Conjugate Acid/Conjugate Base Pair

\[
\text{CH}_3\text{COOH} + \text{CH}_3\text{NH}_2 \rightleftharpoons \text{CH}_3\text{COO}^- + \text{CH}_3\text{NH}_3^+ 
\]

Aceitic acid is an Acid since it donates a proton; the acetate ion is a conjugate base because it can accept a proton to become aceitic acid

\[
\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+ \quad \text{Ka} = 1.75 \times 10^{-5} \\
\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^- \quad \text{Kb} = \frac{\text{Kw}}{\text{Ka}} = 5.71 \times 10^{-10}
\]

Methylamine is a Base because it accepts a proton; methlyammonium ion is a conjugate acid because it can donate a proton to become methylamine

\[
\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^- \quad \text{Kb} = 4.4 \times 10^{-4} \\
\text{CH}_3\text{NH}_3^+ + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_2 + \text{H}_3\text{O}^+ \quad \text{Ka} = \frac{\text{Kw}}{\text{Kb}} = 2.3 \times 10^{-11}
\]

Therefore, \( \text{Kw} = \text{Ka} \times \text{Kb} \)

Equilibrium and Acid/Base Chemistry

1. **Strong Acids and Strong Bases** completely dissociate i.e. large K
   a. \([\text{H}^+] = [\text{HA}]_{\text{before dissociation}} \) since HA completely dissociates
   b. \( \text{pH} = -\log[\text{H}^+] \)

2. We can write \( \text{K} \) for weak acids and bases
   a. \([\text{H}^+] \) not equal to \([\text{HA}] \)
   b. need to find \([\text{H}^+] \) with \( \text{K} \)
   c. \( \text{pH} = -\log[\text{H}^+] \)

REMEMBER:

If you have \([\text{H}^+] \) and need \([\text{OH}^-] \) then
\( [\text{OH}^-] = 1.0 \times 10^{-14}/[\text{H}^+] \)

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If you have pH and need \([\text{H}^+] \) then
\( [\text{H}^+] = 10^{-\text{pH}} \)

If you have pOH and need \([\text{OH}^-] \) then
\( [\text{OH}^-] = 10^{\text{pOH}} \)

\( \text{pH} + \text{pOH} = 14 \)

\( \text{Kw} = \text{Ka} \times \text{Kb} \)
Ka = Kw/Kb

Kb = Kw/Ka

Kw = 1.0 x 10^{-14}

Ka = weak acid; acetic acid; carboxylic acids
1. Weak acid \rightleftharpoons conjugate base + H^+
2. HA + H_2O \rightleftharpoons A^- + H_3O^+
3. \frac{K_a}{[HA]} = [A^-][H_3O^+] ; [H^+] \ll [HA] \text{ NOTE that for a strong } [H^+] = [HA]
4. Brønsted-Lowry Acid = proton donor

Kb = weak base; methylamine; amines
1. Weak base + water \rightleftharpoons conjugate acid + OH^-
2. B + H_2O \rightleftharpoons BH^+ + OH^-
3. \frac{K_b}{[B]} = [BH^+][OH^-] ; [OH^-] \ll [B] \text{ NOTE that for a strong base } [OH^-] = [B]
4. Brønsted-Lowry Base = proton acceptor

Calculate the pH of a 1.0 M solution of acetic acid.

CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+ \text{ Ka} = 1.75 \times 10^{-5}

\text{Start} \quad 1 \text{ M} \quad 0 \quad 0
\text{Change} \quad -x \quad +X \quad +X
\text{End} \quad 1-x \quad X \quad X

1.75 \times 10^{-5} = \frac{x^2}{1-x} ; \quad x \text{ is small so } \frac{x^2}{1} = \frac{1.75 \times 10^{-5}}{1-x}

X = 4.2 \times 10^{-3}

pH = -\log (4.2 \times 10^{-3}) = 2.37

what is the percent dissociation?
% dissociation = \frac{[H_3O^+]}{[CH_3COOH]_{before \, dissociation}} \times 100\% = 0.42\%

SAME IDEA FOR A WEAK BASE!!!!