Types of Weak Bases

- Bases generally have lone pairs or negative charges in order to attack protons.
- Most neutral weak bases contain nitrogen.
- Amines are related to ammonia and have one or more N-H bonds replaced with N-C bonds (e.g., CH₃NH₂ is methylamine).
- Anions of weak acids are also weak bases. Example: OCl⁻ is the conjugate base of HOCl (weak acid): 
  \[
  \text{ClO}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HClO}(aq) + \text{OH}^-\text{(aq)} \quad K_b = 3.3 \times 10^{-7}
  \]

Relationship Between \(K_a\) and \(K_b\)

- We need to quantify the relationship between strength of acid and conjugate base.
- When two reactions are added to give a third, the equilibrium constant for the third reaction is the product of the equilibrium constants for the first two:
  - Reaction 1 + reaction 2 = reaction 3
  - \(K_1 \times K_2 = K_3\).

Relationship Between \(K_a\) and \(K_b\)

- For a conjugate acid-base pair \(K_a \times K_b = K_w\).
- Therefore, the larger the \(K_{act}\), the smaller the \(K_{b}\). That is, the stronger the acid, the weaker the conjugate base.
- Taking negative logarithms:
  \[pK_a + pK_b = pK_w\]

Acid-Base Properties of Salt Solutions

- Nearly all salts are strong electrolytes.
- Therefore, salts exist entirely of ions in solution.
- Acid-base properties of salts are a consequence of the reaction of their ions in solution.
- The reaction in which ions produce H⁺ or OH⁻ in water is called hydrolysis.
- Anions from weak acids are basic.
- Anions from strong acids are neutral.
- Anions with ionizable protons (e.g. HSO₄⁻) are amphoteric.

- To determine whether a salt has acid-base properties we use:
  - Salts derived from a strong acid and strong base are neutral (e.g. NaCl, Ca(NO₃)₂).
  - Salts derived from a strong base and weak acid are basic (e.g. NaOCl, Ba(C₂H₃O₂)₂).
  - Salts derived from a weak base and strong base are acidic (e.g. NH₄Cl, Al(NO₃)₃).
  - Salts derived from a weak acid and weak base can be either acidic or basic. Equilibrium rules apply!
Factors That Affect Acid Strength
Consider H-X. For this substance to be an acid we need:
• H-X bond to be polar with H\(\delta^+\) and X\(\delta^-\) (if X is a metal then the bond polarity is H\(\delta^-\), X\(\delta^+\) and the substance is a base),
• the H-X bond must be weak enough to be broken,
• the conjugate base, X\(^-\), must be stable.

Binary Acids
• Acid strength increases across a period and down a group.
• Conversely, base strength decreases across a period and down a group.
• HF is a weak acid because the bond energy is high.
• The electronegativity difference between C and H is so small that the C-H bond is non-polar and CH\(_4\) is neither an acid nor a base.

Oxyacids
• Oxyacids contain O-H bonds.
• All oxyacids have the general structure Y-O-H.
• The strength of the acid depends on Y and the atoms attached to Y.
  – If Y is a metal (low electronegativity), then the substances are bases.
  – If Y has intermediate electronegativity (e.g. I, EN = 2.5), the electrons are between Y and O and the substance is a weak oxyacid.
Oxyacids
- If Y has a large electronegativity (e.g. Cl, EN = 3.0), the electrons are located closer to Y than O and the O-H bond is polarized to lose H⁺.
- The number of O atoms attached to Y increases the O-H bond polarity and the strength of the acid increases (e.g. HOCl is a weaker acid than HClO₂ which is weaker than HClO₃ which is weaker than HClO₄ which is a strong acid).

Carboxylic Acids
- These are organic acids which contain a COOH group (R is some carbon containing unit):

- Brønsted-Lowry acid is a proton donor.
- Focusing on electrons: a Brønsted-Lowry acid can be considered as an electron pair acceptor.
- Lewis acid: electron pair acceptor.
- Lewis base: electron pair donor.
- Note: Lewis acids and bases do not need to contain protons.
- Therefore, the Lewis definition is the most general definition of acids and bases.

<table>
<thead>
<tr>
<th>Acid</th>
<th>EN of Y</th>
<th>$K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>HClO</td>
<td>3.0</td>
<td>$3.0 \times 10^{-8}$</td>
</tr>
<tr>
<td>HBrO</td>
<td>2.8</td>
<td>$2.5 \times 10^{-9}$</td>
</tr>
<tr>
<td>HIO</td>
<td>2.5</td>
<td>$2.3 \times 10^{-11}$</td>
</tr>
</tbody>
</table>
Lewis acids generally have an incomplete octet (e.g. BF$_3$).
Transition metal ions are generally Lewis acids.
Lewis acids must have a vacant orbital (into which the electron pairs can be donated).
Compounds with $\pi$-bonds can act as Lewis acids:

$$H_2O(l) + CO_2(g) \rightarrow H_2CO_3(aq)$$

**Hydrolysis of Metal Ions**

- Metal ions are positively charged and attract water molecules (via the lone pairs on O).
- The higher the charge, the smaller the metal ion and the stronger the M-OH$_2$ interaction.
- Hydrated metal ions act as acids:

$$Fe(H_2O)_6^{3+}(aq) \rightleftharpoons Fe(H_2O)_5(OH)^{2+}(aq) + H^+(aq) \quad K_a = 2 \times 10^{-3}$$

- The pH increases as the size of the ion increases (e.g. Ca$^{2+}$ vs. Zn$^{2+}$) and as the charge increases (Na$^+$ vs. Ca$^{2+}$ and Zn$^{2+}$ vs. Al$^{3+}$).