Worksheet #4

KOH

pH = 11.5

V = 500.mL

m = ?

$pH + pOH = 14.0$

$pOH = 14.0 - 11.5$

$pOH = 2.5$

$KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$

$3.16\times10^{-3}M$ $\qquad$ $3.16\times10^{-3}M$

$[OH^-_{(aq)}] = 10^{-pOH}$

$[OH^-_{(aq)}] = 10^{-2.5}$

$[OH^-_{(aq)}] = 3.16\times10^{-3}M$

KOH

$C = 3.16\times10^{-3}M$

V = 500.mL

m = ?

$C = \frac{n}{V}$

$n = (3.16\times10^{-3}\,mol/L)(0.500L)$

$n = 1.58\times10^{-3}\,mol$

$1.58\times10^{-3}\,mol\ KOH \times \frac{56.11\,g\ KOH}{1\ mol\ KOH} = \boxed{0.09\,g\ KOH}$

Ionization Constants for Acids

Strong acids - ionizes quantitatively in water to form hydronium ions

$$HCl_{(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Weak acids - ionizes partially in water to form hydronium ions

$$CH_3COOH_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

To describe the equilibrium of acids in water, the equilibrium law is used to calculate the acid ionization constant, K_a.

$$CH_3COOH_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

Ex. Predict the hydronium ion concentration, and pH of a 1.0 mol/L nitrous acid solution at equilibrium.

$$HNO_{2(aq)} + H_2O_{(\ell)} \rightleftharpoons NO^-_{2(aq)} + H_3O^+_{(aq)}$$

1.0 mol/L

$$K_a = \frac{[NO^-_{2(aq)}][H_3O^+_{(aq)}]}{[HNO_{2(aq)}]}, \quad [NO^-_{2(aq)}] = [H_3O^+_{(aq)}]$$

$$K_a = \frac{[H_3O^+_{(aq)}]^2}{[HNO_{2(aq)}]} \Rightarrow [H_3O^+_{(aq)}] = \sqrt{K_a[HNO_2]}$$

$$[H_3O^+_{(aq)}] = \sqrt{(7.2\times10^{-4})[1.0]}$$

$$[H_3O^+_{(aq)}] = 0.027\ M$$

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

$$pH = -\log[0.027]$$

$$\boxed{pH = 1.57}$$

Weak Bases

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

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

0.221M

base dissociation constant

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

$$K_aK_b = K_w$$

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

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

$$K_b = \frac{1.0 \times 10^{-14}}{5.8 \times 10^{-10}}$$

$$[OH^{-}_{(aq)}] = \sqrt{(1.72 \times 10^{-5})[0.221]}$$

$$\boxed{[OH^{-}_{(aq)}] = 1.95 \times 10^{-3} M}$$

$$K_b = 1.72 \times 10^{-5}$$

WEAK ACIDS

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

$$K_a K_b = K_w$$

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

Weak Acids Worksheet