What is a Chemical Reaction?

- A chemical reaction happens when a substance changes into a new substance with different properties.
- This is different from a physical change, where the substance's form might change, but its chemical makeup stays the same (like ice melting into water).

How do we know a chemical reaction has occurred?

Several clues can indicate a chemical reaction:

- **Change in state:** A substance might change from solid to liquid, liquid to gas, etc.
- **Change in color:** The original substance might change color.
- **Evolution of a gas:** Bubbles or fizzing might indicate a new gas is being produced.
- **Change in temperature:** The reaction might release heat (get warmer) or absorb heat (get colder).

Examples of Chemical Reactions in Daily Life:

- **Milk spoiling:** Bacteria in milk cause a chemical change, making it sour and curdled.
- **Iron rusting:** Iron reacts with oxygen and water in the air to form iron oxide (rust).
- **Grapes fermenting:** Yeast converts the sugar in grapes into alcohol and carbon dioxide.
- **Food cooking:** Heat causes chemical changes in the food, altering its texture, taste, and composition.
- **Food digesting:** Your body uses enzymes to break down food into simpler substances.
- **Respiration:** You breathe in oxygen and your body uses it to break down sugars, releasing energy, carbon dioxide, and water.

Activities to Observe Chemical Reactions:

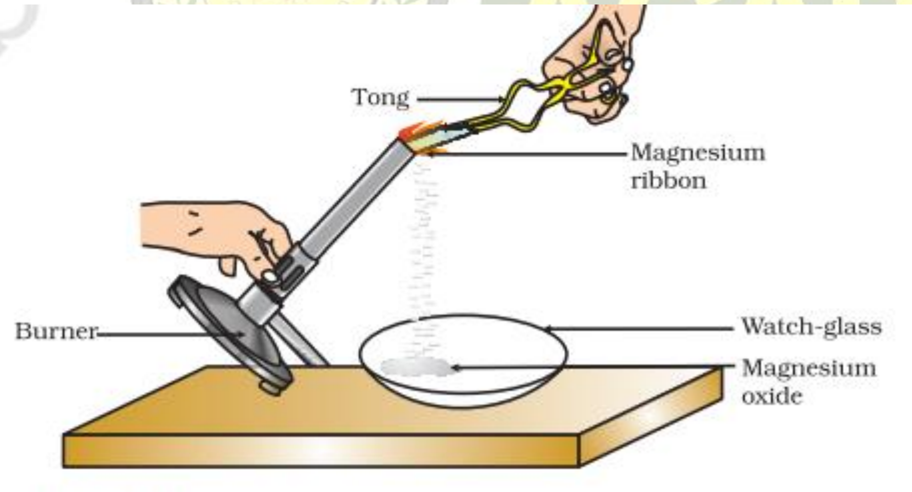


Figure 1.1
Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch-glass

- **Burning magnesium:**
  - Magnesium ribbon burns in air, producing a bright white light and leaving behind a white powder (magnesium oxide). This shows a change in color, state, and temperature.

- **Zinc and acid:**
  - Zinc granules react with acid to produce bubbles of hydrogen gas. This shows the evolution of a gas and a possible change in temperature.

Chemical Equations: A Summary

- **Word Equations:**
  - Describe a chemical reaction using the names of the reactants and products.
  - Example: Magnesium + Oxygen → Magnesium oxide
  - Reactants are on the left side of the arrow, products on the right.
  - The arrow shows the direction of the reaction.
- **Chemical Equations:**
  - Use chemical formulas instead of words to represent reactions.
  - Example: $Mg + O2 \rightarrow MgO$
  - This is a more concise way to show a reaction.
- **Balanced Equations:**
  - A balanced chemical equation has the same number of atoms of each element on both sides of the arrow.
  - This is important because it reflects the Law of Conservation of Mass – matter cannot be created or destroyed in a chemical reaction.
- **Skeletal Equations:**
  - Chemical equations that are not balanced are called skeletal equations.
  - Example: $Mg + O2 \rightarrow MgO$ (This equation is not balanced because there are two oxygen atoms on the left and only one on the right).

Key Points:

- Chemical equations are a shorthand way to represent chemical reactions.
- Balanced equations are essential for accurately describing chemical reactions.
- We need to balance equations to ensure they obey the Law of Conservation of Mass.

Balancing Equation (1.2)

Now, let's balance the equation from earlier: $Mg + O\ 2 \rightarrow MgO$

1. **Draw boxes:** [Mg] + [O 2] → [MgO]
2. **Count atoms:**
   - Mg: 1 on both sides
   - O: 2 on the left, 1 on the right
3. **Balance oxygen:** Add a coefficient of 2 in front of MgO: $Mg + O\ 2 \rightarrow 2\ MgO$
4. **Balance magnesium:** Add a coefficient of 2 in front of Mg: $2\ Mg + O\ 2 \rightarrow 2\ MgO$
5. **Final check:** The equation is now balanced!

Questions on Chemical Reactions and Equations

1. Magnesium Ribbon Cleaning

- **Why?** Magnesium ribbon is often covered with a layer of magnesium oxide formed by its reaction with oxygen in the air. This layer can prevent the magnesium from burning properly. Cleaning it with sandpaper removes the oxide layer, allowing for a clean and efficient reaction.

2. Balanced Chemical Equations

Write balanced equations for these reactions:

- **(i) Hydrogen + Chlorine → Hydrogen chloride**
  - $H_2 + Cl_2 \rightarrow 2\,HCl$
- **(ii) Barium chloride + Aluminium sulphate → Barium sulphate + Aluminium chloride**
  - $3\,BaCl_2 + Al_2(SO_4)_3 \rightarrow 3\,BaSO_4 + 2\,AlCl_3$
- **(iii) Sodium + Water → Sodium hydroxide + Hydrogen**
  - $2\,Na + 2\,H_2O \rightarrow 2\,NaOH + H_2$

3. Balanced Equations with State Symbols

Write balanced equations with state symbols for these reactions:

- **(i) Solutions of barium chloride and sodium sulphate in water react to give insoluble barium sulphate and sodium chloride solution.**
  - $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,NaCl(aq)$
- **(ii) Sodium hydroxide solution (in water) reacts with a hydrochloric acid solution (in water) to produce sodium chloride solution and water.**
  - $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

1.2 Types of Chemical Reactions

- **What happens in a chemical reaction?**
  - Atoms of one element don't change into atoms of another element.
  - Atoms don't disappear or appear out of nowhere.
  - Chemical bonds between atoms are broken and new bonds are formed to create new substances.

1.2.1 Combination Reactions

- **Definition:** Two or more substances combine to form a single new substance.
- **Example:** Calcium oxide (quicklime) reacts with water to produce calcium hydroxide (slaked lime).
  - **Word equation:** Calcium oxide + Water → Calcium hydroxide
  - **Chemical equation:** $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$
- **Observation:** The reaction releases heat (exothermic), making the beaker hot.

- **Applications:** This reaction is used to make mortar and cement.

Key Point: Think of combination reactions like baking a cake. You combine separate ingredients (flour, eggs, sugar, etc.) to create a single product (the cake).

Combination Reaction: Calcium Oxide and Water

- **Reactants:**
    - Calcium oxide (CaO) - also known as quicklime.
    - Water (H_2O)
- **Product:**
    - Calcium hydroxide ($Ca(OH)_2$) - also known as slaked lime.
- **Reaction:** $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$ + Heat
- **Type:** Combination reaction (two reactants combine to form one product).
- **Energy Change:** The reaction releases much heat (exothermic reaction).
- **Observation:** You would observe the beaker getting hot as the reaction proceeds.

Key takeaway: This is a classic example of a combination reaction where two simple compounds react to produce a more complex compound, releasing heat in the process.

More on Combination Reactions and Exothermic Reactions

- **Whitewashing Walls:**
    - Slaked lime (calcium hydroxide) reacts with carbon dioxide in the air to form a thin layer of calcium carbon...

More on Combination Reactions & Exothermic Reactions

- **Whitewashing and Combination Reactions:**
    - Slaked lime (calcium hydroxide) is used for whitewashing.
    - It reacts with carbon diox...

More on Combination Reactions & Exothermic Reactions

- **Whitewashing Walls:**
    - Slaked lime (calcium hydroxide) reacts with carbon dioxide in the air to form a thin layer of calcium carbonat...

More on Combination Reactions & Exothermic Reactions

- **Whitewashing Walls:**

- Slaked lime (calcium hydroxide) reacts with carbon dioxide in the air to form a thin layer of calcium carbonate, giving walls a shiny finish.
- $Ca(OH)_2 (aq) + CO_2 (g) \rightarrow CaCO_3 (s) + H_2O(l)$
- Interestingly, calcium carbonate is also the chemical formula for marble!

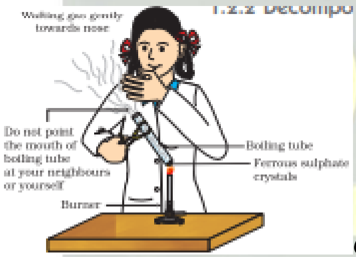
- **Examples of Combination Reactions:**
  - Burning of coal: $C(s) + O_2 (g) \rightarrow CO_2 (g)$
  - Formation of water: $2H_2 (g) + O_2 (g) \rightarrow 2H_2O(l)$
- **Exothermic Reactions:**
  - **Definition:** Reactions that release heat along with the formation of products.
  - **Examples:**
    - Burning of natural gas: $CH_4 (g) + 2O_2 (g) \rightarrow CO_2 (g) + 2H_2O(g)$
    - Respiration: $C_6H_{12}O_6 (aq) + 6O_2 (aq) \rightarrow 6CO_2 (aq) + 6H_2O(l)$ + energy
      - This is the process where glucose combines with oxygen in our cells to provide energy.
    - Decomposition of vegetable matter into compost.
- **Activity 1.1 Revisited:**
  - The burning of magnesium ribbon in Activity 1.1 is also an exothermic combination reaction.

Key Points:

- Combination reactions involve two or more substances combining to form a single product.
- Exothermic reactions release heat, making the surroundings warmer.
- Many everyday processes, like burning fuels and respiration, are exothermic reactions.

1.2.2 Decomposition Reactions

- **Definition:** A single reactant breaks down into two or more simpler products.



Correct way of heating the boiling tube containing crystals of ferrous sulphate and of smelling the odour

- **Activity 1.5 (Heating Ferrous Sulfate)**
  - **What to do:**

        1. Take ferrous sulfate crystals in a dry boiling tube.
        2. Note the initial color (green).
        3. Heat the tube over a flame.
        4. Observe the color change and any odor.
    - **What you'll observe:**
        - The green crystals change color (to reddish-brown).
        - You'll smell the odor of burning sulfur.
    - **Explanation:**
        - Ferrous sulfate ($FeSO_4 \cdot 7H_2O$) loses water when heated.
        - It then decomposes into ferric oxide (Fe_2O_3), sulfur dioxide (SO_2), and sulfur trioxide (SO_3).
        - $2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$
- **Thermal Decomposition:**
    - **Definition:** A decomposition reaction carried out by heating.
    - **Example:** Decomposition of calcium carbonate (limestone) into calcium oxide (quicklime) and carbon dioxide.
        - $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
        - This reaction is important in various industries, including cement production.

Key Point: Decomposition reactions are like taking something apart. You start with one complex substance and break it down into simpler ones.

Thermal Decomposition of Lead Nitrate

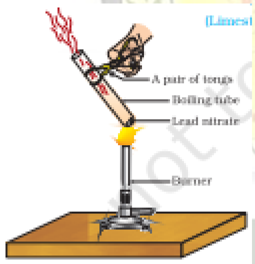


Figure 1.5
Heating of lead nitrate and emission of nitrogen dioxide

- **Activity:**
    1. Take about 2 grams of lead nitrate powder in a boiling tube.
    2. Hold the boiling tube with tongs and heat it over a flame.
- **Observation:**
    - Brown fumes are emitted. These are nitrogen dioxide (NO_2) gas.
- **Reaction:**
    - Lead nitrate decomposes into lead oxide, nitrogen dioxide, and oxygen gas.
    - $2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$

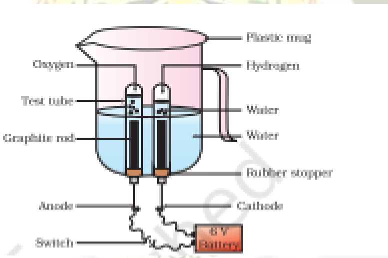
- **Type of Reaction:** Thermal decomposition (decomposition caused by heat).

Key Point: This is another example of a thermal decomposition reaction where a single compound breaks down into multiple products when heated.

1.2.2 Decomposition Reactions (Continued)

- **Activity 1.7 (Silver Chloride in Sunlight)**
  - **What to do:**
    1. Take silver chloride in a china dish.
    2. Note the initial color (white).
    3. Place the dish in sunlight.
    4. Observe the color change.
  - **What you'll observe:**
    - The white silver chloride turns grey.
  - **Explanation:**
    - Sunlight causes silver chloride to decompose into silver and chlorine.
    - $2AgCl(s) \rightarrow 2Ag(s) + Cl_2(g)$
  - **Type of Reaction:** Photochemical decomposition (decomposition caused by light).

- **Activity 1.8 (Electrolysis of Water)**



  - **What to do:**
    0. Set up an electrolysis apparatus with carbon electrodes in water containing a few drops of dilute sulfuric acid.
    1. Connect the electrodes to a 6-volt battery.
    2. Invert water-filled test tubes over the electrodes.
    3. Turn on the current and observe.
  - **What you'll observe:**
    - Bubbles form at both electrodes, displacing water in the test tubes.
    - The volume of gas collected in one test tube is twice that in the other.
  - **Explanation:**
    - Water decomposes into hydrogen and oxygen gases.
    - $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
    - The gas with the larger volume is hydrogen; the other is oxygen.

        - **Caution:** Testing the gases with a burning candle should be done carefully by the teacher, as hydrogen is flammable.
    - **Type of Reaction:** Electrolytic decomposition (decomposition caused by electricity).

Key Points:

- Decomposition reactions can be caused by heat (thermal), light (photochemical), or electricity (electrolytic).
- These activities demonstrate different ways to break down compounds into simpler substances.

Photochemical Decomposition of Silver Chloride

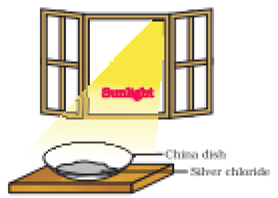


Figure 1.7

- **Activity:**
    1. Take about 2 grams of silver chloride in a china dish.
    2. Observe its initial color (white).
    3. Place the dish in sunlight for some time.
    4. Observe the color change.
- **Observation:** The white silver chloride turns grey.
- **Explanation:**
    - Sunlight provides the energy for silver chloride to decompose into silver and chlorine.
    - $2AgCl(s) \rightarrow 2Ag(s) + Cl_2(g)$
    - Silver bromide ($AgBr$) behaves similarly: $2AgBr(s) \rightarrow 2Ag(s) + Br_2(g)$
- **Application:** These reactions are used in black and white photography.
- **Type of Reaction:** Photochemical decomposition (decomposition caused by light).
- **Endothermic Reaction:**
    - **Definition:** A reaction that absorbs energy from its surroundings.
    - In this case, light energy is absorbed for the decomposition to occur.

Key Points:

- This activity shows how light can cause a chemical reaction.
- Many decomposition reactions are endothermic, meaning they require an input of energy to happen.
- This specific reaction is fundamental to how traditional photography works.

Another Endothermic Reaction

- **Activity:**
  1. Take about 2 grams of barium hydroxide in a test tube.
  2. Add 1 gram of ammonium chloride and mix with a glass rod.
  3. Touch the bottom of the test tube with your palm.
- **Observation:** The test tube feels cold.
- **Explanation:**
  - The reaction between barium hydroxide and ammonium chloride absorbs heat from the surroundings.
  - This is an example of an endothermic reaction.

Questions

1. Whitewashing

- **(i) Substance 'X' and its formula:**
  - The substance 'X' used for whitewashing is calcium hydroxide, $Ca(OH)_2$ (slaked lime).
- **(ii) Reaction with water:**
  - When calcium oxide (quicklime) reacts with water, it forms calcium hydroxide (slaked lime) and releases heat.
  - $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + \text{Heat}$

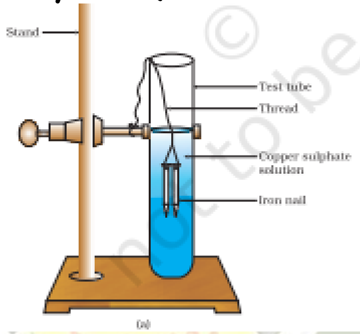
2. Electrolysis of Water (Activity 1.7)

- **Why double the volume of gas in one tube?**
  - Water (H_2O) decomposes into hydrogen (H_2) and oxygen (O_2) gases.
  - The balanced chemical equation is: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
  - This shows that 2 moles of water decompose to produce 2 moles of hydrogen gas and 1 mole of oxygen gas.
  - Since the molar ratio of hydrogen to oxygen is 2:1, the volume of hydrogen collected is double the volume of oxygen.
- **Name of the gas with double the volume:** Hydrogen (H_2)

Key Points:

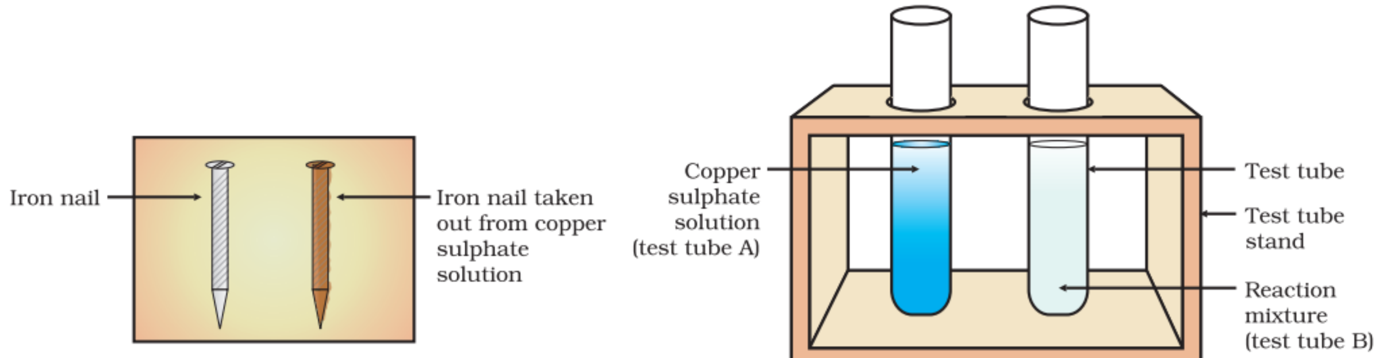
- The barium hydroxide and ammonium chloride reaction is an example of an endothermic reaction.
- Calcium hydroxide is the key component in whitewashing.
- The electrolysis of water produces hydrogen and oxygen gases in a 2:1 ratio by volume.

1.2.3 Displacement Reactions

- **Activity 1.9 (Iron Nails and Copper Sulfate)**
- 

    - **What to do:**
        1. Clean three iron nails with sandpaper.
        2. Take two test tubes (A and B) with 10 mL of copper sulfate solution in each.
        3. Immerse two iron nails tied together in test tube B for 20 minutes. Keep one nail aside.
        4. After 20 minutes, remove the nails from test tube B.
        5. Compare the color intensity of the copper sulfate solutions in both test tubes.
        6. Compare the color of the nails dipped in the solution with the one kept aside.
- **What you'll observe:**
    - The blue color of the copper sulfate solution in test tube B fades.
    - The iron nails dipped in the solution become coated with a reddish-brown substance (copper).
- **Explanation:**
    - Iron is more reactive than copper.
    - Iron displaces copper from copper sulfate solution.
    - The reaction is: $Fe(s) + CuSO4(aq) \rightarrow FeSO\ 4\ (aq) + Cu(s)$
    - This is a displacement reaction where a more reactive element displaces a less reactive element from its compound.

Key Point: Displacement reactions are like a "trading places" game where one element takes the place of another in a compound.



(b)

Displacement Reactions (Continued)

- **Why does the iron nail become brownish?**
  - The brownish color is due to the deposition of copper metal on the iron nail.
  - Iron displaces copper from copper sulfate solution, and the displaced copper atoms coat the nail.
- **Why does the blue color of copper sulfate fade?**
  - Copper sulfate solution is blue due to the presence of Cu²⁺ ions.
  - As iron displaces copper, the concentration of Cu²⁺ ions decreases, causing the blue color to fade.

1.2.4 Double Displacement Reactions

- **Activity 1.10 (Sodium Sulfate and Barium Chloride)**
  - **What to do:**
    1. Take sodium sulfate solution in a test tube.
    2. Take barium chloride solution in another test tube.
    3. Mix the two solutions.
  - **What you'll observe:**
    - A white insoluble substance (precipitate) forms.
- **Explanation:**
  - Sodium sulfate reacts with barium chloride to form barium sulfate (white precipitate) and sodium chloride.
  - $Na_2SO_4 (aq) + BaCl_2 (aq) \rightarrow BaSO_4 (s) + 2NaCl(aq)$
  - This is a double displacement reaction where ions are exchanged between the reactants.
  - Ba²⁺ ions from barium chloride react with SO 4 ²⁻ ions from sodium sulfate to form insoluble barium sulfate.
- **Precipitation Reaction:**
  - Any reaction that produces a precipitate is called a precipitation reaction.

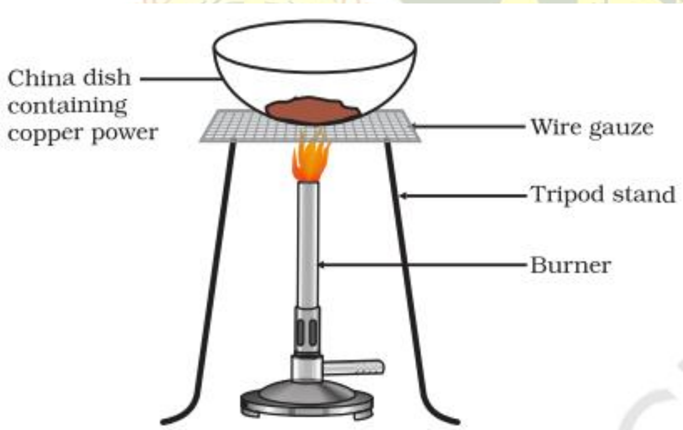
Key Point: In double displacement reactions, think of it as two couples "switching partners" to form two new compounds.

Double Displacement Reaction: Lead Nitrate and Potassium Iodide

- **Recall Activity 1.2:**
  - You mixed solutions of lead(II) nitrate and potassium iodide.
- **(i) Color and name of the precipitate:**
  - The precipitate formed was **yellow**.
  - The compound precipitated is **lead iodide (PbI_2)**.
- **(ii) Balanced chemical equation:**
  - $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$
- **(iii) Is this a double displacement reaction?**
  - **Yes**, this is a double displacement reaction.
  - The positive and negative ions in the reactants switch places to form two new compounds.
  - Lead (Pb²⁺) and potassium (K⁺) exchange their partners (NO_3 ⁻ and I⁻) to form lead iodide and potassium nitrate.

Key Point: This is another classic example of a double displacement reaction that results in the formation of a precipitate. The yellow color of the precipitate is a characteristic of lead iodide.

1.2.5 Oxidation and Reduction



- **Activity 1.11 (Heating Copper Powder)**
  - **What to do:**
    1. Heat a china dish containing copper powder.
  - **What you'll observe:**
    - The surface of the copper powder becomes coated with black copper(II) oxide.
  - **Explanation:**
    - Oxygen is added to copper, forming copper oxide.
    - $2Cu + O_2 \rightarrow 2CuO$
- **Reduction of Copper Oxide:**

    - If hydrogen gas is passed over the heated copper oxide, the black coating turns brown as copper is obtained.
    - $CuO + H_2 \rightarrow Cu + H_2O$
- **Oxidation:**
    - A substance is oxidized if it **gains oxygen** or **loses hydrogen** during a reaction.
    - In the above reaction, hydrogen is oxidized as it gains oxygen.
- **Reduction:**
    - A substance is reduced if it **loses oxygen** or **gains hydrogen** during a reaction.
    - In the above reaction, copper oxide is reduced as it loses oxygen.
- **Redox Reactions:**
    - Reactions where one reactant gets oxidized while the other gets reduced are called oxidation-reduction reactions or redox reactions.
- **Examples of Redox Reactions:**
    - $ZnO + C \rightarrow Zn + CO$ (Carbon is oxidized, and zinc oxide is reduced)
    - $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$ (HCl is oxidized, and MnO_2 is reduced)

Key Points:

- Oxidation and reduction always occur together.
- Redox reactions involve the transfer of oxygen or hydrogen between reactants.
- Many chemical reactions, including combustion and respiration, are redox reactions.