Acids and Bases

Unit 10 Lesson 2
The Brønsted–Lowry Theory of acids and bases is broader than the Arrhenius theory. According to the B–L theory:

**Acids** -- donate $\text{H}^+$ ions

**Bases**—accept $\text{H}^+$ ions

This theory broadens the definition of a base beyond substances that contain the $\text{OH}^-$ ion.
B–L Theory of Acids and Bases

Examples of B–L Acid/Base Reactions:

\[ \text{HCl} \text{(aq)} + \text{OH}^- \text{(aq)} \leftrightarrow \text{Cl}^- \text{(aq)} + \text{H}_2\text{O} \text{(aq)} \]

Acid (H\(^+\) donor) Base (H\(^+\) Acceptor)

\[ \text{NH}_3 \text{(aq)} + \text{HNO}_3 \text{(aq)} \leftrightarrow \text{NH}_4^+ \text{(aq)} + \text{NO}_3^- \text{(aq)} \]

Base (H\(^+\) Acceptor) Acid (H\(^+\) donor)
According to B–L Theory, acids become bases after they have donated an H⁺.

The difference between an acid and its conjugate base is simply an H⁺.

<table>
<thead>
<tr>
<th>Acid</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>HSO₄⁻</td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>HSO₄⁻</td>
</tr>
<tr>
<td>NH₄⁺</td>
<td>NH₃</td>
</tr>
</tbody>
</table>
Strong and Weak Acids

Strong acids ionize 100% while weak acids ionize less than 10% in water.

Strong Acid: \[
\text{HCl} \quad \rightleftharpoons \quad 100\% \quad \rightarrow \quad \text{H}^+ + \text{Cl}^-
\]

Weak Acid: \[
\text{HF} \quad \rightleftharpoons \quad \sim 2.5\% \quad \rightarrow \quad \text{H}^+ + \text{F}^-
\]
Self–Ionization of Water

Water can react with itself, acting as *both* a weak Brønsted–Lowry **acid** and **base**:

\[
\ce{H2O + H2O <=> OH^- + H3O^+}
\]

- **Acid**
- **Base**
- **Conjugate Base**
- **Conjugate Acid**

\(\ce{H3O^+}\) is the *hydronium ion* and is produced when Brønsted–Lowry acids react with water. It is analogous to \(\ce{H^+}\).
**K_w— A special Equilibrium Constant**

Water molecules ionize to a very small degree:

\[
H_2O (l) \rightleftharpoons H^+ (aq) + OH^- (aq) \quad \text{or} \quad H_2O (l) + H_2O (l) \rightleftharpoons H_3O^+ (aq) + OH^- (aq)
\]

The equilibrium constant for the reaction is:

\[
K_{eq} = [H^+][OH^-] = [H_3O^+][OH^-]
\]

This equilibrium constant is so common that it is referred to as *the dissociation constant* of water and is given the symbol, \( K_w \).
The concentration of $H^+$ and $OH^-$ in pure water are both $1 \times 10^{-7}$ M. The $K_w$ in pure water is:

$$K_w = [H^+] \times [OH^-] = [1 \times 10^{-7} \text{ M}] \times [1 \times 10^{-7} \text{ M}]$$

$$= 1 \times 10^{-14}$$

In all aqueous solutions:

$$[H^+] [OH^-] = 1 \times 10^{-14}$$

$[H_3O^+]$ could also be used.
Calculations of H\(^+\) and OH\(^-\) in Aqueous Solutions

If you know the concentration of H\(^+\) in solution, you can always find the concentration of OH\(^-\).

\[ \text{e.g. Find the } [\text{OH}^-] \text{ in a 0.25 M solution of HCl.} \]

\[ \text{Ans: } 1 \times 10^{-14} = [0.25 \text{ M}][\text{OH}^-] ; [\text{OH}^-] = 4.0 \times 10^{-14} \text{ M} \]

In the same way, you can find \([H^+]\) from \([\text{OH}^-]\).
pH Scale

pH is a convenient way to express the amount of acid or base in a solution without having to use exponents.

<table>
<thead>
<tr>
<th>pH from $[H^+]$</th>
<th>$[H^+]$ from pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>$pH = -\log [H^+]$</td>
<td>$[H^+] = 10^{-pH}$</td>
</tr>
</tbody>
</table>

Examples:

Find the pH of a 0.010 M solution of HCl.

Ans: $pH = -\log [0.010M] = 2.00$

Find the $[H^+]$ of a solution with pH $= 3.5$.

Ans: $10^{-3.5} = 0.0032$ M
# pH Scale

<table>
<thead>
<tr>
<th>Concentration of hydrogen ions compared to distilled water</th>
<th>Examples of solutions at this pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>10,000,000</td>
<td>pH = 0 battery acid, strong hydrofluoric acid</td>
</tr>
<tr>
<td>1,000,000</td>
<td>pH = 1 hydrochloric acid secreted by stomach lining</td>
</tr>
<tr>
<td>100,000</td>
<td>pH = 2 lemon juice, gastric acid, vinegar</td>
</tr>
<tr>
<td>10,000</td>
<td>pH = 3 grapefruit, orange juice, soda</td>
</tr>
<tr>
<td>1,000</td>
<td>pH = 4 tomato juice, acid rain</td>
</tr>
<tr>
<td>100</td>
<td>pH = 5 soft drinking water, black coffee</td>
</tr>
<tr>
<td>10</td>
<td>pH = 6 urine, saliva</td>
</tr>
<tr>
<td>1</td>
<td>pH = 7 “pure” water</td>
</tr>
<tr>
<td>1/10</td>
<td>pH = 8 sea water</td>
</tr>
<tr>
<td>1/100</td>
<td>pH = 9 baking soda</td>
</tr>
<tr>
<td>1/1,000</td>
<td>pH = 10 Great Salt Lake, milk of magnesia</td>
</tr>
<tr>
<td>1/10,000</td>
<td>pH = 11 ammonia solution</td>
</tr>
<tr>
<td>1/100,000</td>
<td>pH = 12 soapy water</td>
</tr>
<tr>
<td>1/1,000,000</td>
<td>pH = 13 bleaches, oven cleaner</td>
</tr>
<tr>
<td>1/10,000,000</td>
<td>pH = 14 liquid drain cleaner</td>
</tr>
</tbody>
</table>

The scale is courtesy of The Pacific Institute for the Mathematical Sciences
These four relationships will allow you to solve any problems involving \([H^+], [OH^-], \text{pH}, \text{and pOH}]:

1. \([H^+] \times [OH^-] = 1 \times 10^{-14}\)
2. \(\text{pH} + \text{pOH} = 14\)
3. \(\text{pH} = -\log [H^+]\)
4. \(\text{pOH} = -\log [OH^-]\)