Chapter Outline

• 15.1 Acids and Bases: The Brønsted–Lowry Model
• 15.2 Acid Strength and Molecular Structure
• 15.3 pH and the Autoionization of Water
• 15.4 Calculations Involving pH, $K_a$, and $K_b$
• 15.5 Polyprotic Acids
• 15.6 pH of Salt Solutions
• 15.7 The Common-Ion Effect
• 15.8 pH Buffers
• 15.9 pH Indicators and Acid–Base Titrations
• 15.10 Solubility Equilibria

pH of Salt Solutions

1. Neutral Salts (pH = 7) are from strong electrolytes (100% ionization)

(a) Ionic Compounds:

\[
\begin{align*}
\text{NaCl}\text{(aq)} & \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \\
\text{HCl}\text{(aq)} + \text{H}_2\text{O}(l) & \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})
\end{align*}
\]

\[\text{base} \quad \downarrow \quad \text{conj. acid}\]

\[\text{acid} \quad \downarrow \quad \text{conj. base}\]

Infinitely strong \quad \text{Infinitely weak}
2. Basic Salts (pH > 7) are conjugate bases of weak acids

\[ \text{HClO}(aq) + \text{H}_2\text{O}(l) = \text{H}_3\text{O}^+(aq) + \text{ClO}^-(aq) \]
\[ \text{weak acid} \hspace{1cm} \text{conj. base} \]

\[ \text{ClO}^-(aq) + \text{H}_2\text{O}(l) = \text{OH}^-(aq) + \text{HClO}(aq) \]
\[ \text{conj. base} \hspace{1cm} \text{pH > 7} \]

\[ K_a (\text{HClO}) = 2.9 \times 10^{-8} \quad K_b = \frac{K_w}{K_a} \quad K_b = \frac{1.0 \times 10^{-14}}{2.9 \times 10^{-8}} \]

\[ K_b (\text{ClO}^-) = 3.4 \times 10^{-7} \]

3. Acidic Salts (pH < 7) are conjugate acids of weak bases

\[ \text{NH}_3(aq) + \text{H}_2\text{O}(l) = \text{OH}^-(aq) + \text{NH}_4^+(aq) \]
\[ \text{weak base} \hspace{1cm} \text{conj. acid} \]

\[ \text{NH}_4^+(aq) + \text{H}_2\text{O}(l) = \text{H}_3\text{O}^+(aq) + \text{NH}_3(aq) \]
\[ \text{conj. acid} \hspace{1cm} \text{pH < 7} \]

\[ K_b (\text{NH}_3) = 1.8 \times 10^{-5} \quad K_a = \frac{K_w}{K_b} \quad K_a = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} \]

\[ K_a (\text{NH}_4^+) = 5.6 \times 10^{-10} \]
Is an aqueous solution of NaClO acidic, basic, or neutral?

Calculating the pH of Solutions of Weak Acids and Bases: Use the RICE Table as Before

1. Calculating the pH of a Solution of a Basic Salt

\[
\text{ClO}^-(aq) + H_2O(l) = OH^-(aq) + HClO(aq) \quad K_a = 2.9 \times 10^{-8}
\]

\[
K_b = 3.4 \times 10^{-7} = \frac{x^2}{0.100 M - x}
\]

Since \( K_a \) is < 10^{-5}, assume that \( x \ll 0.100 \text{ M} \)

\[
x = \sqrt{(3.4 \times 10^{-7})(0.100)}
\]

\[
x = [OH^-] = 1.9 \times 10^{-4} \text{ M} \quad \text{Assumption OK}
\]

\[
\text{pH} = -\log (1.9 \times 10^{-4}) = 3.7 \quad \text{pH} = 14 - 3.7 = 10.3
\]
Calculating the pH of Solutions of Weak Acids and Bases:
Use the RICE Table as Before

2. Calculating the pH of a Solution of an Acidic Salt (Ex. 15.8)

What is the pH of a 0.25 M solution of NH₄Cl?

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The Common Ion Effect

A shift in equilibrium caused by the addition of a compound having an ion in common with the dissolved substance.

\[
\text{CH}_3\text{COONa} (s) \rightarrow \text{Na}^+ (aq) + \text{CH}_3\text{COO}^- (aq)
\]

\[
\text{CH}_3\text{COOH} (aq) \leftrightharpoons \text{H}^+ (aq) + \text{CH}_3\text{COO}^- (aq)
\]

The common ion effect can be used to produce a **BUFFER SOLUTION** = a solution of a weak acid or base and it's conjugate, e.g. CH$_3$COOH and CH$_3$COONa

By controlling the ratio of weak acid/base to it's conjugate, we can shift the equilibrium to whatever [H$^+$] and therefore pH we want.

The Henderson-Hasselbach Equation

\[
\text{HA}(aq) + \text{H}_2\text{O}(aq) = \text{H}_3\text{O}^+ (aq) + \text{A}^- (aq)
\]

\[
K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}
\]
1. Solution using the RICE table

2. Solution using the Henderson-Hasselbach equation