

**Dev Academy**  
Yamunanagar (Haryana)

# Chapter-1

## Chemical Reactions and Equations

Class  
10

Revision Notes

# SCIENCE

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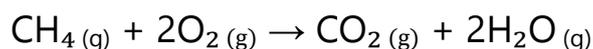


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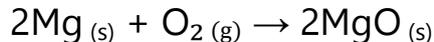
# Part – 1

## Chemical Equations

- Chemical equations are symbolic representations of chemical reactions.
- The general form of a chemical equation is:  
Reactants → Products
- Reactants are the substances present at the start of the reaction, and products are the substances formed as a result of the reaction.
- The arrow represents the direction of the reaction, indicating the conversion of reactants into products.
- Chemical equations follow the **law of conservation of mass**, which states that matter is neither created nor destroyed during a chemical reaction. Therefore, the number and types of atoms must be balanced on both sides of equation.
- The common method used for balancing is **Hit and Trial method**.
  - For example, the combustion of methane can be represented as:



- The chemical equation for burning of magnesium can be represented as:

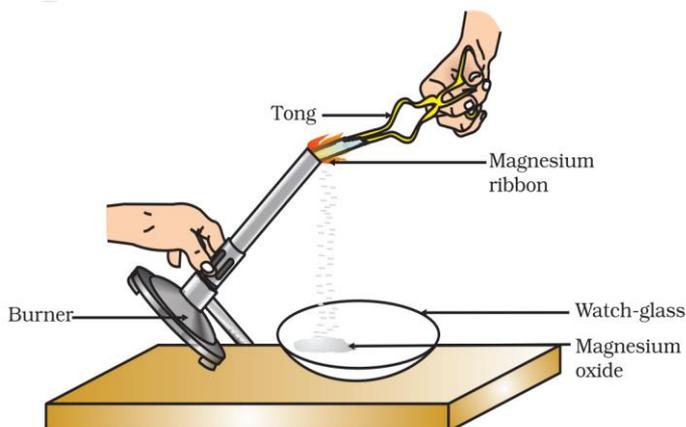


The reaction is exothermic, releasing energy in the form of heat and light.

- Chemical equations can also include additional information such as state symbols (s = solid, l = liquid, g = gas, aq = aqueous) to indicate the physical states of the substances involved, as well as reaction conditions such as temperature, pressure, and catalysts.

### Some common chemical reactions

- **Burning of Magnesium ribbon**

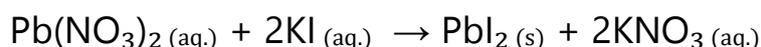


- When the magnesium ribbon is ignited, it starts to burn with a dazzling white flame, giving rise to a white powder known as magnesium oxide (MgO).
- Chemical equation for the burning of magnesium can be represented as:
$$2\text{Mg}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{MgO}_{(s)}$$
- Magnesium ribbon is often rubbed with sandpaper before burning to remove the thin layer of magnesium oxide (MgO) that forms on its surface over time. Rubbing the ribbon with sandpaper exposes fresh, unoxidized magnesium metal, which readily reacts with oxygen when ignited.

- **Reaction between Lead nitrate and Potassium iodide**

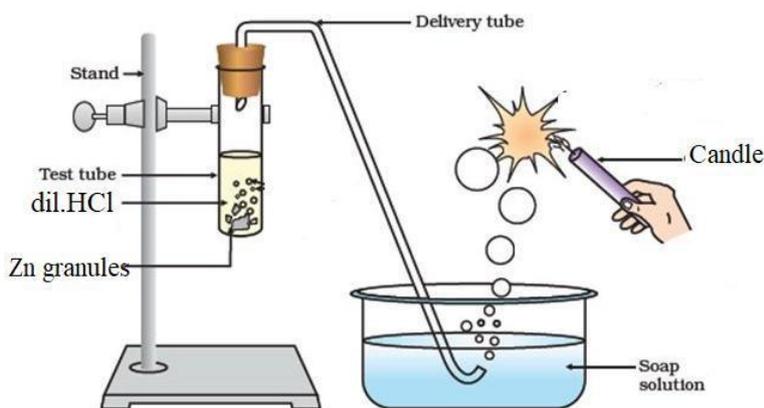


- Imagine two colourless solutions in separate containers. In one container, there is a solution of lead nitrate,  $\text{Pb}(\text{NO}_3)_2$ , and in the other container, there is a solution of potassium iodide, KI.
- When the two solutions are mixed together, a yellow precipitate of lead iodide ( $\text{PbI}_2$ ) starts to form in the solution. The yellow precipitate gradually settles at the bottom of the container.
- The chemical equation for the reaction between lead nitrate and potassium iodide can be represented as:



- The reaction between lead nitrate and potassium iodide is a double displacement reaction.

## • Reaction between Zinc metal and Hydrochloric acid



- As the zinc metal is placed into the test tube of hydrochloric acid, bubbles of  $H_2$  gas start to form and rise to the surface of the liquid.
- The chemical equation for the reaction between zinc metal (Zn) and hydrochloric acid (HCl) can be represented as:
$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$
- To confirm the presence of hydrogen gas, you can perform "**pop test**." Follow these steps:
  - Pass the gas formed during the reaction through the soap solution.
  - Bring the soap bubbles 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.

### NCERT Questions

1. Why should a magnesium ribbon be cleaned before burning in air?
2. Write the balanced equation for the following chemical reactions.
  - (i) Hydrogen + Chlorine  $\rightarrow$  Hydrogen chloride
  - (ii) Barium chloride + Aluminium sulphate  $\rightarrow$  Barium sulphate + Aluminium chloride
  - (iii) Sodium + Water  $\rightarrow$  Sodium hydroxide + Hydrogen
3. Write a balanced chemical equation with state symbols for the following reactions.
  - (i) Solutions of barium chloride and sodium sulphate in water react to give insoluble barium sulphate and the solution of sodium chloride.
  - (ii) Sodium hydroxide solution (in water) reacts with hydrochloric acid solution (in water) to produce sodium chloride solution and water.

## Part – 2

# Types of Chemical Reactions

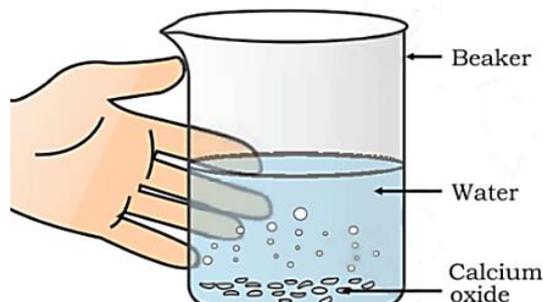
### 2.1 Combination Reaction

- Combination reactions are chemical reactions where two or more substances (reactant molecules) combine to form a single, more complex product. Generally represented as:  $A + B \rightarrow AB$

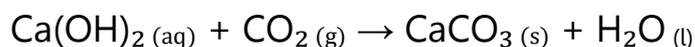
- Examples:

- $2\text{Na}_{(s)} + \text{Cl}_{2(g)} \rightarrow 2\text{NaCl}_{(s)}$
- $2\text{Mg}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{MgO}_{(s)}$
- $\text{CaO}_{(s)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{Ca(OH)}_{2(aq)}$
- $\text{CO}_{2(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{H}_2\text{CO}_{3(aq)}$  (Carbonic acid)
- $\text{N}_{2(g)} + 3\text{H}_{2(g)} \rightarrow 2\text{NH}_{3(g)}$  (Ammonia)
- $\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)}$

- **Reaction between calcium oxide (CaO) and water (H<sub>2</sub>O)**



- Calcium oxide (quicklime; CaO) reacts with water (H<sub>2</sub>O) to form calcium hydroxide (slaked lime; Ca(OH)<sub>2</sub>). The reaction is **exothermic**.
  - The chemical equation for the reaction between calcium oxide and water can be represented as:
- $$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$$
- A solution of Ca(OH)<sub>2</sub> (slaked lime) produced in above reaction is used for **whitewashing** walls. Calcium hydroxide reacts slowly with the carbon dioxide (CO<sub>2</sub>) in air to form a thin layer of calcium carbonate (CaCO<sub>3</sub>) on the walls. Calcium carbonate is formed after two to three days of whitewashing and gives a shiny finish to the walls.

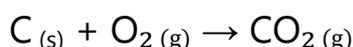


- It is interesting to note that the chemical formula for marble is also CaCO<sub>3</sub>.

- **Burning of Coal (C)**



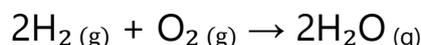
- During the burning of coal, the carbon present in coal undergoes oxidation by oxygen present in air.
- The reaction is **exothermic**.
- The combustion continues as long as there is a sufficient supply of oxygen.
- The chemical equation for the burning of coal can be represented in a simplified form as:



- **Burning of hydrogen (H<sub>2</sub>)**



- The reaction between hydrogen (H<sub>2</sub>) and oxygen (O<sub>2</sub>) is a highly **exothermic** chemical reaction, resulting in the formation of water (H<sub>2</sub>O).
- The balanced chemical equation for the reaction between H<sub>2</sub> and O<sub>2</sub> is:



- The combustion of hydrogen with oxygen is not only a chemical reaction but also a source of clean energy because the only byproduct of its combustion is water (H<sub>2</sub>O).

- **Burning of Magnesium ribbon (Mg)**

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## 2.1.2 Exothermic Nature of Combination Reactions

- Exothermic reactions are chemical reactions that release energy in the form of heat to the surroundings.
- Many combination reactions tend to be **exothermic**. However, it is important to note that not all combination reactions are necessarily exothermic.
- Some common examples of exothermic reactions include:
  - Combustion Reactions: (Burning of natural gas, fossil fuel, and wood etc.)
 
$$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$$
  - Neutralization Reactions: (Reaction between an acid and a base)
 
$$\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$$
  - Oxidation of glucose during cellular respiration in living organisms.
    - The balanced equation for aerobic respiration is as follows:
 
$$\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} + \text{Energy (ATP)}$$
    - The release of energy during respiration can be observed in the form of heat when an organisms respire.
  - The decomposition of vegetable matter into compost is an example of an exothermic reaction. This process is carried out by the action of microorganisms like bacteria and fungi.

## 2.2 Decomposition Reaction

- Decomposition reactions are chemical reactions in which a single reactant molecule breaks down into two or more simpler product molecules. Generally represented as:  $\text{AB} \rightarrow \text{A} + \text{B}$
- These reactions generally occur when the reactant compound is unstable.
- Decomposition reactions can occur due to various factors, such as heat, light, electric current or the presence of catalysts.
- Examples:
  - $2\text{FeSO}_4(\text{s}) \xrightarrow{\text{Heat}} \text{Fe}_2\text{O}_3(\text{s}) + \text{SO}_2(\text{g}) + \text{SO}_3(\text{g})$
  - $\text{CaCO}_3(\text{s}) \xrightarrow{\text{Heat}} \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
  - $2\text{Pb}(\text{NO}_3)_2(\text{s}) \xrightarrow{\text{Heat}} 2\text{PbO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$
  - $2\text{AgBr}(\text{s}) \xrightarrow{\text{Sunlight}} 2\text{Ag}(\text{s}) + \text{Br}_2(\text{g})$
  - $2\text{H}_2\text{O}_2(\text{l}) \xrightarrow{\text{KI}} 2\text{H}_2\text{O}(\text{g}) + \text{O}_2(\text{g})$
  - $2\text{KClO}_3(\text{s}) \xrightarrow{\text{Heat}} 2\text{KCl}(\text{s}) + 3\text{O}_2(\text{g})$

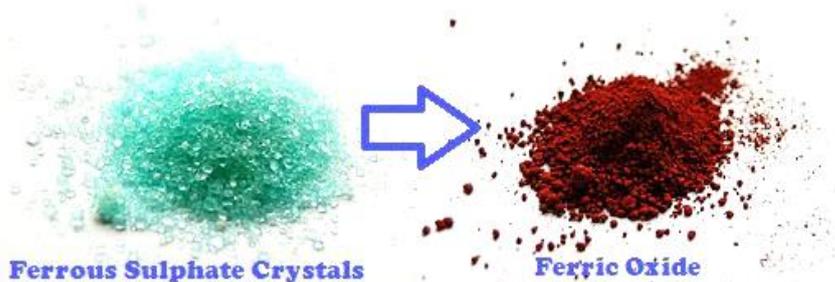
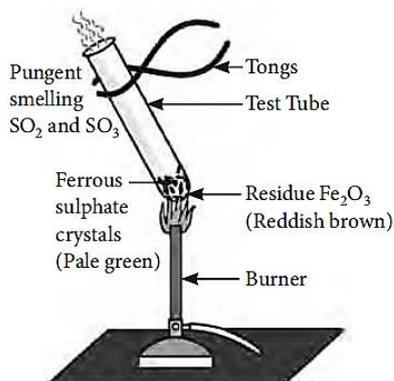
## 2.2.1 Types of Decomposition Reactions

### (a) Thermal Decomposition Reaction

This type of decomposition reaction occurs when a reactant compound breaks down upon heating.

Example:

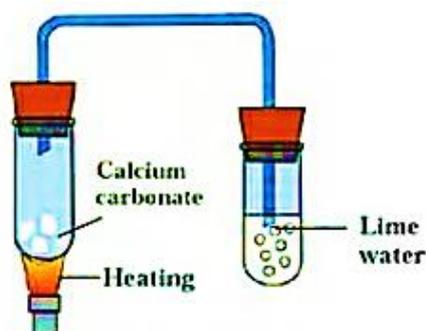
- **Thermal Decomposition of Ferrous sulphate ( $\text{FeSO}_4$ ):**



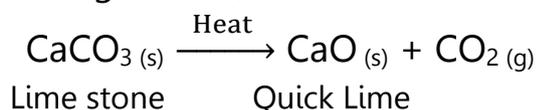
- Initially, the ferrous sulphate crystals are pale green in colour. However, upon heating, they gradually turn into a reddish-brown colour.
- This change in colour is due to the formation of ferric oxide ( $\text{Fe}_2\text{O}_3$ ).
- As the ferrous sulphate ( $\text{FeSO}_4$ ) crystals decompose, the reaction releases  $\text{SO}_2$  and  $\text{SO}_3$  gases. These gases have pungent or suffocating smell.
- This reaction can be represented by the following equation:



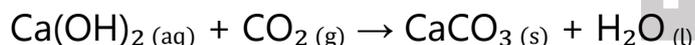
- **Thermal Decomposition of Calcium carbonate ( $\text{CaCO}_3$ ):**



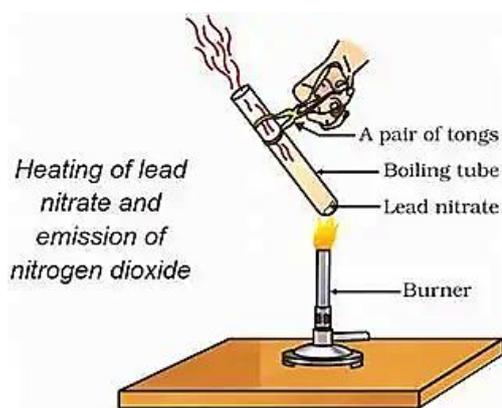
- Upon heating, calcium carbonate ( $\text{CaCO}_3$ ) breaks down into calcium oxide ( $\text{CaO}$ ), and carbon dioxide gas ( $\text{CO}_2$ ).



- To prove that carbon dioxide gas is produced during the decomposition of calcium carbonate, we can perform a simple experiment.
  - ✓ Prepare a solution of lime water ( $\text{Ca}(\text{OH})_2$ ). Now, pass the gas generated during the reaction through this lime solution.
  - ✓ As the gas enters the lime water, it undergoes a chemical reaction with  $\text{Ca}(\text{OH})_2$  to form  $\text{CaCO}_3$ . As a result, lime solution turns milky.



- **Thermal Decomposition of Lead nitrate ( $\text{Pb}(\text{NO}_3)_2$ ):**



- Upon Heating, lead nitrate breaks down into three simpler substances: lead(II) oxide ( $\text{PbO}$ ), nitrogen dioxide gas ( $\text{NO}_2$ ), and oxygen gas ( $\text{O}_2$ ).



- Nitrogen dioxide is a reddish-brown gas with a pungent odor.
- Oxygen is a colourless and odorless gas that supports combustion.
- The production of nitrogen dioxide and oxygen gases can be visually represented by observing the change in colour and the release of fumes during the heating of lead nitrate.

**(b) Photo-decomposition Reaction**

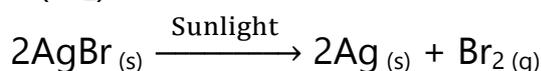
Photodecomposition is a type of decomposition reaction that occurs in the presence of sunlight energy.

Example:

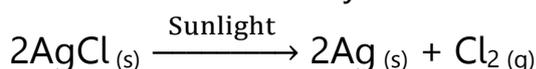
- **Photo-decomposition of silver bromide ( $\text{AgBr}$ ):**
  - Silver bromide is a photosensitive compound.
  - Initially, silver bromide appears as a yellowish-white solid.



- However, upon exposure to light (blue and Uv regions of the spectrum), the silver bromide dissociates or breakdown into black metallic silver (Ag) and elemental bromine gas (Br<sub>2</sub>).



- The photodecomposition of silver bromide plays a crucial role in black and white photography.
- Silver chloride also behave in the same way.

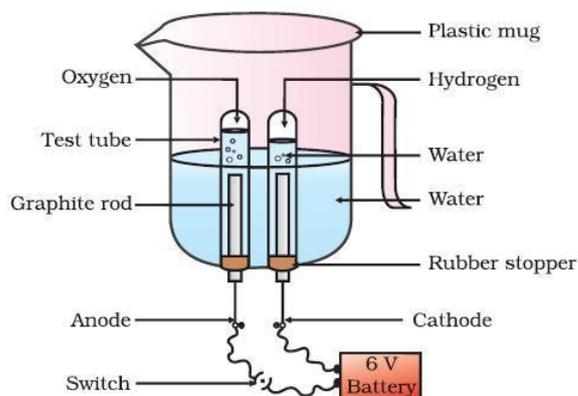


### (c) Electrolytic Decomposition Reaction

Electrolytic decomposition is a process in which an electric current is passed through an electrolyte solution, leading to the decomposition of the compound into its constituent elements or ions.

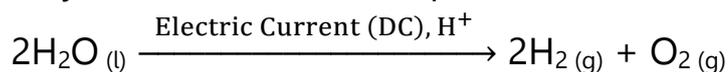
Example:

- **Electrolytic Decomposition of Water (H<sub>2</sub>O):**



- When electrical current (DC) passes through the acidified water, the water (H<sub>2</sub>O) is decomposed into hydrogen gas (H<sub>2</sub>) and oxygen gas (O<sub>2</sub>).
- Oxygen gas (O<sub>2</sub>) is evolved at the anode (+) and hydrogen gas (H<sub>2</sub>) is evolved at the cathode (-).

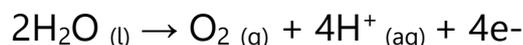
- Overall, the electrolysis of water can be represented as follows:



- The reactions that occur at each electrode are as follows:
  - At cathode (–): The hydrogen gas is evolved at the cathode.

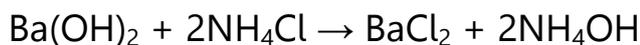


- At anode (+): The oxygen gas is evolved at the anode.



## 2.2.2 Endothermic Nature of Decomposition Reactions

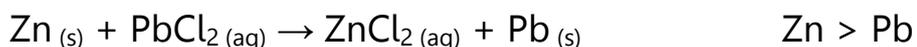
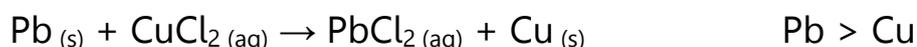
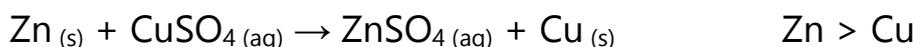
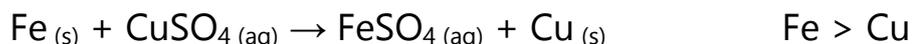
- Reactions in which energy is absorbed are known as endothermic reactions.
- Decomposition reactions are endothermic because they require an input of energy, whether it be in the form of heat, light, or electricity in order to break-down the reactants.
- Other example of endothermic reaction:
  - The reaction between barium hydroxide and ammonium chloride is an endothermic reaction.



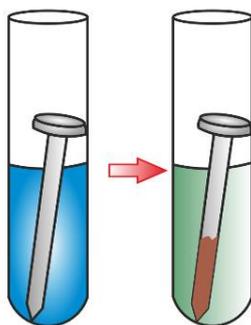
Upon touching the bottom of the test tube, you may feel a sensation of coldness or a decrease in temperature.

## 2.3 Displacement Reaction

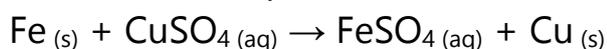
- Displacement or replacement reactions are a type of chemical reaction in which a more reactive element displaces a less reactive element from its compound.
- Generally represented as: **A + BC → AC + B**
- Here is a simplified reactivity series of elements, arranged in order of decreasing reactivity:  
**K > Na > Ca > Mg > Al > Zn > Fe > Pb > Hydrogen (H) > Cu > Ag > Au**
- Examples:



- **Iron (Fe) displaces copper (Cu) from copper sulphate (CuSO<sub>4</sub>) solution**

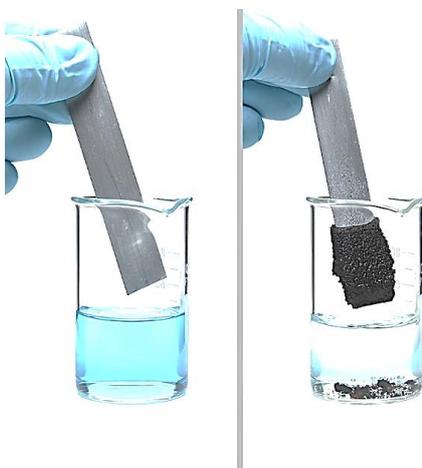


- When an iron nail is added to a **blue** solution of copper sulphate (CuSO<sub>4</sub>), iron displaces copper from the copper sulphate solution. This leads to the formation of **light green** ferrous sulphate (FeSO<sub>4</sub>) and solid copper (Cu).

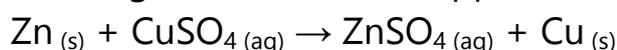


- As the reaction continues, the blue colour fades because the concentration of CuSO<sub>4</sub> decreases, while a light green colour develops due to increase in the concentration of FeSO<sub>4</sub>.
- This reaction demonstrates the higher reactivity of Fe compared to Cu.

- **Zinc displaces copper from copper sulphate solution**



- When zinc is added to aqueous copper sulphate solution, zinc displaces copper from the copper sulphate solution. This results in the formation of aqueous zinc sulphate (ZnSO<sub>4</sub>) and solid copper (Cu).
- The blue colour of the copper sulphate solution fades away as copper is displaced by zinc, indicating the removal of copper ions from the solution.

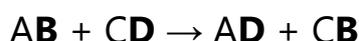


## 2.4 Double Displacement Reaction

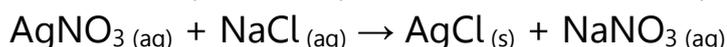
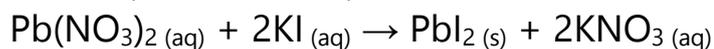
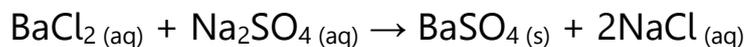
- Double displacement reactions occur when two different compounds exchange their ions when their aqueous solutions are mixed and form new compounds.

Example: Precipitation reactions and Neutralization reactions.

- **Precipitation reactions** involve the formation of a precipitate, which is an insoluble solid that separates from the solution.
- The formation of a precipitate is due to the limited solubility of certain compounds in water, causing it to separate as a solid.
- The general format of a double displacement reaction can be represented as:



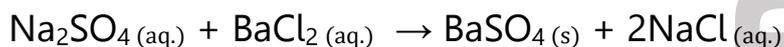
- Example;



- **Reaction between Sodium sulphate ( $Na_2SO_4$ ) and Barium chloride ( $BaCl_2$ )**



- When the two colourless solutions, sodium sulphate ( $Na_2SO_4$ ) and barium chloride ( $BaCl_2$ ) are mixed, a white precipitate of barium sulphate ( $BaSO_4$ ) starts to form in the solution and gradually settles at the bottom of the container while sodium chloride ( $NaCl$ ) remain in solution.
- The chemical equation for the reaction between Sodium sulphate and Barium chloride can be represented as:



- **Reaction between Lead nitrate ( $Pb(NO_3)_2$ ) and Potassium iodide (KI)**

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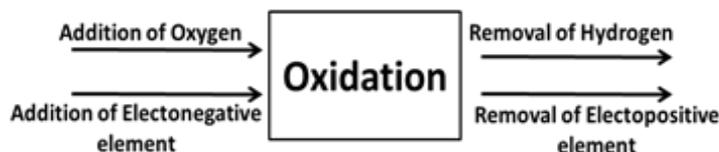
## 2.5 Oxidation and Reduction Reaction

- Oxidation and reduction reactions, commonly known as redox reactions.
- Oxidation and reduction always occur simultaneously and are interconnected.

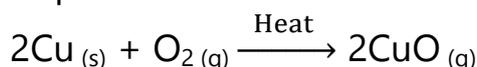
### 2.5.1 Classical Concept of Oxidation and Reduction

- **Oxidation:**

- The chemical reaction which involves in gain of oxygen or electronegative elements and loss of hydrogen or electropositive elements is called oxidation.



- When copper is heated in air, it reacts with oxygen present in the air to form black coloured copper oxide on its surface through oxidation reaction. The reaction can be represented as follows:

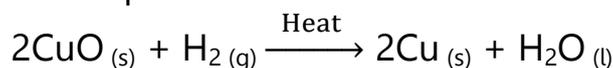


- **Reduction:**

- The chemical reaction which involves in gain of hydrogen or electropositive elements and loss of oxygen or electronegative elements is called reduction.



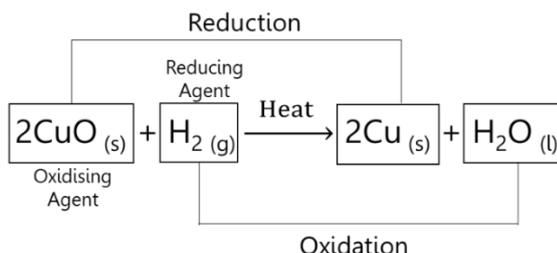
- When hydrogen gas is passed over heated black coloured copper(II) oxide, it turns brown due to formation of copper through reduction reaction. The reaction can be represented as follows:



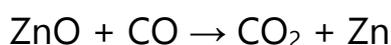
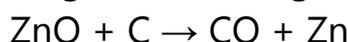
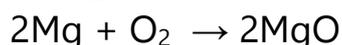
- **Oxidising Agent and Reducing Agent:**

- Oxidising agents or Oxidants are the chemical substances which oxidise the other chemical substance but itself undergo in reduction.
- Reducing agents or Reductants are the chemical substances which reduce the other chemical substance but itself undergo in oxidation.

- Reduction and oxidation are interconnected processes that occur simultaneously.



- Some other examples of redox reactions are:



## 2.5.2 Electronic Transfer Concept of Oxidation and Reduction

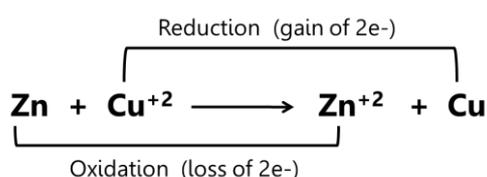
- Oxidation involves the loss of electrons. Hence, oxidation state of the reactant molecule increases.



- Reduction involves the gain of electrons. Hence, oxidation state of the reactant molecule decreases.



- Examples of Redox reaction:



## NCERT Questions

- Why does the colour of copper sulphate solution change when an iron nail is dipped in it?
- Give an example of a double displacement reaction.
- Identify the substances that are oxidised and the substances that are reduced in the following reactions.
  - $4\text{Na}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{Na}_2\text{O}_{(s)}$
  - $\text{CuO}_{(s)} + \text{H}_{2(g)} \rightarrow \text{Cu}_{(s)} + \text{H}_2\text{O}_{(l)}$

## Part – 3

### Effects of Oxidation Reaction in Everyday Life

#### 3.1 Rancidity

- Rancidity refers to the process of deterioration or spoilage of foods containing fats and oils, resulting in undesirable changes in taste, odour, and texture. It is caused by the oxidation of fats and oils.
- To prevent the rancidity, antioxidants are often added to foods that contain fats and oils. These antioxidants inhibit the oxidation reactions. Commonly used antioxidants are BHA (Butylated hydroxy anisole) and BHT (Butylated hydroxy toluene)
- An interesting example of preventing oxidation can be seen in the production of potato chips. Manufacturers flush the bags of chips with gases such as nitrogen. By replacing the air inside the bag with nitrogen, the oxygen content is reduced, minimizing the chances of oxidation, and ensuring that the chips remain crispy and free from the undesirable effects of rancidity.
- The process of rancidity can be slowed down by storing food items in airtight containers, thereby minimizing their exposure to air or oxygen.

#### 3.2 Corrosion

- Corrosion is a natural process that occurs when metals react with substances in their environment, such as moisture, gases, or chemicals.
- Corrosion typically results in the formation of oxides, hydroxides, or sulphides on metal's surface, leading to change in its physical and chemical properties.
- You may have noticed that when iron articles are new, they have a shiny appearance. However, over time, they develop a reddish-brown powder coating. This process is known as rusting of iron.
- Rusting is a specific type of corrosion that occurs when iron or steel reacts with oxygen in the presence of water or moisture, resulting in the formation of iron oxide (commonly known as rust).
- Additionally, some other metals can also undergo a similar process of tarnishing called corrosion.
- If you observe copper and silver objects, you may notice a black coating on silver and the green coating on copper, are examples of corrosion.

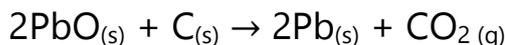
- Black coating on silver is caused by the formation of silver sulphide ( $\text{Ag}_2\text{S}$ ) layer on the surface of silver due to the reaction between silver and sulphur compounds present in air (i.e.,  $\text{H}_2\text{S}$  and  $\text{SO}_2$  etc.)
- The green coating on copper is caused by the formation of layer of copper carbonate, also known as verdigris ( $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$ ).
- Various methods are employed to protect metals from corrosion. Here are some commonly used methods:
  - 1.) **Applying a protective layer** on the surface of the metal is an effective way to prevent corrosion. This can be achieved through methods such as painting, electroplating, or hot-dip galvanizing. The coating acts as a barrier, preventing direct contact between the metal and corrosive agents in the environment.
  - 2.) **Galvanization** is a widely used method for protecting metals, particularly iron or steel, from corrosion. It involves coating the metal with a layer of zinc to provide a protective barrier against environmental factors that can cause rusting and deterioration.
  - 3.) **Anodising** is an electrochemical process in which a metal surface is protected from corrosion by depositing a thick metal oxide layer. Anodising is usually done for non-ferrous metals or alloys. The metals for which anodising can be done are Aluminium, Magnesium, Titanium etc.

### Note

- Several metals exhibit a protective nature against corrosion due to their inherent properties. These metals form a protective layer of metal oxide on their surface, preventing further corrosion.  
Here are some examples: Al, Cr, Mg form their oxides  $\text{Al}_2\text{O}_3$ ,  $\text{Cr}_2\text{O}_3$ ,  $\text{MgO}$  respectively on their surface to prevent further corrosion.

## NCERT Exercise Questions

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



- (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.  $\text{Fe}_2\text{O}_3 + 2\text{Al} \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}$   
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)  $\text{HNO}_3 + \text{Ca}(\text{OH})_2 \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}$
  - (b)  $\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$
  - (c)  $\text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3$
  - (d)  $\text{BaCl}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + \text{HCl}$

7. Write the balanced chemical equations for the following reactions.
- Calcium hydroxide + Carbon dioxide → Calcium carbonate + Water
  - Zinc + Silver nitrate → Zinc nitrate + Silver
  - Aluminium + Copper chloride → Aluminium chloride + Copper
  - Barium chloride + Potassium sulphate → Barium sulphate + Potassium chloride
8. Write the balanced chemical equation for the following and identify the type of reaction in each case.
- Potassium bromide<sub>(aq)</sub> + Barium iodide<sub>(aq)</sub> → Potassium iodide<sub>(aq)</sub> + Barium bromide<sub>(s)</sub>
  - Zinc carbonate<sub>(s)</sub> → Zinc oxide<sub>(s)</sub> + Carbon dioxide<sub>(g)</sub>
  - Hydrogen<sub>(g)</sub> + Chlorine<sub>(g)</sub> → Hydrogen chloride<sub>(g)</sub>
  - Magnesium<sub>(s)</sub> + Hydrochloric acid<sub>(aq)</sub> → Magnesium chloride<sub>(aq)</sub> + Hydrogen<sub>(g)</sub>
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.
- Oxidation
  - 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.
- Corrosion
  - Rancidity



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