Name: Solutions 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	$H_2SO_4 \leftrightarrow H^+ + HSO_4^-$
d. KOH	$KOH \rightarrow K^+ + OH^-$
e. $HC_2H_3O_2$	$HC_2H_3O_2 \leftrightarrow H^+ + C_2H_3O_2$
f. HCN	$HCN \leftrightarrow H^+ + CN^-$
g. $Cu(OH)_2$	$Cu(OH)_2 \leftrightarrow Cu^{2+} + 2OH^-$
h. NH_4OH	$NH_4OH \rightarrow NH_4^+ + OH^-$

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

HCl - $[H^+] = 1 \times 10^{-3}$ M

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

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

NaOH - $[OH^-] = 0.0020$ M

$[H^+] = 5 \times 10^{-12}$ M

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.

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

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.

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

$[H^+] = [CNO^-]$

$$K_a = \frac{[H^+][CNO^-]}{[HCNO]} = \frac{(0.001)^2}{0.10} = 1.0 \times 10^{-5}$$

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.

$HC_6H_5CO_2 \leftrightarrow H^+ + C_6H_5CO_2^-$

$[H^+] = [C_6H_5CO_2^-]$

$[HC_6H_5CO_2] = \frac{1.22g}{1.0L} \times \frac{1 mol}{122.}$

$= 0.10 M$

$$K_a = \frac{[H^+][C_6H_5CO_2^-]}{[HC_6H_5CO_2]} = \frac{(8 \times 10^{-4})^2}{0.10}$$

$K_a = 6.36 \times 10^{-6}$

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

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

$[H^+] = [C_4H_7O_2^-]$

$pH = -\log[H^+]$

$[H^+] = 10^{-4.0}$ M

$$K_a = \frac{[H^+][C_4H_7O_2^-]}{[HC_4H_7O]}$$

$$K_a = \frac{(10^{-4})^2}{0.0050} = 2.0 \times 10^{-6}$$

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

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

$[H^+] = [HCOO^-]$

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

$$1.8 \times 10^{-4} = \frac{[H^+]^2}{0.20}$$

$[H^+]^2 = 3.6 \times 10^{-5}$

$[H^+] = 6.0 \times 10^{-3}$ M

$[OH^-] = 1.\bar{6} \times 10^{-12}$ M

(4) $pH = -\log[H^+]$

$pH = 3.87$

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.)

(1)

	[HCN]	$[H^+]$	$[CN^-]$
I	0.50M	0	0
C	$-x$	$+x$	$+x$
E	$0.5-x$	x	x

(2) $K_a = \frac{[H^+][CN^-]}{[HCN]}$

$3.7\times10^{-8} = \frac{x^2}{.5-x}$

$x^2 + 3.7\times10^{-8}x - 1.85\times10^{-8}$

(3) $x = \frac{-b \pm \sqrt{b^2-4ac}}{2a}$

$x = \frac{-3.7\times10^{-8} \pm \sqrt{(3.7\times10^{-8})^2 - 4(1)(-1.85\times10^{-8})}}{2(1)}$

$x = \frac{-3.7\times10^{-8} \pm 2.72\times10^{-4}}{2}$

$x = 1.36\times10^{-4}\ M = [H^+]$

* 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.)

$K_a = \frac{[H^+][C_2H_5NH_2^+]}{[C_2H_5NH_3]}$

$5.6\times10^{-4} = \frac{x^2}{.024-x}$

$x^2 + 5.6\times10^{-4} - 1.344\times10^{-5} = 0$

$x = \frac{-b \pm \sqrt{b^2-4ac}}{2a}$

$x = \frac{-5.6\times10^{-4} \pm \sqrt{(5.6\times10^{-4})^2 - 4(1)(-1.344\times10^{-5})}}{2(1)}$

$x = \frac{-5.6\times10^{-4} \pm 7.35\times10^{-3}}{2}$

$x = 3.40\times10^{-3} M = [H^+]$

$pH = -\log[H^+]$

$pH = 2.47$

* 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.)

$NH_4OH \rightarrow NH_4^+ + OH^-$

$\frac{0.75\ mol}{0.5L} = 1.5\ M = [NH_4OH]$

$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]}$

$1.8\times10^{-5} = \frac{x^2}{1.5-x}$

$x^2 + 1.8\times10^{-5}x - 2.7\times10^{-5} = 0$

$x = \frac{-b \pm \sqrt{b^2-4ac}}{2a}$

$x = \frac{-1.8\times10^{-5} \pm \sqrt{(1.8\times10^{-5})^2 - 4(1)(-2.7\times10^{-5})}}{2(1)}$

$x = \frac{-1.8\times10^{-5} \pm 1.039\times10^{-2}}{2}$

$x = 5.187\times10^{-3} = [OH^-]$

$pOH = -\log[OH]$

$pOH = 2.285$

$pH = 14 - pOH$

$pH = 11.71$

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?

$pOH = 3.3 \rightarrow [OH^-] = 10^{-3.3} = 5.01\times10^{-4}\ M$

$K_b = \frac{[OH^-][N_2H_5^+]}{[N_2H_4]} = \frac{(5.01\times10^{-4})^2}{0.15}$

$N_2H_5OH \rightarrow N_2H_5^+ + OH^-$

$K_b = 1.67\times10^{-6}$

* Skipped ICE table to save room - DO NOT do this