ACIDS AND BASES

Acids / Bases

**ACID**
- Sour taste, turn litmus paper red, low pH

**BASE**
- Bitter taste, turn litmus paper blue, high pH

Arrhenius Definition

**ACID**
- Donates a hydrogen ion (H\(^+\)) in water

**BASE**
- Donates a hydroxide ion (OH\(^-\)) in water

Limited to substances with those “parts”... **NH\(_3\)** exception!

Bronsted-Lowry Definition

**ACID**
- Donates a proton in water

**BASE**
- Accepts a proton in water

Better definition that explains **NH\(_3\)** as a base!

Lewis Definition

**ACID**
- Accepts an electron pair

**BASE**
- Donates an electron pair

Explains all traditional acids and bases!

\(H^+\) Ion

- Just like a proton in the nucleus with no surrounding valence e-
- \(H^+\) reacts strongly with the lone pairs of e- around the oxygen atom of water
- \(H_3O^+\) (hydronium ion) forms from \(H^+\) reacting with \(H_2O\)

\(H^+\) and \(H_3O^+\) are interchangeable!
**Amphiprotic / Amphoteric**

Molecules or ions that can act as EITHER acids or bases

- Amphiprotic... capable of accepting or donating a proton
- Ex: \( \text{H}_2\text{O} \)

**Conjugate Acid-Base Pairs**

Pair of compounds that differ by the presence of one H\(^+\) unit

\[ \text{HA} \rightleftharpoons \text{H}^+ + \text{A}^- \]

Forward: **Acid** \(\rightarrow\) **Conjugate Base**

Reverse: **Conjugate Acid** \(\rightarrow\) **Base**

**Practice**

- **EXAMPLE:**
  Write the conjugate base for each of the following acids.
  - \( \text{CH}_3\text{CH}_2\text{OH} \)
  - \( \text{H}_2\text{O}^- \)
  - \( \text{H}_2\text{PO}_4^- \)
  - \( \text{NH}_4^+ \)
  - \( \text{HCO}_3^- \)

**Practice**

- **EXAMPLE:**
  Write the conjugate acid for each of the following bases.
  - \( \text{NH}_2^- \)
  - \( \text{OH}^- \)
  - \( \text{H}_2\text{O} \)
  - \( \text{NH}_3 \)
  - \( \text{H}_2\text{PO}_4^- \)

**Practice**

- **EXAMPLE:**
  For the following reaction, identify the **acid** on the left and its **conjugate base** on the right. Similarly, identify the **base** on the left and its **conjugate acid** on the right.

\[ \text{HBr} + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{Br}^- \]
Strengths of Acids and Bases

**STRONG Acids / Bases**
- Dissociate **COMPLETELY** in water
- Very large K values (no equilibrium though)
  
  Ex: \( \text{HCl} \rightarrow \text{H}^+ + \text{Cl}^- \)

**WEAK Acids / Bases**
- Dissociate only **SLIGHTLY** in water
- Very small K values

  Ex: \( \text{HF} \rightarrow \text{H}^+ + \text{F}^- \)

*NOTE ABOUT WEAK ACIDS / BASES AND NET-IONIC EQUATIONS:*

When writing net-ionic equations on the AP Exam, the WEAK acids / bases do NOT rip apart into ions since they are not 100% dissociated!

Ex: Write net-ionic for \( \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{NaOH}(aq) \rightarrow \)

**6 STRONG ACIDS:**
- HCl
- HBr
- HI
- HNO\(_3\)
- H\(_2\)SO\(_4\)
- HClO\(_4\)

**5 STRONG BASES:**
- LiOH
- NaOH
- KOH
- Ba(OH)\(_2\)
- Sr(OH)\(_2\)

**DO NOT CONFUSE STRENGTH (HOW MUCH DISSOCIATION TAKES PLACE) WITH CONCENTRATION (mol/L)!!**

**EXAMPLES:**

If we have 1.25 moles of HCl in water, how many moles of H\(^+\) and Cl\(^-\) will there be?

If we have 0.236 M HNO\(_3\), what is the [H\(^+\)] and [NO\(_3^-\)]?

**STRONGER the acid, the **WEAKER** the conjugate base!**

**STRONGER the base, the **WEAKER** the conjugate acid!**

\( \text{HCl} \rightarrow \text{H}^+ + \text{Cl}^- \)

Conjugate base (Cl\(^-\)) of the strong acid does **NOT** want to grab the H\(^+\).
**Strengths of Acids and Bases**

- In every acid-base reaction, the position of the equilibrium favors the transfer of the proton from the **STRONGER** acid to the **STRONGER** base to form the weaker acid and the weaker base...

**WEAKER ACIDS / BASES ARE FAVORED AT EQUILIBRIUM!!**

**Factors Affecting Strength**

- Not every molecule with an H is an acid or every molecule with OH a base...
- Strong bonds (X–H) don’t want to break to give off H⁺
- High bond POLARITY (H bonded to electronegative atom) tends to **INCREASE** the acidity of the H atom because the H wants “get away”
- Example:

  ![Dimethyl Ether](image)

  **NOT ACIDIC!**

**Autoionization of Water**

\[ \text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^- \]

- Very important because it allows water to act as an acid or base

\[ K_w = [\text{H}^+][\text{OH}^-] \]

\[ K_w = 1 \times 10^{-14} \text{ at } 25^\circ C \]

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

**Factors Affecting Strength**

- Molecules of the form H–O–X:
  - If X has high electronegativity... H tends to be **ACIDIC**
  - If X has low electronegativity... OH tends to come off and the compound is **BASIC**
  - More oxygens around X... the more **ACIDIC**

**pH Scale**

- pH is a logarithmic scale which is an exponential scale
- Ex: pH of 3 is **10 times** stronger than a pH of 4 and **100 times** stronger than a pH of 5

\[ [\text{H}^+] = [\text{OH}^-] \rightarrow \text{NEUTRAL (pH = 7)} \]

\[ [\text{H}^+] > [\text{OH}^-] \rightarrow \text{ACIDIC (pH < 7)} \]

\[ [\text{H}^+] < [\text{OH}^-] \rightarrow \text{BASIC (pH > 7)} \]
**pH Scale**

_How acidic or basic a SOLUTION is_

- Lower pH is more acidic
- Neutral
- Higher pH is more alkaline

**pH Formula**

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

**pOH Formula**

\[ \text{pOH} = -\log [OH^-] \]

**pH + pOH = 14**

**SIG FIGS with pH**

Use as many **DECIMAL PLACES** on pH as there are SF in the **LEAST** accurate measurement...

(Ex: 1.15 M has 3 SF, so pH with 3 decimal places)

**pH Scale**

Need to be able to look at \([H^+] = 1.0 \times 10^{-5}\) and know that the pH is 5 **WITHOUT** using a calculator for the the MC part of the AP Exam!

(If base # is 1, pH = positive value of superscript)

**pH Scale**

- **EXAMPLES:**
  
  In a sample of lemon juice, \([H^+] = 3.8 \times 10^{-4} \text{ M}\). What is the pH?

  A window-cleaning solution has \([OH^-] = 1.9 \times 10^{-6} \text{ M}\). What is the pH?

**pH Scale**

- **EXAMPLES:**

  A solution formed by dissolving an antacid tablet has a pH of 9.18. Calculate the \([H^+]\).

  The pH of a sample of human blood was measured to be 7.41 at 25° C. Calculate the pOH, \([H^+],\) and \([OH^-]\) for the sample.

**Strong Acids and Bases**

- **Exist in solution ENTIRELY as ions**

- **Do NOT exist in equilibrium... reaction lies entirely to the right (side with the ions)**

- **Determine the \([H^+]\) or \([OH^-]\) to find the pH or pOH**
Strong Acids and Bases

- **EXAMPLES:**
  
  Calculate the pH of 0.050 M HNO_3.
  
  What is the pH of 5.0 \times 10^{-2} \ M \text{NaOH}?

  What is the concentration of a solution of KOH for which the pH is 11.89?

Weak Acids and Bases

- Only **PARTIALLY** ionized in solution
- Use equilibrium rules with ICE to determine how much of the acid or base dissociates and find the [ ]
- Use \( K_a \) (**Acid Dissociation Constant**) for weak acids and \( K_b \) (**Base Dissociation Constant**) for weak bases

\[
\begin{align*}
\text{HA} & \rightleftharpoons \text{H}^+ + \text{A}^- \\
\text{B} + \text{H}_2\text{O} & \rightleftharpoons \text{HB}^+ + \text{OH}^- \\
K_a &= \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \\
K_b &= \frac{[\text{HB}^+][\text{OH}^-]}{[\text{B}]} 
\end{align*}
\]

Weak Acids and Bases

- **EXAMPLE:**
  What is the \( K_a \) of 0.10 M of formic acid, HCOOH, if its pH is 2.38?

\[
\begin{align*}
\text{HCOOH} & \rightleftharpoons \text{H}^+ + \text{COOH}^- \\
\text{Initial} & \quad 0.10 \ \text{M} \quad 0 \quad 0 \\
\text{Change} & \quad -4.2 \times 10^{-3} \ \text{M} \quad +4.2 \times 10^{-3} \ \text{M} \quad +4.2 \times 10^{-3} \ \text{M} \\
\text{Equilibrium} & \quad 4.2 \times 10^{-3} \ \text{M} \\
\end{align*}
\]

pH = 2.38

\([\text{H}^+] = 4.2 \times 10^{-3} \ \text{M}\)
Example: What is the $K_a$ of 0.10 M of formic acid, HCOOH, if its pH is 2.38?

\[ \text{HCOOH} \rightleftharpoons \text{H}^+ + \text{COOH}^- \]

<table>
<thead>
<tr>
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<th>HCOOH</th>
<th>H$^+$</th>
<th>COOH$^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.10 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-4.2 x 10$^{-3}$ M</td>
<td>+4.2 x 10$^{-3}$ M</td>
<td>+4.2 x 10$^{-3}$ M</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>9.6 x 10$^{-2}$ M</td>
<td>4.2 x 10$^{-3}$ M</td>
<td>4.2 x 10$^{-3}$ M</td>
</tr>
</tbody>
</table>

\[ \text{pH} = 2.38 \]
\[ [\text{H}^+] = 4.2 \times 10^{-3} \text{ M} \]

$$K_a = \frac{[\text{H}^+] [\text{COOH}^-]}{[\text{HCOOH}]} = \frac{(4.2 \times 10^{-3})^2}{9.6 \times 10^{-2}} = 1.8 \times 10^{-4}$$

Example: What is the pH of 0.20 M HC$_2$H$_3$O$_2$ with a $K_a = 1.8 \times 10^{-5}$?

\[ \text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^- \]

<table>
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<th>C$_2$H$_3$O$_2^-$</th>
</tr>
</thead>
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<tr>
<td>Initial</td>
<td>0.20 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.20 - x M</td>
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$$K_a = \frac{[\text{H}^+] [\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

For these problems, we ALWAYS ignore what is subtracted or added because it is NEGLIGIBLE!
**Weak Acids and Bases**

**EXAMPLE:**

What is the pH of 0.20 M HC₂H₃O₂ with a $K_a = 1.8 \times 10^{-5}$?

$$\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$$

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<td>HC₂H₃O₂</td>
<td>0.20 M</td>
<td>-x</td>
<td>0.20 - x M</td>
</tr>
<tr>
<td>H⁺</td>
<td>0</td>
<td>+x</td>
<td>x M</td>
</tr>
<tr>
<td>C₂H₃O₂⁻</td>
<td>0</td>
<td>+x</td>
<td>x M</td>
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</table>

$$1.8 \times 10^{-5} = \frac{x^2}{0.20 - x} \quad x = [\text{H}^+] = 1.9 \times 10^{-3}$$

$$\text{pH} = -\log [\text{H}^+] = 2.72$$

**Weak Acids and Bases**

- In mixtures of **WEAK** acids, only the acid with the **LARGEST** $K_a$ value will contribute an appreciable $[\text{H}^+]$. Use this acid to determine the pH (same is true for bases too).

- In mixtures of **STRONG** acids, determine the moles of H⁺ from both and **ADD** together to give total moles. Determine [ ] from the new total volume to find the pH.

**Weak Acids and Bases**

- **EXAMPLES:**

  - The hypochlorite ion (OCl⁻) is a strong oxidizing agent often found in household bleaches. Calculate the pH of a 0.10 M solution of hypochlorous acid ($K_a = 3.5 \times 10^{-8}$)?

  - Calculate the pH for a 15.0 M solution of NH₃ ($K_b = 1.8 \times 10^{-5}$).

**% Dissociation / Ionization**

- Describes the strength of an acid just like $K_a$.

  - **GREATER % Dissociation... STRONGER** the acid (remember strong acids are 100% dissociated).

  $$\text{% Dissociation / Ionization} = \frac{[\text{H}^+ \text{ or } [\text{OH}^-] \text{ Ionized}]}{[\text{Acid} \text{ or } [\text{Base}]} \times 100$$

- **EXAMPLE:**

  - What is the % Dissociation for 0.10 M HCOOH which contains $[\text{H}^+] = 4.2 \times 10^{-3}$ M? 4.2%
**% Dissociation / Ionization**

- **EXAMPLES:**
  
  Calculate the % ionization of 0.15 M acetic acid ($K_a = 1.8 \times 10^{-5}$).

  Lactic acid ($\text{H}_3\text{C}_3\text{H}_5\text{O}_3$) is a waste product that accumulates in muscle tissue during exertion. In a 0.10 M solution, lactic acid is 3.7% dissociated. Calculate the value of $K_a$ for this acid.

**Polyprotic Acids**

- **Acids that can donate MORE than one $H^+$**

  
  - **MONOPROTIC:** acids donating ONE $H^+$  
    (Ex: $\text{HNO}_3$)
  
  - **DIPROTIC:** acids donating TWO $H^+$  
    (Ex: $\text{H}_2\text{SO}_4$)
  
  - **POLYPROTIC BASES:** accept MORE than one $H^+$  
    (Ex: $\text{PO}_4^{3-}$)

  - Dissociate in steps one proton at a time with each step having its own $K_a$ value

- **EXAMPLE:**
  
  If $[\text{H}_2\text{CO}_3] = 0.0037 \text{ M}$, what is the pH and $[\text{CO}_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)

  
  \[
  \begin{align*}
  \text{H}_2\text{CO}_3 & \rightleftharpoons \text{H}^+ + \text{HCO}_3^- \\
  \end{align*}
  \]

  - **Initial**
    
    | $\text{H}_2\text{CO}_3$ | $\text{H}^+$ | $\text{HCO}_3^-$ |
    |----------------|---------|----------------|

  - **Change**
    
    $-x$ $x$ $x$

  - **Equilibrium**
    
    | $\text{H}_2\text{CO}_3$ | $\text{H}^+$ | $\text{HCO}_3^-$ |
    |----------------|---------|----------------|

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**Polyprotic Acids**

*Acids that can donate MORE than one H⁺*

- **EXAMPLE:**
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\[
H_2CO_3 ⇌ H^+ + HCO_3^-
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<td>0.0037 M</td>
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<td>-x</td>
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\[x = [H^+] = 4.0 \times 10^{-5}\quad \text{pH} = -\log [H^+] = 4.40\]

---

**Polyprotic Acids**

*Acids that can donate MORE than one H⁺*

- **EXAMPLE:**
  If \([H_2CO_3] = 0.0037 \text{ M}\), what is the pH and \([CO_3^{2-}]\)? (\(K_{a1} = 4.3 \times 10^{-7}\), \(K_{a2} = 5.6 \times 10^{-11}\))

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HCO_3^- ⇌ H^+ + CO_3^{2-}
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*Ignore the + and – x!!*

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**Polyprotic Acids**

*Acids that can donate MORE than one H⁺*

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  If \([H_2CO_3] = 0.0037 \text{ M}\), what is the pH and \([CO_3^{2-}]\)? (\(K_{a1} = 4.3 \times 10^{-7}\), \(K_{a2} = 5.6 \times 10^{-11}\))

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**Common Ion Effect**

*REDUCTION in the solubility when a soluble compound containing one of the SAME ions is added*

- Common ion will affect the pH
- Weak acid or base will dissociate / ionize **LESS** than if it were alone

- **EXAMPLE:**
  Determine the pH of 0.30 moles of \(\text{HC}_2\text{H}_3\text{O}_2\) (\(K_a = 1.8 \times 10^{-5}\)) and 0.30 moles of \(\text{NaC}_2\text{H}_3\text{O}_2\) in 1.0 L of water.

\(\text{C}_2\text{H}_3\text{O}_2^-\) is the **COMMON ION**!
Common Ion Effect

**EXAMPLE:**

Determine the pH of 0.30 moles of HC\(_2\)H\(_3\)O\(_2\) (\(K_a = 1.8 \times 10^{-2}\)) and 0.30 moles of NaC\(_2\)H\(_3\)O\(_2\) in 1.0 L of water.

\[
\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^- 
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\(x = [\text{H}^+] = 1.8 \times 10^{-5}\) \(\text{pH} = -\log [\text{H}^+] = 4.74\)

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\(x = [\text{H}^+] = 1.8 \times 10^{-5}\) \(\text{pH} = -\log [\text{H}^+] = 4.74\)

\(K_a \times K_b = K_w = 1 \times 10^{-14}\)

\(pK_a = -\log K_a\)

\(pK_b = -\log K_b\)

\(pK_a + pK_b = pK_w = 14\)

- \(pK_a / pK_b\): how acidic or basic a hydrogen or hydroxide is in a molecule

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

**Strong Acids / Bases**

- Larger \(K_a\) or \(K_b\) = STRONGER

**Weak Acids / Bases**

- Smaller \(K_a\) or \(K_b\) = STRONGER

- Higher \(pK_a\) or \(pK_b\) = STRONGER

**EXAMPLES:**

Calculate the pH of 0.20 M lactic acid, HC\(_3\)H\(_5\)O\(_3\), with a \(pK_a\) of 3.08.

What is the pH of a 0.045 M solution of KOBr? The \(pK_a\) of HOBr is 8.63. (*Hint: write the weak base reaction of OBr\(^-\) with water first*)
Neutralization Reactions
When an acid and base mix, the acid will donate protons to the base (usually get a salt and water)

FOUR TYPES:
1) Strong Acid and Strong Base
   • Products are always a salt and water
   • Example:
     \[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]
   • Try it: \( \text{HClO}_4 + \text{NaOH} \)

2) Strong Acid and Weak Base
   • Salt forms but water may not be present if no "OH" in the base
   • Example:
     \[ \text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl} \]

3) Weak Acid and Strong Base
   • Salt is the conjugate base of the weak acid and water are products
   • Example:
     \[ \text{HC}_2\text{H}_3\text{O}_2 + \text{NaOH} \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \]
     \( \text{(C}_2\text{H}_3\text{O}_2^-) \)

4) Weak Acid and Weak Base
   • Products are the conjugate base and conjugate acid (or salts of them)
   • Example:
     \[ \text{HC}_2\text{H}_3\text{O}_2 + \text{NH}_3 \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{NH}_4^+ \]

Practice

• EXAMPLES:
  Write the products for the following acid-base reactions.
  - \( \text{H}_2\text{SO}_4 \text{ and NaOH} \)
  - \( \text{H}_3\text{PO}_4 \text{ and KOH} \)
  - \( \text{HF} \text{ and NH}_4\text{OH} \)
  - \( \text{HBr} \text{ and NH}_3 \)

Acid-Base Properties of Salts

• Salts produced in neutralization reactions are NOT always neutral... might contribute to pH!!
• Most salts FULLY dissociate into their ions
• Some can hydrolyze (react with water) to produce acidic or basic solutions
Acid-Base Properties of Salts

**GENERAL STATEMENTS:**

• Salts of strong acids and strong bases (ex: NaCl) do NOT hydrolyze and solution is **NEUTRAL**... pH = 7

• Examples:
  
  NaCl + H₂O → **NO RXN!!**
  
  KI + H₂O → **NO RXN!!**

You MUST look at which acid and base made the salt!!

---

**Acid-Base Properties of Salts**

**GENERAL STATEMENTS:**

• Salts of weak acids and strong bases (ex: NaC₂H₃O₂) do hydrolyze and solution is **BASIC**... pH > 7
  
  (anion acts as a base)

Remember... the salt FULLY dissociates into its ions which might be able to hydrolyze with water!

---

**Acid-Base Properties of Salts**

**GENERAL STATEMENTS:**

• Salts of strong acids and weak bases (ex: NH₄Cl) do hydrolyze and solution is **ACIDIC**... pH < 7
  
  (cation acts as an acid)

• Example:
  
  NH₄⁺ + H₂O ⇌ H₃O⁺ + NH₃

**Acid-Base Properties of Salts**

**GENERAL STATEMENTS:**

• Salts of weak acids and weak bases (Ex: NH₄C₂H₃O₂) do hydrolyze but pH depends on $K_a$ and $K_b$ values (both components hydrolyze)

  - If $K_a$ of acidic ion is greater than $K_b$ of basic ion... **ACIDIC pH**
  
  - If $K_b$ of basic ion is greater than $K_a$ of acidic ion... **BASIC pH**
  
  - If $K_a$ of acidic ion and $K_b$ of basic ion are equal... **NEUTRAL pH**

---

**Acid-Base Properties of Salts**

**GENERAL STATEMENTS:**

• Salts of weak acids and weak bases (Ex: NH₄C₂H₃O₂) do hydrolyze but pH depends on $K_a$ and $K_b$ values (both components hydrolyze)

  - Example:
  
  $K_a = 5.6 \times 10^{-10}$

  BOTH occur:

  - NH₄⁺ + H₂O ⇌ H₃O⁺ + NH₃
  
  - C₂H₃O₂⁻ + H₂O ⇌ OH⁻ + HC₂H₃O₂

  **NEUTRAL!**
Practice

- **EXAMPLES:**
  Predict whether an aqueous solution of the following salts will be acidic, basic, or neutral. Prove with appropriate equations.

  \[
  \text{NH}_4\text{NO}_3, \quad \text{K}_3\text{PO}_4, \quad \text{Na}_2\text{SO}_4, \quad \text{BaI}_2
  \]

Acid-Base Properties of Salts

- **EXAMPLE:**
  Calculate the pH of a 0.10 M \( \text{NH}_4\text{Cl} \) solution. The \( K_b \) value for \( \text{NH}_3 \) is \( 1.8 \times 10^{-5} \).

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

  Because an ACID is forming, we need the \( K_a \) value!!

  \[
  K_a \times K_b = K_w
  \]

  \[
  K_a \times 1.8 \times 10^{-5} = 1 \times 10^{-14}
  \]

  \[
  K_a = 5.6 \times 10^{-10}
  \]
Acid-Base Properties of Salts

**EXAMPLE:**

Calculate the pH of a 0.10 M NH₄Cl solution.

The Kₐ value for NH₃ is 1.8 x 10⁻⁵.

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

<table>
<thead>
<tr>
<th></th>
<th>NH₄⁺</th>
<th>H⁺</th>
<th>NH₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.10 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.10 - x M</td>
<td>x M</td>
<td>x M</td>
</tr>
</tbody>
</table>

\[ x = [\text{H}^+] = 7.5 \times 10^{-6} \]

\[ \text{pH} = -\log [\text{H}^+] = 5.12 \]

Acid-Base Properties of Salts

**EXAMPLE:**

Calculate the pH of a 0.30 M NaF solution.

The Kₐ value for HF is 7.2 x 10⁻⁴.

Titrations

Solution with a known [ ] is added to a solution with an unknown [ ] in order to determine the [ ]

- Neutralization reactions are generally performed this way (acid added to base or vice versa)
- Uses a burette filled with STANDARDIZED (know the exact [ ]) solution
- Exact amount of solution added is important... Proper techniques MUST be followed (you will learn these during lab)

Titrations

**Watch equations for ratio... is it 1:1, etc?**

- **EQUILIBRIUM POINT:** point at which equal stoichiometric amounts of acid and base have been added (moles acid = moles base)
- **HALF-EQUILIBRIUM POINT:** enough substance added to convert exactly half of the acid or base into its conjugate (pH = pKₐ)
- **INDICATORS** (change color at specific pH) are added to identify the equivalence point
Titrations
Solution with a known [ ] is added to a solution with an unknown [ ] in order to determine the [ ].

Be careful when adding the titrant to the ENDPOINT (Point in titration when the indicator undergoes a color change)... important not to go too far over!

Selecting Indicators
- Indicator color changes can be sharp and occur with the addition of a single extra drop of titrant
- Choose a suitable indicator based on the EQUIVALENCE POINT of the reaction
  - Indicators with a $pK_a = pH$ at the equivalence point work best

Titration Curves
Graphs of pH (y-axis) as a function of titrant added (x-axis)
- Can be generated from burette and pH readings (paper or meter)
- pH increases SLOWLY at first and then JUMPS just before the equivalence point

Titration Curves
Graphs of pH (y-axis) as a function of titrant added (x-axis)
- When acid and base are both STRONG... equivalence point pH = 7
- When using STRONG acid and WEAK base... equivalence point pH < 7
- When using WEAK acid and STRONG base... equivalence point pH > 7
  - pH can be found for titrations at different points...

Shape of the Curve
INITIALLY: pH does NOT change much because of the "buffering" action of the analyte (large amount present versus only small amount of titrant... Analyte destroys what is added)
Shape of the Curve

**AT EQUIVALENCE POINT**: pH dramatically changes and increases quickly

**EXCESS ADDED**: pH levels off and does not change much (as long as the titrant remains the same)

---

**Strong Acid / Strong Base**

- **Four Regions of the Curve**:
  1. Initial pH (initial [ ])
  2. Between initial pH and equivalence point
  3. Equivalence point
  4. After equivalence point

- **Endpoint**: point where indicator changes color (usually close to the equivalence point)

---

**Strong Acid / Strong Base**

- **Titration**: Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

Since the titration involves a STRONG acid and STRONG base, we know the equivalence point pH = 7!

---

**Strong Acid / Strong Base**

- **Four Regions of the Curve**:
  1. Initial pH (initial [ ])
  2. Between initial pH and equivalence point
  3. Equivalence point
  4. After equivalence point

- **Endpoint**: point where indicator changes color (usually close to the equivalence point)
**Strong Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **BEFORE** any NaOH is added...
Just use \[ \text{pH} = -\log [\text{H}^+] \]
(strong acid so completely dissociated)

\[ \text{Ans: 1.00} \]

---

**Strong Acid / Strong Base**

**Four Regions of the Curve:**
1) Initial pH (initial \[ \text{[ } \])
2) Between initial pH and equivalence point
3) Equivalence point
4) After equivalence point

**ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

---

**Strong Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

1) Find the moles of NaOH added
2) Find amount of moles of HCl started
3) Determine moles of HCl left
4) Use **TOTAL** volume to determine \([\text{H}^+]\)
   and then find the pH

\[ \text{If 0.0010 mol of NaOH are added, that means 0.0010 mol of HCl are now gone (neutralized) because the reaction is a 1:1 ratio!} \]

---

**Strong Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

\[ \begin{align*}
10.0 \text{ mL} & \times \frac{L}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{L} = 0.0010 \text{ mol NaOH added} \\
25.0 \text{ mL} & \times \frac{L}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{L} = 0.0025 \text{ mol HCl started}
\end{align*} \]

---

**Strong Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

\[ \begin{align*}
0.0025 \text{ mol} & - 0.0010 \text{ mol} = 0.0015 \text{ mol HCl left} \\
\frac{0.0015 \text{ mol}}{35.0 \text{ mL}} & \times \frac{1000 \text{ mL}}{L} = 0.043 \text{ M [H}^+] \\
\text{**[25.0 + 10.0 mL]**} & \text{ Don't forget!} \\
\text{pH} & = -\log [\text{H}^+] \\
\text{-log [0.043]} & = 1.37
\end{align*} \]
Strong Acid / Strong Base

• Four Regions of the Curve:
  1) Initial pH (initial [ ])
  2) Between initial pH and equivalence point
  3) Equivalence point
  4) After equivalence point

• ENDPOINT: point where indicator changes color (usually close to the equivalence point)

Strong Acid / Strong Base

• TITRATION:
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH at the EQUVALENCE POINT...

MOLES ACID = MOLES BASE AT THE EQUIVALENCE POINT!!!!!!!!!

Strong Acid / Strong Base

• TITRATION:
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH AFTER the addition of 35.0 mL NaOH...

1) Find the moles of NaOH added
2) Find amount of moles of HCl started
3) Determine moles of NaOH excess
4) Use TOTAL volume to determine [OH] and then find the pH from pOH

Strong Acid / Strong Base

• TITRATION:
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH AFTER the addition of 35.0 mL NaOH...

35.0 mL \times \frac{L}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{L} = 0.0035 \text{ mol NaOH added}

25.0 mL \times \frac{L}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{L} = 0.0025 \text{ mol HCl started (SAME AS BEFORE)}
**Strong Acid / Strong Base**

- **TITRATION:**
  Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH AFTER the addition of 35.0 mL NaOH...

If 0.0035 mol of NaOH are added, that means ALL of HCl are now gone (neutralized) because the reaction is a 1:1 ratio!

**Strong Acid / Strong Base**

- **TITRATION:**
  Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH AFTER the addition of 35.0 mL NaOH...

0.0035 mol - 0.0025 mol = 0.0010 mol NaOH left

\[
\frac{0.0010 \text{ mol}}{60.0 \text{ mL}} \times \frac{1000 \text{ mL}}{L} = 0.017 \text{ M [OH}^-\text{]}
\]

\[
pOH = -\log [\text{OH}^-]\]

\[
\text{pH} + \text{pOH} = 14
\]

\[
\text{pH} = 14 - 1.77 = 12.23
\]

**Summary**

For Strong Acid / Strong Base Titrations:

- pH initially is just $-\log$ [analyte] because it is 100% dissociated.
- pH = 7 at the equivalence point.
- At any other point, find moles of acid and moles of base and subtract the two... determine \[\text{[ ]}\] of whichever there is more of and use that to determine pH.

**Practice**

- **EXAMPLE:**
  Determine the pH when the following quantities of 0.10 M HNO\(_3\) have been added to 25.0 mL of 0.10 M KOH solution: a) 24.0 mL and b) 27.0 mL.
Practice

- EXAMPLE:
A solution of 0.10 M NaOH is added to a burette and used to titrate 50.0 mL of 0.20 M HNO₃. a) Find the initial pH of the acid solution b) after 65.0 mL of NaOH have been added and c) after 150.0 mL have been added.

Weak Acid / Strong Base

- TITRATION:
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M H₂C₂H₃O₂ (Kₐ = 1.8 x 10⁻⁵).

Get in the habit of writing the neutralization and hydrolysis reactions first!
H₂C₂H₃O₂ + NaOH → NaC₂H₃O₂ + H₂O
Salt is NOT neutral... from strong base and weak acid so needs to form OH!!
C₂H₃O₂⁻ + H₂O ⇌ HC₂H₃O₂ + OH

*Note the difference in the arrows!!

Weak Acid / Strong Base

- Four Regions of the Curve:
1) Initial pH (initial [ ])
2) Between initial pH and equivalence point
3) Equivalence point
4) After equivalence point

• ENDPOINT: point where indicator changes color (usually close to the equivalence point)

Weak Acid / Strong Base

- Four Regions of the Curve:
1) Initial pH (initial [ ])
2) Between initial pH and equivalence point
3) Equivalence point
4) After equivalence point

• ENDPOINT: point where indicator changes color (usually close to the equivalence point)
**Weak Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M H<sub>2</sub>C<sub>2</sub>H<sub>3</sub>O<sub>2</sub> (K<sub>a</sub> = 1.8 x 10<sup>-5</sup>).

To find the pH AFTER the addition of 10.0 mL of NaOH:
1) Find the moles of NaOH added
2) Find amount of moles of acid started
3) Determine moles of acid left
4) Determine moles of weak base formed
5) Determine [ ] of acid and weak base then plug into HH equation to get pH

**Henderson-Hasselbalch Equation**

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

OR

\[
\text{pOH} = \text{p}K_b + \log \left( \frac{[\text{Acid}]}{[\text{Base}]} \right)
\]

*Used for BUFFERS (weak acids / bases and their conjugate bases / acids)... More to come!

We will use it here because with only 10.0 mL of NaOH added, we have a weak acid and its conjugate base remaining!

**To find the pH AFTER the addition of 10.0 mL of NaOH...**

- **10.0 mL x 0.10 mol L x 1000 mL = 0.0010 mol NaOH added**
- **25.0 mL x 0.10 mol L x 1000 mL = 0.0025 mol acid started**

If 0.0010 mol of NaOH are added, that means 0.0010 mol of H<sub>2</sub>C<sub>2</sub>H<sub>3</sub>O<sub>2</sub> are neutralized because the reaction is a 1:1 ratio!

**BUFFER**
Weak Acid / Strong Base

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl (K_a = 1.8 x 10^{-5}).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

\[ \text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]} \]

\[ \text{pK}_a = -\log K_a \]

\[ \text{pK}_a = -\log (1.8 \times 10^{-5}) \]

\[ \text{pK}_a = 4.74 \]

---

**Weak Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl (K_a = 1.8 x 10^{-5}).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

\[ \text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]} \]

\[ \text{pH} = 4.74 + \log \frac{[0.028]}{[0.043]} = 4.55 \]

---

**Weak Acid / Strong Base**

**Four Regions of the Curve:**

1) Initial pH (initial [ ])
2) Between initial pH and equivalence point
3) Equivalence point
4) After equivalence point

**ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

---

**Weak Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl (K_a = 1.8 x 10^{-5}).

To find the pH at the **EQUIVALENCE POINT**...

When 25.0 mL of NaOH is added, all of the weak acid has been neutralized... BUT the conjugate base will hydrolyze to give a pH that is higher than 7

---

**Weak Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl (K_a = 1.8 x 10^{-5}).

To find the pH at the **EQUIVALENCE POINT**...

\[ 25.0 \text{ mL} \times 0.10 \text{ mol L}^{-1} \times \frac{1000 \text{ mL}}{1000 \text{ mL}} = 0.0025 \text{ mol NaOH added} \]

So 0.0025 mol of C_2H_3O_2^- formed!
**Weak Acid / Strong Base**

**TITRATION:**
Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M H\(_2\)C\(_2\)H\(_3\)O\(_2\) (\(K_a = 1.8 \times 10^{-5}\)).

To find the pH at the **EQUIVALENCE POINT**...

\[
\frac{0.0025 \text{ mol}}{50.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.050 \text{ M} \\
\]

\[\text{[C}_2\text{H}_3\text{O}_2^-]\]

\[
\begin{array}{c|c|c}
\text{C}_2\text{H}_3\text{O}_2^- & \text{HC}_2\text{H}_3\text{O}_2^- & \text{OH}^- \\
\hline
\text{Initial} & 0.050 \text{ M} & 0 \\
\text{Change} & -x & +x \\
\text{Equilibrium} & 0.050 \text{ M} - x & +x \\
\end{array}
\]

\[x = \text{[OH}^-] = 5.3 \times 10^{-6} \]

\[\text{pOH} = -\log [\text{OH}^-] = 5.28\]
Weak Acid / Strong Base

• **Four Regions of the Curve:**
  1) Initial pH (initial $[\ldots]$)
  2) Between initial pH and equivalence point
  3) Equivalence point
  4) After equivalence point

• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

---

**Weak Acid / Strong Base**

**TITRATION:**

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH *AFTER* the addition of 40.0 mL NaOH:

1) Find the moles of NaOH added
2) Find amount of moles of acid started
3) Determine the moles of excess NaOH
4) Get the [ ] of excess... pOH to get pH

---

**Weak Acid / Strong Base**

**TITRATION:**

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH *AFTER* the addition of 40.0 mL NaOH:

\[
\begin{align*}
40.0 \text{ mL} & \times \frac{L}{1000 \text{ mL}} \times 0.10 \text{ mol} \text{ L}^{-1} = 0.0040 \text{ mol NaOH added} \\
25.0 \text{ mL} & \times \frac{L}{1000 \text{ mL}} \times 0.10 \text{ mol} \text{ L}^{-1} = 0.0025 \text{ mol acid started} \\
\end{align*}
\]

\[
0.0040 \text{ mol} - 0.0025 \text{ mol} = 0.0015 \text{ mol NaOH excess}
\]

\[
\begin{align*}
0.0015 \text{ mol} & \times \frac{1000 \text{ mL}}{65.0 \text{ mL}} = 0.023 \text{ M [OH\text{^-}]}
\end{align*}
\]

\[
pOH = -\log [\text{OH\text{^-}}] = -\log [0.023] = 1.64
\]

---

**Weak Acid / Strong Base**

**TITRATION:**

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH *AFTER* the addition of 40.0 mL NaOH:

\[
\begin{align*}
40.0 \text{ mL} & \times \frac{L}{1000 \text{ mL}} \times 0.10 \text{ mol} \text{ L}^{-1} = 0.0040 \text{ mol NaOH added} \\
25.0 \text{ mL} & \times \frac{L}{1000 \text{ mL}} \times 0.10 \text{ mol} \text{ L}^{-1} = 0.0025 \text{ mol acid started} \\
\end{align*}
\]

\[
0.0040 \text{ mol} - 0.0025 \text{ mol} = 0.0015 \text{ mol NaOH excess}
\]

\[
\begin{align*}
0.0015 \text{ mol} & \times \frac{1000 \text{ mL}}{65.0 \text{ mL}} = 0.023 \text{ M [OH\text{^-}]}
\end{align*}
\]

\[
pH = 14 - 1.64 = 12.36
\]

---

**Weak Base / Strong Acid**

If a **STRONG** acid is added to a **WEAK** base instead, just do the **OPPOSITE**!

Curve will look different too!!
Polyprotic Acid Titrations

• Neutralization reaction occurs in steps due to MULTIPLE hydrogens
  • Curve has multiple equivalence points... one for each H+

Don’t forget... pH = pKₐ at the midpoint of each!
Also, notice as Kₐ gets smaller (acid gets weaker)... pH change at equivalence point is smaller!

Summary
For Weak / Strong Titrations:
  • pH initially is found by doing equilibrium problem for weak (ICE)
  • If some weak still remains... determine [ ] of weak and [ ] of the conjugate and plug into HH
  • At equivalence point, find [ ] of conjugate salt formed and use ICE table with hydrolysis rxn
  • If excess strong remains... use moles and [ ] of excess to find pH

Practice
  • EXAMPLE:
A solution is formed by adding 0.050 M NaOH to 40.0 mL of 0.025 M benzoic acid, H₇C₅H₅O₂⁻ (Kₐ = 6.3 x 10⁻⁵). Calculate the pH a) of the initial solution *b) after 10.0 mL of NaOH added c) after 20.0 mL added and d) after 35.0 mL added

Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

  • As acid is added to the solution, the base component neutralizes it and vice versa
  • The weak acid or base and its conjugate CANNOT consume each other
  • pH remains stable even if solution is diluted with water or water is lost through evaporation (*WATCH THIS!)

Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

  • TWO WAYS TO MAKE:
  1) Mix weak acid or base with salt of its conjugate
     Ex: Add H₇C₅H₅O₂⁻ to NaC₅H₅O₂
     Add NH₃ to NH₄Cl
Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• TWO WAYS TO MAKE:
  2) Add SOME strong acid or base to a weak base or acid to make the conjugate salt OR vice versa... SOME strong added to a salt to make the conjugate acid or base (just enough strong added to neutralize about half the weak)

  Ex: Add some NaOH to HCOOH (C2H3O2): forms and is with the remaining weak
  Add some HCl to NaF (HF forms and is with the remaining F–)

Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• Buffers are an example of the common ion effect

• Most effective or IDEAL when EQUAL amounts of weak acid or base and conjugate are used

• Resists change in pH until the conjugate acid or base is used up

Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• These are not new problems... can use common ion effect equilibrium or HH to solve

• Any weak acid or base titration is a buffer problem from the time the titrant is added until the equivalence point

Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

Which of the following could be added to a solution of sodium acetate to produce a buffer? Choose any that apply.

Acetic acid
Hydrochloric acid
Potassium acetate
Sodium chloride

Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• EXAMPLE:

• Added H+ gets neutralized by the conjugate base!

• Added OH– gets neutralized by the weak acid!

Buffers
Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• pH of a Buffer

  Adding HCl to NaF

  Adding NaOH to Acetic Acid

  equivalence point

  equivalence point
**Henderson-Hasselbalch Equation**

\[
\text{pH} = pK_a + \log \frac{[\text{Base}]}{[\text{Acid}]}
\]

**OR**

\[
\text{pOH} = pK_b + \log \frac{[\text{Acid}]}{[\text{Base}]}
\]

*NOTE: Technically, HH becomes less accurate as the acid or base becomes STRONGER or the [ ] becomes SMALLER... This could affect some lab data in theory, so just be aware!*

**Buffer Capacity**

*Amount of acid or base the buffer can neutralize before the pH begins to change*

- Depends on the amount of acid and base in the buffer
- **INCREASE** the amount of acid and base... the MORE RESISTANT the buffer is to changes in pH
- pH range: buffers act best at the pH = \( pK_a \) (or \( pK_b \)) of the weak acid / base because that is where the [ ] of the two components are equal... **IDEAL**

**Preparing a Buffer**

*EXAMPLE:*

How many moles of NH\(_4\)Cl must be added to 2.0 L of 0.10 M NH\(_3\) (\( K_b = 1.8 \times 10^{-5} \)) to form a buffer whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

\[
pOH = 14 - \text{pH} \\
14 - 9.00 = 5.00 \\
-\log K_b = pK_b \\
-\log 1.8 \times 10^{-5} = 4.74
\]

\[
\text{Use HH with pOH and } pK_b!!
\]

**Henderson-Hasselbalch Equation**

\[
\text{pH} = pK_a + \log \frac{[\text{Base}]}{[\text{Acid}]}
\]

**OR**

\[
\text{pOH} = pK_b + \log \frac{[\text{Acid}]}{[\text{Base}]}
\]

*EXAMPLE:*

What is the pH of the buffer of a 0.12 M HC\(_2\)H\(_5\)O\(_3\) and a 0.10 M NaC\(_2\)H\(_5\)O\(_3\) where \( K_a = 1.4 \times 10^{-4} \)?

**Buffer Capacity**

*Amount of acid or base the buffer can neutralize before the pH begins to change*

- Buffers with a **1:1 ratio** of [conjugate]:[weak] have the greatest buffer capacity
- Maximum amount of acid that can be added to buffer = amount of conjugate base in the buffer and vice versa
- Buffer is still a buffer IF...
  - pH within \( pK_a \pm 1 \)
  - [base]:[weak] = .10 - 10

**Preparing a Buffer**

*EXAMPLE:*

How many moles of NH\(_4\)Cl must be added to 2.0 L of 0.10 M NH\(_3\) (\( K_b = 1.8 \times 10^{-5} \)) to form a buffer whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

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-\log K_b = pK_b \\
-\log 1.8 \times 10^{-5} = 4.74
\]

\[
\text{Use HH with pOH and } pK_b!!
\]
Preparing a Buffer

**EXAMPLE:**

How many moles of NH₄Cl must be added to 2.0 L of 0.10 M NH₃ (Kb = 1.8 x 10⁻⁵) to form a buffer whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

\[
5.00 = 4.74 + \log \frac{[\text{Acid}]}{[0.10]}
\]

\[
0.26 = \log \frac{[\text{Acid}]}{[0.10]}
\]

\[
10^{-0.26} = \frac{[\text{Acid}]}{[0.10]}
\]

\[
1.82 = \frac{[\text{Acid}]}{[0.10]} \quad [\text{Acid}] = 0.18 \text{ M}
\]

pH Changes in Buffers

**EXAMPLE:**

A buffer is made by adding 0.30 mol HC₂H₃O₂ (Kₐ = 1.8 x 10⁻⁵) and 0.30 mol NaC₂H₃O₂ to enough water to make 1.0 L of solution. The pH of the buffer is 4.74. Calculate the pH of this solution after 0.020 mol of NaOH is added.

0.30 mol acid - 0.020 mol base = 0.28 mol acid
*Added base "attacks" acid part so subtracted

0.30 mol conj. + 0.020 mol base = 0.32 mol conj.
*Acid reacting forms more of the base part so added

\[
\frac{0.28 \text{ mol}}{1.0 \text{ L}} = 0.28 \text{ M} \quad [\text{HC}_2\text{H}_3\text{O}_2^-]
\]

\[
\frac{0.32 \text{ mol}}{1.0 \text{ L}} = 0.32 \text{ M} \quad [\text{C}_2\text{H}_3\text{O}_2^-]
\]
**pH Changes in Buffers**

- **EXAMPLE:**
  A buffer is made by adding 0.30 mol HCH\(_2\)O\(_2\) (\(K_a = 1.8 \times 10^{-5}\)) and 0.30 mol NaC\(_2\)H\(_3\)O\(_2\) to enough water to make 1.0 L of solution. The pH of the buffer is 4.74. Calculate the pH of this solution after 0.020 mol of NaOH is added.

\[
pH = pK_a + \log \frac{[\text{Base}]}{[\text{Acid}]}
\]

\[
pH = 4.74 + \log \frac{0.32}{0.28} = 4.80
\]

**Real World Buffers**

- **Blood pH Levels**

  - Buffers are needed in the body to maintain pH levels in order for certain enzymes to work
  - Human blood **MUST** maintain a pH between 7.35 - 7.45 in order to function normally
  - pH falls below... **ACIDOSIS**
  - pH rises above... **ALKALOSIS**

**Practice**

- **EXAMPLE:**
  A chemist needs a solution buffered at pH 4.30 and can choose from the following acids and their salts. Which should be chosen?

  - Chloroacetic acid \(K_a = 1.35 \times 10^{-3}\)
  - Propanoic acid \(K_a = 1.30 \times 10^{-5}\)
  - Benzoic acid \(K_a = 6.40 \times 10^{-5}\)
  - Hypochlorous acid \(K_a = 3.5 \times 10^{-8}\)

**Practice**

- **EXAMPLES:**
  - What is the pH of a buffer solution that is composed of 0.45 M NH\(_4\)Cl and 0.15 M NH\(_3\)? \((K_b = 1.8 \times 10^{-5})\)

  \[
  \text{Calculate the concentration of sodium benzoate that must be present in 0.20 M HCH\(_2\)O\(_2\) to produce a pH of 4.00. (}K_a = 6.4 \times 10^{-5}\)\]