Acid and Base Reactions
What is an acid?

- **Definitions**
  - **Arrhenius** - an acid is a substance that dissociates in water to increase $\text{H}^+$ concentration.
  
  - **Bronsted-Lowry** - an acid is a proton ($\text{H}^+$) donator.
  
  - **Lewis** – an acid is any species that accepts an electron pair.
What is an base?

• Definitions
  • Arrhenius- an acid is a substance that dissociates in water to increase \( \text{OH}^- \) concentration.
  • Bronsted-Lowry- an acid is a proton (\( \text{OH}^- \)) acceptor.
  • Lewis – an acid is any species that donates an electron pair.
Identification of acids and bases.

- **HCl** + **H₂O** → **H₃O⁺** + **Cl⁻**

  - Acid because it dissociates to form a **H⁺**
  - Base because it accepts an **H**
Conjugate Acids and Bases

• Every acid has a conjugate base that is formed when the proton is donated.
• Every base has a conjugate acid that forms when a proton is accepted.

\[
\text{HCN} + \text{HCO}_2^- \rightarrow \text{CN}^- + \text{HCO}_2\text{H}
\]

• Acid  Base  Conjugate base  Conjugate acid
Examples of Acids and Bases

Acids:
- Hydrochloric Acid - HCl
- Hydrogen Bromide - HBr
- Hydrogen Iodide - HI
- Sulfuric Acid - H₂SO₄
- Nitric Acid - HNO₃
- Perchloric Acid - HClO₄

Bases:
- Lithium Hydroxide - LiOH
- Sodium Hydroxide - NaOH
- Potassium Hydroxide - KOH
- Rubidium Hydroxide - RbOH
- Cesium Hydroxide - CsOH

Water: acts as an acid and a base can accept or donate a proton.
Autoprotolysis of Water and calculation of $K_w$

- $\text{H}_2\text{O} + \text{H}_2\text{O} \rightarrow \text{OH}^- + \text{H}_3\text{O}^-$

- In neutral water, some molecules of water act as bases and some act as bases. This is known as the autoprotolysis of water.
- A constant ($K_w$) can be found which is known as the equilibrium constant.
  \[ K_w = [\text{OH}^-] \times [\text{H}_3\text{O}^-] = 1.0 \times 10^{-14} \]
- This number is constant and can be used to find the concentrations of either component if the other is known.
- $K_w$ is used in many acid/base reactions to help calculate pH.
Relation Between \([H^+]\), \([OH^-]\), and pH

- The pH of a solution is a measure of how acidic or basic a particular solution is.
- Using the concentration of \([H^+]\), you can calculate the pH of a solution.

\[
pH = -\log[H^+]
\]
\[
pOH = -\log[OH^-]
\]
\[
pH + pOH = 14
\]
\[
[H^+] \times [OH^-] = 1.0 \times 10^{-14}
\]
Examples of pH calculations:

Calculate the concentrations of H$^+$ and OH$^-$ in pure water at 25°C.

$$K_w = 1.0 \times 10^{-14} = [\text{OH}^-] \times [\text{H}_3\text{O}^+] = [x] \times [x]$$

$$x = (1.0 \times 10^{-14})^{(1/2)}$$

$$[\text{OH}^-] = [\text{H}_3\text{O}^+] = (1.0 \times 10^{-7})$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = 7$$

Examples:

- $[\text{H}^+] = 1.0 \times 10^{-3}$
  $$\text{pH} = -\log(1.0 \times 10^{-3}) = 3$$

- $[\text{H}^+] = 1.0 \times 10^{-8}$
  $$\text{pH} = -\log(1.0 \times 10^{-8}) = 8$$

- $[\text{H}^+] = 3.8 \times 10^{-8}$
  $$\text{pH} = -\log(3.8 \times 10^{-8}) = 7.42$$
Why is this important?

• Acids and Bases are used for:

  • Titrations- determining the concentration of an acid or base using an acid or base of known concentration.
References


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