tube. lead nitrate is a colourless compound.
- Hold the boiling tube in the hand test tube holder and it should be heat over burner.
- Brown fumes of nitrogen dioxide gas are evolved which fill the boiling tube
- If a glowing splinter is held over the mouth of the boiling tube, it catches fire and starts burning again. this shows that oxygen gas is also evolved during this reaction.
- A yellow solid is left behind in the boiling tube. This is lead monoxide. (lead monoxide is reddish-brown when hot but yellow when cold).



a. Decomposition of $FeSO_4$ crystals.



b. Decomposition of lead nitrate.

- Decomposition of water which are carried out by electricty? (3M)
- Example- When electric current is passed through acidified water, it decomposes to give hydrogen gas and oxygen gas.

$$2H_2O(l) \xrightarrow[\text{Decomposition}]{\text{Electricity}} 2H_2(g) + O_2(g)$$

water → Hydrogen + oxygen

- This is called Electrolysis of water.
- EXPERIMENT-

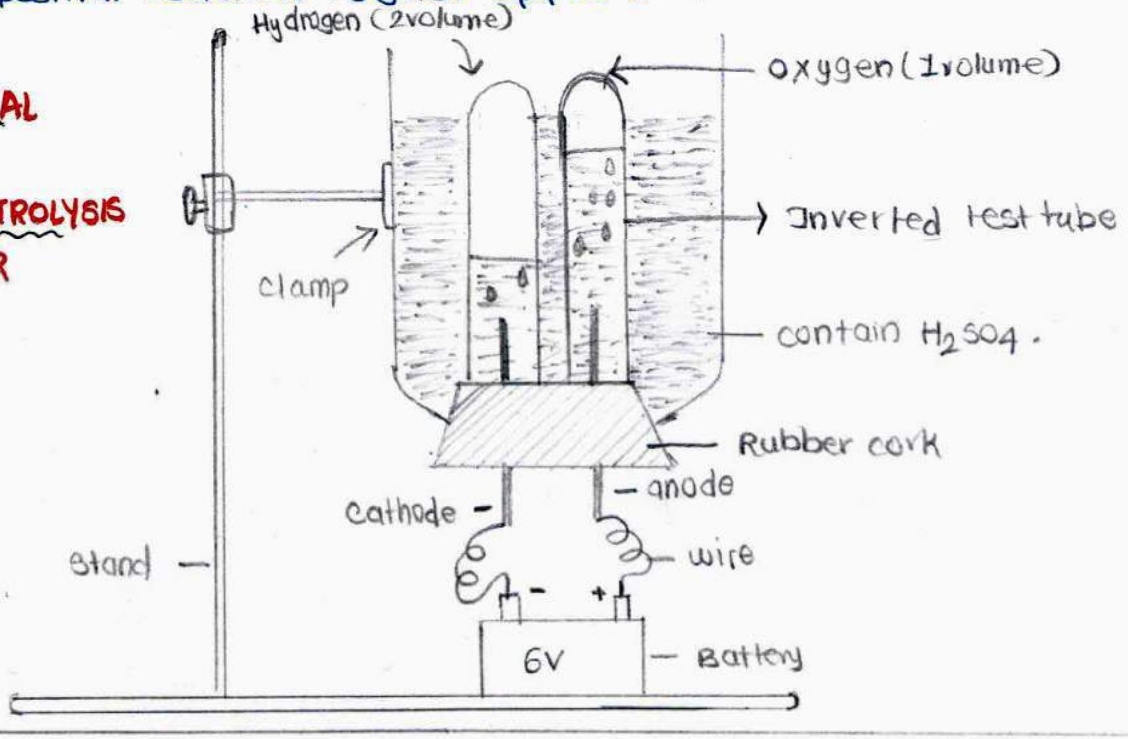
1. Take a wide-mouthed glass bottle. Fix with a inverted position on a stand.
2. A rubber cork having two holes is fitted in the neck of the bottle. Two carbon rods are fixed in two holes of cork tightly.
3. Fill the glass bottle two third with water. Add a few drops of dilute H_2SO_4 to water.
4. Two similar test tube filled with water inverted over the two carbon electrodes by keeping thumb over their mouth.
5. Connect the outer ends of carbon rods to the two terminals of 6 volt battery by wires having a switch.
6. Negative terminal of battery is called cathode (negative terminal). The right side carbon connected to the positive terminal of battery called anode (positive electrode).
7. Pass the electric current through water by turning on the switch and leave the apparatus undistrubed for some time.

8. We can see the bubbles of gases being formed at the both carbon electrodes inside test tube containing water
9. These gases are formed by decomposition of water on passing electricity.
10. The gases formed at two electrode go on collecting in the top parts of inverted test tubes.
11. The volume of gases are not same in both test tube. The volume of gas collected on negative electrode is double the volume of gas collected on positive electrode.
12. Keep on passing electric current till both the test tube completely filled with gases. Then remove the gas - filled test tube and test them one by one by bringing a burning candle close to the mouth of each test tube.
13. Bringing candle close to mouth burns rapidly and make "popupsound".
14. The gas collected in left test tube over negative electrode is hydrogen. and gas near the mouth of right test tube the candle burns brightly
15. We know that oxygen gas makes things burn brightly. so, the gas collected in the right side test tube over positive electrode is oxygen.
16. Since the electrolysis of water produces 2 volumes of hydrogen gas and 1 volume of oxygen gas we conclude that the ratio of hydrogen gas and oxygen gas is 2:1 by volume.
- It shows that water is a compound made up of 2 parts of hydrogen gas and 1 part of oxygen gas by volume.

Que- Why decomposition reaction is just opposite of combination reaction? (2M)
- In a combination reaction two or more substances are combined to form new substance
- In a decomposition reaction one substance decomposes into two or more substance

Hence, decomposition reaction is just opposite of combination rx.

- **EXPERIMENTAL SET-UP FOR THE ELECTROLYSIS OF WATER**



Example - When electric current is passed through molten sodium chloride, it decomposes to give sodium metal and chlorine gas.

$$2NaCl(l) \xrightarrow[\text{Decomposition}]{\text{Electricity}} 2Na(s) + Cl_2(g)$$

Sodium chloride → Sodium; Chlorine gas

- This decomposition reaction is used to obtain sodium metal from sodium chloride. It is called Electrolysis of molten sodium chloride. Sodium chloride is in the form of molten state.

- Decomposition reaction which brought out by light energy?

Example - When silver chloride is exposed to light, it decomposes to form silver metal and chlorine gas.

$$2AgCl(s) \xrightarrow[\text{Decomposition}]{\text{light}} 2Ag(s) + Cl_2(g)$$

Silver chloride (white) → Silver (Greyish white) + Chlorine (Yellowish green)

- Decomposition of 2AgCl by light

Example - When silver bromide is exposed to light, it decomposes to form silver metal and bromine vapours.

$$2AgBr(s) \xrightarrow[\text{Decomposition}]{\text{Light}} 2Ag(s) + Br_2(g)$$

Silver bromide (Pale yellow) → Silver (Greyish white) + Bromine (Red brown)

- It is also used in black and white photography.

- **Uses of Decomposition Reactions -**

1. The decomposition reaction carried out by electricity are used to extract several metals from their naturally occuring compounds like bromide and oxides.
2. When the fused metal chloride or metal oxide is decomposed by passing electricity, then metal is extracted at the cathode.

- **Decomposition reaction in our body -**

- The digestion of food in the body is an example of decomposition reaction.
- When we eat food like wheat, rice or potato, the strach present in them decomposes to give simple sugars like glucose in the body and the proteins decomposes to form amino acid.

- Strach $\xrightarrow[\text{Decom.}]{}$ Glucose $\xrightarrow[\text{Decomp.}]{}$ Proteins $\xrightarrow[\text{Decomp.}]{}$ Amino acid

3. DISPLACEMENT REACTIONS -

- Those reactions in which one element takes place of another element in a comp. are known as displacement reactions.
- A more reactive elements displaces a less reactive element from its compound.
- It is also called as single displacement reaction.

- Some examples of displacement reactions –
- Example 1 – When a strip of zinc metal is placed in copper sulphate solution, then zinc sulphate solution and copper are obtained.

$CuSO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + Cu(s)$

copper sulphate	zinc	zinc sulphate	copper
(Blue solution)	(silvery white)	(colourless solution)	(Red-brown)

- This displacement reaction takes place because zinc is more reactive than copper.

- Example 2 – When a piece of ~~pap~~ magnesium metal is placed in copper sulphate solution, then magnesium sulphate solution and copper metal are formed.

$CuSO_4(aq) + Mg(s) \longrightarrow MgSO_4(aq) + Cu(s)$

copper sulphate	magnesium	Magnesium sulphate	copper
Blue solution	silvery white	colourless solution	Red-brown

- Magnesium is able to displace copper from copper sulphate solution because magnesium is more reactive than copper.

- Example 3 – When a piece of iron metal is placed in a copper sulphate solution, then iron sulphate solution and copper metal are formed.

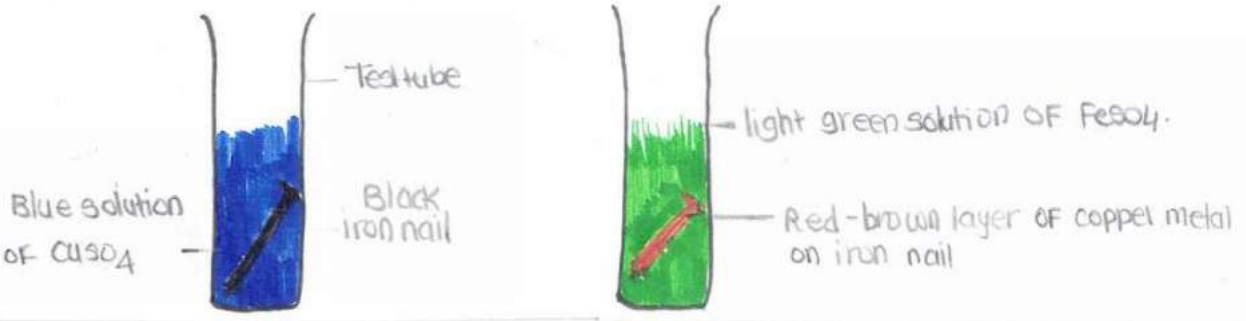
$CuSO_4(aq) + Fe(s) \longrightarrow FeSO_4(aq) + Cu(s)$

copper sulphate	iron	Iron sulphate	copper
(Blue solution)	(Grey)	(Greenish sol)	(Red-brown)

- This is displacement reaction occurs because iron is more reactive than copper.

- We can perform the displacement reaction between iron and copper sulphate solution as follows –

1. Take about 10 mL of copper sulphate solution in a test tube. It is deep blue in colour.
2. Take a big iron nail and clean its surface by rubbing with a sand paper.
3. Put the cleaned iron nail in the test tube containing copper sulphate sol. Allow the iron nail to remain in copper sulphate solution for about half an hour.
4. After half and hour, take out the iron nail from copper sulphate solution. We will find that the iron nail is covered with brown layer of copper metal.
5. If we look at the test tube, we find the original deep blue colour of copper sulphate solution has faded. The solution turns light green due to formation of iron sulphate or ferrous sulphate.



Example 4 - When a strip of lead metal is placed in a solution of copper chloride, then lead chloride solution and copper metal are formed -

$CuCl_2(aq) + Pb(s) \longrightarrow PbCl_2(aq) + Cu(s)$

Copper chloride	lead	lead chloride	copper
Green solution	Bluish grey	colourless sol.	Red brown

- lead is able to displace copper from copper chloride solution because lead is more reactive than copper.
- Here $CuCl_2$ used in this reaction is actually copper(II) chloride.

Example 5 - Iron metal reacts with dilute hydrochloric acid to form iron(II) chloride and hydrogen gas.

- $Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$

Iron	Hydrochloric acid	Iron(II) chloride	Hydrogen

- Here iron is more reactive than hydrogen (Reason)

- Example 6 - Sodium metal reacts with water to form sodium hydroxide and hydrogen gas.

- $2Na(s) + H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$

sodium	water	sodium hydroxide	hydrogen

Here sodium is more reactive than water (hydrogen)

- In case of oxides -

Example 7 - when copper oxide is heated with magnesium powder, then magnesium oxide and copper is formed -

- $CuO(s) + Mg(s) \longrightarrow MgO(s) + Cu(s)$

copper oxide	magnesium	magnesium oxide	copper.

Note - A more reactive metal displaces a less reactive metal from its oxide. Here, magnesium is displacing a less reactive metal, copper, from its oxide copper oxide.

Example 8 - When iron(III) oxide is heated with aluminium powder, then aluminium oxide and iron metal are formed -

$Fe_2O_3(s) + 2Al(s) \longrightarrow Al_2O_3(s) + 2Fe(l)$

Iron(III) oxide (Ferric oxide)	Aluminium	Aluminium oxide	Iron (molten)

Here, a more reactive metal, aluminium, is displacing a less reactive metal, iron, from its oxide, iron(III) oxide.

- All the above examples of displacement reactions are actually 'single displacement reactions' This is because in all these reactions only 'one element' displaces 'another element' from its compound.
- Another type of displacement reactions called 'double displacement reactions'.

4. DOUBLE DISPLACEMENT REACTION -

- These reactions in which two compounds react by an exchange of ions to form two new compounds are called double displacement reactions.
- Some examples of double displacement reactions -

Example 1 - When silver nitrate solution is added to sodium chloride solution, then a white precipitate displaces silver chloride with sodium nitrate solution.

- $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

Silver nitrate — Sodium chloride — Silver chloride (white ppt) — Sodium nitrate

- Here, silver chloride is formed as an insoluble white solid called as white precipitate.

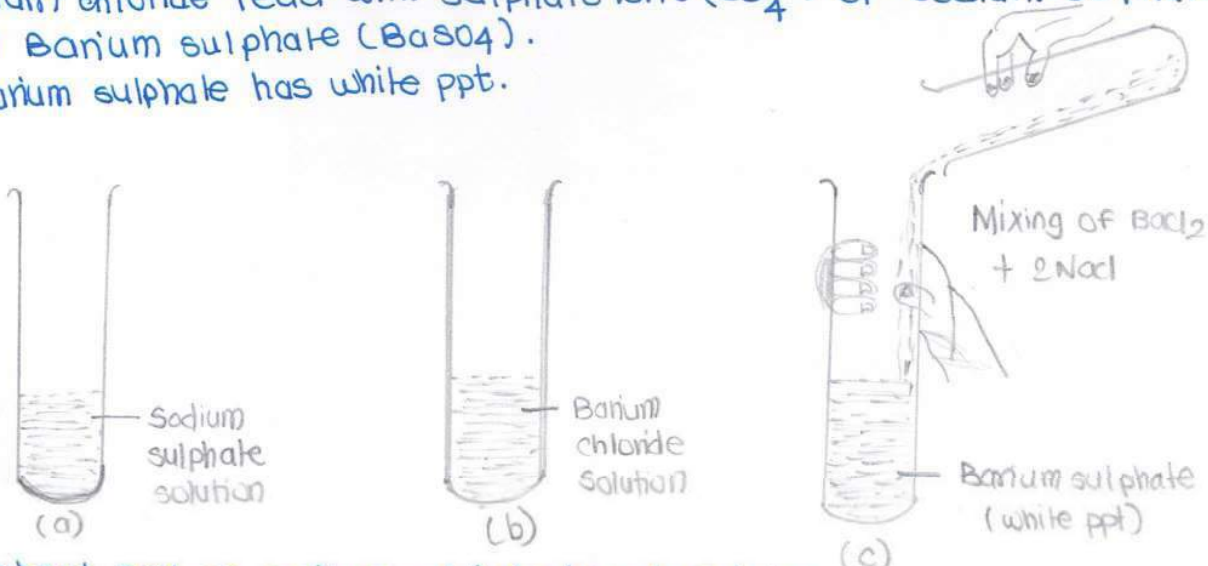
- What is Precipitation reaction?
- Any reaction in which an insoluble solid (called precipitate) is formed that separates from solution.

Example 2 - When barium chloride solution is added to sodium sulphate solution, then a white precipitate of Barium sulphate is formed along with sodium chloride solution.

$$BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NaCl(aq)$$

Barium chloride — Sodium sulphate — Barium sulphate (white ppt) — Sodium chloride

- Here, exchange of ions takes place in this reaction. The barium ions (Ba^{+2}) of Barium chloride react with sulphate ions (SO_4^{2-}) of sodium sulphate to form Barium sulphate ($BaSO_4$).
- Here Barium sulphate has white ppt.



a. Take about 3ml of sodium sulphate in a test tube.
b. In another test tube, take 3ml of Barium chloride solution.
c. Add Barium chloride solution to sodium sulphate solution
d. A white precipitate of Barium sulphate is formed at once.

- Example 3 - If barium chloride solution is added to copper sulphate solution then a white precipitate of barium sulphate is produced along with copper chloride solution.

$$Ba^{+}Cl^{-}_{2}(aq) + Cu^{+}SO^{-}_{4}(aq) \longrightarrow Ba^{+}SO^{-}_{4}(s) + Cu^{+}Cl^{-}_{2}(aq)$$

Barium chloride — copper sulphate — barium sulphate (white ppt) — copper chloride

- Example 4 - When potassium iodide solution is added to lead nitrate sol. then yellow precipitate of lead iodide is produced alongwith potassium nitrate.

$$Pb^{+}(NO_3)^{-}_{2}(aq) + 2K^{+}I^{-}(aq) \longrightarrow Pb^{+}I^{-}_{2}(s) + 2K^{+}NO^{-}_{3}(aq)$$

Lead nitrate — Potassium iodide — lead iodide (yellow ppt) — Potassium nitrate

- lead nitrate, $Pb(NO_3)_2$ is also written as lead(II) nitrate.
- Example 5 - When ammonium hydroxide solution is added to ~~ammonium~~ aluminium chloride solution, then white precipitate of Aluminium hydroxide is formed alongwith ammonium chloride solution.

$$Al^{+}Cl^{-}_{3}(aq) + 3NH^{+}_{4}OH^{-}(aq) \longrightarrow Al^{+}(OH)^{-}_{3}(s) + 3NH^{+}_{4}Cl^{-}(aq)$$

Aluminium chloride — Ammonium hydroxide — Aluminium hydroxide (white ppt) — Ammonium chloride

- These are some few examples of double displacement reaction.

5. OXIDATION AND REDUCTION REACTION -

- Oxidation -
  - The addition of oxygen to a substance is called oxidation.
  - The removal of hydrogen from a substance is called Oxidation.
- Reduction -
  - The addition of hydrogen to a substance is called Reduction.
  - The removal of oxygen from a substance is called Reduction.
- The process of reduction is just opposite of oxidation. (Note)

Morever, oxidation and reduction occurs together. (Note).

• OXIDISING AGENT AND REDUCING AGENT -

- Oxidising agent - 1. The substance which gives oxygen for oxidation is called oxidising agent.
  2. The substance which remove hydrogen is called oxidising agent.
- Reducing agent - 1. The substance which gives hydrogen for reduction is called reducing agent.
  2. The substance which remove oxygen is called reducing agent.

NOTE - The oxidation and reduction reaction are also called REDOX REACTIONS. Redox stands for red - reduction and ox - oxidation.

Example 1 - When copper oxide is heated with hydrogen, then copper metal and water are formed -

$$CuO_{(s)} + H_{2(g)} \xrightarrow{Heat} Cu_{(s)} + H_2O_{(l)}$$

copperoxide Hydrogen copper water

- In this reaction, CuO is changing into Cu. That is, oxygen is being removed from copperoxide so by definition, removal of oxygen from a substance is called reduction. copperoxide is being reduced to copper.
- In this reaction H_2 is changing into H_2O. That is oxygen is being added to hydrogen. By addition of oxygen to a substance is called oxidation. Hydrogen is being oxidised to water.

Removal of Oxygen - Reduction

$$CuO + H_2 \xrightarrow{Heat} Cu + H_2O$$

Addition of oxygen - oxidation

- In these reaction, CuO (copperoxide) is the oxidising agent. and H_2 is reducing agent.
- Substance oxidised - H_2
- substance reduced - CuO
- oxidising agent - CuO
- Reducing agent - H_2
- The reaction between copperoxide and hydrogen to form copper and water is an oxidation-reduction reaction which is also a displacement reaction.

Example 2 - When hydrogen sulphide react with chlorine, then sulphur and hydrogen chloride are formed -

$$H_2S + Cl_2 \xrightarrow{Heat} S + 2HCl$$

Hydrogen sulphide chlorine sulphur hydrogenchloride

- Here, H_2S is changing into S. That is hydrogen is removed from hydrogen sulphide. By def. the removal of hydrogen compound is called oxidation Hydrogen sulphide is being oxidised to sulphur.
- Here, Cl_2 is changing into HCl. That is hydrogen is being added to chlorine. Now by def. the addition of hydrogen to a substance is called reduction. so, chlorine is being reduced to hydrogen chloride.

Removal of hydrogen - Oxidation

$$H_2S + Cl_2 \xrightarrow{Heat} S + 2HCl$$

Addition of hydrogen - Reduction

- Substance oxidised - H_2S
- substance reduced - Cl_2
- Oxidising agent - Cl_2
- Reducing agent - H_2S

a. The substance which gets oxidised is the reducing agent
b. The substance which gets reduced is the oxidising agent.

• Example 3 - When zinc oxide is heated with carbon, then zinc metal and carbon monoxide are formed -

$$ZnO + C \xrightarrow{Heat} Zn + CO$$

zinc oxide — carbon — zinc — carbon monoxide

• zinc oxide (ZnO) is losing oxygen, so it is being reduced to zinc. and carbon (C) is gaining oxygen, so it is being oxidised to carbon monoxide.
• Oxidising agent - ZnO • Reducing agent - C
• It is used in the production of zinc metal in industry.
• Carbon is used in the form of coke for the extraction of zinc metal

• Example 4 - When manganese dioxide reacts with hydrochloric acid, then manganese dichlodride, chlorine and water are formed.

$$MnO_2 + 4HCl \xrightarrow{Heat} MnCl_2 + Cl_2 + 2H_2O$$

Manganese dioxide — hydrochloric acid — Manganese dichloride — chlorine — water

• Oxidising agent - MnO_2 • Reducing agent - HCl
• MnO_2 is losing oxygen to form $MnCl_2$, so magnese dioxide (MnO_2) is reduced to manganese dichloride.
• HCl is losing hydrogen to form Cl_2, so hydrochloric acid (HCl) is being oxidised to chlorine (Cl_2).

• Concept of oxidation and reduction in terms of metal and nonmetal-
• The addition of non-metallic element (or removal of metallic element) is called oxidation.
• The addition of metallic element (or removal of non-metallic element) is called reduction.

• Example 5 - When copper is heated in air, it reacts with the oxygen of air to form a black compound copper oxide.

$$2Cu_{(s)} + O_{2(g)} \xrightarrow{Heat} 2CuO$$

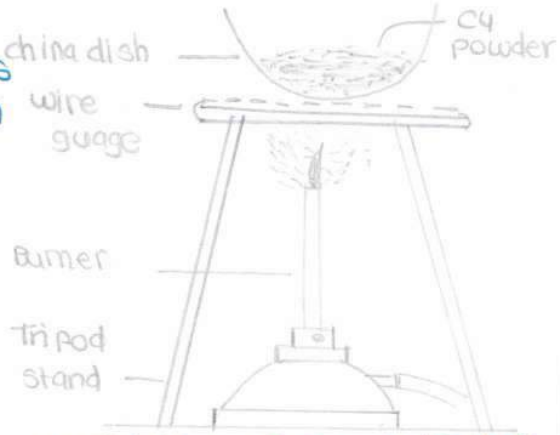
• copper (Red-brown) — • Oxygen (From air) — • copper oxide (Black)

• Cu is changing into CuO. This is addition of oxygen. But addition of oxygen is called oxidation, so copper (Cu) is oxidised to copper oxide (CuO).
• O_2 is changing into CuO. This is addition of copper (Cu) which is a metal. But, addition of metal is called reduction, so in this reaction, oxygen (O_2) is reduced to copper oxide (CuO).
• Oxidising agent - O_2 • Reducing agent - Cu

- Reaction carried out as-

1. Take about 1g of Cu powder in a china dish. It is red-brown in colour.
2. Heat the china dish strongly over a burner.
3. A black substance is formed. The black substance is copper oxide.

- The oxidation of magnesium is similar to the oxidation of copper.
- In case of magnesium ribbon, oxygen is oxidising agent and magnesium is reducing agent.
- When copper metal is heated in air, it gets oxidised to form copper oxide. This reaction can be reversed by hydrogen gas (passed) over heated copper oxide to get back Cu metal.
- If hydrogen gas is passed over heated copper oxide, then black copper oxide is reduced and red brown copper metal is obtained.



- oxidation of copper to CuO.

- $CuO_{(s)} + H_{2(g)} \xrightarrow{Heat} Cu_{(s)} + H_2O_{(l)}$

copper oxide (Black) + hydrogen → copper (Red brown) + water

- Copper oxide reduced to copper metal, whereas hydrogen is oxidised to water.

- What are the effects of oxidation reaction in everyday life?
- Oxidation has damaging effects on metals as well as on food.
- There are two common effects of oxidation reactions which we observe in daily life-

1. Corrosion of metals
2. Rancidity of food

- The oxidation involved in the corrosion of metals as well as rancidity of food is caused naturally by the oxygen present in air.

- **CORROSION** -

- Corrosion is the process in which metal are eaten up gradually by the action of air, moisture or a chemical (such as an acid) on their surface
- It is caused mainly by the oxidation of metals by oxygen of air.
- **Rusting** of iron metal is the most common form of corrosion.
- When an iron object is left in damp air for considerable time, it gets covered with a red-brown flasky substance called **Rust**. and this is called **Rusting** of iron.
- Iron metal is oxidised by the oxygen of air in the presence of water to form hydrated iron (II) oxide called rust.

- $4Fe + 3O_2 + 2xH_2O \longrightarrow 2Fe_2O_3 . xH_2O$ (Rust)

Iron + oxygen + water → Hydrated iron (III) oxide

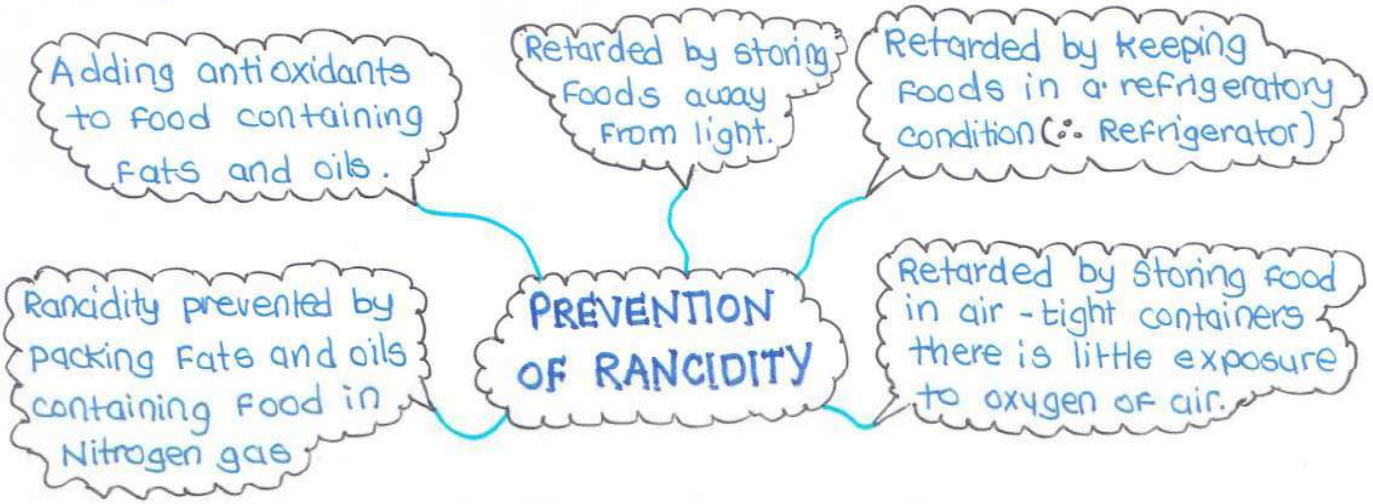
- The number of water molecules (x) in the rust varies, it is not fixed. The rusting of iron is a redox reactions.
- Rusting involves unwanted oxidation of iron metal which occurs in nature on its own.
- Rusting of iron is a continuous process, which, if not prevented in time eats up the whole iron object.
- Corrosion weakens the iron and steel objects and structure such as railings, car bodies, bridges and ships and cuts short their life.

METHODS OF PREVENTION -

- It can be prevented by painting.
- It can be prevented by applying grease or oil.
- It can be prevented by galvanisation.
- It can be prevented by tin plating and chromium plating
- It can be prevented by alloying to make stainless steel.
- It can be prevented by coating or depositing a thin layer of metal.

RANCIDITY -

- Oxidation also has damaging effects on foods containing fats and oils.
- When the food material prepared in fats and oils are kept for a long time, they starts giving unpleasant smell and taste. these are said to be rancid.
- The condition produced by aerial oxidation of fats and oils in foods marked by unpleasant smell and taste. is called Rancidity.
- Rancidity is called 'vikritgandhita' in Hindi.



- The antioxidant which are added to prevent rancidity are -
- BHA - Butylated Hydroxy - Anisole
- BHT - Butylated Hydroxy - Toluene
- By these way Food remains fresh and not become rancid.
- When the packed food is surrounded by unreactive gas nitrogen, there is no oxygen to cause its oxidation and make rancid.

- The refrigerator has a low temperature inside it. when the food is kept in refrigeratory, the oxidation of fats and oils in it slowed down due to low temperature.
- Due to absence of light, the oxidation of fats and oils present in food is slowed down and hence the development of randidity is retarded.
- These are the some prevention methods-

- Is oxidation an exothermic or an endothermic reaction?

- Mostly oxidation reaction is exothermic reactions. but it can be both exo and endothermic reaction.

Ex- Photosynthesis in Plants, digestion of food.