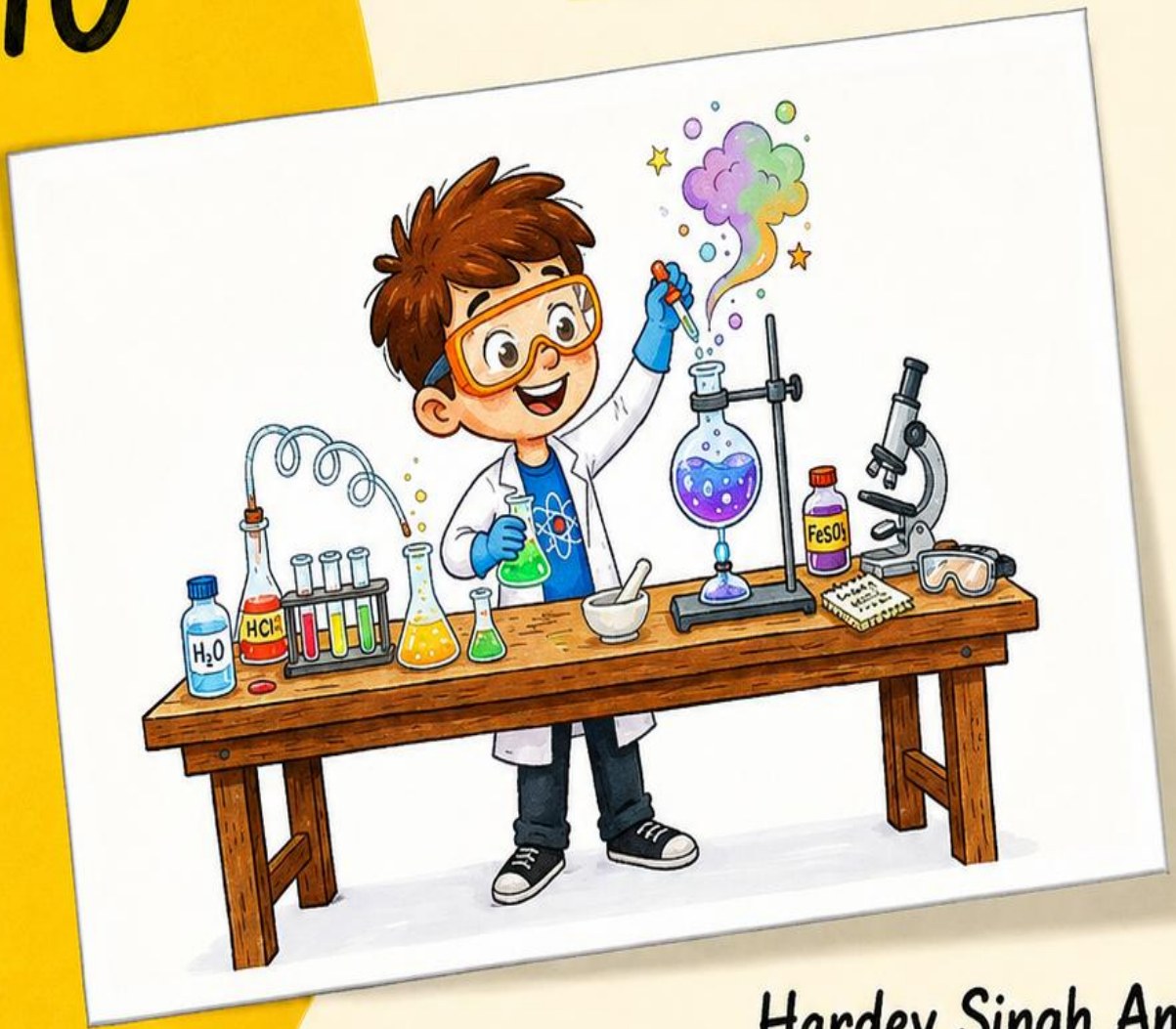
Chemical Reactions and Equations

Class 10

Science



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Introduction

- Everything around us is matter, which is composed of atoms and molecules.
- Atoms and molecules of the same or different matter interact with each other and with energy, resulting in different types of changes.
- These changes are broadly classified into two groups:
    1. Physical Change
    2. Chemical Change

Physical Change

a.) A temporary change where only the physical properties (shape, size, state) of the substance change.

b.) No new substance is formed.

c.) Chemical properties of the substance remain completely unchanged.

d.) These changes are usually reversible.

Examples:

Melting, Boiling, Freezing, Evaporation, Sublimation, Dissolution, Crystallisation, Tearing and Crushing

Chemical Change

a.) A permanent change where chemical properties of the substance change.

b.) New substances are formed.

c.) Chemical properties of the original substances are permanently altered.

d.) These changes are usually irreversible and cannot be undone by simple physical means.

Examples:

Combustion or Burning, Corrosion (Rusting/Patination), Precipitation, Fermentation (i.e., Curdling of Milk)

Chemical Reaction

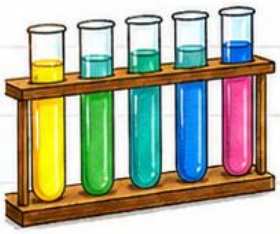
- Whenever a chemical change occurs, we can confidently say that chemical reaction has taken place.
- On an atomic level, chemical reaction actually involves the breaking and making of bonds between atoms to produce new substances.

How to Identify a Chemical Reaction ?

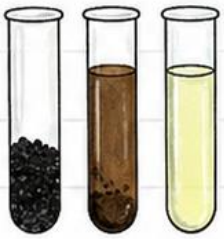
Whenever a chemical reaction takes place, we can observe certain changes in the substance.

The following observations are strong indicators of a chemical reaction:

Change in Colour



Change in Physical State



Evolution of Gas

Change in Temperature

Formation of Precipitate

Practice Questions – A

1. Which among the following are physical or chemical changes?
   (a) Evaporation of petrol
   (b) Burning of Liquefied Petroleum Gas (LPG)
   (c) Heating of an iron rod to red hot
   (d) Curdling of milk
   (e) Sublimation of solid ammonium chloride

2. Grapes hanging on the plant do not ferment but after being plucked from the plant can be fermented. Is this a chemical or a physical change?

3. Which of the following is a physical change?
   (a) Rusting of iron
   (b) Combustion of magnesium ribbon
   (c) Burning of candle
   (d) Melting of wax

Chemical Equation

- A chemical equation is a symbolic representation of a chemical reaction.
- A chemical equation provides a concise way to communicate what occurs during a chemical reaction.

How to Write a Chemical Equation ?

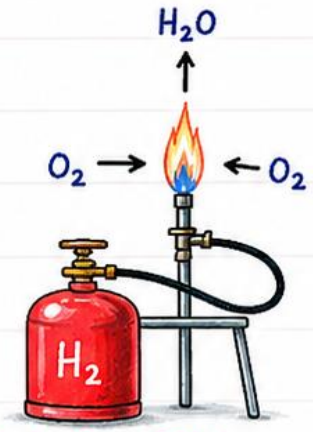
- The general form of a chemical equation is:

Reactants ⟶ Products

- Reactants are the substances present at the start of the reaction.
- Products are the substances formed as a result of the reaction.
- Arrow represents the direction of the reaction, indicating the conversion of reactants into products.

Example : H_2 + O_2 ⟶ H_2O

Hydrogen Oxygen Water



Combustion of Hydrogen

Need for Balancing a Chemical Equation

- Atoms are neither created nor destroyed during an ordinary chemical reaction.
- Therefore, the number and types of atoms present at the start of the reaction must remain the same at the end of the reaction.
- Hence, we can say that chemical equations follow the law of conservation of mass.
- This is why a chemical equation must always be balanced on both sides.

Example : $2\,H_2$ + O_2 ⟶ $2\,H_2O$

Hydrogen Oxygen Water

$2Na$ + $2H_2O$ ⟶ $2NaOH$ + H_2

Sodium Water Sodium Hydroxide Hydrogen

Adding Information to a Chemical Equation

- Chemical equations can include additional information like physical states and reaction conditions.
  - Physical States: Solid, Liquid, Gas, and Aqueous
  - Reaction Conditions: Pressure, Temperature, Catalyst etc.

Adding Physical States

- $2H_{2\,(g)} + O_{2\,(g)} \longrightarrow 2H_2O_{(g)}$
- $Zn_{(s)} + 2HCl_{(aq.)} \longrightarrow ZnCl_{2\,(aq.)} + H_{2\,(g)}$
- $2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq.)} + H_{2\,(g)}$

Solid	=	s
Liquid	=	l
Gas	=	g
Aqueous	=	aq.

Adding Reaction Conditions

If a reaction requires specific conditions (such as: Temperature, Pressure, Use of catalysts) these details are written directly above and/or below the arrow.

- $3H_{2\,(g)} + N_{2\,(g)} \xrightarrow[400\,°C - 450\,°C]{Fe\,,\,200\ atm} 2NH_{3\,(g)}$: Haber Process
- $6CO_{2\,(g)} + 6H_2O_{(l)} \xrightarrow[Sunlight]{Chlorophyll} C_6H_{12}O_{6\,(aq.)} + O_{2\,(g)}$: Photosynthesis

Limitations of Chemical Equations

- Chemical equation does not give any actual idea about the rate of a reaction.

  Fast: $2K_{(s)} + 2H_2O_{(l)} \longrightarrow 2KOH_{(aq.)} + H_{2\,(g)}$

  Slow: $2Li_{(s)} + 2H_2O_{(l)} \longrightarrow 2LiOH_{(aq.)} + H_{2\,(g)}$

- Chemical equation does not provide any information about concentration of products formed during the reaction.
- Chemical equation does not tell us about the feasibility of a reaction (whether a reaction will take place or not.)

  Not Feasible: $AgCl_{(s)} + NaNO_{3\,(aq.)} \not\longrightarrow NaCl_{(aq.)} + AgNO_{3\,(aq.)}$

  Feasible: $NaCl_{(aq.)} + AgNO_{3\,(aq.)} \longrightarrow AgCl_{(s)} + NaNO_{3\,(aq.)}$

Balancing of a Chemical Equation

- The common method used for balancing chemical equations is **Hit and Trial method.**
  - **Start with Complex Molecules:**

    Select the most complex molecule first, and use it to balance the other molecules and atoms on both sides of the equation by adding suitable coefficients.

  - **Handling Odd Numbers:**

    If any atom or molecule is found in an odd number on either side, double the coefficient of the selected complex molecule and attempt to balance both sides of the equation again.

Example 1: $Fe + H_2O \longrightarrow \boxed{Fe_3O_4} + H_2$ (Unbalanced Equation)

$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$ (Balanced Equation)

Example 2: $\boxed{CH_4} + O_2 \longrightarrow CO_2 + H_2O$ (Unbalanced Equation)

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ (Balanced Equation)

Example 3: $\boxed{Fe_2O_3} + HCl \longrightarrow FeCl_3 + H_2O$ (Unbalanced Equation)

$Fe_2O_3 + 6HCl \longrightarrow 2FeCl_3 + 3H_2O$ (Balanced Equation)

Example 4: $\boxed{Pb(NO_3)_2} \longrightarrow PbO + NO_2 + O_2$ (Unbalanced Equation)

$Pb(NO_3)_2 \longrightarrow PbO + 2NO_2 + O_2$ (Unbalanced Equation with odd no. of oxygen at one side)

$Pb(NO_3)_2 \longrightarrow 2PbO + 4NO_2 + O_2$ (Balanced Equation)

Example 5: $BaCl_2 + \boxed{Al_2(SO_4)_3} \longrightarrow BaSO_4 + AlCl_3$ (Unbalanced Equation)

$3BaCl_2 + Al_2(SO_4)_3 \longrightarrow 3BaSO_4 + 2AlCl_3$ (Balanced Equation)

Chemical Formulas of Compounds (From NCERT Class 9)

Understanding Ions

Formation of Cation

$A \longrightarrow A^{+} + e^{-}$

A^{+} : Cation

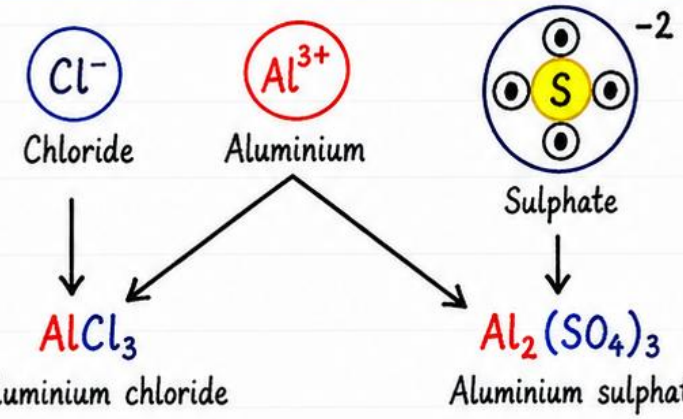
Formation of Anion

$B + e^{-} \longrightarrow B^{-}$

B^{-} : Anion

Formation of Compound

A^{+} (Cation) + B^{-} (Anion) $\longrightarrow$ $A^{+}B^{-}$ (Compound)



List of Common Cations

Univalent Cations (+1)

- Li^{+1} = Lithium ion
- Na^{+1} = Sodium ion
- K^{+1} = Potassium ion
- NH_4^{+1} = Ammonium ion
- Ag^{+1} = Silver ion

Divalent Cations (+2)

- Be^{+2} = Beryllium ion
- Mg^{+2} = Magnesium ion
- Ca^{+2} = Calcium ion
- Ba^{+2} = Barium ion
- Zn^{+2} = Zinc ion

Cations With Variable Charge

- Fe^{+2} = Ferrous ion or Iron (II)
- Fe^{+3} = Ferric ion or Iron (III)
- Cu^{+1} = Cuprous ion or Copper (I)
- Cu^{+2} = Cupric ion or Copper (II)
- Pb^{+2} = Plumbous ion or Lead (II)
- Pb^{+4} = Plumbic ion or Lead (IV)

Trivalent Cation (+3)

- Al^{+3} = Aluminum ion

List of Common Anions

Univalent Anions (−1)

F^{-1}	=	Fluoride
Cl^{-1}	=	Chloride
Br^{-1}	=	Bromide
I^{-1}	=	Iodide
OH^{-1}	=	Hydroxide
CN^{-1}	=	Cyanide
NO_2^{-1}	=	Nitrite
NO_3^{-1}	=	Nitrate
HCO_3^{-1}	=	Hydrogen carbonate
HSO_4^{-1}	=	Hydrogen sulphate
HSO_3^{-1}	=	Hydrogen Sulphite
CH_3COO^{-1}	=	Acetate

Divalent Anions (−2)

O^{-2}	=	Oxide
S^{-2}	=	Sulphide
SO_4^{-2}	=	Sulphate
SO_3^{-2}	=	Sulphite
CO_3^{-2}	=	Carbonate
$S_2O_3^{-2}$	=	Thiosulphate
$C_2O_4^{-2}$	=	Oxalate
$Cr_2O_7^{-2}$	=	Dichromate

Trivalent Anions (−3)

N^{-3}	=	Nitride
P^{-3}	=	Phosphide
PO_4^{-3}	=	Phosphate
PO_3^{-3}	=	Phosphite

Practice Questions – B

1. Write the chemical formulas of the following compounds:

a) Magnesium chloride
b) Sodium carbonate
c) Ammonium sulphate
d) Sodium hydrogen carbonate

2. Write the balanced equation for the following chemical reactions.

(a) Hydrogen + Chlorine ⟶ Hydrogen chloride
(b) Barium chloride + Aluminium sulphate ⟶ Barium sulphate + Aluminium chloride
(c) Sodium + Water ⟶ Sodium hydroxide + Hydrogen

3. Write a balanced chemical equation with state symbols for the following reactions.

(a) Solutions of barium chloride and sodium sulphate in water react to give insoluble barium sulphate and the solution of sodium chloride.
(b) Sodium hydroxide solution (in water) reacts with hydrochloric acid solution (in water) to produce sodium chloride solution and water.

Note : (aq) = aqueous solution , (s) = solid , (ℓ) = liquid

4. Translate the following statements into chemical equations and then balance them.

(a) Hydrogen gas combines with nitrogen to form ammonia.

(b) Hydrogen sulphide gas burns in air to give water and sulphur dioxide.

(c) Barium chloride reacts with aluminium sulphate to give aluminium chloride and precipitate of barium sulphate.

(d) Potassium metal reacts with water to give potassium Hydroxide and hydrogen gas.

5. Write the balanced chemical equations for the following reactions.

(a) Calcium hydroxide + Carbon dioxide ⟶ Calcium carbonate + water

(b) Zinc + Silver nitrate ⟶ Zinc nitrate + Silver

(c) Aluminium + Copper chloride ⟶ Aluminium chloride + Copper

(d) Barium chloride + Potassium sulphate ⟶ Barium sulphate + Potassium chloride

Exothermic and Endothermic Reactions

Exothermic Reactions

- The chemical reactions in which **energy is released** into the surroundings.
- While this energy is most commonly released as **heat**, it can also take the form of **light** and **electricity**.

$$2H_2 + O_2 \longrightarrow 2H_2O + \text{Heat} + \text{Light}$$

$$CaO + H_2O \longrightarrow Ca(OH)_2 + \text{Heat}$$

$$Zn + 2\,HCl \longrightarrow ZnCl_2 + H_2 + \text{Heat}$$

Why Do Some Reactions Tend to Be Exothermic?

- A reaction is exothermic when the energy released by forming the new bonds is **greater** than the energy required to break the old bonds.

$$\underset{\text{Hydrogen}}{2H\text{-}H} + \underset{\text{Oxygen}}{O{=}O} \longrightarrow \underset{\text{Water}}{2\,H\text{-}O\text{-}H}$$

E Bond Break Required < E Bond Form Released

Endothermic Reactions

- Chemical reactions that absorb energy from their surroundings to proceed.
- While this energy is most commonly absorbed as heat, it can also take the form of light or electricity.

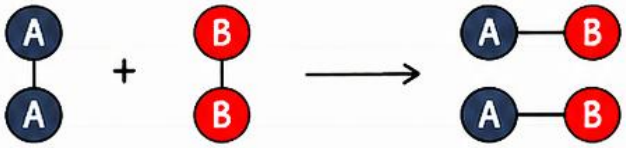
$$\text{Heat} + 2\,Pb(NO_3)_2 \longrightarrow 2\,PbO + 4\,NO_2 + O_2$$

$$\text{Sunlight} + 2\,AgBr \longrightarrow 2\,Ag + Br_2$$

$$\text{Electricity} + 2\,NaCl \longrightarrow 2\,Na + Cl_2$$

Why Do Some Reactions Tend to Be Endothermic?

- A reaction is endothermic when the energy required to break the bonds in the reactants is greater than the energy released when the new bonds form in the products.



$$E_{\text{Bond Break Required}} > E_{\text{Bond Form Released}}$$

Types of Chemical Reactions

Chemical reactions can be broadly classified into five primary types

1. Combination Reactions
2. Decomposition Reactions
   a. Thermal Decomposition
   b. Electrolytic Decomposition
   c. Photodecomposition
3. Displacement Reactions
4. Double-Displacement Reactions
   a. Precipitation Reactions
   b. Neutralisation Reactions
5. Reduction and Oxidation (Redox) Reactions

1. Combination Reactions (Synthesis Reactions)

A reaction in which two or more reactant molecules combine to form a large, complex product molecule.

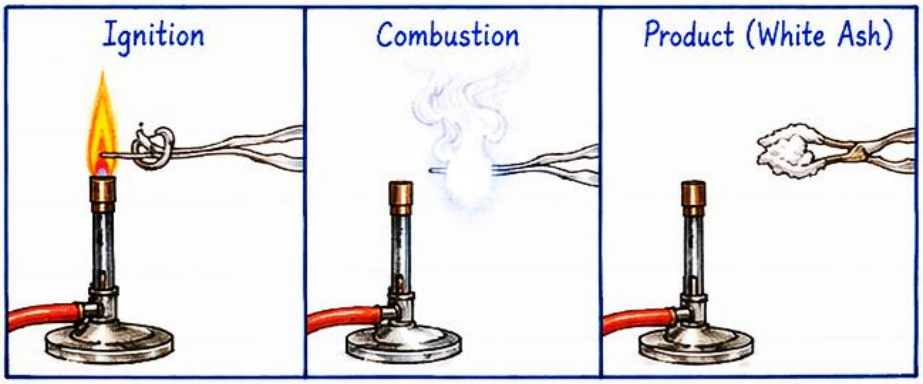
Generally Represented as: $A + B \longrightarrow AB$

Example 1: Burning of Magnesium Ribbon

- When the magnesium ribbon is ignited, it starts to burn with a dazzling white flame, giving rise to a white powder known as magnesium oxide (MgO). The reaction is highly exothermic.

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)} + Energy$$

Magnesium | Oxygen (Air) | Magnesium oxide (White)



- Magnesium ribbon is often rubbed with sandpaper before burning to remove the thin layer of magnesium oxide (MgO) that forms on its surface over time.

Example 2: Quicklime (CaO) in Water

- Calcium oxide (quicklime) reacts with water to form calcium hydroxide (slaked lime). The reaction is exothermic.

$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)} + Heat$$

Calcium oxide (Quicklime) | Water | Calcium hydroxide (Slaked lime)

- A solution of $Ca(OH)_2$ (slaked lime) produced in above reaction is used for whitewashing walls.

Calcium hydroxide reacts slowly with the carbon dioxide (CO_2) in air to form a thin layer of calcium carbonate ($CaCO_3$) on the walls.

Calcium carbonate is formed after 2 to 3 days of whitewashing and gives a shiny finish to the walls.

$$\underset{\text{Calcium hydroxide (Slaked lime)}}{Ca(OH)_{2\,(aq)}} + \underset{\text{Carbon dioxide (Air)}}{CO_{2\,(g)}} \longrightarrow \underset{\text{Calcium carbonate (Lime stone)}}{CaCO_{3\,(s)}} + \underset{\text{Water}}{H_2O_{(l)}}$$

NOTE: It is interesting to note that the chemical formula for **marble** is also $CaCO_3$.

Example 3: Combustion of Hydrogen

$$\underset{\text{Hydrogen}}{2H_{2\,(g)}} + \underset{\text{Oxygen (Air)}}{O_{2\,(g)}} \longrightarrow \underset{\text{Water}}{2H_2O_{(g)}} + \text{Heat}$$

Example 4: Combustion of Coal (C)

$$\underset{\text{Carbon}}{C_{(s)}} + \underset{\text{Oxygen (Air)}}{O_{2\,(g)}} \longrightarrow \underset{\text{Carbon dioxide}}{CO_{2\,(g)}} + \text{Heat}$$

Example 5: Formation of Carbonic acid (H_2CO_3)

$$\underset{\text{Carbon dioxide}}{CO_{2\,(g)}} + \underset{\text{Water}}{H_2O_{(l)}} \longrightarrow \underset{\text{Carbonic acid}}{H_2CO_{3\,(aq)}} + \text{Heat}$$

Exothermic Nature of Combination Reactions:

- Many combination reactions tend to be exothermic.
- However, it is important to note that; not all combination reactions are necessarily exothermic.
  - Here are a few reactions that are Exothermic but not Combination reactions:
    - Burning of Natural Gas: $CH_{4\,(g)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{(g)} + \text{Heat}$
    - Respiration: $C_6H_{12}O_{6\,(aq.)} + 6O_{2\,(aq.)} \longrightarrow 6CO_{2\,(aq.)} + 6H_2O_{(l)} + \text{Energy}$

      The release of energy during respiration can be observed in the form of heat when an organism exhale.

- Thermite process: $Fe_2O_{3\,(s)} + 2Al_{(s)} \longrightarrow Al_2O_{3\,(s)} + 2Fe_{(l)} + Heat$
- Neutralisation Reaction: $HCl_{(aq.)} + NaOH_{(aq.)} \longrightarrow NaCl_{(aq.)} + H_2O_{(g)} + Heat$

- Here are a few reactions that are Combination reactions but not Exothermic:
  - Nitric Oxide formation: $N_{2\,(g)} + O_{2\,(g)} + Heat \longrightarrow 2NO_{(g)}$ (Endothermic)
  - Hydrogen Iodide formation: $H_{2\,(g)} + I_{2\,(g)} \longrightarrow 2HI_{(g)}$ (Endothermic)

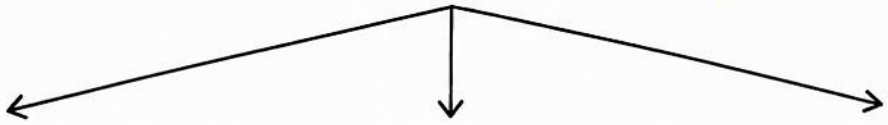
2. Decomposition Reactions

A reaction in which a single reactant molecule breaks down into two or more simpler product molecules.

Generally Represented as: $AB \longrightarrow A + B$

➢ These reactions almost always require an input of energy (such as heat, light, or electricity), making them endothermic processes.

Decomposition reactions can be broadly classified into **three types**



Thermal Decomposition | Photodecomposition | Electrolytic Decomposition

Thermal Decomposition

It occurs when a reactant compound breaks down upon heating.

Example 1: Thermal Decomposition of Calcium Carbonate ($CaCO_3$)

$$CaCO_{3\,(s)} \xrightarrow{Heat} CaO_{(s)} + CO_{2\,(g)}$$

Calcium carbonate (Lime stone) | Calcium oxide (Quick lime) | Carbon dioxide

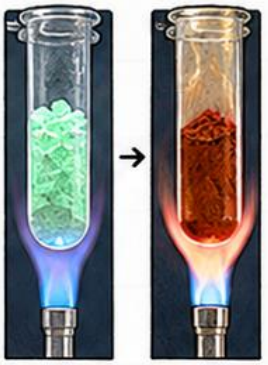
To confirm the presence of CO_2, pass the gas into lime water.

The CO_2 will turn the lime water milky.

Example 2: Thermal Decomposition of Ferrous Sulphate ($FeSO_4$)

$$2FeSO_{4\,(s)} \xrightarrow{\text{Heat}} Fe_2O_{3\,(s)} + SO_{2\,(g)} + SO_{3\,(g)}$$

Ferrous Sulphate (Pale Green) → Ferric oxide (Reddish-Brown) + Sulphur dioxide + Sulphur trioxide

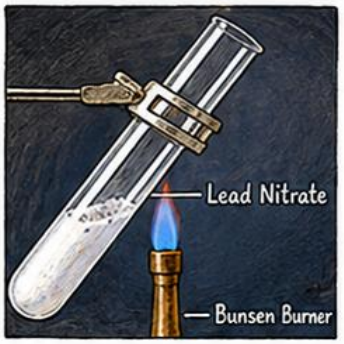


SO_2 and SO_3 gases have a pungent, suffocating smell. These are toxic, pollutant gases with an acidic nature. They turn Blue litmus solution to Red.

Example 3: Thermal Decomposition of Lead Nitrate ($Pb(NO_3)_2$)

$$2Pb(NO_3)_{2\,(s)} \xrightarrow{\text{Heat}} 2PbO_{\,(s)} + 4NO_{2\,(g)} + O_{2\,(g)}$$

Lead Nitrate (Colourless) → Lead (II) oxide (Yellow) + Nitrogen dioxide (Reddish-Brown) + Oxygen



The white crystalline powder of lead nitrate turns into a yellow residue of lead oxide (PbO).

A pungent, reddish-brown gas is evolved, which is nitrogen dioxide (NO_2).

Oxygen gas (O_2) is also released.

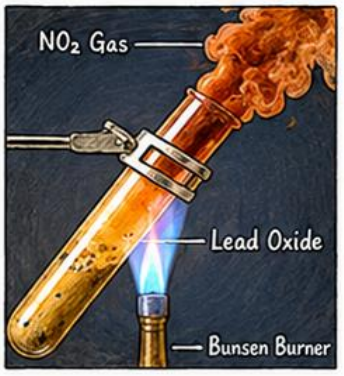


Photo-decomposition

It occurs when a reactant compound breaks down in presence of Sunlight.

Example: Photo-decomposition of Silver Bromide and Silver Chloride

$$2AgBr_{\,(s)} \xrightarrow{\text{Sunlight}} 2Ag_{\,(s)} + Br_{2\,(g)}$$

Silver Bromide (Yellow) → Silver (Grey) + Bromine

$$2AgCl_{\,(s)} \xrightarrow{\text{Sunlight}} 2Ag_{\,(s)} + Cl_{2\,(g)}$$

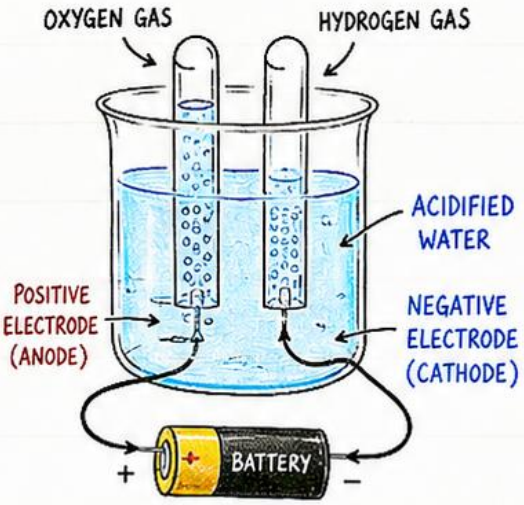
Silver Chloride (White) → Silver (Grey) + Chlorine

These reactions are widely used in **black-and-white photography.**

AgCl and AgBr are stored in **dark-coloured bottles** to protect them from sunlight.

Electrolytic Decomposition

It occurs when a reactant compound **breaks down** on passing **electric current** through its **molten** or **aqueous** state.



Example : **Electrolytic decomposition of Water**

$$2H_2O_{(\ell)} \xrightarrow[\text{Dilute Acid or Base}]{\text{Electricity (D.C.)}} 2H_{2(g)} + O_{2(g)}$$

Water → Hydrogen (at Cathode) + Oxygen (at Anode)

Pure water is **bad conductor** of electricity, so to make it conductor, we have to **add a few drops of dilute acid or base.**

- The Volume of gases liberated at cathode and anode are: $H_2 : O_2 = 2 : 1$

Endothermic Nature of Decomposition Reactions:

- Most of decomposition reactions tend to be **endothermic.**
- However, it is important to note that; **not all decomposition reactions are necessarily endothermic.**
  - Here are a few reactions that are **Endothermic** but not Decomposition reactions:
    - $Ba(OH)_{2(s)} + 2NH_4Cl_{(s)} + \text{Energy} \longrightarrow BaCl_{2(aq)} + 2H_2O_{(\ell)} + 2NH_{3(g)}$

      If you perform this reaction in a test tube, you will record **three distinct observations:**
      - The bottom of the test tube becomes **ice-cold.**
      - A **strong, choking or pungent smell** of ammonia gas (NH_3) is evolved.
      - The two solid powders mix and turn into a **liquid-like paste** due to the formation of water during the reaction.

3. Displacement Reactions

A chemical reaction in which a more reactive element displaces a less reactive element from its salt solution.

Generally Represented as: $A + BC \longrightarrow AC + B$ $[A > B]$

(Solution of 'B')

(where element 'A' is more reactive than element 'B')

➢ Reactivity Series:

Based on displacement reactions, we can compare the reactivities of different elements and organize them into a sequence called the reactivity series.

$$K > Na > Ca > Mg > Al > Zn > Fe > Pb > [H] > Cu > Hg > Ag > Au$$

Note: The reactivity of a metal depends on its tendency to lose valence electrons. A metal that loses electrons more easily is considered more reactive.

Displacement of Copper (Cu) from aqueous solutions (i.e., $CuSO_4$ or $CuCl_2$)

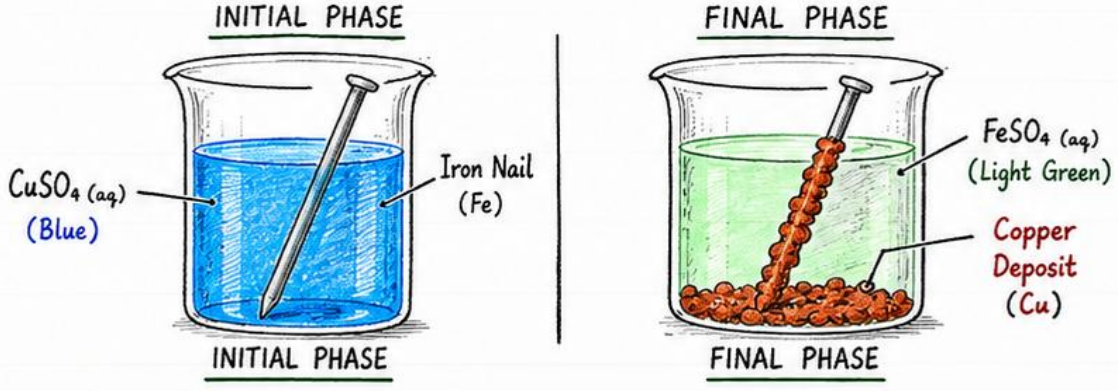
Copper (Cu) can be displaced from Copper sulphate ($CuSO_4$) solution by elements that have higher reactivity than copper, such as Fe, Zn, or Pb.

Example 1: Displacement of Cu from $CuSO_4$ using Iron (Fe)

$$CuSO_{4\,(aq)} + Fe_{(s)} \longrightarrow FeSO_{4\,(aq)} + Cu_{(s)}$$

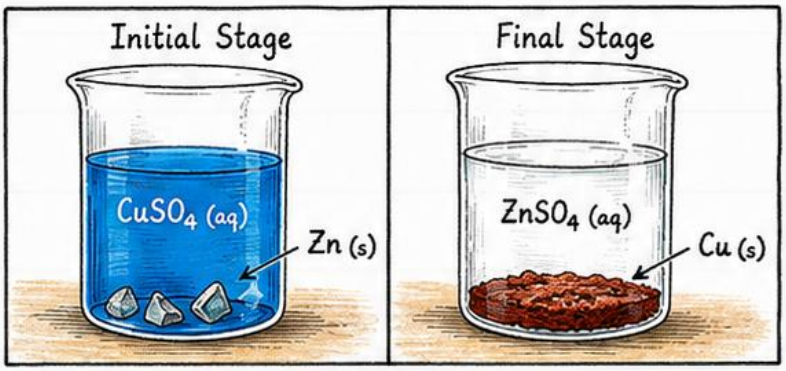
Copper Sulphate (Blue) | Iron | Ferrous Sulphate (Light Green) | Copper

When an iron nail is dipped in a Blue copper sulphate ($CuSO_4$) solution, iron displaces copper from the solution because iron is more reactive. This reaction results in the formation of a light green ferrous sulphate ($FeSO_4$) solution and a reddish-brown deposit of solid copper (Cu) on iron nail.



Example 2: Displacement of Cu from $CuSO_4$ using Zinc (Zn)

$$\underset{\text{Copper Sulphate (Blue)}}{CuSO_4\ (aq)} + \underset{\text{Zinc}}{Zn\ (s)} \longrightarrow \underset{\text{Zinc Sulphate (Colourless)}}{ZnSO_4\ (aq)} + \underset{\text{Copper}}{Cu\ (s)}$$



Example 3: Displacement of Cu from $CuCl_2$ using Lead (Pb)

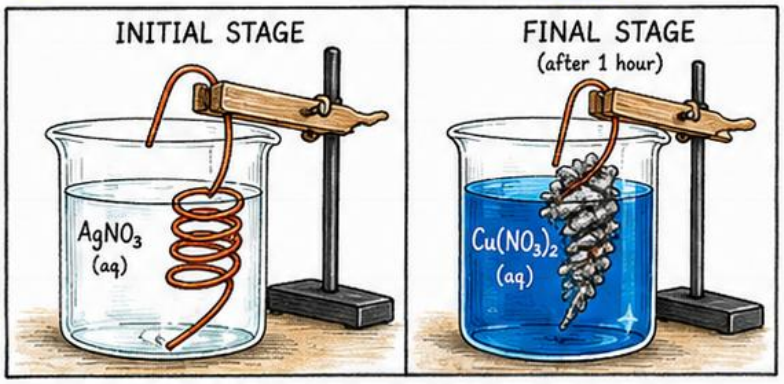
$$\underset{\text{Copper Chloride (Blue-green)}}{CuCl_2\ (aq)} + \underset{\text{Lead}}{Pb\ (s)} \longrightarrow \underset{\text{Lead Chloride (Colourless)}}{PbCl_2\ (aq)} + \underset{\text{Copper}}{Cu\ (s)}$$

Displacement of Silver (Ag) from $AgNO_3$

Silver (Ag) can be displaced from Silver nitrate ($AgNO_3$) solution by elements that have higher reactivity than silver, such as Cu or Zn.

Example : Displacement of Ag from $AgNO_3$ using Copper (Cu)

$$\underset{\text{Silver Nitrate (Colourless)}}{2AgNO_3\ (aq)} + \underset{\text{Copper}}{Cu\ (s)} \longrightarrow \underset{\text{Copper(II) Nitrate (Blue)}}{Cu(NO_3)_2\ (aq)} + \underset{\text{Silver}}{2Ag\ (s)}$$



Left: 0.1 M $AgNO_3$ | Right: Final Solution $Cu(NO_3)_2$ and Silver Deposit.

Displacement of Hydrogen (H_2) from Acid Solutions (i.e., HCl or H_2SO_4)

Metals placed above hydrogen in the reactivity series (such as K, Na, Ca, Mg, Al, Zn, Fe, Pb) can displace hydrogen gas from dilute acids like hydrochloric acid or sulphuric acid.

Example : Displacement of H_2 from HCl solution using Zinc (Zn)

$$2HCl_{(aq)} + Zn_{(s)} \longrightarrow ZnCl_{2\,(aq)} + H_{2\,(g)} + Heat$$

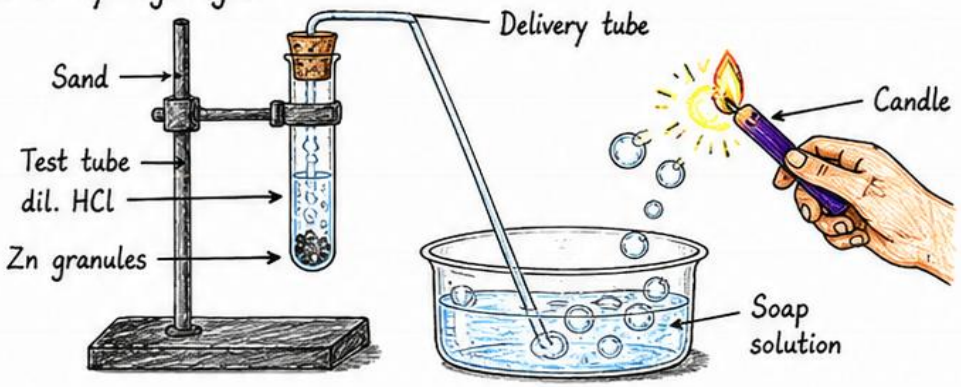
Hydrochloric acid (Colourless)	Zinc	Zinc Chloride (Colourless)	Hydrogen (Bubble formation)

As the zinc metal is placed into the test tube of hydrochloric acid, bubbles of H_2 gas start to form.

To confirm the presence of hydrogen gas formed during the reaction, you can perform "pop test."

Follow these steps:

- Pass the gas formed during the reaction through the soap solution.
- Bring the soap bubble close to a flame or a lit candle.
- If the gas is hydrogen, the soap bubble will ignite with a small "pop" sound, indicating the presence of hydrogen gas.



Solid-State Displacement of Iron (Fe) using Aluminium powder (Al)

$$Fe_2O_{3\,(s)} + 2Al_{(s)} \longrightarrow Al_2O_{3\,(s)} + 2Fe_{(l)} + Heat$$

Ferric Oxide	Aluminium	Aluminium oxide	Iron

This is an exothermic displacement reaction known as the Thermite Reaction, where the heat evolved is so immense that the displaced iron is obtained in a molten state, which is used to join cracked railway tracks.

4. Double Displacement Reactions

Double displacement reactions occur when two different compounds exchange their ions when their aqueous solutions are mixed and form new compounds.

Generally Represented as: $AB + CD \longrightarrow AD + BC$

Double displacement reactions can be broadly classified into two types

- Precipitation Reactions
- Neutralisation Reactions

Precipitation Reactions

Precipitation reactions involve the formation of a precipitate, which is an insoluble solid that separates from the solution.

Example 1: Formation of Lead iodide (PbI_2 ; Yellow Precipitate)

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

Lead Nitrate	Potassium iodide	Lead iodide	Potassium Nitrate
(Colourless)	(Colourless)	(Yellow ppt.)	(Colourless)

Example 2: Formation of Barium Sulphate ($BaSO_4$; White Precipitate)

$$BaCl_{2\,(aq)} + Na_2SO_{4\,(aq)} \longrightarrow BaSO_{4\,(s)} + 2NaCl_{\,(aq)}$$

Barium chloride	Sodium sulphate	Barium sulphate	Sodium chloride
(Colourless)	(Colourless)	(White ppt.)	(Colourless)

Example 3: Formation of Silver Chloride (AgCl ; White Precipitate)

$$\underset{\text{Silver Nitrate (Colourless)}}{AgNO_3\ (aq)} + \underset{\text{Sodium Chloride (Colourless)}}{NaCl\ (aq)} \longrightarrow \underset{\text{Silver chloride (White ppt.)}}{AgCl\ (s)} + \underset{\text{Sodium Nitrate (Colourless)}}{NaNO_3\ (aq)}$$

Neutralisation Reactions

Double displacement reactions in which an acid neutralises a base and results in the formation of salt and water.

Example 1: Neutralisation of HCl with NaOH

$$\underset{\text{Hydrochloric Acid}}{HCl\ (aq)} + \underset{\text{Sodium Hydroxide}}{NaOH\ (aq)} \longrightarrow \underset{\text{Sodium chloride}}{NaCl\ (aq)} + \underset{\text{Water}}{H_2O\ (l)} + \text{Energy}$$

Example 2: Reaction between Barium hydroxide and Ammonium chloride

$$\underset{\text{Barium hydroxide}}{Ba(OH)_2\ (aq)} + \underset{\text{Ammonium chloride}}{2NH_4Cl\ (aq)} + \text{Energy} \longrightarrow \underset{\text{Barium chloride}}{BaCl_2\ (aq)} + \underset{\text{Water}}{2H_2O\ (l)} + \underset{\text{Ammonia}}{2NH_3\ (g)}$$

5. Oxidation and Reduction (Redox) Reactions

Oxidation:

The chemical reaction which involves the gain of oxygen or the loss of hydrogen is called oxidation.

Example: When copper is heated in air, it reacts with oxygen present in the air to form black-coloured copper oxide on its surface through an oxidation reaction.

$$\underset{\text{Copper (Reddish-brown)}}{2Cu\ (s)} + \underset{\text{Oxygen}}{O_2\ (g)} \xrightarrow{\text{Heat}} \underset{\text{Copper(II) oxide (Black)}}{2CuO\ (s)}$$

Reduction:

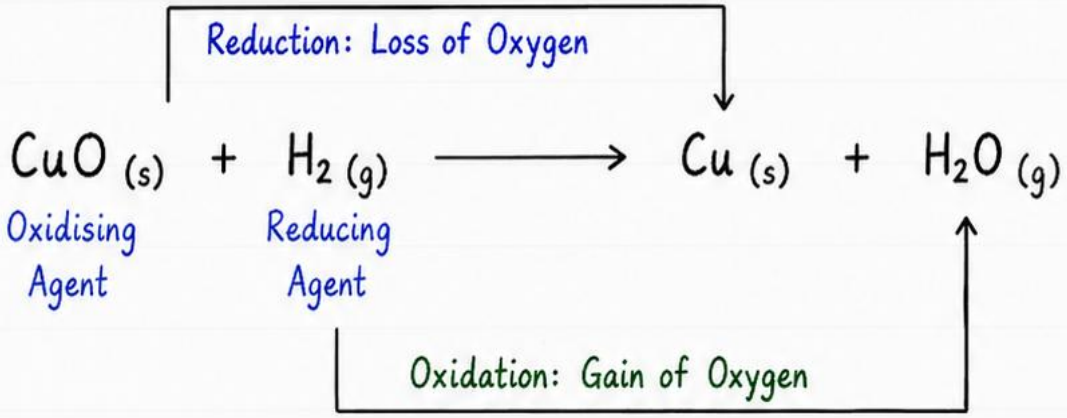
The chemical reaction which involves the gain of hydrogen or the loss of oxygen is called reduction.

Example: When hydrogen gas is passed over heated black-coloured CuO, it turns brown due to the formation of copper through a reduction reaction.

$$\underset{\text{Copper(II) oxide (Black)}}{CuO\ (s)} + \underset{\text{Hydrogen}}{H_2\ (g)} \xrightarrow{\text{Heat}} \underset{\text{Copper (Brown)}}{Cu\ (s)} + \underset{\text{Water}}{H_2O\ (l)}$$

NOTE:

Reduction and oxidation are interconnected processes that occur simultaneously. As one molecule is oxidized during a reaction, another molecule undergoes reduction at the same time.



Oxidising Agent and Reducing Agent

- **Oxidising agents (Oxidants):**

  Chemical substances that oxidise another chemical substance but themselves undergo reduction.

- **Reducing agents (Reductants):**

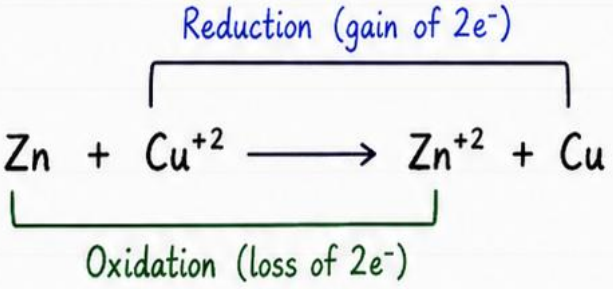
  Chemical substances that reduce another chemical substance but themselves undergo oxidation.

NOTE:

Electronic Concept of Oxidation and Reduction:

Oxidation involves the loss of electrons.

Reduction involves the gain of electron.



Oxidative Effects in Everyday Life

1. Rancidity

Rancidity refers to the process of deterioration or spoilage of food containing fats and oils, resulting in undesirable changes in taste, odour, and texture. It is primarily caused by the oxidation of fats and oils.

How to prevent Rancidity ?

a.) Storing food in airtight containers:

This minimizes exposure to atmospheric oxygen, thereby slowing down the oxidation process.

b.) Nitrogen flushing:

Replacing the air inside food packaging with an inert gas like nitrogen protects the contents. For example, potato chips manufacturers flush packaging with nitrogen gas to displace oxygen, ensuring the chips remain crispy and free from rancid flavours.

c.) Using antioxidants:

Adding chemical compounds inhibits the oxidation. Commonly used synthetic antioxidants include BHA (Butylated Hydroxyanisole) and BHT (Butylated Hydroxytoluene).

2. Corrosion

Corrosion is a natural process that occurs when metals react with substances in their environment, such as moisture, gases, or chemicals. This process typically results in the formation of oxides, hydroxides, or sulphides on the metal's surface, leading to alteration of its physical and chemical properties.

Examples of Corrosion

a.) Rusting of Iron:

When iron or steel reacts with oxygen in the presence of water or moisture, it forms a flaky, reddish-brown layer of hydrated iron oxide ($Fe_2O_3 \cdot xH_2O$).

b.) Patination of Copper:

When copper reacts with carbon dioxide and moisture in the air, it forms a **bluish-green** coating of basic copper carbonate ($CuCO_3.Cu(OH)_2$).

c.) Black-Tarnishing of Silver:

When silver reacts with sulphur compounds in the air (such as H_2S and SO_2) it forms a **black** coating of silver sulphide (Ag_2S)

How to protect metals from Corrosion ?

a.) Barrier Protection:

This can be achieved through painting, oiling/greasing, or electroplating. The coating acts as a physical barrier, preventing direct contact between the metal and environmental corrosive agents.

b.) Galvanization:

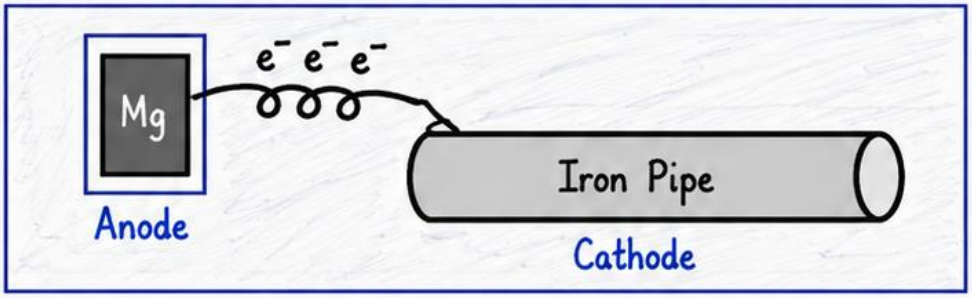
This is a widely used method for protecting iron or steel. It involves coating the metal with a thin layer of zinc.

c.) Anodising:

Anodising is typically performed on non-ferrous metals and their alloys, such as Aluminium, Magnesium, and Titanium. It is an electrochemical process that thickens the natural oxide layer on a metal's surface to make it corrosion-resistant.

d.) Cathodic Protection:

This method protects submerged or buried iron structures, such as pipelines, storage tanks, and ship hulls. The iron object is connected via a wire to a more electropositive, sacrificial metal (such as **Mg**)



NCERT Exercise Questions

1. Which of the statements about the reaction below are incorrect?

$$2PbO_{(s)} + C_{(s)} \longrightarrow 2Pb_{(s)} + CO_{2(g)}$$

(a) Lead is getting reduced.
(b) Carbon dioxide is getting oxidised.
(c) Carbon is getting oxidised.
(d) Lead oxide is getting reduced.

Select Answer:

(i) (a) and (b)
(ii) (a) and (c)
(iii) (a), (b) and (c)
(iv) all

2. $Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe$

The above reaction is an example of a

(a) combination reaction.
(b) double displacement reaction.
(c) decomposition reaction.
(d) displacement reaction.

3. What happens when dilute hydrochloric acid is added to iron fillings? Tick the correct answer.

(a) Hydrogen gas and iron chloride are produced.
(b) Chlorine gas and iron hydroxide are produced.
(c) No reaction takes place.
(d) Iron salt and water are produced.

4. What is a balanced chemical equation? Why should chemical equations be balanced?

5. Translate the following statements into chemical equations and then balance them.

(a) Hydrogen gas combines with nitrogen to form ammonia.
(b) Hydrogen sulphide gas burns in air to give water and sulphur dioxide.
(c) Barium chloride reacts with aluminium sulphate to give aluminium chloride and a precipitate of barium sulphate.
(d) Potassium metal reacts with water to give potassium hydroxide and hydrogen gas.

6. Balance the following chemical equations.

   (a) $HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + H_2O$

   (b) $NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$

   (c) $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

   (d) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + HCl$

7. Write the balanced chemical equations for the following reactions.

   (a) Calcium hydroxide + Carbon dioxide $\longrightarrow$ Calcium carbonate + Water

   (b) Zinc + Silver nitrate $\longrightarrow$ Zinc nitrate + Silver

   (c) Aluminium + Copper chloride $\longrightarrow$ Aluminium chloride + Copper

   (d) Barium chloride + Potassium sulphate $\longrightarrow$ Barium sulphate + Potassium chloride

8. Write the balanced chemical equation for the following and identify the type of reaction in each case.

   (a) Potassium bromide$_{(aq)}$ + Barium iodide$_{(aq)}$ $\longrightarrow$ Potassium iodide$_{(aq)}$ + Barium bromide$_{(aq)}$

   (b) Zinc carbonate$_{(s)}$ $\longrightarrow$ Zinc oxide$_{(s)}$ + Carbon dioxide$_{(g)}$

   (c) Hydrogen$_{(g)}$ + Chlorine$_{(g)}$ $\longrightarrow$ Hydrogen chloride$_{(g)}$

   (d) Magnesium$_{(s)}$ + Hydrochloric acid$_{(aq)}$ $\longrightarrow$ Magnesium chloride$_{(aq)}$ + Hydrogen$_{(g)}$

9. What does one mean by exothermic and endothermic reactions? Give examples.

10. Why is respiration considered an exothermic reaction? Explain.

11. Why are decomposition reactions called the opposite of combination reactions? Write equations for these reactions.

12. Write one equation each for decomposition reactions where energy is supplied in the form of heat, light or electricity.

13. What is the difference between displacement and double displacement reactions? Write equations for these reactions.

14. In the refining of silver, the recovery of silver from silver nitrate solution involved displacement by copper metal. Write down the reaction involved.

15. What do you mean by a precipitation reaction? Explain by giving examples.

16. Explain the following in terms of gain or loss of oxygen with two examples each.

    (a) Oxidation (b) Reduction

17. A shiny brown coloured element 'X' on heating in air becomes black in colour. Name the element 'X' and the black-coloured compound formed.

18. Why do we apply paint on iron articles?

19. Oil and fat containing food items are flushed with nitrogen. Why?

20. Explain the following terms with one example each.

    (a) Corrosion (b) Rancidity

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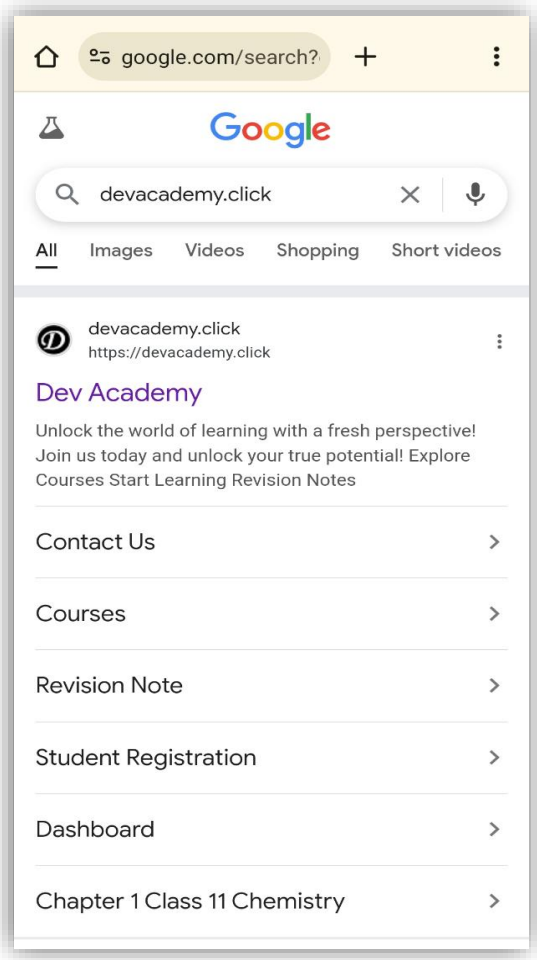

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