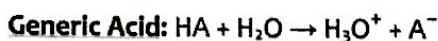


Totally Epic AP Chem Review: Acids and Bases in a Day!

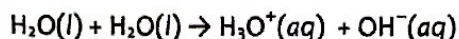
Various Ways to Describe Acid Strength		
Property	Strong Acid	Weak Acid
Ionization constant (K_a) value	K_a is large	K_a is small
Percent Ionization	% Ionization = 100%	% Ionization \ll 100%
Position of the dissociation (ionization) equilibrium	Far to the <u>right</u> (product-favored)	Far to the <u>left</u> (reactant-favored)
Equilibrium concentration of H^+ when compared to original $[HA]$	$[H^+] \approx [HA]_0$	<u>RICE!</u> $[H^+] \ll [HA]_0$



$$K_a = \frac{[x][x]}{[HA]_i - x} \approx \frac{[x][x]}{[HA]_i} \text{ where } [H_3O^+] = x \ll [HA]_i$$

$$K_b = \frac{[x][x]}{[A^-]_i - x} \approx \frac{[x][x]}{[A^-]_i} \text{ where } [OH^-] = x \ll [A^-]_i$$

Self-ionization of Water: About 2 out of 1 billion water molecules self-ionize!



$$K_w = [H_3O^+][OH^-] = K_a \times K_b = 1.0 \times 10^{-14} \quad (\text{at } 25^\circ\text{C})$$

1) Strong Acids/ Strong Bases	2) Weak Acids/ Weak Bases
<p><u>You MUST memorize:</u></p> <p>Strong Acids: HBr, HI, HCl, H_2SO_4, HNO_3, $HClO_4$</p> <p><i>Hint:</i> BriCl-SO-NO-ClO ("Brickle-So-No-Clo")</p> <p>Strong Bases: Groups IA and IIA metal hydroxides</p> <p>100% Dissociation! Easy life:</p> $pH = -\log[H^+] = -\log[HA]_0$ $pOH = -\log[OH^-] = -\log[B]_0$ $pH + pOH = 14$	<p><u>If it's not strong, it's weak!</u></p> <p>$< 1\%$ Dissociation \rightarrow Equilibrium!</p> <p>Time saver:</p> <ul style="list-style-type: none"> Since acids ionize 1 H^+ at a time, $[H_3O^+] = [A^-]$, and $[OH^-] = [BH^+]$. For weak acids and bases, make the assumption $[HA]_0 - x \approx [HA]_0$ and $[B]_0 - x \approx [B]_0$. <p><u>Weak Acids:</u></p> $K_a = \frac{[x][x]}{[HA]_i - x} \approx \frac{[x][x]}{[HA]_i} \text{ where } [H_3O^+] = x \ll [HA]_i$ <p><u>Weak Bases:</u></p> $K_b = \frac{[x][x]}{[B]_i - x} \approx \frac{[x][x]}{[B]_i} \text{ where } [OH^-] = x \ll [B]_i$

Percent Ionization

Percent Ionization: percentage of acid molecules that dissociate (ionize) when dissolved in water

$$\% \text{ Ionization} = \frac{\text{molarity of ionized acid}}{\text{initial molarity of acid}} \times 100 = \frac{[H_3O^+]_{\text{equil}}}{[HA]_0} \times 100$$

Important Notes

- The favored direction of the reaction is the one in which the weaker acid/base are produced.
- The stronger an acid is, the weaker its conjugate base (and vice versa).
- Diluting an acid (\downarrow concentration) will \uparrow pH and \uparrow percent ionization.

Practice, practice, practice!

HCl = strong acid!

1. Calculate the pH of a 0.020 M solution of hydrochloric acid.

$$\text{pH} = -\log [\text{H}^+] = -\log [\text{HCl}] = -\log (0.020) = \boxed{1.70}$$

2 s.f. 2 s.f. after decimal in pH!

2. Calculate the pH and percent ionization of a 0.020 M solution of acetic acid (
- $K_a = 1.8 \times 10^{-5}$
-).

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{x^2}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{x^2}{0.020 - x} \approx \frac{x^2}{0.020} = 1.8 \times 10^{-5}$$

Assume $x \ll 0.020$

$$x = \sqrt{(1.8 \times 10^{-5})(0.020)} = 6.0 \times 10^{-4} \text{ M} = [\text{H}_3\text{O}^+] \quad \left. \begin{array}{l} \text{pH} = -\log(6.0 \times 10^{-4}) \\ = \boxed{3.22} \end{array} \right\}$$

$$\% \text{ Ion.} = \frac{[\text{H}_3\text{O}^+]}{[\text{HC}_2\text{H}_3\text{O}_2]_i} \times 100 = \frac{6.0 \times 10^{-4}}{0.020} \times 100 = \boxed{3.0\%}$$

3. Calculate the pH for a
- 1.7×10^{-2}
- M solution of KOH.

$$\text{pOH} = -\log [\text{OH}^-] \quad \text{strong base!}$$

$$= -\log [1.7 \times 10^{-2}] = 1.77$$

$$\text{pH} = 14 - \text{pOH} = 14 - 1.77 = \boxed{12.23}$$

4. Calculate the pH and percent ionization for a
- 1.7×10^{-2}
- M solution of
- NH_3
- (
- $K_b = 1.8 \times 10^{-5}$
-).

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{x^2}{[\text{NH}_3]} = \frac{x^2}{(1.7 \times 10^{-2}) - x} \approx \frac{x^2}{1.7 \times 10^{-2}} = 1.8 \times 10^{-5}$$

assume $x \ll 1.7 \times 10^{-2}$

$$x = \sqrt{(1.8 \times 10^{-5})(1.7 \times 10^{-2})} = 5.5 \times 10^{-4} \text{ M} = [\text{OH}^-] \quad \left. \begin{array}{l} \text{pOH} = -\log(5.5 \times 10^{-4}) = 3.26 \\ \text{pH} = 14 - \text{pOH} \\ = 14 - 3.26 = \boxed{10.74} \end{array} \right\}$$

$$\% \text{ Ion.} = \frac{[\text{OH}^-]}{[\text{NH}_3]_i} \times 100 = \frac{5.5 \times 10^{-4}}{1.7 \times 10^{-2}} \times 100 = \boxed{3.3\%}$$