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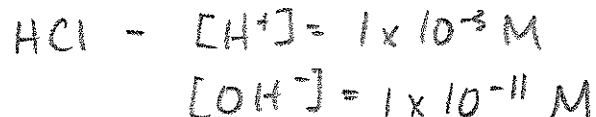
## K<sub>a</sub> and K<sub>b</sub> Calculations Worksheet

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.

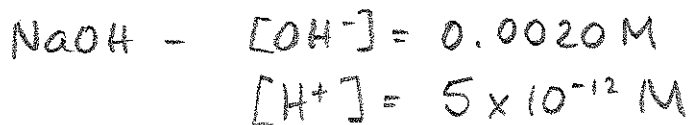
1. Write chemical equations which represent the dissociation of each of these acids or bases in aqueous solution. Use a single arrow in the case of a strong acid or base, and a double arrow to represent the equilibrium condition that exists in the solution of a weak acid or base.

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

2. 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.



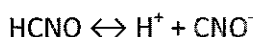
3. Calculate the [OH<sup>-</sup>] and the [H<sup>+</sup>] of a 0.0020 M solution of NaOH, a strong base.



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  $K_a$ .

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

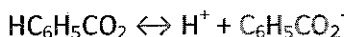
5. The  $[\text{H}^+]$  of a 0.10 M solution of cyanic acid ( $\text{HCNO}$ ) is found to be 0.0010 M. Calculate the  $K_a$  for cyanic acid.



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

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

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.



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

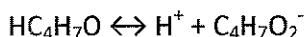
$$K_a = \frac{[\text{H}^+][\text{C}_6\text{H}_5\text{CO}_2^-]}{[\text{HC}_6\text{H}_5\text{CO}_2]} = \frac{(8 \times 10^{-4})^2}{0.10}$$

$$[\text{HC}_6\text{H}_5\text{CO}_2] = \frac{1.22 \text{ g}}{1.0 \text{ L}} \times \frac{1 \text{ mol}}{122 \text{ g}}$$

$$= 0.10 \text{ M}$$

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

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



$$[\text{H}^+] = [\text{C}_4\text{H}_7\text{O}_2^-]$$

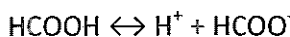
$$K_a = \frac{[\text{H}^+][\text{C}_4\text{H}_7\text{O}_2^-]}{[\text{HC}_4\text{H}_7\text{O}_2]}$$

$$\text{pH} = -\log [\text{H}^+]$$

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

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

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



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

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

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

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

$$[\text{OH}^-] = 1.6 \times 10^{-12} \text{ M}$$

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

①  $pH = -\log [H^+]$   
 $pH = 3.87$

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.)

|   | [HCN] | [H <sup>+</sup> ] | [CN <sup>-</sup> ] |
|---|-------|-------------------|--------------------|
| I | 0.50M | 0                 | 0                  |
| C | -x    | +x                | +x                 |
| E | 0.5-x | x                 | x                  |

②  $K_a = \frac{[H^+][CN^-]}{[HCN]}$   
 $3.7 \times 10^{-8} = \frac{x^2}{0.5-x}$   
 $x^2 + 3.7 \times 10^{-8}x - 1.85 \times 10^{-8} = 0$

③  $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$   
 $x = \frac{-3.7 \times 10^{-8} \pm \sqrt{(3.7 \times 10^{-8})^2 - 4(1)(-1.85 \times 10^{-8})}}{2(1)}$   
 $x = \frac{-3.7 \times 10^{-8} \pm 2.72 \times 10^{-4}}{2}$   
 $x = 1.36 \times 10^{-4} M = [H^+]$

\* 10. Ethylamine ( $C_2H_5NH_2$ ) is a weak Bronsted-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.)

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

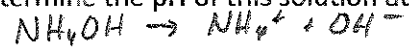
$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$   
 $x = \frac{-5.6 \times 10^{-4} \pm \sqrt{(5.6 \times 10^{-4})^2 - 4(1)(-1.344 \times 10^{-5})}}{2(1)}$   
 $x = \frac{-5.6 \times 10^{-4} \pm 7.35 \times 10^{-3}}{2}$   
 $x = 3.40 \times 10^{-3} M = [H^+]$

$5.6 \times 10^{-4} = \frac{x^2}{0.024-x}$

$x^2 + 5.6 \times 10^{-4}x - 1.344 \times 10^{-5} = 0$

$pH = -\log [H^+]$   
 $pH = 2.47$

\* 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  $1.8 \times 10^{-5}$ . Determine the pH of this solution at equilibrium. (Hint: This is an ICE problem.)



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

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

$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$   
 $x = \frac{-1.8 \times 10^{-5} \pm \sqrt{(1.8 \times 10^{-5})^2 - 4(1)(-2.7 \times 10^{-5})}}{2(1)}$   
 $x = \frac{-1.8 \times 10^{-5} \pm 1.039 \times 10^{-2}}{2}$   
 $x = 5.187 \times 10^{-3} = [OH^-]$

$1.8 \times 10^{-5} = \frac{x^2}{1.5-x}$

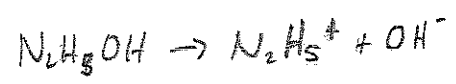
$x^2 + 1.8 \times 10^{-5}x - 2.7 \times 10^{-5} = 0$

$pOH = -\log [OH^-]$   
 $pOH = 2.285$   
 $pH = 14 - pOH$   
 $pH = 11.71$

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?

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

$K_b = \frac{[OH^-][N_2H_5^+]}{[N_2H_4]} = \frac{(5.01 \times 10^{-4})^2}{0.15}$



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

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