Acids, Bases and Salts

1 Acids

Acids furnish H^+ ions or H_3O^+ ions when dissolved in water. Acids have one or more replaceable H atoms.

- Acids generally have sour taste.
- Acids change Blue litmus Red.
- They are colorless with **phenolphthalein** and pink with **methyl orange**.
- Acids show acidic nature in their aqueous form.

1.1 Dissociasion of Acids

Acid is capable of producing hydrogen ion H^+ by dissociating in aqueous solution. This reaction can be represented by

$$HA\,(aq) \longrightarrow H^+\,(aq) + A^-\,(aq)$$

For example: Hydrochloric Acid (HCl)

$$HCl\,(aq) \longrightarrow H^+\,(aq) + Cl^-\,(aq)$$

The proton or hydrogen ion binds itself to a water molecule to form a **hydronium ion** (H_3O^+)

$$\underset{\text{Hydrogen Ion}}{H^+} + \underset{\text{Water}}{H_2O} \longrightarrow \underset{\text{Hydronium Ion}}{H_3O^+}$$

The **hydronium ion** is also known as **oxonium ion** or **hydroxonium ion**.

Note:

- H^+ ions are protons.
- Metal oxides usually are **basic** in nature whereas Non metal oxides usually are **acidic** in nature.
- Aqueous Solutions of Acids are good conductors of electricity because the **hydronium ions** produced help in conducting electricity.

On the basis of extent of dissociation of acids, they are classified as Strong and

Weak Acids.

- The Acids which completely dissociate in water are called **Strong Acids**.

$$HNO_3\,(aq) \longrightarrow H^+\,(aq) + NO_3{}^-\,(aq)$$

Nitric Acid completely dissociates in water.

- There are only seven strong acids.
    - H_2SO_4 - Sulphuric Acid
    - HCl - Hydrochloric Acid
    - HNO_3 - Nitric Acid
    - HBr - Hydrobromic Acid
    - HI - Hydroiodic Acid
    - $HClO_4$ - Perchloric Acid
    - $HClO_3$ - Chloric Acid
- The Acids which dissociate partially in water are **weak acids**.
- All organic acids are weak.
- Since their dissociation is only partial, itis denoted be a reversible reaction.

$$HF\,(aq) \rightleftharpoons H^+\,(aq) + F^-\,(aq)$$

Hydrofluoric Acid dissociates partially in water.

$$CH_3COOH \longrightarrow CH_3COO^- + H^+$$

Acetic acid is an organic acid and a weak acid.

- The double arrows indicate that:
    - The aqueous solution of hydrofluoric acid not only contains H^+ and F^- ions, but also the undissociated acid HF.
    - There is an equilibrium between the undissociated acid HF and the ions furnished by it, H^+ and F^-.
- Examples of Weak Acids are
    - CH_3COOH - Ethanoic (Acetic) acid
    - HF - Hydrofluoric acid
    - HCN - Hydrocynic acid
    - C_6H_5COOH - Benzoic acid

- All hydrogen containing compounds are not acids.

  Although Ethyl alcohol (C_2H_5OH) and glucose $(C_6H_{12}O_6)$

contain hydrogen, they do not produce H^+ ions on dissolving in water. Their solutions are not acidic.

- Although acetic acid being an organic acid is a weak acid, but concentrated acetic acid is corrosive and can damage the skin if poured over it.

1.2 Classification of Acids

- **Based on Source:**
    - **Organic Acids** are present in plants and animals (living beings).

      Eg:

        - $HCOOH$ (Formic Acid/Methanoic Acid, found in **stings of bees/ants**)
        - CH_3COOH (Acetic Acid/Ethanoic Acid, found in **Vinegar**)

    - **Inorganic Acids** are found from rocks and minerals.

      Eg: HCl (Hydrochloric Acid), HNO_3 (Nitric Acid), H_2SO_4 (Sulphuric Acid)

- **Based on their Basicity**

  Basicity = The number of H atoms replaceable by a base in a particular acid.

    - **Monabasic Acid** gives one H^+ ion per molecule of the acid in solution.

      Eg: HCl, HNO_3

    - **Dibasic Acid** gives two H^+ ions per molecule of the acid in the solution.

      Eg: H_2SO_4, H_2CO_3

    - **Tribasic Acid** gives three H^+ ions per molecule of the acid in the solution.

      Eg: H_3PO_4

- **Based on Concentration**
    - **Concentrated Acid** has a relatively high percentage of acid in its aqueous solution.
    - **Dilute Acid** has a relatively low percentage of acid in its aqueous solution.

1.3 Chemical Properties of Acids

- **Reaction of acids with Metals**

  Acids give hydrogen gas along with respective salt when they react with a metal.

  $$\text{Metal} + \text{Acid} \longrightarrow \text{Salt} + \text{Hydrogen}$$

  Examples:

  - $Zn + 2\,HCl \longrightarrow ZnCl_2 + H_2 \uparrow$
  - $2\,Na + 2\,HCl \longrightarrow 2\,NaCl_2 + H_2 \uparrow$
  - $Fe + 2\,HCl \longrightarrow FeCl_2 + H_2 \uparrow$
  - $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$

- **Reaction of acids with Metal Carbonates**

  Acids react with metal carbonates to give respective salt, carbon dioxide and water.

  $$\text{Metal Carbonate} + \text{Acid} \longrightarrow \text{Salt} + \text{Carbon Dioxide} + \text{Water}$$

  Examples:

  - $CaCO_3 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + CO_2 \uparrow$
  - $Na_2CO_3 + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O + CO_2 \uparrow$
  - $CaCO_3 + 2\,HCl \longrightarrow CaCl_2 + H_2O + CO_2 \uparrow$
  - $Na_2CO_3 + 2\,HCl \longrightarrow 2\,NaCl + H_2O + CO_2 \uparrow$
  - $MgCO_3 + 2\,HCl \longrightarrow MgCl_2 + H_2O + CO_2 \uparrow$
  - $Na_2CO_3 + 2\,HNO_3 \longrightarrow NaNO_3 + H_2O + CO_2 \uparrow$

- **Reaction of acids with Metal Hydrogen Carbonates (Bicarbonates)**

  Acids give CO_2 gas, respective salt and water when they react with metal hydrogen carbonates.

  $$\text{Metal Bicarbonate} + \text{Acid} \longrightarrow \text{Salt} + \text{Water} + \text{Carbon Dioxide}$$

  Examples:

  - $NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2 \uparrow$
  - $2\,NaHCO_3 + H_2SO_4 \longrightarrow Na_2SO_4 + 2\,H_2O + 2\,CO_2 \uparrow$

Notes:

- Sodium Bicarbonate $(NaHCO_3)$ is also known as Sodium Hydrogen Carbonate, **Baking Soda** and **Baking Powder**

- The gas evolved in the reaction of acid and metal hydrogen carbonate or bicarbonate, **turns lime water milky**. This indicates that the gas is Carbon Dioxide (CO_2).

  This is due to the formation of white ppt of **Calcium Carbonate**, $CaCO_3$

  $$CaOH_2 + CO_2 \longrightarrow H_2O + CaCO_3 \downarrow$$
  Carbon Dioxide turns lime water milky.

- But when excess CO_2 is passed through lime water, it makes the milky colour disappear.

  This happens because of formation of **calcium hydrogen carbonate**. As it is soluble in water, the milky colour dissapears.

  $$CaCO_3 + H_2O + CO_2 \longrightarrow CaHCO_3\,(aq)$$
  Excess carbon dioxide makes the milky colour disappear.

- **Calcium Carbonate**, $CaCO_3$ is a salt found in **eggshells**, **chalk powder** and **marble**.

- **Reaction of acids with Metallic oxides**

  Metal oxides are basic in nature. Thus, when an acid reacts with a metal oxide, both neutralise each other. In this reaction, respective salt and water is formed.

  $$\text{Acid} + \text{Metal Oxide} \longrightarrow \text{Salt} + \text{Water}$$

  Examples:

  - $2\,HCl + CaO \longrightarrow CaCl_2 + H_2O$
  - $H_2SO_4 + ZnO \longrightarrow ZnSO_4 + H_2O$
  - $6\,HCl + Al_2O_3 \longrightarrow 2\,AlCl_3 + 3\,H_2O$

1.4 Corrosive nature of Acids

The ability of acids to attack various substances like metals, metal oxides and hydroxides is referred to as their corrosive nature. Acids are corrosive in nature as they can attack a variety of substances.

Strengh and Corrosive Action of Acids

Corrosive action of acids is not related to their strength. It is related to the negatively charged ion of the acid.

Example: Hydrofluoric Acid, a weak acid can attack and dissolve glass.

$$4\,HF\,(aq) + SiO_2 \longrightarrow SiF_4 + H_2O$$

Weak Acid Hydrogen Fluoride corrodes glass.

2 Bases

Bases release hydroxide ions when dissolved with water.

- Bases are **bitter** in taste and **soapy** to touch.
- They change red litmus blue.
- They are pink with phenolphthalein and yellow with methyl orange.
- **Alkalis** are water-soluble bases.

2.1 Dissociation of Bases

Similar to acids, **aqueous solutions of bases** conduct electricity due to the formation of *hydroxyl ions*.

$$NaOH\,(aq) \longrightarrow Na^+\,(aq) + OH^-\,(aq)$$

On the basis of the extent of dissociation occuring in their solution, bases are classified as strong or weak and their characteristics are as follows:

- **Strong bases** completely dissociate in water to form a cation and hydroxide ion (OH^-).

  Eg: $KOH\,(aq) \longrightarrow K^+\,(aq) + OH^-\,(aq)$

- There are only 8 strong bases. They are

  - $LiOH$ - Lithium Hydroxide
  - $NaOH$ - Sodium Hydroxide
  - KOH - Potassium Hydroxide
  - $RbOH$ - Rubidium Hydroxide
  - $CsOH$ - Caesium Hydroxide
  - $Ca(OH)_2$ - Calcium Hydroxide
  - $Sr(OH)_2$ - Strontium Hydroxide
  - $Ba(OH)_2$ - Barium Hydroxide

- **Weak Bases** do not furnish OH^- ions by dissociation. They react with water to furnish OH^- ions.

Eg: Ammonia

$$NH_3(g) + H_2O(l) \longrightarrow NH_4OH$$

$$NH_4OH(aq) \rightleftharpoons NH_4^+ + OH^-$$

- The reaction resulting in the formation of OH^- ions does not go to completion and the solution contains relatively low concentration of OH^- ions.

- The double sided arrows indicate that equilibrium is reached before the reaction is completed.

- Examples of Weak Bases are

  - NH_4OH - Ammonium Hydroxide
  - $Cu(OH)_2$ - Copper Hydroxide
  - $Cr(OH)_3$ - Chromium Hydroxide
  - $Zn(OH)_2$ - Zinc Hydroxide

2.2 Classification of Bases

- **Based on their Acidity**

  The number of ionizable hydroxide $(OH-)$ ions present in one molecule of base is called the acidity of bases.

  - **Monaacidic Base** gives one OH^- ion per molecule of the base in solution.

    Eg: $NaOH, KOH$

  - **Diacidic Base** gives two OH^- ions per molecule of the base in the solution.

    Eg: $Ca(OH)_2, Mg(OH)_2$

  - **Triacidic Base** gives three OH^- ions per molecule of the base in the solution.

    Eg: $Al(OH)_2, Fe(OH)_2$

- **Based on Concentration**

  - **Concentrated Alkali** has a relatively high percentage of alkali in its aqueous solution.
  - **Dilute Alkali** has a relatively low percentage of alkali in its aqueous solution.

2.3 Chemical Properties of Bases

- **Reaction of Base with Metals**

  When alkali (base) reacts with metal, it produces salt and hydrogen gas.

  $$\text{Alkali} + \text{Metal} \longrightarrow \text{Salt} + \text{Hydrogen}$$

  - Sodium hydroxide gives hydrogen gas and **sodium zincate** on reacting with zinc metal.

  $$2\,NaOH + Zn \longrightarrow Na_2ZnO_2 + H_2$$

  Such reactions are not possible with all metals.

- **Reaction of Base with Non Metallic Oxides**

  Non metallic oxides are acidic in nature. For example, CO_2 dissolved in water gives carbonic acid.

  When a base reacts with a non-metal oxide, both neutralise each other resulting in production of respective salt and water.

  $$\text{Base} + \text{Non-Metal Oxide} \longrightarrow \text{Salt} + \text{Water}$$

  - $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$
  - $2\,NaOH + CO_2 \longrightarrow Na_2CO_3 + H_2O$

3 Indicators

Substances which show the acidic or basic behaviour of other substances by change are indicators.

Types of Indicators:

- Natural Indicators
- Synthetic Indicators
- Olfactory Indicators
- Universal Indicator

Natural Indicators:

Indicators obtained from natural sources are natural indicators. Eg: Litmus, Turmeric, Red Cabbage, China Rose

- **Litmus**

  Litmus is obtained from Lichens. The solution of litmus is purple in colour. Litmus paper comes in red and blue colors.

  Acids turns blue litmus paper red and bases turn red litmus paper blue.

- **Turmeric**

  Turmeric is yellow in colour. Turmeric solution or paper turns reddish brown with base.

- **Red Cabbage**

  The juice of red cabbage is purple in colour. It turns reddish with acids and greenish with bases.

Olfactory Indicators

Substances which change their smell when mixed with acid or base are olfactory indicators.

- **Onion**

  Paste or juice of onion loses its smell when added with base.

- **Vanilla**

  The smell of vanilla vanishes with base, but its smell does not vanish with acid.

Synthetic Indicators

Indicators that are synthesized in laboratory are synthetic indicators.

Phenolphthalein is a colourless liquid. It turns pink with a base.

Methyl Orange is originally orange in colour. It turns red with acid and yellow with base.

4 Dilution and Neutralizarion

Dilution of Acids and Bases

The concentration of hydrogen ion in an acid and hydroxide ion in a base, per unit volume, shows the concentration of an acid or base.

By mixing acid to water, the concentration of H^+ per unit volume decreases. Similarly, by addition of base to water, the concentration of OH^- per unit volume decreases. **The process of adding acid or base to water in order to decrease its concentration is called dilution.**

> The dilution of acids and bases is **exothermic**. Thus acid/base is

always added to water, and not the *vice versa*.

If water is added to a concentrated acid or base, a lot of heat is generated, which may cause **splashing out of acids and bases** and may cause severe damage as concentrated acid and base are highly corrosive.

Neutralisation Reaction

An acid neutralises a base when they react with each other, and respective salt and water is formed.

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

$$H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + 2\,H_2O$$

When an acid reacts with a base, the **hydrogen ion** (H^+) of the acid combines with the **hydroxide ion** (OH^-) of the base and forms water.

As these ions combine together and form water, instead of remaining free, thus both neutralise each other.

Creation of Salt: Salts can be of three types: Neutral, Acidic or Basic.

Acid	**Base**	**Salt**
Strong	Strong	Neutral
Strong	Weak	Acidic
Weak	Strong	Basic
Weak	Weak	NA

5 Strength of Acid and Base

Acid-Base indicators canbe used to distinguish between an acid and a base.

The classification of Acids and Bases ad **Strong and Weak** can be done by checking the extent of ionisation of the acid/base.

A scale for measuring hydrogen ion concentration in a solution, called **pH Scale** has been developed.

The **pH** value of a solution represents the concentration of H^+ (aq) ions in the solution, in a logarithmic scale.

$$\mathrm{pH} = \log \frac{1}{[H^+]}$$

Or

$$\mathrm{pH} = -\log [H^+]$$

Here, $[H^+]$ denotes the **molar concentration of** H^+ (aq) **ions**.

For example, in neutral solutions,

$$[H^+] = 1.0 \times 10^{-7} \text{ mol L}^{-1}$$

Conc. of H^+ ions in pure water is 10^{-7} moles per liter.

Because of the negative sign in the expression $-\log[H^+]$, **as the** H^+ **conc. increases,** pH **decreases**, and vice verca.

The pH value of a neutral solution is:

$$\begin{aligned}\mathrm{pH} &= -\log_{10}\left(10^{-7}\right) \\ &= -(-7) \\ &= 7\end{aligned}$$

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