

Short Notes on Chemical Reactions and Equations

1. Chemical Reactions and Identification

A chemical reaction is said to have taken place whenever a chemical change occurs.

Observations that indicate a chemical reaction:

- Change in state
- Change in colour
- Evolution of a gas
- Change in temperature

Examples:

Milk turning sour, rusting of iron, fermentation of grapes, and respiration.

1.1 Chemical Equations

1. Representation:

A chemical reaction can be written in a shorter form using a word equation:

Reactants → Products

Reactants are written on the Left-Hand Side (LHS) and products on the Right-Hand Side (RHS). Using chemical formulae instead of words makes the equation concise and useful.

2. Skeletal vs. Balanced Equations:

- A *skeletal* chemical equation is unbalanced (mass not same on both sides).
- A *balanced* chemical equation obeys the Law of Conservation of Mass (total mass and number of atoms remain the same before and after the reaction).

3. Balancing Method:

Balancing is done by the hit-and-trial method.

4. Informativeness (Physical States):

Physical states are mentioned to make equations more informative:

(g) – gas, (l) – liquid, (aq) – aqueous solution, (s) – solid.

Reaction conditions (like temperature, pressure, catalyst) are written above or below the arrow.

1.2 Types of Chemical Reactions

1. Combination Reactions

A single product is formed from two or more reactants (elements or compounds).

- **Exothermic Reactions:** Heat is released during the reaction.
Example: Calcium oxide reacts with water to form calcium hydroxide (slaked lime).
Respiration is also an exothermic process.

2. Decomposition Reactions

A single substance breaks down into two or more simpler substances. These are generally endothermic.

- **Thermal Decomposition:** Decomposition by heating (e.g., limestone, ferrous sulphate).
 - **Electrolytic Decomposition:** Decomposition by electricity (e.g., electrolysis of water).
 - **Photolytic Decomposition:** Decomposition by light (e.g., silver chloride turning grey in sunlight).
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3. Displacement and Double Displacement Reactions

- **Displacement Reaction:**
A more reactive element displaces a less reactive element from its compound.
(e.g., Zinc displaces copper from copper sulphate.)
 - **Double Displacement Reaction:**
Exchange of ions between two compounds.
 - **Precipitation Reaction:**
Formation of an insoluble solid (precipitate). Example: Reaction between sodium sulphate and barium chloride.
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4. Oxidation and Reduction (Redox Reactions)

- **Oxidation:** Gain of oxygen or loss of hydrogen.
 - **Reduction:** Loss of oxygen or gain of hydrogen.
 - **Redox Reaction:** Both oxidation and reduction occur simultaneously.
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1.3 Effects of Oxidation in Everyday Life

1. Corrosion:

Metals are attacked by substances like moisture or acids.

- Iron → rust (reddish-brown)
- Silver → black coating
- Copper → green coating

2. Rancidity:

Fats and oils get oxidised, changing their smell and taste.

- Prevented by using antioxidants or flushing with nitrogen gas (as in chip packets).
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Important Reactions

Combination Reactions

1. Formation of Slaked Lime (Exothermic):
 $\text{CaO (s)} + \text{H}_2\text{O (l)} \rightarrow \text{Ca(OH)}_2 \text{ (aq)} + \text{Heat}$
 2. Reaction of Slaked Lime with Air (CO_2):
 $\text{Ca(OH)}_2 \text{ (aq)} + \text{CO}_2 \text{ (g)} \rightarrow \text{CaCO}_3 \text{ (s)} + \text{H}_2\text{O (l)}$
 3. Burning of Coal:
 $\text{C (s)} + \text{O}_2 \text{ (g)} \rightarrow \text{CO}_2 \text{ (g)}$
 4. Formation of Water:
 $2\text{H}_2 \text{ (g)} + \text{O}_2 \text{ (g)} \rightarrow 2\text{H}_2\text{O (l)}$
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Exothermic Reactions

1. Burning of Natural Gas:
 $\text{CH}_4 \text{ (g)} + 2\text{O}_2 \text{ (g)} \rightarrow \text{CO}_2 \text{ (g)} + 2\text{H}_2\text{O (g)}$
 2. Respiration:
 $\text{C}_6\text{H}_{12}\text{O}_6 \text{ (aq)} + 6\text{O}_2 \text{ (aq)} \rightarrow 6\text{CO}_2 \text{ (aq)} + 6\text{H}_2\text{O (l)} + \text{Energy}$
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Decomposition Reactions

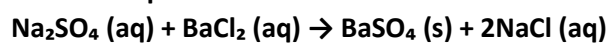
1. Thermal Decomposition of Ferrous Sulphate:
 $2\text{FeSO}_4 \text{ (s)} \rightarrow \text{Fe}_2\text{O}_3 \text{ (s)} + \text{SO}_2 \text{ (g)} + \text{SO}_3 \text{ (g)}$
 2. Thermal Decomposition of Limestone:
 $\text{CaCO}_3 \text{ (s)} \rightarrow \text{CaO (s)} + \text{CO}_2 \text{ (g)}$
 3. Thermal Decomposition of Lead Nitrate:
 $2\text{Pb(NO}_3)_2 \text{ (s)} \rightarrow 2\text{PbO (s)} + 4\text{NO}_2 \text{ (g)} + \text{O}_2 \text{ (g)}$
 4. Photolytic Decomposition of Silver Chloride:
 $2\text{AgCl (s)} \rightarrow 2\text{Ag (s)} + \text{Cl}_2 \text{ (g)}$
 5. Photolytic Decomposition of Silver Bromide:
 $2\text{AgBr (s)} \rightarrow 2\text{Ag (s)} + \text{Br}_2 \text{ (g)}$
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Displacement Reactions

1. Iron displacing Copper:
 $\text{Fe (s)} + \text{CuSO}_4 \text{ (aq)} \rightarrow \text{FeSO}_4 \text{ (aq)} + \text{Cu (s)}$
2. Zinc displacing Copper:
 $\text{Zn (s)} + \text{CuSO}_4 \text{ (aq)} \rightarrow \text{ZnSO}_4 \text{ (aq)} + \text{Cu (s)}$
3. Lead displacing Copper:
 $\text{Pb (s)} + \text{CuCl}_2 \text{ (aq)} \rightarrow \text{PbCl}_2 \text{ (aq)} + \text{Cu (s)}$

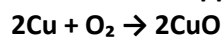
Double Displacement and Precipitation Reaction

1. Sodium Sulphate and Barium Chloride:

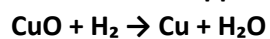


Oxidation and Reduction (Redox) Examples

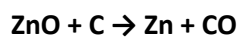
1. Oxidation of Copper:



2. Reduction of Copper Oxide:



3. Zinc Oxide and Carbon:



4. Manganese Dioxide and Hydrochloric Acid:

