Metals and Non-metals

Physical Properties of Metals

- **Metallic Luster:** Metals, in their pure state, have a shining surface.
- **Hardness:** Metals are generally hard, although the hardness varies from metal to metal.
- **Malleability:** Some metals can be beaten into thin sheets. Gold and silver are the most malleable metals.
- **Ductility:** Metals can be drawn into thin wires. Gold is the most ductile metal.
- **Heat Conductivity:** Metals are good conductors of heat and have high melting points. Silver and copper are the best conductors of heat.
- **Examples:** Iron, copper, aluminum, magnesium, sodium, lead, zinc, gold, and silver are all examples of metals.

Why are metals used for making cooking vessels?

Metals are used for making cooking vessels because they are good conductors of heat, meaning they can transfer heat efficiently to cook food. They also have high melting points, so they can withstand the high temperatures required for cooking.

Electrical Conductivity and Sonority in Metals

- **Electrical Conductivity:** Metals are good conductors of electricity. This is why they are used to make wires.
- **Insulation:** Electric wires are coated with materials like PVC or rubber because these materials are insulators. Insulators prevent the flow of electricity, protecting us from electric shock.
- **Sonority:** Metals are sonorous, meaning they produce a sound when struck. This is why school bells are made of metal.

Non-metals

- **Definition:** Non-metals are elements that generally lack the characteristics of metals.
- **States of Matter:** They exist as solids or gases at room temperature, except for bromine, which is a liquid.
- **Examples:** Carbon, sulfur, iodine, oxygen, hydrogen, and bromine.

Physical Properties (and Exceptions)

While metals and non-metals have distinct general properties, there are exceptions:

- **Melting Point:**
  - Most metals have high melting points.
  - **Exception:** Gallium and cesium (metals) have very low melting points and melt in your hand.
- **Luster:**

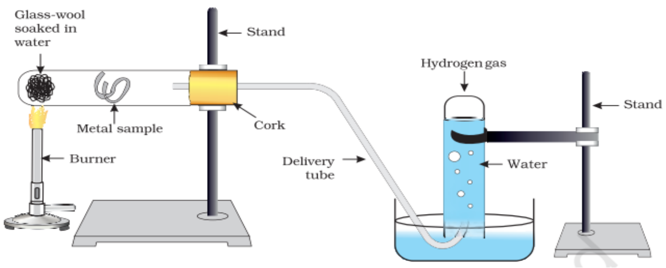
- Metals are usually lustrous (shiny).
- **Exception:** Iodine (a non-metal) is lustrous.

- **Hardness:**
  - Metals are generally hard.
  - **Exception:** Alkali metals (lithium, sodium, potassium) are soft and can be cut with a knife.
- **Allotropes:**
  - Carbon (a non-metal) has allotropes (different forms) with varying properties:
    - Diamond: Hardest natural substance.
    - Graphite: Conducts electricity.

Chemical Properties of Metals

- **Reaction with Oxygen:**
  - Most metals react with oxygen to form metal oxides.
  - Example: Copper reacts with oxygen to form copper(II) oxide (black oxide).
  - $2Cu + O_2 \rightarrow 2CuO$
  - Reactivity with oxygen varies:
    - Potassium and sodium react vigorously and catch fire in the open, so they are stored in kerosene oil.
    - Magnesium, aluminum, zinc, and lead form a protective oxide layer that prevents further reaction.
    - Iron burns when in the form of filings.
    - Copper forms a black oxide layer when heated.
    - Silver and gold do not react with oxygen.
- **Metal Oxides:**
  - Most metal oxides are basic in nature.
  - Some metal oxides, like aluminum oxide and zinc oxide, are amphoteric (react with both acids and bases).
  - $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$
  - $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$
  - Some metal oxides dissolve in water to form alkalis.
  - $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$
- **Reactivity Series:**
  - Metals exhibit different reactivities.
  - Based on the reactions with oxygen, sodium is the most reactive among the metals mentioned.

Reactions of Metals with Water



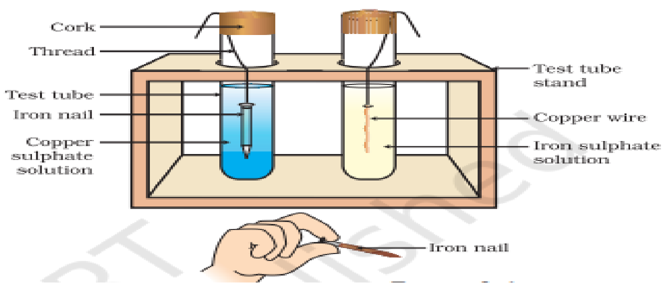
- **General Reaction:**
  - Some metals react with water to form metal oxide and hydrogen gas.
  - If the metal oxide is soluble in water, it further reacts to form metal hydroxide.

- Metal + Water → Metal oxide + Hydrogen
- Metal oxide + Water → Metal hydroxide
- **Reactivity Varies:**
  - **Highly Reactive (Potassium and Sodium):** React violently with cold water, producing enough heat to ignite the hydrogen gas.
  - $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$ + heat energy
  - **Moderately Reactive (Calcium):** Reacts less violently with cold water; hydrogen gas produced but doesn't ignite. Calcium floats due to hydrogen bubbles.
  - $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$
  - **Less Reactive (Magnesium):** Reacts with hot water, producing hydrogen gas. Magnesium floats due to hydrogen bubbles.
  - **Unreactive with Water (Aluminum, Iron, Zinc):** React with steam to form metal oxide and hydrogen gas.
  - $2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$
  - **Inert (Lead, Copper, Silver, Gold):** Do not react with water or steam.

Reactions of Metals with Acids

- **General Reaction:**
  - Metals react with dilute acids to form salt and hydrogen gas.
- Metal + Dilute acid → Salt + Hydrogen
- **Reactivity with Acids:**
  - Reactivity varies with different metals and acids.
  - Magnesium reacts most vigorously with dilute hydrochloric acid (HCl), followed by aluminum, zinc, and then iron.
  - Copper does not react with dilute HCl.
- **Reaction with Nitric Acid:**
  - Nitric acid (HNO_3) is a strong oxidizing agent, so it usually does not produce hydrogen gas when reacting with metals.
  - Instead, it oxidizes the hydrogen gas to water and gets reduced to nitrogen oxides.
  - Magnesium and manganese are exceptions and can react with very dilute nitric acid to produce hydrogen gas.

Reactions of Metals with Salt Solutions



• **Displacement Reactions:** More reactive metals can displace less reactive metals from their salt solutions.

- This means a more reactive metal will "kick out" a less reactive metal from a compound and take its place.

Metal A + Salt solution of B → Salt solution of A + Metal B

- Example: If metal A displaces metal B from its salt solution, it means metal A is more reactive than metal B.

 Determining Reactivity Order:

- Reactions with oxygen, water, and acids help us understand metal reactivity, but not all metals react with these substances.
- Displacement reactions provide a clearer way to establish the reactivity order of metals.

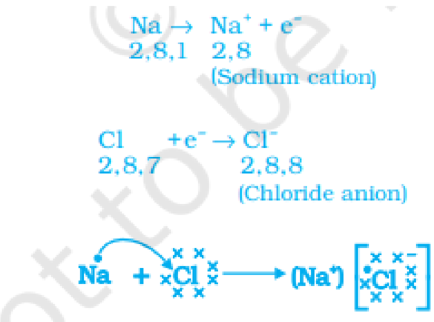
 Key Idea: The ability of a metal to displace another metal from its salt solution indicates its relative reactivity. The more easily it displaces others, the more reactive it is.

The Reactivity Series

The reactivity series is a list of metals arranged in the order of their decreasing activities.

K	Potassium	Most reactive
Na	Sodium	
Ca	Calcium	
Mg	Magnesium	
Al	Aluminium	
Zn	Zinc	Reactivity decreases
Fe	Iron	
Pb	Lead	
[H]	[Hydrogen]	
Cu	Copper	
Hg	Mercury	
Ag	Silver	
Au	Gold	Least reactive

How Metals and Non-metals React



 Noble Gases and Reactivity:

- Noble gases have a full outermost shell of electrons (a stable octet).
- They are chemically unreactive because they already have a stable electron configuration.

- Other elements react to try to achieve a similar stable electron configuration.

Metals:

- Tend to lose electrons to achieve a stable octet.
- Example: Sodium (Na) loses one electron to become a positively charged sodium ion (Na+).

 Non-metals:

- Tend to gain electrons to achieve a stable octet.
- Example: Chlorine (Cl) gains one electron to become a negatively charged chloride ion (Cl-).

 Give and Take:

Metals and non-metals can react with each other.

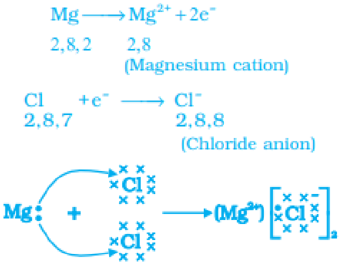
- Metals give electrons, and non-metals take them.
- This creates oppositely charged ions that attract each other, forming ionic compounds (like sodium chloride or table salt).

Formation of Ionic Compounds

- **Electrostatic Attraction:** Oppositely charged ions, like sodium ions (Na+) and chloride ions (Cl-), attract each other strongly.
- **Ionic Compounds:** This attraction forms ionic compounds, such as sodium chloride (NaCl) or table salt.
- **No Individual Molecules:** Ionic compounds exist as a vast network of ions, not as separate molecules.

Example: Magnesium Chloride (MgCl2)

- Magnesium (Mg) loses two electrons to become a stable magnesium ion (Mg2+).
- Two chlorine atoms each gain one electron to become chloride ions (Cl-).
- The Mg2+ ion attracts two Cl- ions to form magnesium chloride.



Key Points

- **Ionic Compounds (Electrovalent Compounds):** Formed by the transfer of electrons from a metal to a non-metal.

- **Cations:** Positively charged ions (formed by metals losing electrons).
- **Anions:** Negatively charged ions (formed by non-metals gaining electrons).
- **In MgCl2:**
  - Cation: Mg2+ (Magnesium ion)
  - Anion: Cl- (Chloride ion)

Properties of Ionic Compounds

Ionic compound	Melting point (K)	Boiling point (K)
NaCl	1074	1686
LiCl	887	1600
$CaCl_2$	1045	1900
CaO	2850	3120
$MgCl_2$	981	1685

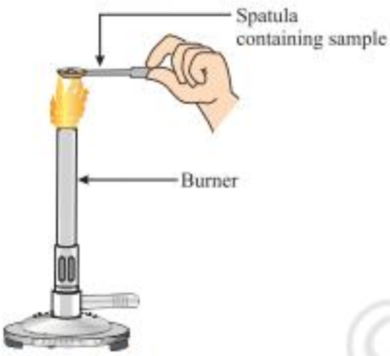


Figure 3.7
Heating a salt sample on a spatula

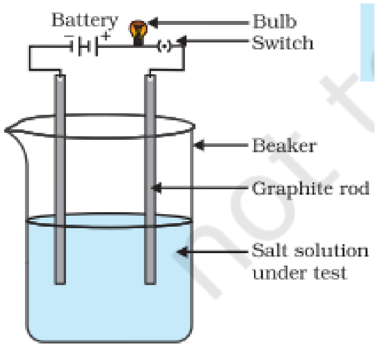


Figure 3.8
Testing the conductivity of a salt solution

- **Physical Nature:**
  - Solid and hard due to strong attraction between positive and negative ions.
  - Brittle, tend to break into pieces under pressure.

 Melting and Boiling Points:

- High melting and boiling points.
- A lot of energy is needed to overcome the strong attraction between ions.

 Solubility:

- Generally soluble in water.
- Insoluble in nonpolar solvents like kerosene and petrol.

 Electrical Conductivity:

- **Do not conduct electricity in the solid state:** Ions are held tightly in a rigid structure and cannot move.

- **Conduct electricity in the molten (liquid) state:** Heat overcomes the attraction between ions, allowing them to move freely and carry electric current.
- **Conduct electricity when dissolved in water:** Ions become free to move in the solution and can carry electric current.

Occurrence and Extraction of Metals

- **Sources of Metals:**
  - Earth's crust (main source)
  - Seawater (contains soluble metal salts)
- **Minerals and Ores:**
  - Minerals: Naturally occurring elements or compounds in the earth's crust.
  - Ores: Minerals with a high concentration of a specific metal, making extraction profitable.
- **Metal Reactivity and Extraction:**
  - **Low Reactivity Metals (e.g., gold, silver, platinum, copper):** Often found in a free state (native state) due to their low reactivity.
  - **Medium Reactivity Metals (e.g., zinc, iron, lead):** Found as oxides, sulfides, or carbonates. Oxygen's high reactivity and abundance explains why many ores are oxides.
  - **High Reactivity Metals (e.g., potassium, sodium, calcium, magnesium, aluminum):** Never found in a free state due to their high reactivity.
- **Extraction Techniques:**
  - The choice of extraction technique depends on the metal's reactivity.
- **Enrichment of Ores:**
  - **Gangue:** Ores contain impurities like soil and sand, collectively called gangue.
  - **Ore Enrichment:** Gangue needs to be removed before metal extraction.
- **Steps in Metal Extraction:**
  - Multiple steps are involved in extracting pure metals from ores.
  - These steps may include enrichment, converting the ore to a suitable form, extracting the metal, and refining it.

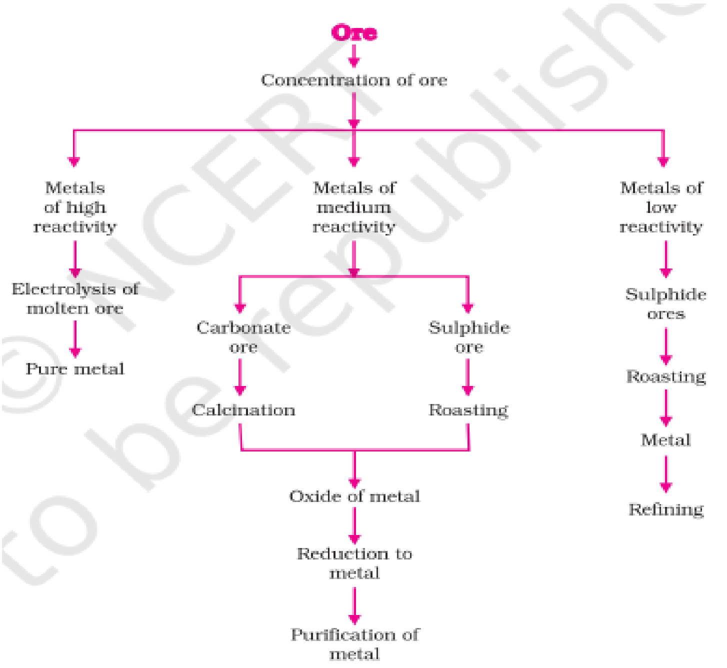


Figure 3.10 *Steps involved in the extraction of metals from ores*

Extracting Metals from Ores

1. Enrichment of Ores

- **Gangue:** Ores have impurities like soil and sand called gangue.
- **Removal of Gangue:** Impurities are removed based on differences in physical or chemical properties between the ore and gangue.

2. Extracting Less Reactive Metals

- **Found in Native State:** Metals like mercury and copper are often found in their pure form.
- **Extraction by Heating:** Heating ores like cinnabar (HgS) converts them to oxides, which can then be reduced to pure metals.
- $2HgS(s) + 3O_2(g) \rightarrow 2HgO(s) + 2SO_2(g)$
- $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

3. Extracting Moderately Reactive Metals

- **Conversion to Oxides:**
  - These metals (e.g., iron, zinc, lead) are often found as sulfides or carbonates.
  - **Roasting:** Sulfide ores are heated in excess air to convert them to oxides.
  - $2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$
  - **Calcination:** Carbonate ores are heated in limited air to convert them to oxides.
  - $ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$
- **Reduction:**
  - **Using Carbon:** Metal oxides are reduced to metals using carbon (coke) as a reducing agent.
  - $ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$

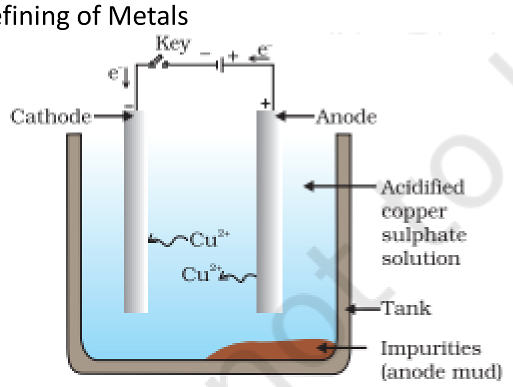
- **Displacement Reactions:** Highly reactive metals (like sodium, calcium, aluminum) can displace less reactive metals from their compounds. This reaction is highly exothermic (releases a lot of heat).
- $3MnO_2(s) + 4Al(s) \rightarrow 3Mn(l) + 2Al_2O_3(s) + Heat$
- **Thermit Reaction:** The reaction of iron(III) oxide (Fe_2O_3) with aluminum is highly exothermic and used to join railway tracks or repair cracked machine parts.

  $Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(l) + Al_2O_3(s) + Heat$

Extracting Highly Reactive Metals

- **Electrolytic Reduction:**
  - Metals high in the reactivity series (e.g., sodium, magnesium, calcium, aluminum) are very reactive.
  - They cannot be extracted by heating with carbon because they have a stronger attraction to oxygen.
  - Instead, they are extracted using electrolysis of their molten chlorides (or oxides in the case of aluminum).
- **Electrolysis Process:**
  - **Cathode (negative electrode):** Metal ions gain electrons and are deposited as pure metal.
  - $Na^+ + e^- \rightarrow Na$
  - **Anode (positive electrode):** Chlorine gas is liberated.
  - $2Cl^- \rightarrow Cl_2 + 2e^-$
- **Example: Aluminum Extraction**
  - Aluminum is obtained by electrolytic reduction of aluminum oxide (alumina).

Refining of Metals



- **Why Refining?** Metals obtained from ores often contain impurities that need to be removed to get pure metal.

- **Electrolytic Refining:**

  - A widely used method to purify metals like copper, zinc, tin, nickel, silver, and gold.

- **Process:**

1. **Setup:**
   - **Anode:** Impure metal is used as the anode (positive electrode).
   - **Cathode:** A thin strip of pure metal is used as the cathode (negative electrode).
   - **Electrolyte:** A solution of the metal salt is used as the electrolyte.

2. **Passing Current:** When electricity is passed through the electrolyte:
    - Pure metal from the anode dissolves into the electrolyte.
    - An equal amount of pure metal from the electrolyte gets deposited on the cathode.

- **Dealing with Impurities:**

    - **Soluble impurities:** Dissolve in the electrolyte.
    - **Insoluble impurities:** Settle at the bottom of the anode as "anode mud".

Corrosion

- **What is Corrosion?**
    - Corrosion is the gradual destruction of materials (usually metals) by reacting with their environment.
- **Examples of Corrosion:**
    - **Silver:** Turns black when exposed to air due to the formation of silver sulfide.
    - **Copper:** Develops a green coating (basic copper carbonate) when exposed to moist air containing carbon dioxide.
    - **Iron:** Forms a brown flaky coating called rust when exposed to moist air.
- **Conditions for Iron Rusting:**
    - **Experiment:**
        - Test tube A: Iron nails exposed to air and water → Rusting occurs.
        - Test tube B: Iron nails exposed to water only → No rusting.
        - Test tube C: Iron nails exposed to dry air only → No rusting.
    - **Conclusion:** Both air (oxygen) and water (moisture) are necessary for iron to rust.

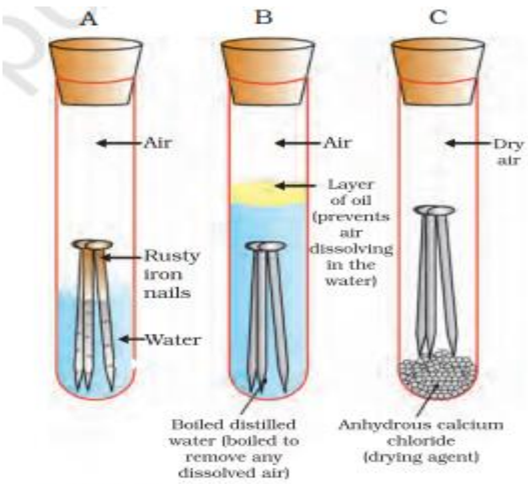


Figure 3.13
Investigating the conditions under which iron rusts. In tube A, both air and water are present. In tube B, there is no air dissolved in the water. In tube C, the air is dry.

Preventing Corrosion

- **Methods to Prevent Iron Rusting:**
    - **Painting:** Creates a barrier between iron and air/moisture.
    - **Oiling/Greasing:** Similar to painting, provides a protective layer.
    - **Galvanization:** Coating iron with a thin layer of zinc. Zinc is more reactive and corrodes first, protecting the iron even if the coating is scratched.
    - **Chrome Plating:** A thin layer of chromium provides a shiny, protective layer.

    - **Anodizing:** An electrochemical process that forms a protective oxide layer on the metal surface.
    - **Alloying:** Mixing iron with other metals to create corrosion-resistant alloys like stainless steel.
- **Alloying:**
    - **What are Alloys?** Homogeneous mixtures of two or more metals, or a metal and a non-metal.
    - **Purpose:** Improves the properties of metals (e.g., strength, hardness, corrosion resistance).
    - **How are they Made?** Melting the primary metal and dissolving other elements in it in specific proportions.
    - **Examples:**
        - **Stainless steel:** Iron with nickel and chromium, resists rusting.
        - **Brass:** Copper and zinc.
        - **Bronze:** Copper and tin.
        - **Solder:** Lead and tin, has a low melting point, used for joining electrical wires.
- **Amalgams:** Alloys containing mercury.
- **Properties of Alloys:**
    - Often have lower electrical conductivity and melting points than pure metals.