4. Calculate the hydroxide ion concentration of 0.022 M hydrochloric acid solution.

$HCl + H_2O \xrightarrow{100\%} H_3O^+ + Cl^-$

.022 → .022 .022

$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1 \times 10^{-14}}{.022} = 4.5 \times 10^{-13}\ M\ OH^-$

5. The pH of canned soda was measured with a pH meter and determined to be 3.5. What is the $[H_3O^+]$?

$[H_3O^+] = 10^{-pH}$

$= 10^{-3.5} = 3 \times 10^{-4}\ M\ H_3O^+$

6. What is the pH, $[H_3O^+]$, and $[OH^-]$ of a solution made by dissolving 750.0g of HI gas in a 4.5 liter container of water?

$\frac{750.0\text{g HI}}{4.5\text{L}} \times \frac{1\text{ mole}}{127.91\text{g}} = 1.3\ mol/L$

$HI \xrightarrow{100\%} H^+ + I^-$

1.3 → 1.3M 1.3M

$pH = -\log 1.3$

$= -.11$

$[H_3O^+] = 1.3M$

$[OH^-] = \frac{1 \times 10^{-14}}{1.3} = 7.7 \times 10^{-15}\ M$

$pOH = 14 - pH$

$= 14.11$

$[OH^-] = 10^{-14.11}$

INTRODUCTORY Ka and Kb

Since weak acids and weak bases do not dissociate to any great extent in water, an equilibrium is reached. You can write an equilibrium expression for the dissociation. The equilibrium constant for an acid is called K_a and the equilibrium constant for a base is called K_b.

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

(H+ moves from CH_3COOH to H_2O; H+ moves from H_2O to CH_3COO^-)

weak acid

Strong conjugate base

$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$

Weak acids have strong conjugate bases. A strong conjugate base means that the particle is a * better base than water. A strong conjugate base is a better proton acceptor than water.

Strong acids have weak conjugate bases. → spectators

If a strong conjugate acid is formed in water, it will have a reaction with water. This reaction is called a hydrolysis reaction.

$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

(H+ moves from H_2O to CH_3COO^-)

SCB

$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$

$K_w = K_a \times K_b = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} \times \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]} = [H_3O^+][OH^-]$

K_a K_b K_w

1×10^{-14}

Introductory K_a and K_b Practice Problems

1. The acid ionization constant (K_a) for propanoic acid (C_2H_5COOH) is 1.3×10^{-5}.
   a.) Write the balanced dissociation equation for propanoic acid in water.
   b.) Write the equilibrium expression.
   c.) Calculate the hydrogen ion concentration in a 0.200 M solution of propanoic acid. (Hydronium)
   d.) Calculate the pH, pOH, and $[OH^-]$ in the solution.
   e.) Calculate the K_a for the reverse reaction.
   f.) Calculate the percent ionization of propanoic acid.

(a) $C_2H_5COOH \rightleftharpoons C_2H_5COO^- + H^+$

(b) $K_a = \dfrac{[C_2H_5COO^-][H^+]}{[C_2H_5COOH]}$

(c) $1.3\times10^{-5} = \dfrac{(x)(x)}{.200-x}$

$x = .0016\ M\ H^+$

(d) $pH = -\log .0016 = 2.80$

$pOH = 14 - 2.80 = 11.20$

$[OH^-] = \dfrac{1\times10^{-14}}{.0016} = 6.3\times10^{-12}\ M\ OH^-$

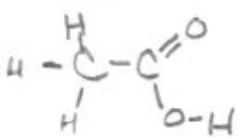
(e) $K_{rev} = \dfrac{1}{K_a} = \dfrac{1}{1.3\times10^{-5}} = 7.7\times10^{4}$

(f) $\dfrac{.0016}{.200}\times100 = .80\%$

2. The K_a for acetic acid is 1.8×10^{-5}.
   a.) Write two different representations for the molecular formula of acetic acid.
   b.) Write the dissociation equation for acetic acid.
   c.) Write the equilibrium expression.
   d.) Calculate the pH of a 0.750 M solution of acetic acid.
   e.) Calculate the $[OH^-]$.
   f.) If 0.200 M HCl is added to the solution, which direction will the equilibrium shift?

$HC_2H_3O_2$

CH_3COOH



(b) $HC_2H_3O_2 \rightleftharpoons C_2H_3O_2^- + H^+$

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

(c) $K_a = \dfrac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$

(d)

$HC_2H_3O_2 \rightleftharpoons$	$C_2H_3O_2^-$	$+ H^+$
.750	0	0
$-x$	$+x$	$+x$
$.750-x$	x	x

$1.8\times10^{-5} = \dfrac{(x)(x)}{.750-x}$

$x = .0037\ M\ H^+$

$\rightarrow pH = -\log .0037 = 2.43$

(e) $[OH^-] = \dfrac{1\times10^{-14}}{.0036} = 2.7\times10^{-12}\ M$

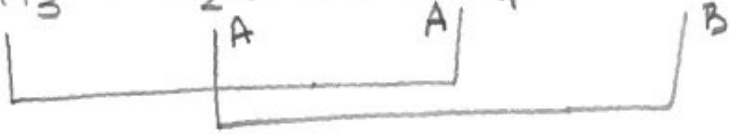
(f) $HCl \xrightarrow{100\%} H^+ + Cl^-$

$HC_2H_3O_2 \rightleftharpoons C_2H_3O_2^- + H^+$

$\leftarrow$ shift Left

3. The K_b value for ammonia is 1.8×10^{-5}.
   a.) Write the equation for the reaction of ammonia with water.
   b.) Identify the conjugate pairs.
   c.) Explain how the ammonium ion is a Bronsted-Lowry acid.
   d.) Calculate the $[OH^-]$ if 1.5 moles of ammonia are dissolved in 750.0 ml of solution.
   e.) Calculate the pH of the solution.
   f.) Calculate the K_b for the reverse reaction. → not the hydrolysis reaction

(a) $\underset{\text{WB}}{NH_3} + H_2O \rightleftharpoons \underset{\text{SCA}}{NH_4^+} + OH^-$

(b) 

(c) NH_4^+ is a proton donor in the reverse rxn.

(d) $\frac{1.5 \text{ moles } NH_3}{.7500 \text{ L}} = 2.0 M$

$NH_3 + H_2O \rightleftharpoons$	NH_4^+	OH^-
2.0	0	0
$-x$	$+x$	$+x$
$2.0-x$	x	x

$$1.8 \times 10^{-5} = \frac{(x)(x)}{2.0-x}$$

$$x = .0060 M\ OH^-$$

(e) $pOH = -\log .0060 = 2.22$

$pH = 14 - 2.22 = 11.78$

(f) $K_{rev} = \frac{1}{K_b} = 5.5 \times 10^4$