**Acid-Base Chemistry**

**Arrhenius acid:** Substance that dissolves in water and provides $H^+$ ions

**Arrhenius base:** Substance that dissolves in water and provides $OH^-$ ions

**Examples:**
- $HCl \rightarrow H^+ \text{ and } Cl^-$ **Acid**
- $NaOH \rightarrow Na^+ \text{ and } OH^-$ **Base**

**Bronsted Acid:** Substance that donates proton to another substance

**Bronsted base:** Substance that accepts proton from another substance

**Example:** $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

- $HCl$ acts as acid; $H_2O$ acts as base

  In the Reverse Reaction,
  - $H_3O^+$ acts as an acid; $Cl^-$ acts as a base

**Note:** ($H_3O^+$ = hydronium ion = $H^+$ = proton)

**Conjugate acid:** Species formed after base accepts a proton

**Conjugate base:** Species remaining after an acid donates its proton

**Conjugate acid-base pair:** an acid and base on opposite sides of the equation that correspond to each other

**Examples:**
- $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$ **acid**
- $H_2O \rightarrow H^+ \text{ and } OH^-$ **base**

**Conjugate pairs: $HNO_3$ and $NO_3^-$**

**Conjugate pairs: $H_2O$ and $H_3O^+$**

**Practice:**
- $HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$

- What are the conjugate pairs?
  - $HClO_4$ and $ClO_4^-$
  - $H_2O$ and $H_3O^+$

**Strengths of Acids and Bases**

**Strong acids/bases:** dissociate completely when dissolved in solution

**Weak acids/bases:** dissociate only partially when dissolved in solution

**Examples:**
- **Strong Acid:** $HCl \rightarrow H^+ \text{ and } Cl^-$ (100% dissociation)
- **Strong Base:** $NaOH \rightarrow Na^+ \text{ and } OH^-$ (100% dissociation)
- **Weak Acid:** $CH_3COOH \rightarrow H^+ \text{ and } CH_3COO^-$ (1.3% dissociation)
- **Weak Base:** $NH_3 + H^+ \rightarrow NH_4^+$

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**Water can act as both an acid and a base (amphiprotic)!!**

- $HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$ (base)
- $NH_3 + H_2O \rightarrow OH^- + NH_4^+$ (acid)
Conjugate Base

**Naming Acids**

**Binary Acids**: hydo + root of anion + ic + “acid”

ex. HCl hydrochloric acid, HBr hydrobromic acid

HNO₃ Hydroiodic acid

**Polyatomic-based Acids**: root of polyatomic ion + ic + “acid”

ex. H₂SO₄ sulfuric acid, H₃PO₄ phosphoric acid

H₂CO₃ carbonic acid

HNO₃ nitric acid

**The Self-Ionization of Water**

\[ H₂O + H₂O \rightarrow H₃O⁺ + OH⁻ \]

Pure water: \([H₃O⁺] = [OH⁻] = 10⁻⁷ \text{ M(at 25°C)} \)

**Neutral Solution**: Any solution in which the concentrations of H₃O⁺ and OH⁻ ions are equal (10⁻⁷ M)

**Acidic Solution**: Solutions having a greater concentration of H₃O⁺ than OH⁻ ions ([H₃O⁺] greater than 10⁻⁷ M)

**Example**: A solution with [H₃O⁺] = 10⁻⁵ M

**Basic Solution**: solution having a greater concentration of OH⁻ than H₃O⁺ ions ([H₃O⁺] less than 10⁻⁷ M)

**Example**: A solution with [H₃O⁺] = 10⁻¹² M

**The pH Scale**

• pH is a measure of acidity

• Scale ranges from 0-14

pH = 7 Neutral
pH < 7 Acidic
pH > 7 Basic

pH represents the concentration of H⁺ ions in solution

**Pure water**: 1 x 10⁻⁷ moles H⁺ per liter and 1 x 10⁻⁷ moles OH⁻ per liter
Solutions with equal concentrations of $H^+$ and $OH^-$ ions are called **Neutral**

Solutions with more than $1 \times 10^{-7}$ moles $H^+$ per liter are **Acidic**

Solutions with less than $1 \times 10^{-7}$ moles $H^+$ per liter are **Basic**

Note: $[H^+] \times [OH^-] = 10^{-14}$ (always!)

**pH Scale Summary**

- pH scale refers to amount of $H^+$ ions in solution
- pH 7 is neutral, less than 7 is acidic, greater than 7 is basic
- Lower pH = more acidic = more $H^+$ ions
- Higher pH = more basic = less $H^+$ ions

Each pH unit represents a 10-fold change in $H^+$ ion concentration!

pH 4 has 10 times more $H^+$ ions than pH 5

pH 9 has 10 times fewer $H^+$ ions than pH 8

Mathematical equation for pH:

$\text{pH} = -\log [H_3O^+]$

Any number can be expressed as 10 raised to some exponent: $y = 10^x$

**Examples:**

- $100 = 10^2$
- $1000 = 10^3$
- $0.10 = 10^{-1}$

The log is that exponent!

- $100 = 10^2$; Log of 100 = 2
- $1000 = 10^3$; Log of 1000 = 3
- $0.10 = 10^{-1}$; Log of 0.10 = -1

We can also take the log of non-whole numbers, but we use our calculators for this.

**Example:** Find the log of $2.4 \times 10^{-3}$

- Enter $2.4 \times 10^{-3}$ into calculator
- Press the “log” key

$0.0024 \log = -2.62$

Therefore, $10^{-2.62} = 2.4 \times 10^{-3}$

**Calculating pH from $[H_3O^+]$**

$\text{pH} = -\log [H_3O^+]$

- Enter $[H_3O^+]$ into calculator
- Press the “log” key
- Change the sign

**Example:** $[H_3O^+] = 1.0 \times 10^{-7}$ M

$\text{pH} = -\log [H_3O^+]$

$\text{pH} = -\log [1 \times 10^{-7}] = 7$
Example:

\[ [H_3O^+] = 1 \times 10^{-11} M \]

\[ pH = -\log [H_3O^+] \]

\[ pH = -\log [1 \times 10^{-11}] = 11 \]

Example:

\[ [H_3O^+] = 1 \times 10^{-9} M \]

\[ pH = -\log [H_3O^+] \]

\[ pH = -\log [1 \times 10^{-9}] = 9 \]

Example:

\[ [H_3O^+] = 4.2 \times 10^{-5} \]

\[ pH = -\log [H_3O^+] \]

\[ pH = -\log [4.2 \times 10^{-5}] \]

• Enter \([H_3O^+]\) into calculator (4.2 x 10^-5)

• Press the “log” key (-4.3767507)

• Change the sign (4.3767507)

\[ pH = 4.3767507 = 4.4 \]

Example:

\[ [H_3O^+] = 8.1 \times 10^{-9} \]

\[ pH = -\log [H_3O^+] \]

• Enter \([H_3O^+]\) into calculator (8.1 x 10^-9)

• Press the “log” key (-8.091515)

• Change the sign (8.091515)

\[ pH = 8.091515 = 8.1 \]

Reactions Between Acids and Bases

Neutralization: reaction between an acid and a base; always produces salt and water

Titration: Process by which acid or base of known concentration is used to neutralize a solution of unknown concentration, to determine its concentration

Buffer: Solution that resists changing pH when acids or bases are added; solution that maintains constant pH

Buffers contain 2 compounds:

• Compound with the ability to react with H⁺ ions
• Compound with the ability to react with OH⁻ ions

Example: \( HCO_3^- + H^+ \rightarrow H_2CO_3 \)

If acids (H⁺) are added, react with HCO₃⁻:

\( HCO_3^- + OH^- \rightarrow HCO_3^- + H_2O \)

If OH⁻ ions are added, react with H₂CO₃:

\( H_2CO_3 \) is unstable: \( H_2CO_3 \rightarrow H_2O + CO_2 \)

Acid-Base Titration

• Titration is a laboratory procedure used to determine the molarity of an acid.
• In a titration, a base such as NaOH is added to a specific volume of an acid.
Indicator

• A few drops of an indicator is added to the acid in the flask.
• The indicator changes color when the base (NaOH) has neutralized the acid.

End Point of Titration

• At the end point, the indicator has a permanent color.
• The volume of the base used to reach the end point is measured.
• The molarity of the acid is calculated using the neutralization equation for the reaction.