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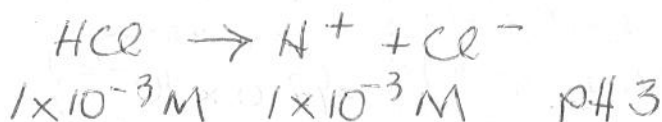
## K<sub>a</sub> and K<sub>b</sub> Calculations Worksheet

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.

- 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	$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$
b. NaOH	$\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$
c. H <sub>2</sub> SO <sub>4</sub>	$\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^-$
d. KOH	$\text{KOH} \rightarrow \text{K}^+ + \text{OH}^-$
e. HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	$\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{C}_2\text{H}_3\text{O}_2^- + \text{H}^+$
f. HCN	$\text{HCN} \rightleftharpoons \text{H}^+ + \text{CN}^-$
g. Cu(OH) <sub>2</sub>	$\text{Cu(OH)}_2 \rightleftharpoons \text{Cu}^{2+} + 2\text{OH}^-$
h. NH <sub>4</sub> OH	$\text{NH}_4\text{OH} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

- Calculate the [H<sup>+</sup>] and [OH<sup>-</sup>] of a 1.0 x 10<sup>-3</sup> M solution of HCl, a strong acid.

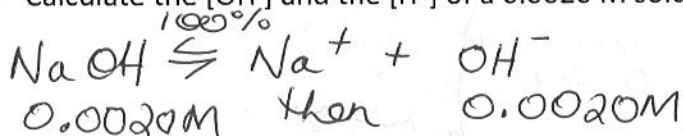


at 25°C

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

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

- Calculate the [OH<sup>-</sup>] and the [H<sup>+</sup>] of a 0.0020 M solution of NaOH, a strong base.

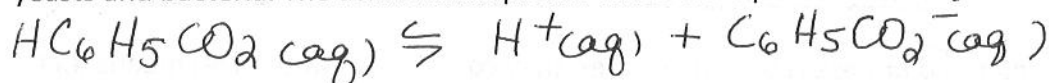


$$[\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$$

$$[\text{OH}^-] = 0.0020 \text{ M}$$

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

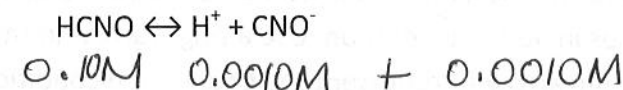
4. Benzoic acid,  $\text{HC}_6\text{H}_5\text{CO}_2$ , is an organic acid whose sodium salt,  $\text{NaC}_6\text{H}_5\text{CO}_2$ , has long been used as a safe food additive to protect beverages and many foods against harmful yeasts and bacteria. The acid is monoprotic. Write the equation for its  $K_a$ .



$$K_A = \frac{[\text{H}^+][\text{C}_6\text{H}_5\text{CO}_2^-]}{[\text{HC}_6\text{H}_5\text{CO}_2]}$$

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

$$\begin{aligned} \% \text{ ion} &= \frac{[\text{H}^+]}{[\text{HA}]} \times 100 \\ &= \frac{0.0010\text{M}}{0.10\text{M}} \times 100 \\ &= 1\% \text{ ion} \\ &\therefore \text{no RICE} \end{aligned}$$



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

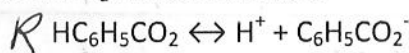
$$K_A = \frac{(0.0010)^2}{(0.10)}$$

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

6. If 1.22 grams of benzoic acid,  $\text{HC}_6\text{H}_5\text{CO}_2$ , is dissolved in 1.0 L of water, the  $[\text{H}^+]$  is found to be  $8.0 \times 10^{-4}$  M. Calculate the  $K_a$  for benzoic acid.

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

$$\begin{aligned} \% \text{ ion} &= \frac{8.0 \times 10^{-4}\text{M}}{0.010\text{M}} \times 100 \\ &= 8\% \text{ ion} \\ &\therefore \text{RICE} \end{aligned}$$



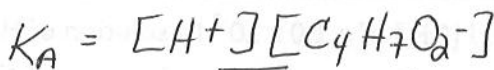
$$\begin{array}{l} \text{I} \quad 0.010 \quad \emptyset \quad \emptyset \\ \text{C} \quad -8.0 \times 10^{-4} \quad +8.0 \times 10^{-4} \quad +8.0 \times 10^{-4} \\ \text{E} \quad 0.0092 \quad 8.0 \times 10^{-4} \quad 8.0 \times 10^{-4} \end{array}$$

$$K_A = \frac{[\text{H}^+][\text{C}_6\text{H}_5\text{CO}_2^-]}{[\text{HC}_6\text{H}_5\text{CO}_2]} = \frac{(8.0 \times 10^{-4})^2}{0.0092} = 7.0 \times 10^{-5}$$

7. A 0.0050 M solution of butyric acid,  $\text{HC}_4\text{H}_7$ , has a pH = 4.0, calculate  $K_a$ .

$$\text{pH} \rightarrow 4.0 \rightarrow [\text{H}^+] = 1.0 \times 10^{-4} \quad \text{HC}_4\text{H}_7\text{O} \leftrightarrow \text{H}^+ + \text{C}_4\text{H}_7\text{O}_2^-$$

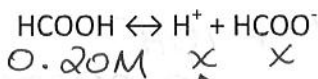
$$\begin{aligned} \% &= \frac{1.0 \times 10^{-4}\text{M}}{0.0050\text{M}} \times 100 \\ &= 2\% \text{ ion} = \text{yes!} \\ &\therefore \text{no RICE} \end{aligned}$$



$$= \frac{(1.0 \times 10^{-4})^2}{0.0050} = 2.0 \times 10^{-6}$$

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

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

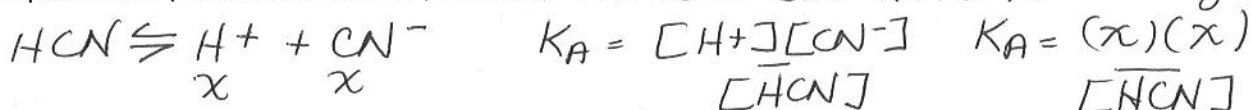


$$[\text{H}^+] = \sqrt{K_A [\text{HCOOH}]}$$

$$= \sqrt{(1.8 \times 10^{-4})(0.20\text{M})} = 6.0 \times 10^{-3}\text{M} = [\text{H}^+] \rightarrow \therefore [\text{OH}^-] = 1.67 \times 10^{-11}$$

- A: pH < 7

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

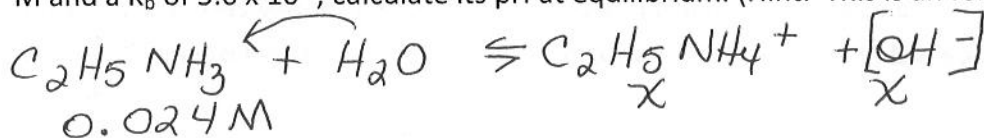


$$x = 1.36 \times 10^{-4} \text{ M}$$

$\text{p}^+$  acceptor

pH = 3.87

10. Ethylamine ( $\text{C}_2\text{H}_5\text{NH}_3$ ) is a weak Bonsted-Lowry base. If it has an initial molarity of 0.024 M and a  $K_b$  of  $5.6 \times 10^{-4}$ , calculate its pH at equilibrium. (Hint: This is an ICE Problem.)



$$K_B = \frac{[\text{C}_2\text{H}_5\text{NH}_4^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_3]} = 5.6 \times 10^{-4}$$

$$[\text{OH}^-] = 3.79 \times 10^{-3} \text{ M}$$

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

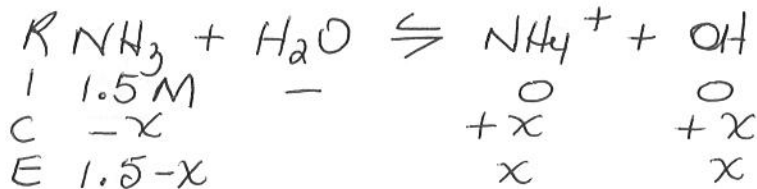
$$K_B = (x)(x) = (5.6 \times 10^{-4})(0.024 - x) = x^2$$

$$(0.024 - x) \quad x^2 + 5.6 \times 10^{-4}x - 1.34 \times 10^{-5} = 0$$

pH = 11.53

11. A chemist adds 0.75 moles of  $\text{NH}_3$  to enough water to make 0.50 liters of solution.  $K_b$  of ammonia is  $1.8 \times 10^{-5}$ . Determine the pH of this solution at equilibrium. (Hint: This is an ICE problem.)

$$[\text{NH}_3] = \frac{0.75 \text{ mol}}{0.50 \text{ L}} = 1.5 \text{ M}$$



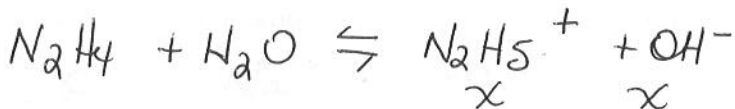
$$K_B = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \quad K_B = \frac{(x)(x)}{(1.5 - x)} = \frac{x^2}{1.5 - x} = 1.8 \times 10^{-5}$$

$$x^2 = (2.7 \times 10^{-5})(1.5 - x) - 1.8 \times 10^{-5}$$

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

12. Hydrazine,  $\text{N}_2\text{H}_4$ , 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? ( $1.67 \times 10^{-6}$ )

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



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

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

$$K_b = \frac{[\text{N}_2\text{H}_5^+][\text{OH}^-]}{[\text{N}_2\text{H}_4]} = \frac{(5.00 \times 10^{-4})(5.00 \times 10^{-4})}{(0.15)} = 0.33\%$$

K<sub>b</sub> = 1.67 × 10<sup>-6</sup>

∴ Not RICE