# ACIDS & BASES

- Many common substances in our daily lives are **acids and bases**. **Oranges, lemons and vinegar** are examples of **acids**. In addition, our stomachs contain acids that help digest foods. **Antacid** tablets taken for heartburn and **ammonia** cleaning solutions are examples of **bases**.

- General properties associated with acids include the following:

  - sour taste
  - change color of litmus from blue to red
  - react with metals to produce $H_2$ gas
  - react with bases to produce salt and water

- General properties associated with bases include the following:

  - bitter taste
  - slippery soapy feeling
  - change color of litmus from red to blue
  - react with acids to produce salt and water

- The most common definition of acids and bases was formulated by the Swedish chemist Svante **Arrhenius** in 1884.

- According to the **Arrhenius** definition,

**Acids are substances that produce hydronium ion ($H_3O^+$) in aqueous solution**

$$HCl\ (g) + H_2O\ (l) \longrightarrow H_3O^+\ (aq) + Cl^-\ (aq)$$

Commonly written as

$$HCl\ (g) \xrightarrow{H_2O} H^+\ (aq) + Cl^-\ (aq)$$

polar covalent

**Bases are substances that produce hydroxide ion ($OH^-$) in aqueous solution**

$$NaOH\ (s) \xrightarrow{H_2O} Na^+\ (aq) + OH^-\ (aq)$$

ionic compound

$$NH_3\ (aq) + H_2O\ (l) \longrightarrow NH_4^+\ (aq) + OH^-\ (aq)$$

# ACID & BASE STRENGTH

- According to the **Arrhenius** definition, the **strength** of acids and bases is based on the amount of their **ionization** in water.

- **Strong** acids and bases are those that **ionize completely** in water.
- Strong acids and bases are **strong electrolytes**.

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

$$NaOH\ (s) \xrightarrow{H_2O} Na^+\ (aq) + OH^-\ (aq)$$

- **Weak** acids and bases are those that **ionize partially** in water.
- Weak acids and bases are **weak electrolytes**.

$$HC_2H_3O_2 \overset{H_2O}{\rightleftharpoons} H^+\ (aq) + C_2H_3O_2^-\ (aq)$$

$$NH_3\ (aq) + H_2O\ (l) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$$

| COMMON STRONG ACIDS & BASES | | | |
|---|---|---|---|
| HCl | Hydrochloric acid | LiOH | Lithium hydroxide |
| HBr | Hydrobromic acid | NaOH | Sodium hydroxide |
| HI | Hydroiodic acid | KOH | Potassium hydroxide |
| $HNO_3$ | Nitric acid | $Ba(OH)_2$ | Barium hydroxide |
| $H_2SO_4$ | Sulfuric acid | | |
| **COMMON WEAK ACIDS & BASES** | | | |
| $HC_2H_3O_2$ | Acetic acid | $NH_3$ | Ammonia |
| $H_2CO_3$ | Carbonic acid | $CO(NH_2)_2$ | Urea |
| HF | Hydrofluoric acid | | |
| HCN | Hydrocyanic acid | | |
| $H_2S$ | Hydrosulfuric acid | | |

# BRØNSTED-LOWRY ACIDS & BASES

- The **Arrhenius** definition of acids and bases is **limited to aqueous solutions**.
- A broader definition of acids and bases was developed by **Brønsted and Lowry** in the early 20th century.
- According to **Brønsted-Lowry** definition, an **acid** is a **proton donor**, and a **base is a proton acceptor**.

$$HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)$$
(Acid) (Base)

$$NH_3\,(aq) + H_2O\,(l) \longrightarrow NH_4^+\,(aq) + OH^-\,(aq)$$
(Base) (Acid)

- A substance that can act as a **Brønsted-Lowry acid and base** (such as **water**) is called **amphiprotic**.
- In Brønsted-Lowry definition, any pair of molecules or ions that can be **interconverted by transfer of a proton** is called **conjugate acid-base pair**.

$$HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)$$
(Acid) (Base) (Conjugate acid of water) (Conjugate base of HCl)

$$NH_3\,(aq) + H_2O\,(l) \longrightarrow NH_4^+\,(aq) + OH^-\,(aq)$$
(Base) (Acid) (Conjugate acid of $NH_3$) (Conjugate base of water)

# BRØNSTED-LOWRY ACIDS & BASES

**Examples:**

1. Identify the conjugate acid-base pairs for each reaction shown below:

$$H_2O + Cl^- \rightarrow HCl + OH^-$$

$$C_6H_5OH + C_2H_5O^- \rightarrow C_6H_5O^- + C_2H_5OH$$

2. Write the formula for the conjugate acid for each base shown:

$HS^-$

$NH_3$

$CO_3^{2-}$

3. Write the formula for the conjugate base for each acid shown:

$HI$

$CH_3OH$

$HNO_3$

# IONIZATION OF WATER

- As noted previously, **water** can act both as an **acid and a base**.

- In pure water, one **water** molecule **donates a proton** to another water molecule to produce **ions**.

$$\underset{\begin{bmatrix}\text{Proton}\\\text{acceptor}\end{bmatrix}}{\underset{\text{Base}}{H_2O}} + \underset{\begin{bmatrix}\text{Proton}\\\text{donor}\end{bmatrix}}{\underset{\text{Acid}}{H_2O}} \rightleftharpoons \underset{\begin{bmatrix}\text{Proton}\\\text{donor}\end{bmatrix}}{\underset{\text{Acid}}{H_3O^+}} + \underset{\begin{bmatrix}\text{Proton}\\\text{acceptor}\end{bmatrix}}{\underset{\text{Base}}{OH^-}}$$

- In pure water, the **transfer of protons** between water molecules produces **equal numbers of $H_3O^+$ and $OH^-$ ions**. However, the number of ions produced in pure water is very small, as indicated below:

**Pure water** $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M

- When the **concentrations of $H_3O^+$ and $OH^-$ are multiplied** together, the **ion-product constant ($K_w$)** is formed.

$$K_w = [H_3O^+] \times [OH^-]$$

$$= (1.0 \times 10^{-7} \text{ M}) \times (1.0 \times 10^{-7} \text{ M}) = 1.0 \times 10^{-14}$$

- All aqueous solutions have **$H_3O^+$ and $OH^-$ ions**. An **increase in the concentration** of one of the ions will cause an equilibrium shift that causes a **decrease in the other one**.

# ACIDIC & BASIC SOLUTIONS

- When **[$H_3O^+$]** and **[$OH^-$] are equal** in a solution, it is **neutral**.
- When **[$H_3O^+$]** is **greater** than **[$OH^-$]** in a solution, it is **acidic**.

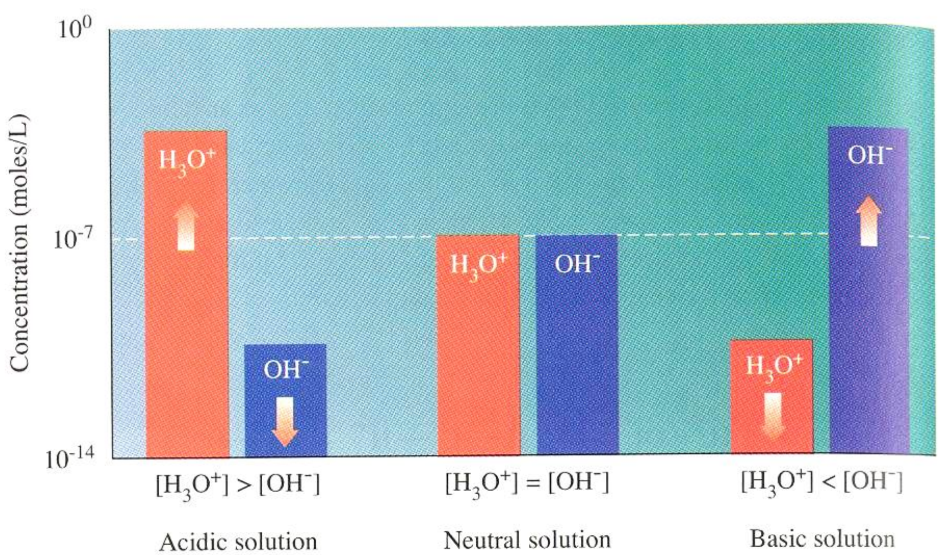
For example, if [$H_3O^+$] is 1.0 x $10^{-4}$ M, then [$OH^-$] would be 1.0 x $10^{-10}$ M.

$$\left[OH^-\right]=\frac{K_w}{\left[H_3O^+\right]}=\frac{1.0 \text{ x } 10^{-14}}{1.0 \text{ x } 10^{-4}}=1.0 \text{ x } 10^{-10} \text{ M}$$

- When **[$OH^-$]** is **greater** than **[$H_3O^+$]** in a solution, it is **basic**.

For example, if [$OH^-$] is 1.0 x $10^{-6}$ M, then [$H_3O^+$] would be 1.0 x $10^{-8}$ M.

$$\left[H_3O^+\right]=\frac{K_w}{\left[OH^-\right]}=\frac{1.0 \text{ x } 10^{-14}}{1.0 \text{ x } 10^{-6}}=1.0 \text{ x } 10^{-8} \text{ M}$$



# ACIDIC & BASIC SOLUTIONS

**Examples:**

1. Calculate the $[OH^-]$ in a solution with $[H_3O^+] = 2.3 \times 10^{-4}$ M. Classify the solution as acid or basic.

$$[OH^-] = \frac{K_w}{[H_3O^+]} =$$

2. Calculate the $[H_3O^+]$ in a solution with $[OH^-] = 2.3 \times 10^{-4}$ M. Classify the solution as acid or basic.

$$[H_3O^+] = \frac{K_w}{[OH^-]} =$$

3. Calculate the $[OH^-]$ in a solution with $[H_3O^+] = 5.8 \times 10^{-8}$ M. Classify the solution as acid or basic.

4. Calculate the $[H_3O^+]$ in a solution with $[OH^-] = 1.3 \times 10^{-2}$ M. Classify the solution as acid or basic.

# THE pH SCALE

- The **acidity** of a solution is commonly measured on a **pH scale**.

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

- The **pH scale** ranges from **0-14**, where **acidic** solutions are **less than 7** and **basic** solutions are **greater than 7**.

| | | |
|---|---|---|
| **Acidic solutions** | **pH < 7** | $[H_3O^+] > 1.0 \times 10^{-7}$ |
| **Neutral solutions** | **pH = 7** | $[H_3O^+] = 1.0 \times 10^{-7}$ |
| **Basic solutions** | **pH > 7** | $[H_3O^+] < 1.0 \times 10^{-7}$ |

**Examples:**

1. The $[H_3O^+]$ of a liquid detergent is $1.4 \times 10^{-9}$ M. Calculate its pH.

   **$pH = -\log [H_3O^+] = -\log [1.4 \times 10^{-9}] = -(-8.85) = 8.85$**

2. The pH of black coffee is 5.3. Calculate its $[H_3O^+]$.

   **$[H_3O^+] = \text{antilog}(-pH) = 10^{-pH} = 10^{-5.3} = 5 \times 10^{-6}$**

3. The $[H_3O^+]$ of a solution is $3.5 \times 10^{-3}$ M. Calculate its pH.

4. The pH of tomato juice is 4.1. Calculate its $[H_3O^+]$.

5. The $[OH^-]$ of a cleaning solution is $1.0 \times 10^{-5}$ M. What is the pH of this solution?

   $[H_3O^+] =$ pH=

6. The pH of a solution is 11.50. Calculate the $[H_3O^+]$ for this solution.

# THE pOH SCALE

- Similar to pH, a convenient way to designate the **basicity** of a solution is the **pOH scale**.

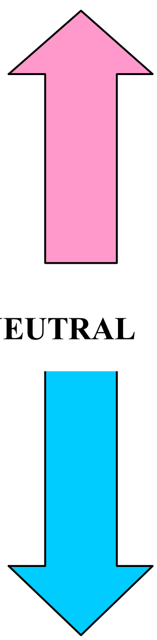
$$\mathbf{pOH = -\log [OH^-]}$$

- In aqueous solutions, the **sum of pH and pOH** of a solution equal to **14**.

$$\mathbf{pH + pOH = 14}$$

## Comparison of $[H_3O^+]$, $[OH^-]$, pH and pOH Values

ACIDIC



NEUTRAL

BASIC

| $[H_3O^+]$ | pH | pOH | $[OH^-]$ |
|---|---|---|---|
| $10^{0}$ | 0 | 14 | $10^{-14}$ |
| $10^{-1}$ | 1 | 13 | $10^{-13}$ |
| $10^{-2}$ | 2 | 12 | $10^{-12}$ |
| $10^{-3}$ | 3 | 11 | $10^{-11}$ |
| $10^{-4}$ | 4 | 10 | $10^{-10}$ |
| $10^{-5}$ | 5 | 9 | $10^{-9}$ |
| $10^{-6}$ | 6 | 8 | $10^{-8}$ |
| $10^{-7}$ | 7 | 7 | $10^{-7}$ |
| $10^{-8}$ | 8 | 6 | $10^{-6}$ |
| $10^{-9}$ | 9 | 5 | $10^{-5}$ |
| $10^{-10}$ | 10 | 4 | $10^{-4}$ |
| $10^{-11}$ | 11 | 3 | $10^{-3}$ |
| $10^{-12}$ | 12 | 2 | $10^{-2}$ |
| $10^{-13}$ | 13 | 1 | $10^{-1}$ |
| $10^{-14}$ | 14 | 0 | $10^{0}$ |

# REACTIONS OF ACIDS & BASES

- The most important reaction of acids and bases is called **neutralization**. In these reactions an acid combines with a base to form a **salt and water**. For example:

$$\mathbf{HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)}$$

**acid** **base** **salt** **water**

- Acids also react with **carbonates and bicarbonates** to produce **salt, carbon dioxide gas and water**. For example:

$$\mathbf{2\ HCl\ (aq) + Na_2CO_3\ (aq) \longrightarrow 2\ NaCl\ (aq) + CO_2\ (g) + H_2O\ (l)}$$

$$\mathbf{HCl\ (aq) + NaHCO_3\ (aq) \longrightarrow NaCl\ (aq) + CO_2\ (g) + H_2O\ (l)}$$

- Acids also react with **active metals** to produce a **salt and hydrogen gas**. For example:

$$\mathbf{Mg\ (s) + 2\ HCl\ (aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)}$$

**<u>Examples:</u>**
Complete and balance each of the following equations:

1. $ZnCO_3$ (aq) + HCl (aq) →

2. $HNO_3$ (aq) + $Mg(OH)_2$ (aq) →

3. Al (s) + HCl (aq) →

# TITRATION

- Determination of concentration of an acid or base from the known concentration of the other is called **titration**.

| $HCl$ | + | $NaOH$ | $\rightarrow$ | $NaCl$ | + | $H_2O$ |
|---|---|---|---|---|---|---|
| **acid** | | **base** | | salt | | |
| **known conc.** | | **unknown conc.** | | | | |

- When solving titration problems,

1. Write a balanced equation.
2. Determine moles of known from concentration and volume.
3. Determine moles of unknown based on reaction stoichiometry.
4. Determine the concentration of unknown from moles and volume titrated.

**Examples:**

1. If 32.6 mL of 1.85 M NaOH is required to titrate 25.0 mL of an HCl solution to the end point, what is the molarity of HCl?

| Bal. Eq. | $HCl + NaOH \rightarrow NaCl + H_2O$ |
|---|---|
| Vol (mL) | |
| Conc. (M) | |
| mol NaOH | |
| mol HCl | |
| Conc. HCl | |

**Examples:**

2. How many mL of 2.15 M KOH are required to titrate 25.0 mL of 0.300 M $HC_2H_3O_2$?

| | |
|---|---|
| Bal. Eq. | |
| Vol (mL) | |
| Conc. (M) | |
| mol HAc | |
| mol NaOH | |
| Vol NaOH (mL) | |

3. If 50.0 mL of 0.300 M NaOH solution is required to titrate 20.0 mL of sulfuric acid, what is the molarity of the acid?

| | |
|---|---|
| Bal. Eq. | |
| Vol (mL) | |
| Conc. (M) | |
| mol NaOH | |
| mol $H_2SO_4$ | |
| Conc. $H_2SO_4$ | |