Name:______________________________________ Period:_________

K_a and K_b Calculations Worksheet

When a strong acid or base is placed in water, they completely ionize. This means that approximately 100% of the acid or base forms products (or the arrow in the chemical equation points one direction). In the case of a weak acid or base, the substance only partially ionizes. This means equilibrium is established in an aqueous solution of a weak acid or base. Using your understanding of acid/base chemistry, complete the following problems.

1. Write chemical equations which represent the dissociation of each of these acids or bases in aqueous solution. Use a single arrow in the case of a strong acid or base, and a double arrow to represent the equilibrium condition that exists in the solution of a weak acid or base.

a. HCl	$HCl \rightarrow H^+ + Cl^-$
b. NaOH	$NaOH \rightarrow Na^+ + OH^-$
c. H_2SO_4	$H_2SO_4 \rightarrow H^+ + HSO_4^-$
d. KOH	$KOH \rightarrow K^+ + OH^-$
e. $HC_2H_3O_2$	$HC_2H_3O_2 \rightleftharpoons C_2H_3O_2^- + H^+$
f. HCN	$HCN \rightleftharpoons H^+ + CN^-$
g. $Cu(OH)_2$	$Cu(OH)_2 \rightleftharpoons Cu^{2+} + 2OH^-$
h. NH_4OH	$NH_4OH \rightleftharpoons NH_4^+ + OH^-$

at 25°C

2. Calculate the $[H^+]$ and $[OH^-]$ of a 1.0×10^{-3} M solution of HCl, a strong acid.

$K_w = [H^+][OH^-] = 1 \times 10^{-14}$

$HCl \rightarrow H^+ + Cl^-$

1×10^{-3} M $\quad 1 \times 10^{-3}$ M $\quad$ pH 3

$\therefore [OH^-] = 1 \times 10^{-11}$ M

3. Calculate the $[OH^-]$ and the $[H^+]$ of a 0.0020 M solution of NaOH, a strong base.

100%

$NaOH \rightleftharpoons Na^+ + OH^-$ $\qquad$ 100%

so if 0.0020M then 0.0020M

$[H^+][OH^-] = 1 \times 10^{-14}$ $\qquad$ $[OH^-] = 0.0020$ M

&

$[H^+] = \dfrac{1 \times 10^{-14}}{0.0020 \text{ M}} = 5.0 \times 10^{-12}$ M $= [H^+]$

4. Benzoic acid, $HC_6H_5CO_2$, is an organic acid whose sodium salt, $NaC_6H_5CO_2$, has long been used as a safe food additive to protect beverages and many foods against harmful yeasts and bacteria. The acid is monoprotic. Write the equation for its K_a.

$$HC_6H_5CO_2\,(aq) \rightleftharpoons H^+(aq) + C_6H_5CO_2^-(aq)$$

$$K_A = \frac{[H^+][C_6H_5CO_2^-]}{[HC_6H_5CO_2]}$$

5. The $[H^+]$ of a 0.10 M solution of cyanic acid (HCNO) is found to be 0.0010 M. Calculate the K_a for cyanic acid.

$$\% \text{ ion} = \frac{[H^+]}{[HA]} \times 100 = \frac{0.0010M}{0.10M} \times 100 = 1\% \text{ ion}$$

∴ no RICE

$$HCNO \leftrightarrow H^+ + CNO^-$$

$$0.10M \quad 0.0010M \quad + \quad 0.0010M$$

$$K_A = \frac{[H^+][CNO^-]}{[HCNO]}$$

$$K_A = \frac{(0.0010)^2}{(0.10)}$$

$$K_A = 1.0 \times 10^{-5}$$

6. If 1.22 grams of benzoic acid, $HC_6H_5CO_2$, is dissolved in 1.0 L of water, the $[H^+]$ is found to be 8.0×10^{-4} M. Calculate the K_a for benzoic acid.

$$\frac{1.22g}{1.0L} \times \frac{1mol}{122g} = 0.010M$$

$$\% \text{ ion} = \frac{8.0 \times 10^{-4}M}{0.010m} \times 100 = 8\% \text{ ion}$$

∴ RICE

R	$HC_6H_5CO_2$	$\leftrightarrow$	H^+	+	$C_6H_5CO_2^-$
I	0.010		Ø		Ø
C	-8.0×10^{-4}		$+8.0 \times 10^{-4}$		$+8.0 \times 10^{-4}$
E	0.0092		8.0×10^{-4}		8.0×10^{-4}

$$K_A = \frac{[H^+][C_6H_5CO_2^-]}{[HC_6H_5CO_2]} = \frac{(8.0 \times 10^{-4})^2}{0.0092} = 7.0 \times 1$$

7. A 0.0050 M solution of butyric acid, HC_4H_7, has a pH =4.0, calculate K_a.

$$pH \rightarrow 4.0 \rightarrow [H^+] = 1.0 \times 10^{-4}$$

$$HC_4H_7O \leftrightarrow H^+ + C_4H_7O_2^-$$

$$\% = \frac{1.0 \times 10^{-4}M}{0.0050m} \times 100 = 2\% \text{ ion}$$

= yeehaw no RICE

$$K_A = \frac{[H^+][C_4H_7O_2^-]}{[HC_4H_7O]} = \frac{(1.0 \times 10^{-4})^2}{0.0050} = 2.0 \times 10^{-6}$$

8. Determine the $[OH^-]$ and the $[H^+]$ of a 0.20 M solution of formic acid. The $K_a = 1.8 \times 10^{-4}$

$$HCOOH \leftrightarrow H^+ + HCOO^-$$

$$0.20M \quad x \quad x$$

$$K_A = \frac{[H^+][HCOO^-]}{[HCOOH]}$$

$$[H^+] = \sqrt{K_A [HCOOH]}$$

$$= \sqrt{(1.8 \times 10^{-4})(0.20M)} = 6.0 \times 10^{-3} M = [H^+] \rightarrow \therefore [OH^-] = 1.67 \times 1$$

$-A \therefore pH < 7$

9. HCN has an initial molarity of 0.50 M, with a K_a value of 3.7×10^{-8}. Calculate its pH at equilibrium. ~~(Hint: This is an ICE problem.)~~ No bec on 2.72% ionized

$HCN \rightleftharpoons \underset{x}{H^+} + \underset{x}{CN^-}$ $\quad K_A = \dfrac{[H+][CN^-]}{[HCN]}$ $\quad K_A = \dfrac{(x)(x)}{[HCN]}$

$x = 1.36 \times 10^{-4} M$

$pH = 3.87$

p^+ acceptor

10. Ethylamine ($C_2H_5NH_3$) is a weak Bonsted-Lowry base. If it has an initial molarity of 0.024 M and a K_b of 5.6×10^{-4}, calculate its pH at equilibrium. (Hint: This is an ICE Problem.)

$\underset{0.024M}{C_2H_5NH_3} + H_2O \rightleftharpoons \underset{x}{C_2H_5NH_4^+} + \underset{x}{[OH^-]}$

$K_B = \dfrac{[C_2H_5NH_4^+][OH^-]}{[C_2H_5NH_3]} = 5.6 \times 10^{-4}$ $\quad [OH^-] = 3.39 \times 10^{-3} M$

$\therefore [H+] = 2.95 \times 10^{-12} M$

$K_B = \dfrac{(x)(x)}{(0.024-x)} = (5.6 \times 10^{-4})(0.024 - x) = x^2$

$x^2 + 5.6 \times 10^{-4} x - 1.34 \times 10^{-5} \therefore$ $pH = 11.53$

11. A chemist adds 0.75 moles of NH_3 to enough water to make 0.50 liters of solution. K_b of ammonia is 1.8×10^{-5}. Determine the pH of this solution at equilibrium. (Hint: This is an ICE problem.) (pH 11.72)

$[NH_3] = \dfrac{0.75 mol}{0.50 L} = 1.5 M$

	R NH_3	$+ H_2O \rightleftharpoons$	NH_4^+	$+ OH^-$
I	1.5 M	—	0	0
C	$-x$		$+x$	$+x$
E	$1.5 - x$		x	x

$K_B = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$ $\quad K_B = \dfrac{(x)(x)}{(1.5-x)} = \dfrac{x^2}{1.5-x} = 1.8 \times 10^{-5}$

$x^2 = (2.7 \times 10^{-5}) - 1.8 \times 10^{-5}x$

$0 = x^2 + 1.8 \times 10^{-5} x - 2.7 \times 10^{-5}$

$\therefore x = 5.19 \times 10^{-3} M = [OH^-] \rightarrow [H^+] = 1.93 \times 10^{-12} \therefore pH$

12. Hydrazine, N_2H_4, has been used as a rocket fuel. Like ammonia, it is a Bronsted base. A 0.15 M solution has a pH of 10.70. What is the K_b for hydrazine? (1.67×10^{-6})

$\hookrightarrow [H+] = 2.00 \times 10^{-11} M \therefore OH^- = 5.00 \times 10^{-4} M$

$x = 5.00 \times 10^{-4}$

$N_2H_4 + H_2O \rightleftharpoons \underset{x}{N_2H_5^+} + \underset{x}{OH^-}$

$\% ion = \dfrac{5.00 \times 10^{-4} M}{0.15 M} \times 100 = 0.33\%$

$\therefore$ Not RICE

$K_b = \dfrac{[N_2H_5^+][OH^-]}{[N_2H_4]} = \dfrac{(5.00 \times 10^{-4})(5.00 \times 10^{-4})}{(0.15)}$

$K_b = 1.67 \times 10^{-6}$