SCH4U Name: ______________________

8.2 Calculations- The Relationship between Kw, Ka and Kb

Derive auto ionization of water From Bronsted- Lowry Theory :

$2\ H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$ amphiprotic can act like an acid or base.

Kw= always must equal= $[H^+_{(aq)}][OH^-_{(aq)}] = 1.0 \times 10^{-14}$

Ka and Kb relationship $K_w = K_a K_b$ (acid/base)

A neutral solution, relationship between $[H^+_{(aq)}] = [OH^-_{(aq)}]$

An acidic solution $[H^+_{(aq)}] > [OH^-]$

A basic solution $[OH^-_{(aq)}] > [H^+]$

Do and apply p. 502, # 1- 2 Standards = $1.0 \times 10^{-14} = [H^+][OH^-]$.

pH and pOH

Define: pH (p. 502) ; pOH (p. 502)

Given pH= 6.8 ; find pOH

pH + pOH = 14 ∴ pOH + 6 = 14
pOH = 14 − 6 = 8

This supports pH + pOH = -log K_w

$[OH^-][H^+] = K_w$

Use pH to determine the formula for H^+ concentration:

Convert pH = -log $[H^+]$ pOH= - log $[OH^-]$

$10^{-pH} = [H^+]$ $10^{-pOH} = [OH^-]$

Do and apply: p. 508, # 1-4

Kw ion-product constant for water - the equilibrium constant for the autoionization of water. (p. 100).

Kb = base ionization constant - the equilibrium constant for the ionization of a base; also called the base dissociation (p. 498).

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons [NH_4^{-}]^{+}[OH]^{-}$

$K_b = 1.8 \times 10^{-5} = [NH_4^{-}]^{+}$