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Class – 10 / Science/ Notes Of Ch – 01 Chemical Reaction and Equation

Chemical Reaction: When two or more substances react and form some new substance, it is called a chemical reaction.

Characteristics of Chemical Reactions

When a chemical reaction takes place, it often shows **observable changes** that indicate a **new substance** (product) has been formed from the original substances (reactants). These observable features are called **characteristics of chemical reactions**.

Below are the **five important characteristics**:

(i) Evolution of a Gas
When a gas is released during a chemical reaction, it is observed as bubbles or effervescence.

Examples: **Zinc + Dilute HCl → Zinc chloride + Hydrogen gas**

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Hydrogen gas bubbles out.

Sodium bicarbonate + Acetic acid → Sodium acetate + Carbon dioxide + Water

$$NaHCO_3 + CH_3COOH \rightarrow CH_3COONa + CO_2 + H_2O$$

Carbon dioxide is evolved.

Formation of a Precipitate : When two aqueous solutions react and form an **insoluble solid**, this solid is called a **precipitate**.

Examples:

Barium chloride + Sodium sulphate → Barium sulphate (white precipitate) + Sodium chloride

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

Lead nitrate + Potassium iodide → Lead iodide (yellow precipitate) + Potassium nitrate

$$Pb(NO_3)_2 + 2KI \rightarrow PbI_2\downarrow + 2KNO_3$$

Change in Colour

A noticeable change in **colour** of the substances involved indicates a chemical reaction.

Examples:

1. **Copper + Silver nitrate → Copper nitrate (blue) + Silver (grey)**

   $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

   Colourless solution turns blue, and silver gets deposited.

2. **Potassium permanganate (purple) is decolourised by oxalic acid in presence of dilute H_2SO_4**

   $2KMnO_4(aq) + 5H_2C_2O_4(aq) + 3H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2MnSO_4(aq) + 10CO_2(g) + 8H_2O$

   Purple solution becomes colourless.

Change in Temperature

Chemical reactions often involve **release** or **absorption** of heat.

- **Exothermic Reaction**: Heat is **released**, temperature **increases**.
- **Endothermic Reaction**: Heat is **absorbed**, temperature **decreases**.

Examples:

1. **Exothermic**:
   Quick lime + Water → Slaked lime + Heat

   $CaO + H_2O \rightarrow Ca(OH)_2 + \text{heat}$

   The container becomes hot.

2. **Endothermic**:
   Barium hydroxide + Ammonium chloride → Barium chloride + Ammonia + Water

   $Ba(OH)_2 + 2NH_4Cl \rightarrow BaCl_2 + 2NH_3 + 2H_2O$

   Temperature falls; the container becomes cold.

Change in State

During chemical reactions, the **physical state** of substances (solid, liquid, gas) may change.

Examples:

Burning of wax

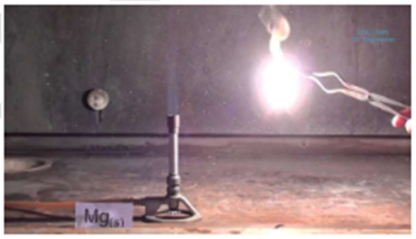
Wax (solid) melts to liquid and then burns to form **CO_2 (gas)** and **water vapour**.

Summary Table

Characteristic	Description	Example Reaction
Evolution of a gas	Gas bubbles or effervescence seen	$Zn + HCl \rightarrow H_2$
Precipitate formation	Insoluble solid forms from solution	$Pb(NO_3)_2 + KI$
Change in colour	Visible colour change in solution/substance	$Cu + AgNO_3$
Change in temperature	Heat is released or absorbed	$CaO + H_2O$
Change in state	Physical state of reactants/products change	Burning wax

Let us see few activities related to chemical reaction:

When **magnesium ribbon burns in air**, it undergoes a **chemical reaction** with oxygen to form **magnesium oxide**.



When you ignite a magnesium ribbon:

- It **burns with a bright white flame**.
- Produces **white ash** of **magnesium oxide (MgO)**.
- The reaction is **highly exothermic** (releases heat and light).

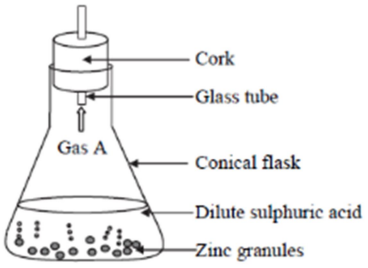
Word Equation: **Magnesium + Oxygen → Magnesium oxide**

Chemical Equation: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

The magnesium ribbon we bring has a layer of carbonate on it so before burning it needs to be cleaned with sandpaper.
Products : magnesium oxide (white color_MgO)
Identification property of reaction: white dazzling flame

2.Reacting zinc granules with dilute sulphuric acid

Chemical reaction : $H_2SO_4 + Zn$—------→ $ZnSo_4 + H_2$
Products : zinc sulphate and hydrogen gas
Identification: beaker becomes hot and hydrogen gas is produced which makes match stick burn with pale blue flame and popping sound.



3.Reacting barium iodide with lead chloride

Chemical reaction : $BaI + PbCl_2 \rightarrow BaCl_2 + Pb$
Products : barium chloride white colour and yellow colour.

What is a Chemical Equation?

A **chemical equation** is a **symbolic representation** of a chemical reaction using chemical formulas of **reactants** and **products**.

It shows:

- **Reactants**: Substances that take part in the reaction (**left side**)
- **Products**: Substances formed after the reaction (**right side**)
- An **arrow (→)** indicates the direction of the reaction

Example:

Word Equation:
Magnesium + Oxygen → Magnesium oxide

Chemical Equation (Unbalanced):

$Mg + O_2 \rightarrow MgO$

Unbalanced Chemical Equation

An **unbalanced chemical equation** is one in which the **number of atoms** of each element on the **reactant side is not equal** to the number of atoms on the **product side**.

➡ It **violates the Law of Conservation of Mass**.

Example:

$Mg + O_2 \rightarrow MgO$

(This is **unbalanced** because there are 2 oxygen atoms on the left but only 1 on the right.)

Balanced Chemical Equation

A **balanced chemical equation** is one in which the **number of atoms of each element** is **equal** on both sides of the equation.

➡It **follows the Law of Conservation of Mass**.

Example:

$Mg + O_2 \rightarrow MgO$ (Now both sides have 2 magnesium atoms and 2 oxygen atoms.)

Why Balance Chemical Equations?

- To obey the **Law of Conservation of Mass**
- To show a **correct and realistic reaction**
- Important for **calculations in chemistry** (stoichiometry)

To Make Equations More Informative

A **chemical equation** becomes more meaningful and helpful when it includes additional details such as **physical states, heat changes**, and **conditions required** for the reaction. These enhanced equations are called **"information-giving" or "more informative" equations**.

1. Indicating the Physical States of Reactants and Products

The physical state of each substance involved in the reaction is shown using **symbols** in brackets:

- **(s)** → solid
- **(l)** → liquid
- **(g)** → gas
- **(aq)** → aqueous (substance dissolved in water)

Example:

$$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

2. Indicating Heat Changes in the Reaction

Chemical reactions are often accompanied by heat changes:

- If **heat is released**, it is an **exothermic reaction** (temperature increases).
- If **heat is absorbed**, it is an **endothermic reaction** (temperature drops).

This can be indicated by writing **Δ (delta)** above the arrow or mentioning heat explicitly.

Examples:

Exothermic Reaction: $C(s) + O_2(g) \rightarrow CO_2(g) + \text{Heat}$

Endothermic Reaction: $N_2(g) + O_2(g) \xrightarrow{\text{Heat}} 2NO(g)$

Indicating the Conditions for the Reaction

Some chemical reactions need specific **conditions** to proceed, such as:

- **Temperature**
- **Pressure**
- **Catalysts**
- **Light (photochemical reactions)**

These are written **above or below the arrow** in the chemical equation.

Examples:

- $N_2 + 3H_2 \xrightarrow{\text{Fe catalyst, } 450^\circ C,\ 200\ atm} 2NH_3$
- $H_2 + Cl_2 \xrightarrow{\text{Sunlight}} 2HCl$

Note: By adding details about **physical states, heat changes**, and **reaction conditions**, a chemical equation becomes more **descriptive**, **informative**, and **scientifically complete**.

Types of chemical reactions

1. **Combination Reaction**
2. **Decomposition Reaction**
3. **Displacement Reaction**
4. **Double Displacement Reaction**
5. **Redox Reaction** (Oxidation and Reduction)
6. **Exothermic Reaction**
7. **Endothermic Reaction**

1. Combination Reaction

A **combination reaction** is a chemical reaction in which **two or more substances** (elements or compounds) **combine** to form a **single product**.

$A + B \rightarrow AB$

These reactions are usually **exothermic**, i.e., they **release heat**.

Types of Combination Reactions:

Element + Element → Compound

In this type, **two elements combine** to form a **compound**.

Examples:

$Mg + O_2 \rightarrow MgO$

(Magnesium + Oxygen → Magnesium oxide)

2. $S + O_2 \rightarrow SO_2$

   (Sulphur + Oxygen → Sulphur dioxide)

3. $H2 + Cl_2 \rightarrow 2HCl$
   (Hydrogen + Chlorine → Hydrogen chloride)

Element + Compound → New Compound

In this type, an **element combines with a compound** to form a **new compound**.

Examples:

1. $2CO + O_2 \rightarrow 2CO_2$
   (Carbon monoxide + Oxygen → Carbon dioxide)

2. $C + CO_2 \rightarrow 2CO$
   (Carbon + Carbon dioxide → Carbon monoxide)

3. $H_2 + O_2 \rightarrow H_2O$
   (Hydrogen + Oxygen → Water)
   (Note: this reaction usually needs a spark to start)

Compound + Compound → New Compound

In this type, **two compounds react** to form **a new compound**.

Examples:

1. $CaO + H_2O \rightarrow Ca(OH)_2$
   (Quick lime + Water → Slaked lime)

2. $SO_3 + H_2O \rightarrow H_2SO_4$
   (Sulphur trioxide + Water → Sulphuric acid)

3. $NH_3 + HCl \rightarrow NH_4Cl$
   (Ammonia + Hydrochloric acid → Ammonium chloride)

Key Features of Combination Reactions:

- Always **form one product**.
- Mostly **exothermic** (release heat).
- Can involve **elements or compounds** as reactants.

2. Decomposition Reaction

A **Decomposition Reaction** is a type of chemical reaction in which **a single compound breaks down** into **two or more simpler substances** (elements or compounds), usually with the help of **heat, light, or electricity**.

General form: $AB \rightarrow A + B$

Decomposition reactions are usually **endothermic**, meaning they **absorb energy** to proceed.

Types of Decomposition Reactions

Thermal Decomposition Reaction

(**Decomposition by heat**)

In this type, a compound **decomposes on heating** to form simpler products.

Examples:

$CaCO_3 \rightarrow CaO + CO_2$
(Calcium carbonate → Quick lime + Carbon dioxide)

2. $2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$
(Lead nitrate → Lead oxide + Nitrogen dioxide + Oxygen)

3. $FeSO4 \rightarrow Fe_2O_3 + SO_2 + SO_3$
(Ferrous sulphate → Ferric oxide + Sulphur dioxide + Sulphur trioxide)

Photolytic Decomposition Reaction

(**Decomposition by light**)

In this type, a compound **breaks down in the presence of sunlight**.

Examples:

1. $2AgCl \rightarrow 2Ag + Cl_2$
(Silver chloride → Silver + Chlorine)
Turns grey in sunlight (used in photography)

2. $2AgBr \rightarrow 2Ag + Br_2$
(Silver bromide → Silver + Bromine)

Electrolytic Decomposition Reaction

(**Decomposition by electricity**)

In this type, a compound **decomposes when electric current** is passed through it.

Examples:

1. **Electrolysis of Water**

$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

(Water → Hydrogen + Oxygen)

2. $2NaCl(l) \rightarrow 2Na(s) + Cl_2(g)$
(Molten sodium chloride → Sodium + Chlorine)

3. Displacement Reaction

A **displacement reaction** is a type of chemical reaction in which **a more reactive element displaces a less reactive element** from its **salt solution or compound**.

General Form: **A + BC → AC + B**
(where A is more reactive than B)

Types of Displacement Reaction:

a) Metal Displacement Reaction:

A more reactive metal displaces a less reactive metal from its salt solution.

Examples:

1. **$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$**
   (Iron displaces copper from copper sulphate)
2. **$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$**
   (Zinc displaces copper)
3. **$Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$**
   (Aluminium displaces iron from iron oxide – *Thermite reaction*)

(b) Non-metal Displacement Reaction:

A more reactive non-metal (like chlorine) displaces a less reactive non-metal (like bromine or iodine) from its salt solution.

Examples:

1. **$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$**
   (Chlorine displaces bromine from potassium bromide)
2. **$Cl_2 + 2KI \rightarrow 2KCl + I_2$**
   (Chlorine displaces iodine)
3. **$Br_2 + 2KI \rightarrow 2KBr + I_2$**
   (Bromine displaces iodine)

(c) Displacement in Acid Reactions:

Metals can displace hydrogen from acids.

Examples:

1. **$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$**
2. **$Fe + 2HCl \rightarrow FeCl_2 + H_2\uparrow$**
3. **$Mg + 2HCl \rightarrow MgCl_2 + H_2\uparrow$**

Note:

- Reactivity series helps to predict displacement reactions.
- **No displacement** occurs if the reacting element is **less reactive** than the one in the compound.
- Displacement reactions are often **exothermic** in nature.

Reactivity Series (Most Reactive to Least Reactive):

Potassium (K)

Sodium (Na)

Calcium (Ca)

Magnesium (Mg)

Aluminium (Al)

Zinc (Zn)

Iron (Fe)

Lead (Pb)

(Hydrogen - non-metal, used for comparison)

Copper (Cu)

Mercury (Hg)

Silver (Ag)

Gold (Au)

Platinum (Pt)

Reactivity Series of Non-Metals

Like metals, **non-metals** also have different reactivities. The **reactivity series of non-metals** is a list of **non-metals arranged in decreasing order of their tendency to gain electrons** (i.e., **oxidizing power**).

Reactivity Series of Common Non-Metals (Most Reactive to Least Reactive):

Fluorine (F_2)

Chlorine (Cl_2)

Bromine (Br_2)

Iodine (I_2)

Sulphur (S)

Phosphorus (P)

Carbon (C)

Hydrogen (H)

Explanation:

- **Halogens (F_2, Cl_2, Br_2, I_2)** are the most reactive non-metals because they have **7 electrons in their outermost shell** and readily gain 1 electron to become stable.
- The reactivity **decreases down the group** in halogens.
- **Fluorine** is the most reactive non-metal.
- **Chlorine** can displace bromine and iodine from their salt solutions.

Examples of Displacement by Non-Metals:

1. **$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$**
   (Chlorine displaces bromine from potassium bromide.)
2. **$Cl_2 + 2KI \rightarrow 2KCl + I_2$**
   (Chlorine displaces iodine from potassium iodide.)
3. **$Br_2 + 2KI \rightarrow 2KBr + I_2$**
   (Bromine displaces iodine from potassium iodide.)

Note:

The reactivity of non-metals is **not as widely extended** as that of metals in the Class 10 syllabus. The **focus is mostly on halogens** and **oxidizing nature**.

Double Displacement Reaction

A **double displacement reaction** is a chemical reaction in which **two compounds react by exchanging ions or radicals** to form two new compounds.

General form: **AB + CD → AD + CB**

Here, positive ions (A and C) and negative ions (B and D) of two reactants are exchanged.

Types of Double Displacement Reactions:

Precipitation Reaction

- When two aqueous solutions react and an **insoluble solid (precipitate)** is formed.

Example:

(a) $BaCl_2$ (aq) + Na_2SO_4 (aq) → $BaSO_4$ (↓) + 2NaCl (aq)
White precipitate of **Barium sulphate** is formed.

(b) $AgNO_3$ (aq) + NaCl (aq) → AgCl (↓) + $NaNO_3$ (aq)
White precipitate of **Silver chloride**.

(c) $Pb(NO_3)_2$ (aq) + KI (aq) → PbI_2 (↓) + $2KNO_3$ (aq)
Yellow precipitate of **Lead(II) iodide**.

(d) $CuSO_4$ (aq) + 2NaOH (aq) → $Cu(OH)_2$ (↓ blue ppt) + Na_2SO_4 (aq)

Copper(II) hydroxide (**$Cu(OH)_2$**) forms a **blue precipitate**.

(e) $FeSO_4$ (aq) + 2NaOH (aq) → $Fe(OH)_2$ (↓ green ppt) + Na_2SO_4 (aq)

Ferrous hydroxide (**$Fe(OH)_2$**) forms a **green precipitate**.

Neutralization Reaction

- When an **acid reacts with a base** to form **salt and water**.

Example:
HCl (aq) + NaOH (aq) → NaCl (aq) + H_2O (l)

H_2SO_4 (aq) + 2KOH (aq) → K_2SO_4 (aq) + $2H_2O$ (l)

$2HNO_3$ (aq) + $Ca(OH)_2$ (aq) → $Ca(NO_3)_2$ (aq) + $2H_2O$ (l)

CH_3COOH (aq) + NH_4OH (aq) → CH_3COONH_4 (aq) + H_2O (l)

Gas Formation Reaction

- When one of the products is a **gas** that escapes from the reaction mixture.

Example:
Na_2CO_3 (aq) + 2HCl (aq) → 2NaCl (aq) + H_2O (l) + CO_2 (g)
Carbon dioxide gas is released.

$NaHCO_3$ (aq) + HCl (aq) → NaCl (aq) + H_2O (l) + CO_2 (↑)

Gas formed: Carbon dioxide (CO_2)
Observation: Effervescence (bubbling) due to gas evolution

$CaCO_3$ (s) + 2HCl (aq) → $CaCl_2$ (aq) + H_2O (l) + CO_2 (↑)

Gas formed: Carbon dioxide (CO_2)
Observation: Brisk effervescence
Colour: All reactants and products are colourless except $CaCO_3$ (white solid)

H_2SO_4 (aq) + Na_2SO_3 (aq) → Na_2SO_4 (aq) + H_2O (l) + SO_2 (↑)

Gas formed: Sulphur dioxide (SO_2) — pungent smelling
Observation: Colorless gas with sharp choking smell

Conditions for Double Displacement to Occur:

At least one of the following must happen after ion exchange:

1. Formation of **precipitate**
2. Formation of **gas**
3. Formation of **weakly ionized compound** (like water in neutralization)

When Double Displacement Will NOT Occur:

Double displacement reactions will **not** occur if:

1. **Both products are soluble salts** and remain in solution → no visible change
2. **No precipitate, gas, or water** is formed → no driving force

Example (no reaction):
NaCl (aq) + KNO_3 (aq) → no visible reaction
All ions stay dissolved – **no precipitate**, **no gas**, **no water**

Redox Reaction (Oxidation and Reduction)

A chemical reaction in which **oxidation** and **reduction** occur **simultaneously** is called a **Redox Reaction**.

- **Oxidation** is the **loss of electrons** or **addition of oxygen** or **removal of hydrogen**.
- **Reduction** is the **gain of electrons** or **removal of oxygen** or **addition of hydrogen**.

Examples of Redox Reactions:

Example 1:

$CuO (s) + H2(g) \rightarrow Cu (s) + H2O (l)$

- **Oxidation**: $H_2 \rightarrow H_2O$ (Hydrogen is oxidized – gains oxygen)
- **Reduction**: CuO → Cu (Copper oxide is reduced – loses oxygen)

Redox Reaction because oxidation and reduction occur simultaneously.

Example 2:

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

- **Cu^{2+}** is **reduced** to Cu (gains electrons)

Example 3:

$2Na (s) + Cl_2(g) \rightarrow 2NaCl$

Na is **oxidized** (loses electrons to form Na^+)

Cl_2 is **reduced** (gains electrons to form Cl^-)

Oxidizing Agent and Reducing Agent:

Oxidizing Agent (Oxidant):

- The substance which **causes oxidation** (i.e., gains electrons or provides oxygen).
- It **gets reduced** in the process.

Example: In $CuO + H_2 \rightarrow Cu + H_2O$,
CuO is the **oxidizing agent** (gives oxygen to H_2 and gets reduced).

Reducing Agent (Reductant):

- The substance which **causes reduction** (i.e., loses electrons or removes oxygen).
- It **gets oxidized** in the process.

Example: In the same reaction,
H_2 is the **reducing agent** (removes oxygen from CuO and gets oxidized to H_2O).

Some Common Redox Reactions

1. **Thermite Reaction:** $Fe_2O_3+2Al \rightarrow 2Fe+Al_2O_3$

   Fe_2O_3 is reduced & **Al** is oxidized

2. **Reaction of Magnesium with HCl:** $Mg + 2HCl \rightarrow MgCl_2+H_2$

   Mg is oxidized to Mg^{2+} & H^+ is reduced to H_2

3. **Burning of Carbon:** $C + O_2 \rightarrow CO_2$

   Carbon is oxidized (gains oxygen)

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

   Lead nitrate on heating decomposes to give:

   Lead oxide (PbO) – yellow solid

   Nitrogen dioxide (NO_2) – brown gas

   Oxygen (O_2) – colorless gas

This reaction is **both thermal decomposition** and **redox**.

6. Corrosion

Corrosion is the slow destruction of metals due to their reaction with substances present in the environment like oxygen, water, acids, etc.

Example:

Iron (Fe) reacts with **oxygen (O_2)** and **moisture (H_2O)** in the air to form **hydrated iron(III) oxide**, commonly called **rust**.

$$4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3 \rightarrow \mathrm{Fe_2O_3 \cdot} xH_2O$$

This reddish-brown flaky substance is **rust**.

Effects of Corrosion:

- Iron articles weaken and break.
- Bridges, vehicles, iron railings, etc., become structurally unsafe.
- Wastage of resources and money.

Prevention of Corrosion:

Method	Description
Galvanization	Coating iron with zinc
Painting	Applying paint to form a protective layer
Oiling/Greasing	Prevents moisture contact with metal
Alloying	Making alloys like stainless steel
Electroplating	Coating with another metal like chromium

7. Rancidity

Rancidity is the condition produced by the oxidation of fats and oils in food materials that results in a bad taste and smell.

Examples of Rancidity:

- Chips kept in open air for a long time develop a bad smell.

- Ghee or oil left outside gets spoiled and smells unpleasant.
- Old butter smelling sour.

Prevention of Rancidity:

Method	Explanation
Refrigeration	Slows down oxidation by keeping food in cool environment
Airtight containers	Limits contact with oxygen
Adding antioxidants	Substances like **BHA (Butylated hydroxyanisole)** & **BHT** delay oxidation
Vacuum Packing	Removing air from packaging
Flushing with Nitrogen	Chips packets are filled with **Nitrogen gas** to prevent rancidity