Acids and Bases

- Acids are sour and turn blue litmus paper red.
- Bases are bitter and turn red litmus paper blue.
- Acids and bases can neutralize each other.
- Litmus paper is a natural indicator that can be used to test for acids and bases.
- Other natural indicators include turmeric, red cabbage leaves, and the colored petals of some flowers.
- Synthetic indicators such as methyl orange and phenolphthalein can also be used to test for acids and bases.

Remedy for Acidity

- If someone is suffering from acidity after overeating, you can suggest a remedy that is basic in nature to neutralize the stomach acid.
- Some possible remedies include baking soda solution, vinegar, or lemon juice.
- Baking soda solution is the most basic of the three options, and it is often used as an antacid.
- Vinegar and lemon juice are also acidic, but they are less acidic than baking soda solution.

Testing for Acids and Bases

- To test for an acid, you can dip a piece of litmus paper into the solution. If the paper turns red, the solution is acidic.
- To test for a base, you can dip a piece of litmus paper into the solution. If the paper turns blue, the solution is basic.
- You can also use other natural indicators to test for acids and bases. For example, turmeric will turn yellow in the presence of an acid and red in the presence of a base.

Understanding the Chemical Properties of Acids and Bases

This section focuses on exploring how to identify acids and bases using different types of indicators.

- **Indicators:** Substances that change color in acidic or basic solutions, helping us identify them.
  - **Examples:** Litmus paper (red and blue), phenolphthalein, methyl orange.
- **Olfactory Indicators:** Substances that change odor in acidic or basic solutions.
  - **Examples:** Onion, vanilla essence, clove oil.

How do Acids and Bases React with Metals?

Key Concepts:

- **Displacement Reaction:** Metals displace hydrogen from acids, forming salt and hydrogen gas.
- **General Equation:** Acid + Metal → Salt + Hydrogen gas
- **Example:** $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

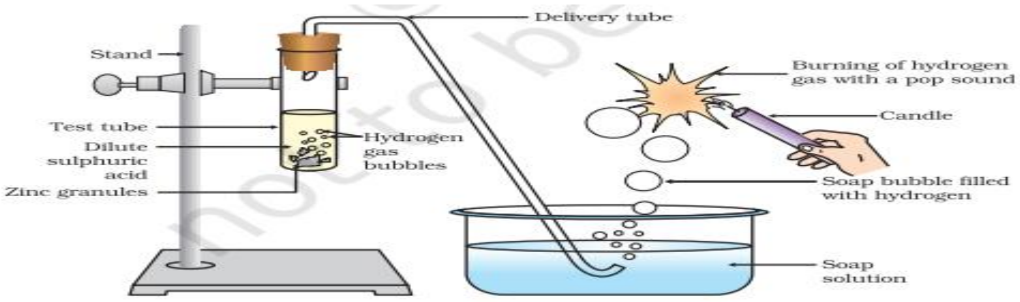


Figure 2.1 *Reaction of zinc granules with dilute sulphuric acid and testing hydrogen gas by burning*

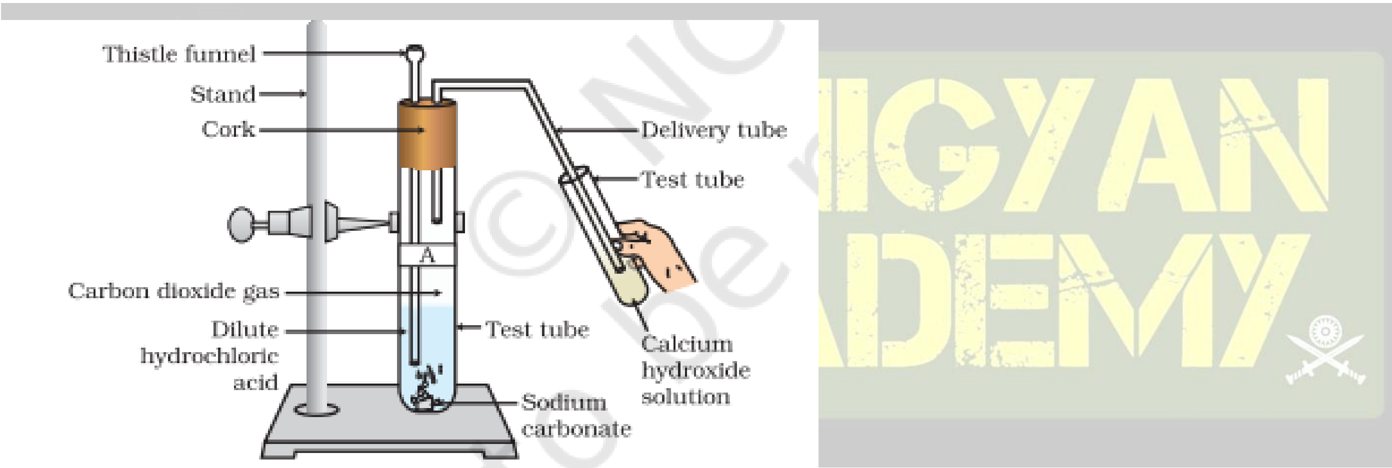


Figure 2.2
Passing carbon dioxide gas through calcium hydroxide solution

Key Concepts:

- **Reaction with Base:** Some metals react with bases to produce hydrogen gas.
- **Example:** $2NaOH(aq) + Zn(s) \rightarrow Na_2ZnO_2(s) + H_2(g)$
- **Important Note:** Not all metals react with bases.

How do Metal Carbonates and Metal Hydrogencarbonates React with Acids?

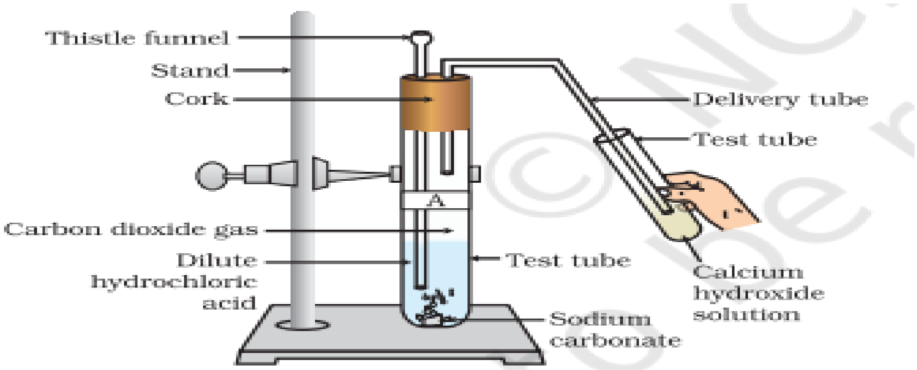


Figure 2.2
Passing carbon dioxide gas through calcium hydroxide solution

This section focuses on the reactions of metal carbonates and metal hydrogen carbonates with acids.

Observing the Reaction

1. **Gather materials:**
   - Sodium carbonate (Na_2CO_3)
   - Sodium hydrogencarbonate ($NaHCO_3$)
   - Dilute hydrochloric acid (HCl)
   - Test tubes (labeled A and B)
   - Limewater (calcium hydroxide solution)
   - Apparatus as shown in Fig. 2.2 (refer to your textbook)
2. **Set up and observe:**
   - Add sodium carbonate to test tube A and sodium hydrogencarbonate to test tube B.
   - Add dilute HCl to both test tubes.
   - Observe what happens in each test tube.
   - Pass the gas produced in each test tube through limewater.
   - Record your observations.

Key Concepts:

- **Reaction with Acid:** Metal carbonates and hydrogencarbonates react with acids to produce salt, carbon dioxide, and water.
- **General Equation:**
  - Metal carbonate + Acid → Salt + Carbon dioxide + Water
  - Metal hydrogencarbonate + Acid → Salt + Carbon dioxide + Water
- **Examples:**
  - $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$
  - $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

Limewater Test for Carbon Dioxide:

- **Initial Reaction:** $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$ (Limewater turns milky due to the formation of white precipitate of calcium carbonate)

- **Reaction with Excess CO2:** $CaCO_3(s) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$ (The white precipitate dissolves as soluble calcium bicarbonate forms)

Important Notes:

- Limestone, chalk, and marble are different forms of calcium carbonate.
- This activity demonstrates a neutralization reaction where an acid and a base (carbonate or hydrogencarbonate) react to form salt and water.

Reaction of a Non-metallic Oxide with Base

- **Non-metallic oxides are acidic in nature.** This means they react with bases in a similar way to how acids react with bases.
- **Example:** Carbon dioxide (CO_2) reacts with calcium hydroxide ($Ca(OH)_2$), a base, to produce calcium carbonate ($CaCO_3$), a salt, and water (H_2O).
- **Neutralization:** This reaction is similar to a neutralization reaction between an acid and a base.
- **Key takeaway:** The reaction between a non-metallic oxide and a base further reinforces the concept of neutralization, where an acidic substance reacts with a basic substance to form a salt and water.

What do all acids and all bases have in common?

This section explores the common properties of acids and bases, particularly their ability to

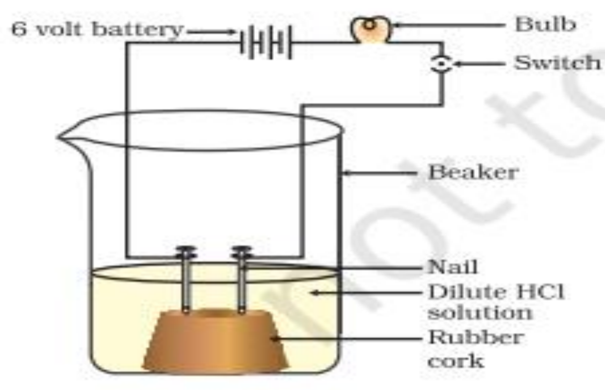


Figure 2.3
Acid solution in water conducts electricity

conduct electricity

Activity 2.6: Testing for Electrical Conductivity

1. **Gather materials:**
   - Solutions of glucose, alcohol, hydrochloric acid (HCl), sulphuric acid (H_2SO_4)
   - Two nails
   - Cork
   - 100 mL beaker

- 6-volt battery
- Light bulb
- Switch
- Wires to connect the circuit (refer to Fig. 2.3 in your textbook)

2. **Set up the circuit:**
   - Fix the nails to the cork and place it in the beaker.
   - Connect the nails to the battery terminals through the bulb and switch.
3. **Test the solutions:**
   - Pour dilute HCl into the beaker and switch on the current. Observe if the bulb glows.
   - Repeat with dilute sulphuric acid.
   - Repeat with glucose and alcohol solutions.

Observations:

- The bulb glows when acids (HCl and H_2SO_4) are used.
- The bulb does not glow when glucose and alcohol solutions are used.

Key Concepts:

- **Electrical Conductivity:** Acids conduct electricity. This is because they produce ions (charged particles) in solution, which allow the flow of electric current.
- **Hydrogen Ions (H^+):** Acids produce hydrogen ions (H^+) in solution. These ions are responsible for the acidic properties of acids.
- **Ions in Acids:** Acids contain H^+ ions as cations (positively charged ions) and anions (negatively charged ions) like Cl^- (in HCl), NO_3^- (in HNO_3), SO_4^{2-} (in H_2SO_4), and CH_3COO^- (in CH_3COOH).

Further Investigation:

- Repeat the activity with alkalis (bases) like sodium hydroxide and calcium hydroxide.
- Observe if the bulb glows, indicating the presence of ions and electrical conductivity in bases.
- **What Happens to an Acid or a Base in a Water Solution?**
- This section explores how acids and bases behave when dissolved in water.

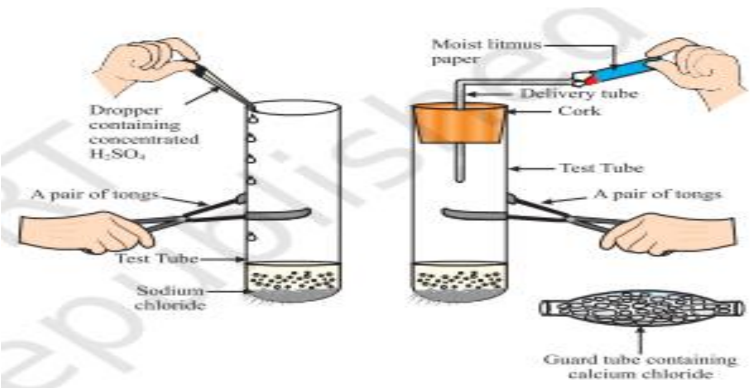


Figure 2.4 *Preparation of HCl gas*

Acids and Water

1. **Gather materials:**
    - Solid sodium chloride ($NaCl$)
    - Concentrated sulphuric acid (H_2SO_4)
    - Test tube
    - Apparatus as shown in Fig. 2.4 (refer to your textbook)
    - Dry and wet blue litmus paper
2. **Set up and observe:**
    - Add solid $NaCl$ to the test tube.
    - Add concentrated sulphuric acid to the test tube.
    - Observe if any gas is produced.
    - Test the gas with dry and wet blue litmus paper.
    - Note in which case the litmus paper changes color.

Observations:

- HCl gas is produced.
- The wet blue litmus paper turns red, indicating the presence of an acid.
- The dry blue litmus paper does not change color.

Key Concepts:

- **Acids and Water:** Acids produce hydrogen ions (H^+) only in the presence of water.
- **Hydronium Ion:** Hydrogen ions (H^+) combine with water molecules to form hydronium ions (H_3O^+). This is why hydrogen ions in solution are represented as $H^+(aq)$ or H_3O^+.
- **Example:** $HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

Bases and Water

- **Hydroxide Ions:** Bases dissolve in water to produce hydroxide ions (OH^-).
- **Alkalis:** Bases that are soluble in water are called alkalis.
- **Examples:**
    - $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
    - $KOH(s) \rightarrow K^+(aq) + OH^-(aq)$
    - $Mg(OH)_2(s) \rightarrow Mg^{2+}(aq) + 2OH^-(aq)$

Neutralization Revisited

- **Formation of Water:** The neutralization reaction between an acid and a base can be viewed as the combination of H^+ and OH^- ions to form water.
- **Equation:** $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Mixing Acid/Base with Water

This section explains the important safety precautions when mixing acids or bases with water.

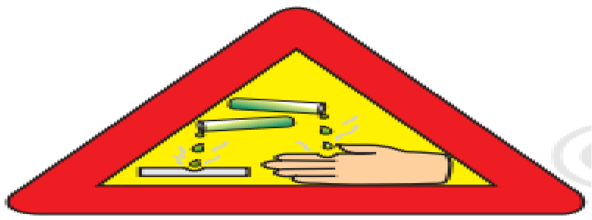


Figure 2.5
Warning sign displayed on containers containing concentrated acids and bases

Observing Heat Changes

1. **Gather materials:**
   - Water
   - Beaker
   - Concentrated sulphuric acid (H2SO4)
   - Sodium hydroxide pellets
2. **Procedure with acid:**
   - Pour 10 mL of water into the beaker.
   - Carefully add a few drops of concentrated H2SO4 to the water and swirl gently.
   - Touch the base of the beaker to feel the temperature change.
3. **Procedure with base:**
   - Repeat the above steps using sodium hydroxide pellets instead of concentrated H2SO4.

Key Concepts:

- **Exothermic Process:** Mixing an acid or base with water releases heat, making the process exothermic. This is because the dissociation of acids and bases in water releases energy.
- **Safety Precautions:**
  - **Always add acid to water, not the reverse.** Adding water to acid can cause a rapid release of heat, leading to splashing and potential burns.
  - **Add acid slowly and stir constantly.** This helps to distribute the heat evenly and prevent localized overheating.
  - **Be aware of warning signs.** Concentrated acids and bases often have warning labels indicating the dangers of improper handling.

Dilution

- **Decreasing Concentration:** Adding water to an acid or base solution decreases the concentration of ions (H3O+ or OH-) per unit volume. This is called dilution.

This breakdown emphasizes the safety aspects of handling concentrated acids and bases while explaining the concept of dilution.

How Strong Are Acid or Base Solutions?

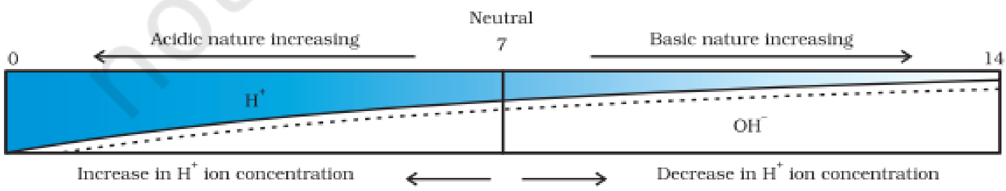


***Figure 2.6** Variation of pH with the change in concentration of $H^+(aq)$ and $OH^-(aq)$ ions*

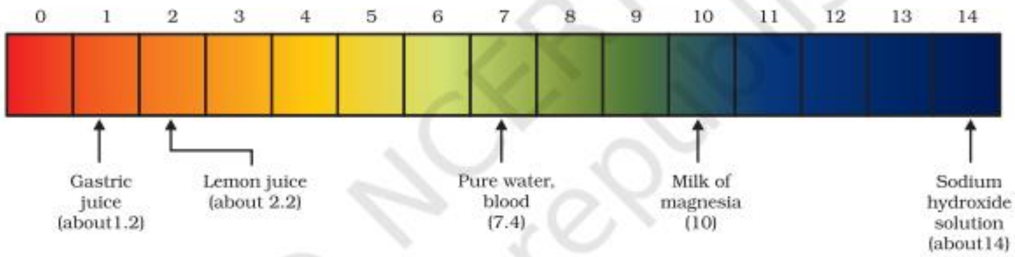


***Figure 2.7** pH of some common substances shown on a pH paper (colours are only a rough guide)*

This section explains how the strength of acids and bases is measured using the pH scale.

Measuring Acid and Base Strength

- **pH Scale:** A scale that measures the concentration of hydrogen ions (H+) in a solution. It ranges from 0 to 14.
  - **0-6:** Acidic (lower pH means stronger acid)
  - **7:** Neutral
  - **8-14:** Basic/Alkaline (higher pH means stronger base)
- **Universal Indicator:** A mixture of indicators that shows different colors at different pH values, allowing us to estimate the acidity or alkalinity of a solution.
- **Strong vs. Weak Acids/Bases:**
  - **Strong acids** produce more H+ ions in solution.
  - **Weak acids** produce fewer H+ ions in solution.
  - **Strong bases** produce more OH- ions in solution.
  - **Weak bases** produce fewer OH- ions in solution.

Key Concepts:

- **pH and H+ Concentration:** The lower the pH value, the higher the concentration of H+ ions and the stronger the acid.
- **pH and OH- Concentration:** The higher the pH value, the higher the concentration of OH- ions and the stronger the base.
- **Example:** Hydrochloric acid (HCl) is a strong acid because it releases a high concentration of H+ ions when dissolved in water. Acetic acid (CH_3COOH) is a weak acid because it releases a lower concentration of H+ ions.

This breakdown explains the pH scale and how it's used to determine the strength of acids and bases.

Importance of pH in Everyday Life

This section highlights how pH plays a crucial role in various aspects of our daily lives and the environment.

pH Sensitivity of Living Organisms

- **Humans and Animals:** Our bodies function within a narrow pH range (7.0 to 7.8). Any significant deviation from this range can be harmful.
- **Plants:** Plants also require a specific pH range for healthy growth. Different plants thrive in different pH conditions.

pH and the Environment

- **Acid Rain:** Rainwater with a pH less than 5.6 is considered acid rain. It can harm aquatic life and damage buildings and monuments.
- **Acids on Other Planets:** The presence of strong acids, like sulfuric acid in the atmosphere of Venus, makes it impossible for life to exist.

pH in Our Digestive System

- **Stomach Acid:** Our stomach produces hydrochloric acid to aid in digestion.
- **Indigestion:** Excess stomach acid can cause pain and irritation.
- **Antacids:** Mild bases called antacids, such as milk of magnesia (magnesium hydroxide), are used to neutralize excess stomach acid.

pH and Tooth Decay

- **Enamel Erosion:** Tooth enamel is corroded when the pH in the mouth falls below 5.5.
- **Bacteria and Acid:** Bacteria in the mouth produce acids that can erode enamel.
- **Prevention:** Maintaining good oral hygiene, including brushing with toothpaste (which is basic), helps neutralize acids and prevent tooth decay.

pH in Self-Defense

- **Bee Stings:** Bee stings release an acid that causes pain. Baking soda, a mild base, can provide relief.
- **Nettle Stings:** Nettle leaves inject methanoic acid, causing a burning sensation.

This breakdown provides a clear and concise overview of the importance of pH in everyday life.

Table 2.3 Some naturally occurring acids

Natural source	Acid	Natural source	Acid
Vinegar	Acetic acid	Sour milk (Curd)	Lactic acid
Orange	Citric acid	Lemon	Citric acid
Tamarind	Tartaric acid	Ant sting	Methanoic acid
Tomato	Oxalic acid	Nettle sting	Methanoic acid

Do You Know?

Nature provides neutralisation options

Nettle is a herbaceous plant which grows in the wild. Its leaves have stinging hair, which cause painful stings when touched accidentally. This is due to the methanoic acid secreted by them. A traditional remedy is rubbing the area with the leaf of the dock plant, which often grows beside the nettle in the wild. Can you guess the nature of the dock plant? So next time you know what to look out for if you accidentally touch a nettle plant while trekking. Are you aware of any other effective traditional remedies for such stings?

More About Salts

This section explores the concept of salt families and how salts are derived from acids and bases.

Identifying Salt Families

1. **Write Chemical Formulas:**
   - Write the chemical formulas for the following salts:
     - Potassium sulphate
     - Sodium sulphate
     - Calcium sulphate
     - Magnesium sulphate
     - Copper sulphate
     - Sodium chloride
     - Sodium nitrate

        - Sodium carbonate
        - Ammonium chloride
2. **Identify Parent Acids and Bases:**
    - Determine the acids and bases that react to form each of the salts listed.
    - For example, sodium chloride (NaCl) is formed from the reaction of sodium hydroxide (NaOH), a base, and hydrochloric acid (HCl), an acid.
3. **Group into Families:**
    - Identify the families of salts based on common positive or negative radicals (ions).
    - **Examples:**
        - NaCl and Na2SO4 belong to the family of sodium salts.
        - NaCl and KCl belong to the family of chloride salts.

Key Concepts:

- **Salt Families:** Salts with the same positive ion (cation) or negative ion (anion) are grouped into families. This helps in understanding their properties and potential reactions.
- **Salts from Acids and Bases:** Salts are formed by the neutralization reaction between an acid and a base.

This breakdown simplifies the activity and highlights the key concepts related to salt families and their origins.

pH of Salts - This activity helps you understand that salts can have different pH values (acidic, basic, or neutral) depending on the acids and bases they are made from.

What to Do:

1. **Gather Your Materials:**
    - Salt samples:
        - Sodium chloride (table salt)
        - Potassium nitrate
        - Aluminum chloride
        - Zinc sulfate
        - Copper sulfate
        - Sodium acetate
        - Sodium carbonate (washing soda)
        - Sodium hydrogen carbonate (baking soda)
        - Any other salts you can find
    - Distilled water (important: regular tap water can affect the results)
    - Test tubes or small beakers
    - Red and blue litmus paper
    - pH paper (gives you a more precise pH number)
2. **Test Each Salt:**

- **Solubility:** Put a small amount of each salt in a separate test tube. Add distilled water and see if it dissolves completely.
- **Litmus Test:**
  - Dip a piece of red litmus paper into the salt solution. Does it turn blue (basic)?
  - Dip a piece of blue litmus paper into the salt solution. Does it turn red (acidic)?
  - If neither paper changes color, the salt is neutral.
- **pH Measurement:**
  - Dip a piece of pH paper into the salt solution.
  - Compare the color of the paper to the chart that comes with the pH paper to find the pH number.

3. **Record Your Results:**
   - Make a table to keep track of your findings.
   - For each salt, write down:
     - Whether it dissolved in water
     - The results of the litmus test (acidic, basic, or neutral)
     - The pH number

Figuring Out the Acid and Base:

- Once you know if a salt is acidic, basic, or neutral, you can make an educated guess about the acid and base that created it. Here's how:
  - **Neutral Salts:** Usually come from a strong acid and a strong base.
    - Example: Sodium chloride (NaCl) is made from hydrochloric acid (HCl - strong acid) and sodium hydroxide (NaOH - strong base).
  - **Acidic Salts:** Usually come from a strong acid and a weak base.
    - Example: Ammonium chloride (NH4Cl) is made from hydrochloric acid (HCl - strong acid) and ammonium hydroxide (NH4OH - weak base).
  - **Basic Salts:** Usually come from a weak acid and a strong base.
    - Example: Sodium acetate (CH3COONa) is made from acetic acid (CH3COOH - weak acid) and sodium hydroxide (NaOH - strong base).

Chemicals from Common Salt

This section explains how common salt (sodium chloride) is a valuable raw material for producing other important chemicals.

Common Salt (Sodium Chloride)

- **Neutral Salt:** Formed from the reaction of hydrochloric acid (HCl) and sodium hydroxide (NaOH).
- **Sources:**
  - Seawater: Extracted by evaporating seawater.
  - Rock salt: Mined from underground deposits.

Sodium Hydroxide Production (Chlor-Alkali Process)

- **Electrolysis of Brine:** Passing electricity through an aqueous solution of sodium chloride (brine) produces sodium hydroxide, chlorine gas, and hydrogen gas.
- **Chemical Equation:** $2NaCl(aq) + 2H_2O(l) \rightarrow 2NaOH(aq) + Cl_2(g) + H_2(g)$
- **Products and their uses:**
    - **Sodium hydroxide (NaOH):** Used in the manufacture of soap, paper, and textiles.
    - **Chlorine gas (Cl_2):** Used for water purification and in the production of plastics.
    - **Hydrogen gas (H_2):** Used as a fuel and in the production of fertilizers.

Key Concepts:

- **Importance of Common Salt:** Common salt is a readily available and essential raw material used in the production of various chemicals.
- **Chlor-Alkali Process:** An important industrial process that uses electrolysis to produce valuable chemicals from common salt.

This breakdown provides a concise overview of the importance of common salt and its role in the chlor-alkali process.

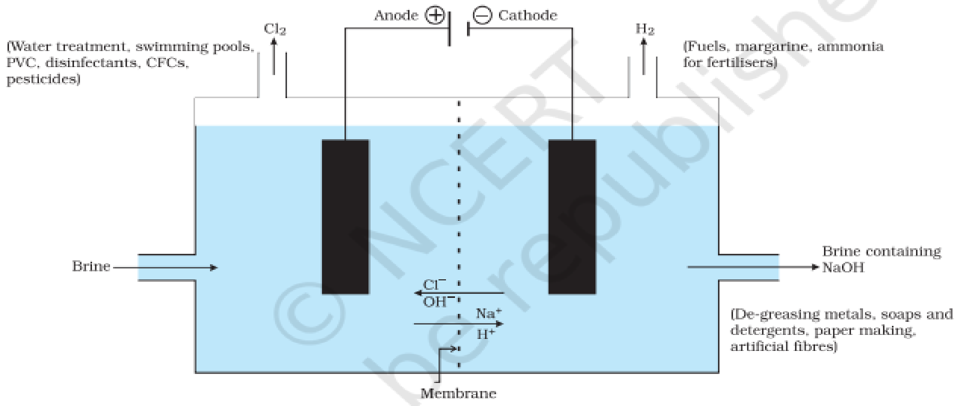


Figure 2.8 *Important products from the chlor-alkali process*

Bleaching Powder

- **Production:**
    - Chlorine gas (from the chlor-alkali process) reacts with dry slaked lime [$Ca(OH)_2$] to produce bleaching powder.
    - Chemical Equation: $Ca(OH)_2 + Cl_2 \rightarrow CaOCl_2 + H_2O$
- **Uses:**
    - Bleaching: Used in the textile industry, paper factories, and laundries for bleaching.
    - Oxidizing agent: Used in various chemical industries.
    - Water purification: Used to disinfect drinking water.

Baking Soda

- **Chemical Name:** Sodium hydrogencarbonate ($NaHCO_3$)
- **Production:**
  - Sodium chloride is used as a raw material in the production of baking soda.
  - Chemical Equation: $NaCl + H_2O + CO_2 + NH_3 \rightarrow NH_4Cl + NaHCO_3$
- **Properties:**
  - Mild, non-corrosive base.
  - Neutralizes acids.
- **Uses:**
  - Baking powder: Reacts with acids to release carbon dioxide, which makes cakes and bread rise.
  - Antacid: Neutralizes excess stomach acid.
  - Fire extinguisher: Used in soda-acid fire extinguishers.

Key Concepts:

- **Baking Soda's Basicity:** Baking soda is slightly basic which allows it to neutralize acids.
- **Heating Baking Soda:** When heated, baking soda decomposes to produce sodium carbonate, water, and carbon dioxide.
- **Versatility of Baking Soda:** Baking soda has a variety of uses in the household and in various industries due to its mild basic properties.

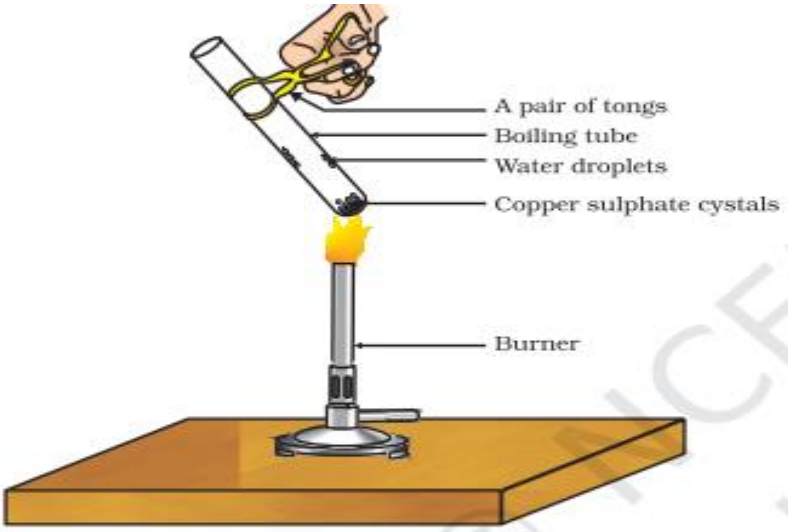
Washing Soda

- **Chemical Name:** Sodium carbonate decahydrate ($Na_2CO_3.10H_2O$)
- **Production:**
  - Heating baking soda produces sodium carbonate (Na_2CO_3).
  - Recrystallization of sodium carbonate with water forms washing soda ($Na_2CO_3.10H_2O$).
  - Chemical Equation: $Na_2CO_3 + 10H_2O \rightarrow Na_2CO_3.10H_2O$
- **Properties:**
  - Basic salt.
  - The "$10H_2O$" signifies that there are 10 water molecules associated with each Na_2CO_3 molecule within the crystal structure. This is called water of crystallization. It doesn't make the salt "wet" in the usual sense, but the water molecules are a part of the crystal structure.
- **Uses:**
  - Industrial uses: Glass, soap, and paper industries.
  - Manufacturing: Used to make other sodium compounds like borax.
  - Cleaning agent: Used for domestic cleaning purposes.
  - Water softening: Used to remove permanent hardness of water.

Key Concepts

- **Water of Crystallization:** Washing soda contains water molecules as part of its crystal structure. These water molecules are loosely bound within the crystal.
- **Importance of Washing Soda:** Washing soda is an important industrial chemical with various applications in manufacturing and cleaning.

Are the Crystals of Salts Really Dry?



his section explores the concept of water of crystallization in salts.

Water of Crystallization

- **Definition:** A fixed number of water molecules present in one formula unit of a salt. These water molecules are part of the crystal structure.
- **Example:** Copper sulfate pentahydrate ($CuSO_4.5H_2O$) contains five water molecules per formula unit of copper sulfate.

Hydrated vs. Anhydrous Salts

- **Hydrated salts:** Salts that contain water of crystallization. They often appear dry, but the water molecules are trapped within the crystal structure.
- **Anhydrous salts:** Salts that do not contain water of crystallization.

Heating Hydrated Salts

- When hydrated salts are heated, they lose their water of crystallization, often changing color and becoming anhydrous.
- Example: When blue copper sulfate pentahydrate ($CuSO_4.5H_2O$) is heated, it loses water and turns into white anhydrous copper sulfate ($CuSO_4$).

Gypsum and Plaster of Paris

- **Gypsum:** A hydrated salt with the chemical formula $CaSO_4.2H_2O$.

- **Plaster of Paris:** Formed by heating gypsum at 373 K, causing it to lose some water molecules and become calcium sulfate hemihydrate ($CaSO_4.1/2H_2O$).
- **Uses of Plaster of Paris:** Used as plaster for supporting fractured bones, making casts, and in construction materials.
- **Setting of Plaster of Paris:** When mixed with water, Plaster of Paris converts back to gypsum, forming a hard solid mass.

Key Concepts:

- **Water of crystallization is not just surface moisture.** The water molecules are integrated into the crystal structure of the salt.
- **Hydrated salts can appear dry.** Even though they contain water molecules, hydrated salts often look and feel dry.
- **Heating and rehydration:** Hydrated salts can lose and regain water of crystallization through heating and exposure to water.
- **Plaster of Paris is a useful material.** Its ability to harden when mixed with water makes it valuable in various applications.