10. Acids, Bases, and Salts

Acids and bases

For centuries people have known acids are in vinegar, lemon juice and many other foods taste sour. Only few hundred years ago that it was discovered that acids taste sour and change litmus red, and they could be identified as acids. In Latin term acere, which means "sour." While bases feel slippery, change litmus blue. There are three different definitions of acids and bases, in this section we will introduce the fundamentals of acid/base chemistry.

- **Acids** taste sour, make metals corrode, change litmus (a dye from plants) red, and get neutralized when bases are added.
- **Bases** feel slippery, change litmus blue, and get neutralized when mixed with acids.

**Names and formulas of common acids and bases**

<table>
<thead>
<tr>
<th>Binary Acids</th>
<th>Ternary (oxo) Acids</th>
<th>Hydroxy Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrofluoric Acid - HF</td>
<td>Nitric Acid - HNO₃</td>
<td>Sodium Hydroxide - NaOH</td>
</tr>
<tr>
<td>Hydrochloric Acid - HCl</td>
<td>Sulfuric Acid - H₂SO₄</td>
<td>Potassium Hydroxide - KOH</td>
</tr>
<tr>
<td>Hydrobromic Acid - HBr</td>
<td>Phosphoric Acid - H₃PO₄</td>
<td>Ammonium Hydroxide - NH₄OH</td>
</tr>
<tr>
<td>Hydroiodic Acid - HI</td>
<td>Carbonic Acid - H₂CO₃</td>
<td>Calcium Hydroxide - Ca(OH)₂</td>
</tr>
<tr>
<td><strong>Organic Acids</strong></td>
<td>Hypochlorous Acid - HClO</td>
<td>Magnesium Hydroxide - Mg(OH)₂</td>
</tr>
<tr>
<td>Acetic Acid - HC₂H₃O₂</td>
<td>Chlorous Acid - HClO₂</td>
<td>Barium Hydroxide - Ba(OH)₂</td>
</tr>
<tr>
<td>Benzoic Acid - C₆H₅CO₂H</td>
<td>Choloric Acid - HClO₃</td>
<td>Aluminum Hydroxide - Al(OH)₃</td>
</tr>
<tr>
<td>Oxalic Acid –H₂C₂O₄</td>
<td>Perchlorlic Acid - HClO₄</td>
<td>Zinc Hydroxide - Zn(OH)₂</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Lithium Hydroxide - LiOH</td>
</tr>
</tbody>
</table>

10.1 Arrhenius Acid-Base Theory

**Arrhenius Definition:**
Arrhenius's theory explains why all acids have similar properties to each other (and, conversely, why all bases are similar): because all acids release H⁺ (proton) or H₃O⁺ (hydronium ions) into solution (and all bases release hydroxide ions, OH⁻).

**Arrhenius Acid:**
A substance that produces H⁺, or (protons) H₃O⁺, (hydronium ion) in an aqueous solution.

**Arrhenius Base:**
A substance that produces OH⁻, or hydroxide ion in an aqueous solution.

This is the first acid/base concept to be developed to describe typical acid/base reactions.

E.g. Arrhenius Acid- HCl (acid), Arrhenius Base-NaOH (base).

10.2 Bronsted-Lowry Acid-Base Theory

Brønsted (Denmark) and Lowry (England) came up with an alternative acids and base bases definition to Arrhenius. There solved the problems associated with non hydroxide bases, especially ammonia which Arrhenius' definition could not include as a
According to Brønsted-Lowry: Acids and bases are substances that are capable of donating and accepting protons (hydrogen ions, H⁺), respectively. An acid-base reaction consists of the transfer of a proton from an acid to a base. Acid and bases are considered as proton transfer agents.

- An acid is a "proton donor."
- A base is a "proton acceptor."

Lowry proposed the use of hydronium ion H₃O⁺ in the place of H⁺ that is commonly used today. He pointed out acidity is a relative thing comparing proton donor ability of two pure compounds. Even hydrogen chloride only becomes an acid when mixed with water.

\[
\text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Cl}^- 
\]

This reaction proceeds to right to a large extent:

- **HCl** - this is a Brønsted-Lowry acid, because it has a proton available to be transferred.
- **H₂O** - this is a Brønsted-Lowry base, since it gets the proton that the Brønsted-Lowry acid lost.

**Conjugate Acid-Base Pairs**

Since this an equilibrium reaction the reverse reaction could also be considered as an acid/base reaction:

\[
\text{H}_3\text{O}^+ + \text{Cl}^- \rightleftharpoons \text{HCl} + \text{H}_2\text{O} 
\]

Acid and bases involved in the reverse reaction is call **conjugate acid** and **conjugate base**

**Conjugate acid:** H₃O⁺ - this is a Brønsted-Lowry **conjugate acid**, because it can give a proton.

**Conjugate base:** Cl⁻ - this is a Brønsted-Lowry **conjugate base**, since it has the capacity to receive a proton.

A conjugate pair is an acid-base pair that differs by one proton in their formulas (remember: proton, hydrogen ion, etc.).

A conjugate pair is always one acid and one base.

\[
\text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Cl}^- 
\]

Here is the one conjugate pair (acid/conjugate base) from the first example reaction:

**HCl and Cl⁻**

The other conjugate pair is:

**H₂O and H₃O⁺**

**Some more conjugate acid-base pairs to look for:**

- H₂O and OH⁻
- HCO₃⁻ and CO₃²⁻
- H₂PO₄⁻ and HPO₄²⁻
- HSO₄⁻ and SO₄²⁻
- NH₄⁺ and NH₃
- CH₃NH₃⁺ and CH₃NH₂
- HC₂H₃O₂ and C₂H₅O₂⁻
Identify the Bronsted-Lowery acid/conjugate base and base/conjugate acid pairs in the equilibrium reactions given below

a) \( \text{HCl(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq) \)

b) \( \text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HSO}_4^-(aq) \)

c) \( \text{H}_2\text{O(l)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \)

d) Error! Bookmark not defined. \( \text{NH}_3(aq) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-(aq) \)

The concept of acid/conjugate base pair and base/conjugate acid pair came out of Bronsted-Lowery definition describing proton transfer reactions.

a) \( \text{HCl/Cl}^- \) is an acid/conjugate base pair
\( \text{H}_2\text{O/H}_3\text{O}^+ \) is a base/conjugate acid pair in this equilibrium.

b) \( \text{H}_2\text{SO}_4/\text{HSO}_4^- \) is an acid/conjugate base pair and \( \text{H}_2\text{O/H}_3\text{O}^+ \) is a base/conjugate acid pair in this equilibrium.
\( \text{HSO}_4^-/\text{SO}_4^{2-} \) is an acid/conjugate base pair, if the second dissociation of \( \text{HSO}_4^- \) (aq) took place.

c) \( \text{H}_2\text{O/OH}^- \) is an acid/conjugate base pair in this equilibrium.
\( \text{H}_2\text{O/H}_3\text{O}^+ \) is a base/conjugate acid pair and \( \text{H}_2\text{O}/\text{OH}^- \) is an acid/conjugate base pair in this equilibrium.
\( \text{H}_2\text{O/H}_3\text{O}^+ \) is a base/conjugate acid pair and \( \text{HC}_2\text{H}_3\text{O}_2/\text{C}_2\text{H}_3\text{O}_2^- \) an acid/conjugate base pair in this equilibrium.

d) \( \text{NH}_3/\text{NH}_4^+ \) is a base/conjugate acid pair and \( \text{H}_2\text{O/OH}^- \) is an acid/conjugate base pair in this equilibrium.

10.3 Mono-, Di-, and Triprotic Acids

Acidity of a compound is mainly created by certain types of hydrogen present in the acid molecule. Not all hydrogen in an acid molecule break off to produce \( \text{H}^+ \) ions: the acidic protons. Most common family acids are known as monoprotic acids have only one acidic proton. Each of these acids has a single \( \text{H}^+ \) ion, or proton, it can donate when it acts as a Brnsted acid. Hydrochloric acid (HCl), acetic acid (CH\(_3\)CO\(_2\)H or HOAc), nitric acid (HNO\(_3\)), and benzoic acid (C\(_6\)H\(_5\)CO\(_2\)H) are all monoprotic acids.

Polyprotic acids

Several important acids can be classified as polyprotic acids, which can lose more than one \( \text{H}^+ \) ion when they act as Brnsted acids.

Diprotic acids: Have two acidic protons as in the sulfuric acid (H\(_2\)SO\(_4\)), carbonic acid (H\(_2\)CO\(_3\)), hydrogen sulfide (H\(_2\)S), chromic acid (H\(_2\)CrO\(_4\)), and oxalic acid (H\(_2\)C\(_2\)O\(_4\)).

Triprotic acids: Have three acidic protons as in such as phosphoric acid (H\(_3\)PO\(_4\)) and citric acid (C\(_6\)H\(_8\)O\(_7\)).
10.4 Strengths of Acids and Bases

Strong acids and bases

In strong acids the first acidic hydrogen is completely separated or ionized (100%) exists as an hydronium ions, $\text{H}_3\text{O}^+$ in aqueous solutions.

$$\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ (\text{aq}) + \text{Cl}^- (\text{aq}) \ (100\% \text{ ions})$$

Strong bases also show complete separation (100%) into ions in aqueous solutions.

$$\text{NaOH(s)} + \text{H}_2\text{O} \rightarrow \text{Na}^+ (\text{aq}) + \text{OH}^-(\text{aq}) \ (100\% \text{ ions})$$

Most bases in Groups 1A and 2A are strong bases.

<table>
<thead>
<tr>
<th>Strong acid</th>
<th>Strong bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Name- formula</td>
<td>Name- formula</td>
</tr>
<tr>
<td>Hydriodic acid HI</td>
<td>Lithium hydroxide LiOH</td>
</tr>
<tr>
<td>Hydrobromic acid HBr</td>
<td>Sodium hydroxide NaOH</td>
</tr>
<tr>
<td>Perchloric acid HClO₄</td>
<td>Potassium hydroxide KOH</td>
</tr>
<tr>
<td>Hydronchloric acid HCl</td>
<td>Rubidium hydroxide RbOH</td>
</tr>
<tr>
<td>Chloric acid HClO₃</td>
<td>Cesium hydroxide CsOH</td>
</tr>
<tr>
<td>Sulfuric acid H₂SO₄</td>
<td>Calcium hydroxide Ca(OH)₂</td>
</tr>
<tr>
<td>Nitric acid HNO₃</td>
<td></td>
</tr>
</tbody>
</table>

In weak acids the first acidic hydrogen in some molecules are not separated or ionized (< 1-10%) only exists as an hydronium ions, $\text{H}_3\text{O}^+$ in aqueous solutions. Acetic acid, $\text{CH}_3\text{COOH}$ in water contains mostly molecules and a few $\text{CH}_3\text{COO}^-$ ions and it called **partial dissociation**.

$$\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ (\text{aq}) + \text{CH}_3\text{COO}^- (\text{aq}) \ (< 1\% \text{ ions})$$

Weak does not bases also show complete dissociation (< 1%) into ions in aqueous solutions.

$$\text{NH}_4\text{OH} + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \ (< 1\% \text{ ions})$$

<table>
<thead>
<tr>
<th>Weak acid</th>
<th>Weak bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Name- formula</td>
<td>Name- formula</td>
</tr>
<tr>
<td>Hydrofluoric acid HF</td>
<td>Magnesium hydroxide Mg(OH)₂</td>
</tr>
<tr>
<td>Formic acid HCOOH</td>
<td>Calcium hydroxide Ca(OH)₂</td>
</tr>
<tr>
<td>Acetic acid CH₂COOH</td>
<td>Strontium hydroxide Sr(OH)₂</td>
</tr>
<tr>
<td>Nitrous acid HNO₂</td>
<td>Barium hydroxide Ba(OH)₂</td>
</tr>
<tr>
<td>Hydrocyanic acid HCN</td>
<td></td>
</tr>
</tbody>
</table>
10.5 Ionization Constants for Acids and Bases

The equilibrium constants equilibrium expression for the ionization of strong and weak acids is useful in predicting the strength of acids and bases.

**For a strong acid** such as HCl in water:

\[ \text{HCl(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+ (\text{aq}) + \text{Cl}^- (\text{aq}) \quad (100 \text{ % ions}) \]

Equilibrium constant expression for the above equilibrium:

\[
K_a = \frac{[\text{H}_3\text{O}^+][\text{Cl}^-]}{[\text{HCl}]} = 1 \times 10^7 \quad \text{large for strong acids}
\]

**For a strong** bases such as NaOH which show complete separation (100%) into ions in aqueous solutions.

\[ \text{NaOH(s)} + \text{H}_2\text{O} \rightarrow \text{Na}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \quad (100 \text{ % ions}) \]

\[
K_b = \frac{[\text{Na}^+][\text{OH}^-]}{[\text{NaOH}]} = 1 \times 10^2 \quad \text{very large for strong bases}
\]

**For a weak acid** that does not show complete dissociation (< 1%) into ions in aqueous solutions.

\[ \text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ (\text{aq}) + \text{CH}_3\text{COO}^- (\text{aq}) \quad (<1 \text{ % ions}) \]

\[
K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 1 \times 10^{-8} \quad \text{small for weak acids}
\]

**For a weak acid** that does not show complete dissociation (< 1%) into ions in aqueous solutions.

\[ \text{NH}_4\text{OH} + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \quad (<1 \text{ % ions}) \]

\[
K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_4\text{OH}]} = 1 \times 10^{-8} \quad \text{small for weak bases}
\]

The definition of \( K_a \) is: The \( pK_a \) value is defined from \( K_a \) and can be calculated from the \( K_a \) value and from the equation

\[ pK_a = -\log (K_a) \]

**Comparing \( K_a \) & \( K_b \)**

- Any compound with a larger \( K_a \) value is the stronger acid.
- Any compound with a larger \( K_b \) value is the stronger base.
### Strong Acid - Strong Bases

<table>
<thead>
<tr>
<th>Name- formula</th>
<th>$K_a$</th>
<th>$pK_a$</th>
<th>Name- formula</th>
<th>$K_b$</th>
<th>$pK_b$</th>
</tr>
</thead>
<tbody>
<tr>
<td>HI</td>
<td>$\approx 10^{11}$</td>
<td>-11</td>
<td>LiOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HBr</td>
<td>$\approx 10^9$</td>
<td>-9</td>
<td>NaOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HClO₄</td>
<td>$\approx 10^7$</td>
<td>-7</td>
<td>KOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HCl</td>
<td>$\approx 10^7$</td>
<td>-7</td>
<td>RbOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HClO₃</td>
<td>$\approx 10^3$</td>
<td>-3</td>
<td>CsOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>$\approx 10^2$</td>
<td>-2</td>
<td>Ca(OH)₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HNO₃</td>
<td>$\approx 20$</td>
<td>-1.3</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Name- formula</th>
<th>$K_a$</th>
<th>$pK_a$</th>
<th>Name- formula</th>
<th>$K_b$</th>
<th>$pK_b$</th>
</tr>
</thead>
<tbody>
<tr>
<td>HF</td>
<td>$6.6 \times 10^{-4}$</td>
<td>3.18</td>
<td>Mg(OH)₂</td>
<td>0.01-0.1</td>
<td>-2 to -1</td>
</tr>
<tr>
<td>HCOOH</td>
<td>$1.77 \times 10^{-4}$</td>
<td>1.75</td>
<td>Ca(OH)₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH₃COOH</td>
<td>$1.76 \times 10^{-5}$</td>
<td>4.75</td>
<td>Sr(OH)₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HNO₂</td>
<td>$4.6 \times 10^{-4}$</td>
<td>3.34</td>
<td>Ba(OH)₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HCN</td>
<td>$6.17 \times 10^{-10}$</td>
<td>9.21</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Which acid is weaker? Answer:** HCN

HNO₂ ; $K_a = 4.0 \times 10^{-4}$.
HOCl₂ ; $K_a = 1.2 \times 10^{-2}$.
HOCI ; $K_a = 3.5 \times 10^{-8}$.
HCN ; $K_a = 4.9 \times 10^{-10}$. (Any compound with a smaller $K_a$ value is the weaker acid)

### Polyprotic Acids have more than one $K_a$

H₃PO₄ is triportic therefore has three $K_{a1}$, $K_{a2}$, and $K_{a3}$.

\[
\begin{align*}
\text{H}_3\text{PO}_4 + \text{H}_2\text{O} & \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{H}_2\text{PO}_4^- (aq); \quad K_{a1} = 7.1 \times 10^{-3} \\
\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} & \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{HPO}_4^{2-} (aq); \quad K_{a2} = 6.3 \times 10^{-8} \\
\text{HPO}_4^{2-} + \text{H}_2\text{O} & \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{PO}_4^{3-} (aq); \quad K_{a3} = 4.2 \times 10^{-13}
\end{align*}
\]

### Acid-Dissociation Equilibrium Constants for Common Polyprotic Acids

<table>
<thead>
<tr>
<th>Acid</th>
<th>$K_{a1}$</th>
<th>$K_{a2}$</th>
<th>$K_{a3}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>sulfuric acid (H₂SO₄)</td>
<td>$1.0 \times 10^{-3}$</td>
<td>$1.2 \times 10^{-2}$</td>
<td></td>
</tr>
<tr>
<td>chromic acid (H₂CrO₄)</td>
<td>9.6</td>
<td>3.2 $\times 10^{-7}$</td>
<td></td>
</tr>
<tr>
<td>oxalic acid (H₂C₂O₄)</td>
<td>$5.4 \times 10^{-2}$</td>
<td>$5.4 \times 10^{-5}$</td>
<td></td>
</tr>
<tr>
<td>sulfuric acid (H₂SO₃)</td>
<td>$1.7 \times 10^{-2}$</td>
<td>$6.4 \times 10^{-8}$</td>
<td></td>
</tr>
<tr>
<td>phosphoric acid (H₃PO₄)</td>
<td>$7.1 \times 10^{-3}$</td>
<td>$6.3 \times 10^{-8}$</td>
<td>$4.2 \times 10^{-13}$</td>
</tr>
<tr>
<td>glycine (C₂H₄NO₂)</td>
<td>$4.5 \times 10^{-3}$</td>
<td>$2.5 \times 10^{-10}$</td>
<td></td>
</tr>
<tr>
<td>citric acid (C₆H₈O₇)</td>
<td>$7.5 \times 10^{-4}$</td>
<td>$1.7 \times 10^{-7}$</td>
<td>$4.0 \times 10^{-7}$</td>
</tr>
<tr>
<td>carbonic acid (H₂CO₃)</td>
<td>$4.5 \times 10^{-7}$</td>
<td>$4.7 \times 10^{-11}$</td>
<td></td>
</tr>
<tr>
<td>hydrogen sulfide (H₂S)</td>
<td>$1.0 \times 10^{-7}$</td>
<td>$1.3 \times 10^{-13}$</td>
<td></td>
</tr>
</tbody>
</table>
Citric acid (C₆H₈O₇ or H₃C₆H₅O₇) is triprotic. Write three dissociation equilibriums for it:

\[
\begin{align*}
\text{H}_3\text{C}_6\text{H}_5\text{O}^+ + \text{H}_2\text{O} & \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{H}_2\text{C}_6\text{H}_5\text{O}^- (aq); \quad K_{a1} = 7.5 \times 10^{-4} \\
\text{H}_2\text{C}_6\text{H}_5\text{O}^- + \text{H}_2\text{O} & \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{HC}_6\text{H}_5\text{O}_2^- (aq); \quad K_{a2} = 1.7 \times 10^{-5} \\
\text{HC}_6\text{H}_5\text{O}_2^- + \text{H}_2\text{O} & \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{C}_6\text{H}_5\text{O}_3^- (aq); \quad K_{a3} = 4.0 \times 10^{-7}
\end{align*}
\]

**10.6 Salts**

Salts are formed when an acid reacts with a base to from water.

E.g. HCl(acid) + NaOH(base) --> NaCl (salt) + H₂O.

Salts normally don't have acidic protons or OH⁻ ions. One can get the parent acid and base from a salt by breaking the compound into metal (cation) and anion (non-metal or polyatomic anions) and adding H⁺ to anion and OH⁻ to metal.

Identify acids bases and salts among the following list:

<table>
<thead>
<tr>
<th>Acids</th>
<th>Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₃PO₄, HCl, NaOH, HCl, NaNO₃, H₂SO₄, HC₂H₃O₂, KNO₃, HNO₃, H₂SO₃</td>
<td>Mg(OH)₂, NaOH, KOH, H₂SO₃, NH₄OH</td>
</tr>
</tbody>
</table>

**Salts**: AgNO₃(Salt), NaCl(Salt), NaNO₃(salt), KNO₃(salt),

**Acids**: H₃PO₄(acid), HCl(acid), H₂SO₄(acid), HC₂H₃O₂(acid), HNO₃(acid), H₂SO₃(acid)

**Bases**: Mg(OH)₂(base), NaOH(base), KOH(base), NH₄OH(base).

**10.7 Acid-Base Neutralization Reactions**

Neutralization: This idea, that a base (or acid) can make an acid (or base) weaker by converting them to water is called neutralization. As you can see from the equations, acids release H⁺ into solution and bases release OH⁻. If we were to mix an acid and base together, the H⁺ ion would combine with the OH⁻ ion to make the molecule H₂O, or plain water: The neutralization reaction of an acid with a base will always produce water and a salt, as shown below:

\[
\begin{align*}
\text{H}^+(aq) + \text{OH}^-(aq) & \rightarrow \text{H}_2\text{O}(l) \\
\text{Acid} + \text{Base} & \rightarrow \text{Salt} + \text{H}_2\text{O}(l)
\end{align*}
\]

<table>
<thead>
<tr>
<th>Acid Name</th>
<th>Formula Name</th>
<th>Sodium salt Name</th>
<th>Sodium salt Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetic acid</td>
<td>H₃C₂H₃O₂</td>
<td>Sodium acetate</td>
<td>NaC₂H₃O₂</td>
</tr>
<tr>
<td>Hydrogen chloride</td>
<td>HCl</td>
<td>Sodium chloride</td>
<td></td>
</tr>
<tr>
<td>Nitric acid</td>
<td>HNO₃</td>
<td>Sodium nitrate</td>
<td>NaNO₃</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H₃PO₄</td>
<td>Sodium phosphate</td>
<td>Na₃PO₄</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>H₂SO₄</td>
<td>Sodium sulfate</td>
<td>Na₂SO₄</td>
</tr>
</tbody>
</table>
### 10.8 Self-Ionization of Water

A water molecule can react with another molecule of water to produce ions and the process is called autoionization.

\[
\text{H}_2\text{O(l)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3^+\text{O}(\text{aq}) + \text{OH}^-\text{(aq)}
\]

This equilibrium is called **autoionization** or **self-ionization** of water. The equilibrium constant for this reaction is called \( K_w \).

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-]; \quad K_w = 1 \times 10^{-14} \text{ is very small}
\]
and water there are only \( 1 \times 10^{-7} \) moles of \( \text{H}_3\text{O}^+ \) and \( 1 \times 10^{-7} \) moles of \( \text{OH}^- \) in one liter (1L) of water.

### 10.9 The pH Concept

In aqueous solutions both acids and bases alter the autoinization equilibrium of the water and change the concentration of hydrogen ions. Acids increase the concentration of hydrogen ions, while bases decrease the concentration of hydrogen ions (by converting them to water). The acidity or basicity of an aqueous solution, therefore, can be measured by its hydrogen ion concentration.

In 1909, the Danish biochemist Sörensen invented the **pH** scale for measuring acidity. The pH scale is described by the formula:

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

Note: concentration is commonly abbreviated by using square brackets, thus \([\text{H}^+] = \text{hydrogen ion concentration}\).

When measuring pH, \([\text{H}^+]\) is in units of moles of \(\text{H}^+\) per liter of solution.

**Neutral pH**

For example, pure water solution with \([\text{H}^+] = 1 \times 10^{-7}\) moles/liter has a pH equal to 7 (a simpler way to think about pH is that it equals the exponent on the \(\text{H}^+\) or \(\text{H}_3\text{O}^+\) concentration (written \([\text{H}^+]\) or \([\text{H}_3\text{O}^+]\)), ignoring the minus sign). The pH scale ranges from 0 to 14.

**Acidic pH**

Substances with a pH between 0 and less than 7 are acids (pH and \([\text{H}^+]\) are inversely related - lower pH means higher \([\text{H}^+]\)).

**Basic pH**

<table>
<thead>
<tr>
<th>Base Name</th>
<th>Formula Name</th>
<th>Chloride salt Name</th>
<th>Chloride salt Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium hydroxide</td>
<td>NaOH</td>
<td>Sodium chloride</td>
<td>NaCl</td>
</tr>
<tr>
<td>Barium oxide</td>
<td>BaO</td>
<td>Barium chloride</td>
<td>BaCl₂</td>
</tr>
<tr>
<td>Calcium oxide</td>
<td>CaO</td>
<td>Calcium chloride</td>
<td>CaCl₂</td>
</tr>
</tbody>
</table>
Substances with a pH greater than 7 and up to 14 are bases (higher pH means lower \([H^+])\). Right in the middle, at pH = 7, are neutral substances, for example, pure water. The relationship between \([H^+]\) and pH is shown in the table below alongside some

<table>
<thead>
<tr>
<th>pH</th>
<th>0</th>
<th>2</th>
<th>4</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>10</th>
<th>12</th>
<th>14</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>(1 \times 10^0)</td>
<td>(1 \times 10^{-2})</td>
<td>(1 \times 10^{-4})</td>
<td>(1 \times 10^{-6})</td>
<td>(1 \times 10^{-7})</td>
<td>(1 \times 10^{-8})</td>
<td>(1 \times 10^{-10})</td>
<td>(1 \times 10^{-12})</td>
<td>(1 \times 10^{-14})</td>
</tr>
</tbody>
</table>

What is the pH of a solution that has a hydrogen ion concentration of \(1.0 \times 10^{-3}\) M?

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

\[
pH = -\log (1 \times 10^{-3})
\]

\[
pH = (-1) \times \log 1 \times 10^{-3}
\]

\[
\log 1 = 0
\]

\[
\log 10^{-3} = -3
\]

\[
pH = (-1) \times (0 + -3) = (-1) \times -3
\]

\[
pH = 3
\]

What is the hydrogen ion concentration of a solution that has a pH of 5?

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

\[
5 = -\log [H^+]
\]

TRICK: Get minus sign with the number

\[
-5 = \log [H^+]
\]

\[
\log [H^+] = -5
\]

take the antilog

\[
[H^+] = \text{antilog} -5
\]

\[
[H^+] = 10^{-5}\text{ M}
\]

10.10 The pK\(_a\) Method for Expressing Acid Strength

**K\(_a\) and pK\(_a\) Values for acid dissociations**
The K\(_a\) value is a value used to describe the tendency of compounds or ions to dissociate and is also called the dissociation constant, the ionization constant and the acid constant.

The pK\(_a\) value is a value related to the K\(_a\) value in a logic manner. pK\(_a\) values are easier to remember than K\(_a\) values and pK\(_a\) values are in many cases easier to use for fast approximations of concentrations of compounds and ions in equilibriums than K\(_a\) values.
**Definition of pK\textsubscript{a}**

The definition of \( K_a \) is: The \( pK_a \) value is defined from \( K_a \) and can be calculated from the \( K_a \) value and from the equation

\[
pK_a = -\log(K_a)
\]

**Example on how pK\textsubscript{a} is calculated**

Given \( K_a = 1.8 \times 10^{-5} \) what is \( pK_a \)? Answer 4.74

\( pK_a = -\log K_a \) means \( pK_a = 4.74 \) do this in steps take \( \log 1.8 \times 10^{-5} = -4.74 \) and change sign (+/-) give 4.74

Any time you see a "p" in terms of acid-base chemistry, it automatically signifies \(-\log\) (whatever is after p).

**Comparing pK\textsubscript{a} & pK\textsubscript{b}**

- Any compound with a smaller \( pK_a \) value is the stronger acid.
- Any compound with a smaller \( pK_b \) value is the stronger base.

**Which acid is Stronger? Answer: propanoic acid**

<table>
<thead>
<tr>
<th>Acid</th>
<th>( pK_a )</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCOOH (formic acid)</td>
<td>3.75</td>
</tr>
<tr>
<td>CH(_3)COOH (acetic acid)</td>
<td>4.76</td>
</tr>
<tr>
<td>CH(_3)CH(_2)COOH (propanoic acid)</td>
<td>4.86</td>
</tr>
</tbody>
</table>

*Any compound with a smaller \( pK_a \) value is the stronger acid.*

**10.11 The pH of Aqueous Salt Solutions**

Some salts dissolved in water make the solution acidic or basic depending on the type of salt. For example, when ammonium chloride, \( \text{NH}_4\text{Cl} \), is dissolved in water the solution become acidic. Since \( \text{NH}_4\text{Cl} \) is produced by the reaction of \( \text{HCl} \) (a strong acid) with \( \text{NH}_4\text{OH} \) (a weak base) which is combination to produce acidic solutions.

1) Salt of **strong acid** and **strong base** forms ions which are **neutral**.
2) Salt of **weak acid** and **strong base** forms ions which are **basic**.
3) Salt of **strong acid** and **weak base** forms ions which are **acidic**.
4) Salt of **weak acid** and **weak base** forms ions which are either neutral/ basic/acidic depending on the relative strengths of the acid and the base.

**Which of the following salt solutions in water would produce ions which are acidic, basic or neutral ions?**

a) \( \text{NaCl} \)  b) \( \text{NaC}_2\text{H}_3\text{O}_2 \)  c) \( \text{NaHSO}_4 \)  d) \( \text{NH}_4\text{Cl} \)
**Answer:** a) NaCl (neutral)  b) NaC₂H₃O₂ (basic)  c) Na₂SO₄ (neutral)  d) NH₄Cl (acidic)

**Chemistry at a Glance: Acids and Acidic Solutions**

**10.12 Buffers**

A buffer solution are usually combinations of weak acid/soluble-salt or weak base/soluble-salt combinations. Buffers resist changes in pH when small quantities of an acid or an alkali are added to it.

**Acidic buffer solutions**

An acidic buffer (weak base/soluble-salt: NH₃/NH₄Cl) solution is simply one which has a pH less than 7. Acidic buffer solutions are commonly made from a weak acid and one of its salts - often a sodium salt.

Acidic buffer: weak acid/soluble-salt of weak base; NH₃/NH₄Cl

The main hydrolysis equilibrium that creates the acidity is:

\[
\text{H}_2\text{O} + \text{NH}_4^+ \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3
\]

The common ion NH₄⁺ will increase the H⁺ or [H₃O⁺] decreasing pH of the solution. Similar to [HC₂H₃O₂/NaC₂H₃O₂] this equilibrium can buffer changes in pH.

**Basic buffer solutions**

Basic buffer solutions buffer (weak acid/soluble-salt: HC₂H₃O₂/NaC₂H₃O₂) solution is simply one which has a pH more than 7. Basic buffer solutions are commonly made from a weak acid and one of its salts - often a sodium salt.

Regardless of acidic buffer or basic buffers, you can change the pH of the buffer solution by changing the ratio of acid to salt, or by choosing a different acid and one of its salts.

Basic buffer: weak acid/soluble-salt of weak acid; HC₂H₃O₂/NaC₂H₃O₂.

The main hydrolysis equilibrium that creates the basicity is:

\[
\text{H}_2\text{O} + \text{C}_2\text{H}_3\text{O}_2^- \rightleftharpoons \text{HO}^- \text{(basicity)} + \text{H}_2\text{C}_2\text{H}_3\text{O}_2 \text{ (acid dissociation)}
\]

The common ion C₂H₃O₂⁻ will decrease the [HO⁻] increasing pH of the solution. This solution can act as a buffer for changes in pH. If the pH of the solution is increased (adding more HO⁻) this equilibrium will shift to left to maintain the original pH. If the pH of solution is decreased (adding H⁺) the equilibrium will shift to right to remove extra H⁺.

**10.13 The Henderson-Hasselbalch Equation**
The Henderson-Hasselbalch equation can be used to prepare buffer solutions and to estimate charges on ionizable species in solution, such as amino acid side chains in proteins. Caution must be exercised in using this equation because pH is sensitive to changes in temperature and salt concentration in the solution being prepared. This equation can then be rearranged to give the **Henderson-Hasselbalch** equation:

\[
pH = pK_a + \log \left( \frac{[\text{Base}]}{[\text{Acid}]} \right)
\]

### Calculate the pH of a solution in which concentrations of NH₄Cl is 0.9346 M and NH₃ is 1.9072 M. (\(K_a(\text{NH}_4\text{Cl}) = 5.56 \times 10^{-10}\))

\[
\begin{align*}
[\text{Base}] = \text{[NH}_3]\text{ = 1.9072 M and [Acid] = [NH}_4\text{Cl]}\text{ = 0.9346 M.}
\end{align*}
\]

\[
pK_a = -\log K_a = -[\log (5.56 \times 10^{-10})] = -(9.26) = 9.26
\]

\[
\begin{align*}
\text{pH} &= pK_a + \log \left( \frac{[\text{Base}]}{[\text{Acid}]} \right) \\
\text{pH} &= 9.26 + \log \left( \frac{1.9072}{0.9346} \right) \\
\text{pH} &= 9.26 + 0.3098 = 9.57
\end{align*}
\]

**Chemistry at a Glance: Buffer Systems**

### 10.14 Electrolytes

**Electrolytes**

Certain substances that are called **electrolytes** produce ions when they dissolve in solution. Because these ions are free to move in the solution, the solution **conducts electricity strongly**. Ions can be produced in solution in either of two ways. Electrolytes can be either ionic compounds (i.e. sodium hydroxide, potassium nitrate) that dissolve in water, giving solutions of ions, or they may be covalent compounds that react with water and form ions in solution as a result. Strong electrolytes fall into three categories: **strong acids**, **strong bases**, and **salts**. (Salts are sometimes also called ionic compounds, but really strong bases are ionic compounds as well.)

Strong Electrolytes are strong acids such as HCl, HBr, HI, HNO₃, HClO₃, HClO₄, and H₂SO₄.
Strong bases NaOH, KOH, LiOH, Ba(OH)$_2$, and Ca(OH)$_2$
Soluble salts NaCl, KBr, and MgCl$_2$, nitrate and acetate salts.

**Weak Electrolytes**

Weak electrolytes are compounds that partially ionize when they dissolve in water. Because there are fewer ions that are free to move in the solution, the solution **does not conduct electricity**. The weak electrolytes include weak acids and weak bases. HF, HC$_2$H$_3$O$_2$ (acetic acid), H$_2$CO$_3$ (carbonic acid), H$_3$PO$_4$ (phosphoric acid), and weak bases such as NH$_3$ (ammonia), C$_5$H$_5$N (pyridine)

**Non-Electrolytes**

Non-electrolytes are compounds that don't ionize when they dissolve in water. Because there are no ions that are free to move in the solution, the solution **conducts electricity weakly**. Non-electrolytes are limited to **covalent compounds**. Many molecular compounds of carbon such as methane CH$_4$, benzene C$_6$H$_6$, ethanol C$_2$H$_5$OH, ether (C$_2$H$_5$)$_2$O, and formaldehyde CH$_2$O, are non-electrolytes.

**10.15 Acid-Base Titrations**

A **acid/base titration** is a procedure used in analytical chemistry to determine the amount or unknown concentration of a acid or base. In a titration one reagent, the **titrant**, is added to another slowly using a **burette** to a second solution in a conical flask. As it is added a chemical stoichiometric reaction occurs until one of the reagents is exhausted, and some process or device signals that this has occurred. The purpose of a titration is generally to determine the quantity or concentration of one of the reagents, that of the other being known beforehand. In any titration at the **end point** there must be a rapid quantitative reaction taking place as the titrant is added, and in acid-base titrations this is a stoichiometric **neutralization**.

**Acid-Base Indicators**

Many substances, including **litmus**, the one dye almost everyone associates with acids and bases, change color in response to acid or base. The pigment in **red cabbage** is another natural substance very commonly used to show color change. **Phenolphthalein** is one of
the most common indicators used for beginning chemistry, because its color change is very obvious which makes it easy to use. There are many other indicators that change colors at different pH's, and so are useful for different purposes. pH paper commonly contains a mixture of different indicators that change colors at different pH's. The mixture is applied to paper, and then compared to a color chart to see what the pH of a solution is, approximately.

Chemical Connections: Excessive Acidity Within the Stomach: Antacids and Acid Inhibitors; Acid Rain: Excess Acidity; Blood Plasma pH and Hydrolysis; Buffering Action in Human Blood; Electrolytes and Body Fluids