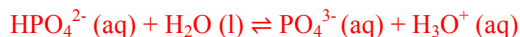
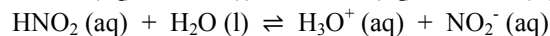
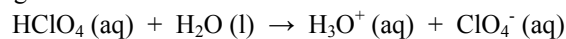


## CHM152 – Chapter 15: Acids and Bases – Homework - KEY

- What is the conjugate base of  $\text{HSO}_4^-$ ?  $\text{SO}_4^{2-}$
- Write a balanced equation for the Bronsted-Lowry acid  $\text{HPO}_4^{2-}$  in water.



- From the following chemical reactions determine the relative Bronsted-Lowry acid strengths (strongest to weakest).



- $\text{H}_3\text{O}^+ > \text{HNO}_2 > \text{HClO}_4$
  - $\text{H}_3\text{O}^+ > \text{HClO}_4 > \text{HNO}_2$
  - $\text{HClO}_4 > \text{HNO}_2 > \text{H}_3\text{O}^+$
  - $\text{HClO}_4 > \text{H}_3\text{O}^+ > \text{HNO}_2$
- Calculate the pH of an aqueous solution that contains  $2.50 \times 10^{-4} \text{ M } [\text{H}_3\text{O}^+]$ .  
 $\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(2.50 \times 10^{-4}) = 3.602$
  - Calculate the pH of an aqueous solution that contains  $3.50 \times 10^{-3} \text{ M } [\text{OH}^-]$ .  
 $\text{pOH} = -\log[\text{OH}^-] = -\log(3.50 \times 10^{-3}) = 2.456$        $\text{pH} + \text{pOH} = 14.00$      $\text{pH} = 11.544$
  - What is the pH of a solution prepared by mixing 100.00 mL of 0.020 M  $\text{Ca}(\text{OH})_2$  with 50.00 mL of 0.100 M NaOH? Assume that the volumes are additive.

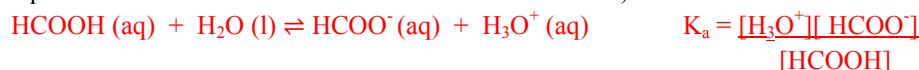
$$0.10000\text{L} \left( \frac{0.020 \text{ mol Ca}(\text{OH})_2}{1 \text{ L}} \right) \left( \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca}(\text{OH})_2} \right) = 4.0 \times 10^{-3} \text{ mol OH}^-$$

$$0.05000\text{L} \left( \frac{0.100 \text{ mol NaOH}}{1 \text{ L}} \right) \left( \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} \right) = 5.00 \times 10^{-3} \text{ mol OH}^-$$

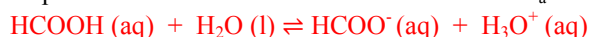
$$\text{M of OH}^- = \frac{4.0 \times 10^{-3} \text{ mol OH}^- + 5.00 \times 10^{-3} \text{ mol OH}^-}{0.15000 \text{ L}} = 9.0 \times 10^{-3} \text{ mol OH}^- / 0.15000 \text{ L} = 0.060 \text{ M OH}^-$$

$$\text{pOH} = -\log 0.060 = 1.22 \quad \text{SO } \text{pH} = 12.78$$

- What is the equilibrium constant expression ( $K_a$ ) for the acid dissociation of formic acid ( $\text{HCOOH}$ )? (Hint: write a balanced equation for the dissociation of formic acid in water first.)



- What is the pH of a 0.100 M formic acid solution with a  $K_a = 1.8 \times 10^{-4}$ ?



$$\text{I} \quad 0.100\text{M}$$

$$\text{C} \quad -x \qquad \qquad \qquad +x \qquad \qquad +x$$

$$\text{E} \quad 0.100 - x \qquad \qquad \qquad x \qquad \qquad x$$

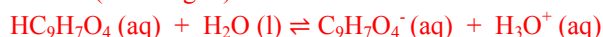
$$K_a = 1.8 \times 10^{-4} = \frac{x^2}{(0.100 - x)} \quad \text{Assume } x \text{ is negligible so } x = 0.0042 \text{ M} = [\text{H}_3\text{O}^+]$$

$$(0.100 - x) \quad (\text{Assumption is ok since } (0.0042/0.100) \times 100\% = 4.2\%)$$

$$\text{pH} = -\log(0.0042) = 2.38$$

- A tablet containing 500.0 mg of aspirin (acetylsalicylic acid or  $\text{HC}_9\text{H}_7\text{O}_4$ ) was dissolved in enough water to make 100.0 mL of solution. Given that  $K_a = 3.0 \times 10^{-4}$  for aspirin, what is the pH of the solution?

$$0.50000 \text{ g ACA} \left( \frac{1 \text{ mol ACA}}{180.17 \text{ g}} \right) = 2.775 \times 10^{-3} \text{ mol ACA} / 0.1000\text{L} = 0.02775 \text{ M ACA}$$



$$\text{I} \quad 0.02775 \text{ M}$$

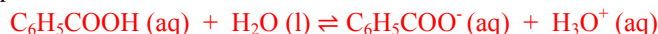
$$\text{C} \quad -x \qquad \qquad \qquad +x \qquad \qquad +x$$

$$\text{E} \quad 0.02775 \text{ M} - x \qquad \qquad \qquad x \qquad \qquad x$$

$$K_a = 3.0 \times 10^{-4} = \frac{x^2}{(0.02775 - x)} = \text{solve for } x \text{ using the quadratic formula } x = 0.0027 \text{ M}$$

$$\text{pH} = -\log(0.0027) = 2.56$$

10. What is the percent dissociation of a benzoic acid ( $\text{C}_6\text{H}_5\text{COOH}$ ) solution with  $\text{pH} = 2.59$ ? The acid dissociation constant for this monoprotic acid is  $6.5 \times 10^{-5}$ .



I

C

$$\text{E} \quad x - 0.0026 \qquad \qquad \qquad x \qquad \qquad \qquad x = 10^{-2.59} = 0.0026 \text{ M}$$

$$K_a = 6.5 \times 10^{-5} = \frac{(0.0026)^2}{(x - 0.0026)} = \text{solve for } x = 0.106 \text{ M} = 0.11 \text{ M}$$

$$\text{So } [\text{C}_6\text{H}_5\text{COOH}]_{\text{initial}} = x - 0.0026 = 0.106 \text{ M} + 0.0026 \text{ M} = 0.11 \text{ M}$$

$$\% \text{ ionization} = (0.0026 \text{ M} / 0.11 \text{ M}) \times 100\% = 2.4 \%$$

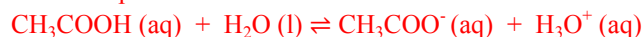
11. Vinegar is a 5.0% m/m% solution of acetic acid in water. Given that  $K_a = 1.8 \times 10^{-5}$  for acetic acid and assuming the density of vinegar to be 1.00 g/mL, what is the pH of this vinegar solution?

5.0% m/m basically means if you assume you have 100 g of vinegar solution, 5.0 g is solute (vinegar)

SO, 100.0 g of vinegar solution with a density of 1.00 g/mL gives 100.0 mL of solution.

5.0 g of solute (acetic acid) and using molar mass of acetic acid (60.06 g/mol) gives 0.083 mol of acetic acid. So the total molarity of acetic acid is  $0.083 \text{ mol} / 0.1000 \text{ L} = 0.83 \text{ M}$

Now set up ICE table:



I 0.83 M

C -x +x +x

E 0.83 M - x x x

$$K_a = 1.8 \times 10^{-5} = \frac{x^2}{(0.83 - x)} = \text{solve for } x \text{ using the assumption that } x \text{ is negligible and } x = 0.0039 \text{ M}$$

$$\text{pH} = -\log(0.0039) = 2.41$$

12. Calculate the concentration of bicarbonate ion,  $\text{HCO}_3^-$ , in a 0.010 M  $\text{H}_2\text{CO}_3$  solution that has the stepwise dissociation constants  $K_{a1} = 4.3 \times 10^{-7}$  and  $K_{a2} = 5.6 \times 10^{-11}$ .



I 0.010 M

C -x +x +x

E 0.010 M - x x x

$$K_{a1} = 4.3 \times 10^{-7} = \frac{x^2}{(0.010 - x)} = \text{solve for } x \text{ using the assumption that } x \text{ is negligible and } x = 6.6 \times 10^{-5} \text{ M}$$



I  $6.6 \times 10^{-5} \text{ M}$

C -x +x +x

E  $6.6 \times 10^{-5} \text{ M} - x$  x x

$$K_{a2} = 5.6 \times 10^{-11} = \frac{x^2}{(6.6 \times 10^{-5} - x)} = \text{solve for } x \text{ using the assumption that } x \text{ is negligible and } x = 6.1 \times 10^{-8} \text{ M}$$

$$[\text{HCO}_3^-] = 6.6 \times 10^{-5} \text{ M} - x = 6.6 \times 10^{-5} \text{ M} - 6.1 \times 10^{-8} \text{ M} = 6.6 \times 10^{-5} \text{ M}$$

13. Ammonia has a  $K_b$  of  $1.8 \times 10^{-5}$ . What is the conjugate acid of ammonia and what is its  $K_a$ ?  
The conjugate acid is  $\text{NH}_4^+$  and  $K_a$  is  $K_w/K_b = 5.6 \times 10^{-10}$

