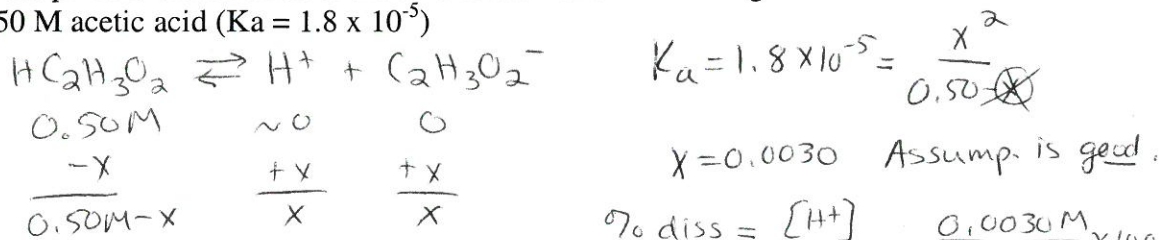


Answers

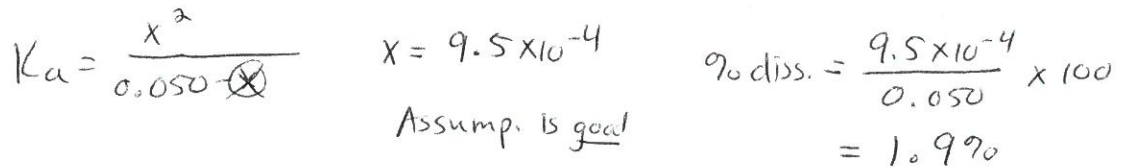
Acid/Base Equilibrium Practice

1. Calculate the percent dissociation of the acid in each of the following solutions:

a. 0.50 M acetic acid ($K_a = 1.8 \times 10^{-5}$)



b. 0.050 M acetic acid



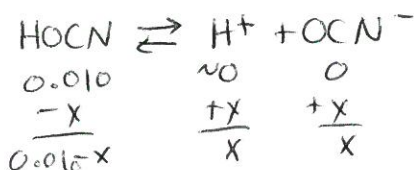
c. Use Le Chatelier's principle to explain why percent dissociation increases as the concentration of a weak acid decreases. [H⁺]

↓ [H⁺], shift right, ↑ dissociation.

2. The pH of a 0.010 M solution of cyanic acid (HOCN) is 2.77 at 25°C. Calculate the K_a for HOCN from this result.

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

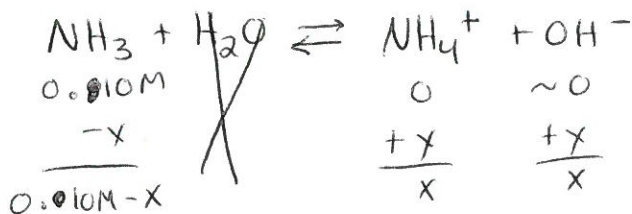
$$10^{-2.77} = 1.7 \times 10^{-3} \text{ M} = [\text{H}^+] = x$$



$$K_a = \frac{(1.7 \times 10^{-3} \text{ M})(1.7 \times 10^{-3} \text{ M})}{(0.010 - (1.7 \times 10^{-3}))}$$

$$\boxed{K_a = 3.5 \times 10^{-4}}$$

3. Calculate the percent ionization of 0.10 M NH_3 ($K_b = 1.8 \times 10^{-5}$).



$$K_b = 1.8 \times 10^{-5} = \frac{x^2}{0.10 \text{ M}}$$

$$x = 1.3 \times 10^{-3}$$

Assump. is good.

$$\% \text{ diss} = \frac{1.3 \times 10^{-3} \text{ M}}{0.10 \text{ M}} \times 100 = \boxed{1.3\%}$$

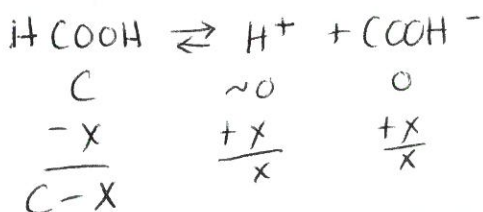
4. A solution of formic acid (HCOOH , $K_a = 1.8 \times 10^{-4}$) has a pH of 2.70. Calculate the initial concentration of formic acid in this solution.

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

$$10^{-2.70} = 2.0 \times 10^{-3} \text{ M}$$

$$K_a = 1.8 \times 10^{-4} = \frac{x^2}{C - x} = \frac{(2.0 \times 10^{-3} \text{ M})^2}{C - (2.0 \times 10^{-3} \text{ M})}$$

$$\boxed{C = 2.4 \times 10^{-2} \text{ M}}$$



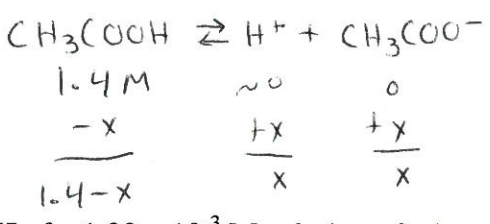
$$x = [\text{H}^+] = 2.0 \times 10^{-3} \text{ M}$$



5. A solution with a total volume of 250.0 mL is prepared by diluting 20.0 mL of glacial acetic acid with water. Calculate the $[H^+]$ and the pH of this solution. Assume that glacial acetic acid is pure liquid acetic acid with a density of 1.05 g/cm^3 . ($K_a = 1.8 \times 10^{-5}$).

$$\frac{1.05 \text{ g}}{\text{mL}} \times \frac{20.0 \text{ mL}}{1 \text{ mol}} = 0.35 \text{ mol}$$

$$\frac{0.35 \text{ mol}}{0.25 \text{ L}} = 1.4 \text{ M}$$



$$1.8 \times 10^{-5} = \frac{x^2}{1.4-x}$$

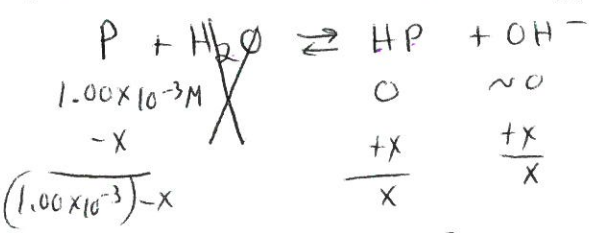
$$x = 5.0 \times 10^{-3} \text{ M} = [H^+]$$

Assump. is good.

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

$\text{pH} = 2.30$

6. The pH of a $1.00 \times 10^{-3} \text{ M}$ solution of a base, pyrrolidine, is 10.82. Calculate the K_b .

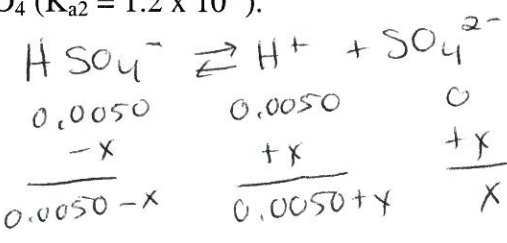
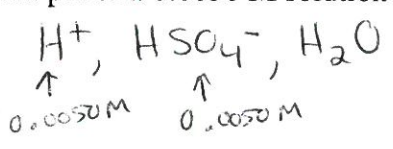


pH = 10.82
 $\text{pOH} = 14 - 10.82 = 3.18$
 $10^{-\text{pOH}} = [OH^-] = 6.6 \times 10^{-4} = x$

$$K_b = \frac{(6.6 \times 10^{-4})^2}{(1.00 \times 10^{-3}) - (6.6 \times 10^{-4})}$$

$K_b = 1.3 \times 10^{-3}$

7. Calculate the pH of a 0.0050 M solution of H_2SO_4 ($K_{a2} = 1.2 \times 10^{-2}$).



$$K_{a2} = \frac{(0.0050 + x)(x)}{(0.0050 - x)} = 1.2 \times 10^{-2}$$

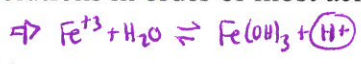
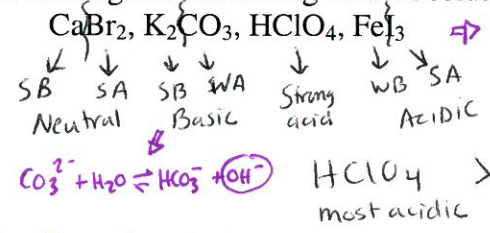
Assump. Not good.

$$(6 \times 10^{-5}) - (1.2 \times 10^{-2})x = 0.0050x + x^2$$

$$x^2 + 0.017x - (6 \times 10^{-5}) = 0$$

$$x = \frac{(-0.017) \pm \sqrt{(0.017)^2 - 4(1)(-6 \times 10^{-5})}}{2(1)}$$

8. Arrange the following 0.10 M solutions in order of most acidic to most basic:



$$x = 3.0 \times 10^{-3} \text{ M}$$

$$[H^+] = (0.0050) + (3.0 \times 10^{-3} \text{ M})$$

$$[H^+] = 8.0 \times 10^{-3} \text{ M}$$

$\text{pH} = 2.10$

9. Given that the K_a value for acetic acid is 1.8×10^{-5} and the K_a value for hypochlorous acid is 3.5×10^{-8} , which is the stronger base, OCl^- or $C_2H_3O_2^-$?

Calculate K_b for each compound & compare. $\uparrow K_b = \uparrow \text{base}$

OR
 Bigger K_a for acetic acid means stronger acid.
 Stronger acid means weaker conj. base

Answers: 1.a. 0.60%

b. 1.9%

c. decrease $[H^+]$ = shift right = more dissociation

2. $K_a = 3.5 \times 10^{-4}$

3. 1.3%

4. $[HCOOH] = 2.4 \times 10^{-2} \text{ M}$

5. pH = 2.30

6. $K_b = 1.3 \times 10^{-3}$

7. pH = 2.20

8. $HClO_4 > FeI_3 > CaBr_2 > K_2CO_3$

9. OCl^-