#22, 23 p. 610

What is the pH of a 0.150 mol/L hypochlorous acid solution at equilibrium?

$$HA + H_2O_{(l)} \rightleftharpoons A^-_{(aq)} + H_3O^+_{(aq)}$$

$$HClO_{(aq)} + H_2O_{(l)} \rightleftharpoons ClO^-_{(aq)} + H_3O^+_{(aq)}$$

0.150 mol/L

$$K_a = \frac{[ClO^-_{(aq)}][H_3O^+_{(aq)}]}{[HClO_{(aq)}]}, \quad [ClO^-_{(aq)}] = [H_3O^+_{(aq)}]$$

$$K_a = \frac{[H_3O^+_{(aq)}]^2}{[HClO_{(aq)}]}$$

$$2.9 \times 10^{-8} = \frac{[H_3O^+_{(aq)}]^2}{[0.150]}$$

$$[H_3O^+_{(aq)}] = \sqrt{(2.9 \times 10^{-8})(0.150)}$$

$$\underline{[H_3O^+_{(aq)}] = 6.59 \times 10^{-5}\,M}$$

$$pH = -\log[H_3O^+_{(aq)}]$$

$$pH = -\log[6.59 \times 10^{-5}]$$

$$\boxed{pH = 4.181}$$

Weak Bases

Weak bases react with water to form the hydroxide ion and conjugate acid of the base.

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^{+}_{4(aq)} + OH^{-}_{(aq)}$$

$$B^{-}_{(aq)} + H_2O_{(l)} \rightleftharpoons HB_{(aq)} + OH^{-}_{(aq)}$$

$$K = \frac{[NH^{+}_{4(aq)}][OH^{-}_{(aq)}]}{[NH_{3(aq)}][H_2O_{(l)}]}$$

$$K_b = \frac{[NH^{+}_{4(aq)}][OH^{-}_{(aq)}]}{[NH_{3(aq)}]}$$

base dissociation constant

Calculate the pH of a 0.221 mol/L solution of $NH_{3(aq)}$ at equilibrium.

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$$

$$K_b = \frac{[NH^+_{4(aq)}][OH^-_{(aq)}]}{[NH_{3(aq)}]}, \quad [NH^+_{4(aq)}] = [OH^-_{(aq)}]$$

$$K_b = \frac{[OH^-_{(aq)}]^2}{[NH_{3(aq)}]}$$

$$K_aK_b = K_w$$

$$K_b = \frac{K_w}{K_a}$$

$$\frac{K_w}{K_a} = \frac{[OH^-_{(aq)}]^2}{[NH_{3(aq)}]}$$

$$\frac{1.0\times10^{-14}}{5.8\times10^{-10}} = \frac{[OH^-_{(aq)}]^2}{0.221}$$

$$\sqrt{(1.72\times10^{-5})(0.221)} = [OH^-_{(aq)}]$$

$$[OH^-_{(aq)}] = 0.00195\,M$$

$$pOH = -\log[OH^-_{(aq)}]$$

$$pOH = -\log[0.00195]$$

$$pOH = 2.710$$

$$pH + pOH = 14.000$$

$$pH = 14.000 - 2.710$$

$$\boxed{pH = 11.290}$$

Worksheet