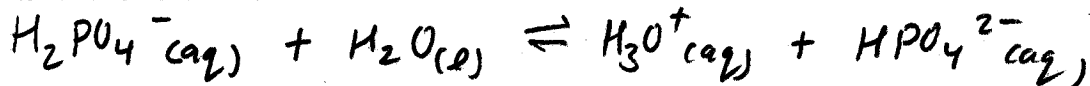


Name KEY

Chemistry 11
pH and Buffers - Advance Study Assignment
Due on December 9, 2002 at noon

1. The pH of a 0.10 M solution of NaH_2PO_4 is 4.5. Write a net ionic equation that accounts for this observation.



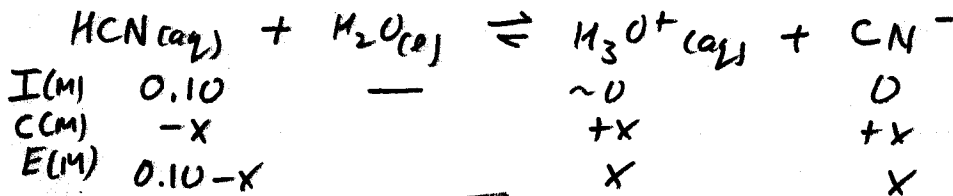
2. HCN is a weak acid. The pH of a 0.10 M HCN solution is 5.2.

- a. Calculate the molarity of H_3O^+ in this solution.

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \Rightarrow [\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-5.2} = 6.3 \times 10^{-6} \text{ M}$$

(technically 1 sig fig)

- b. Calculate the molarity of CN^- and HCN in this same solution.



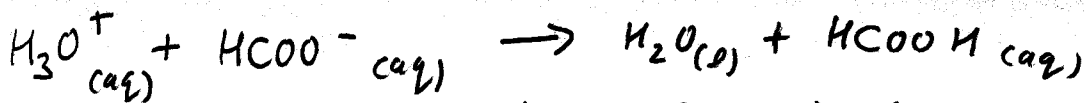
and $x = \boxed{6.3 \times 10^{-6} \text{ M}} = [\text{CN}^-] \Rightarrow [\text{HCN}] = (0.10 - 6.3 \times 10^{-6}) \text{ M}$
 $= \boxed{0.10 \text{ M}}$

- c. Use the above results to calculate a numerical value for K_a for HCN. Compare your value with that given in your text.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = \frac{(6.3 \times 10^{-6})^2}{0.10} = 4 \times 10^{-10}$$

vs. Silberberg p. 762 $K_a = 6.2 \times 10^{-10}$

So overestimate is $\frac{(4 - 6.2) \times 10^{-10}}{6.2 \times 10^{-10}} \times 100\% \approx \boxed{-35\% \text{ off}}$
(not bad!)



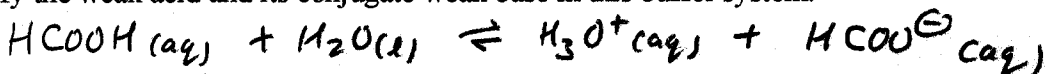
Moreover, since we still have a fair bit of HCOO^- in sol'n, we suppress (by Le Chatelier's principle) the rxn of HCOOH w/ H_2O to (re-)generate H_3O^+ .

Name KEY

pH and Buffers ASA (cont.)

3. A student is asked to use 0.10 M formic acid, HCOOH , and 0.10 M sodium formate, HCOONa to prepare a buffer solution having a pH of 3.40. K_a for HCOOH is 1.8×10^{-4} . $\Rightarrow pK_a = -\log K_a = 3.745$

- a. Identify the weak acid and its conjugate weak base in this buffer system.



weak acid

weak conjugate base

- b. Calculate the volumes of the two solutions that need to be mixed to make 1.00 liter of this buffer.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{HCOO}^-]_0}{[\text{HCOOH}]_0} \quad (\text{Henderson-Hasselbalch} - \text{neglect the rxns of both HCOOH} + \text{HCOO}^- \text{ w/ H}_2\text{O})$$

$$\log \frac{[\text{HCOO}^-]_0}{[\text{HCOOH}]_0} = \text{pH} - \text{p}K_a \Rightarrow \frac{[\text{HCOO}^-]_0}{[\text{HCOOH}]_0} = 10^{(\text{pH} - \text{p}K_a)} = 10^{(3.40 - 3.745)} = 0.452 \quad (\text{only 2 sig figs})$$

(continued...)

- c. When 5 drops of 0.10 M HCl were added to 20 mL of this buffer, the pH went from 3.40 to 3.37. Write a net ionic equation that explains why the pH change was much smaller than it would have been if this volume of HCl were added to 20 mL of distilled water.

Let $V_a \equiv$ vol of HCOOH sol'n used and $V_b \equiv$ vol of HCOO^- sol'n used.

$$V_a + V_b = 1.00 \text{ L} \Rightarrow V_b = 1.00 \text{ L} - V_a$$

$$\text{In buffer, } [\text{HCOOH}]_0 = \frac{(0.10 \text{ M}) V_a}{1.00 \text{ L}} \quad \text{and} \quad [\text{HCOO}^-]_0 = \frac{(0.10 \text{ M}) V_b}{1.00 \text{ L}}$$

(ie the 2 individual sol'ns get diluted when they're mixed)

$$\text{so } \frac{[\text{HCOO}^-]_0}{[\text{HCOOH}]_0} = \frac{\frac{(0.10 \text{ M}) V_b}{1.00 \text{ L}}}{\frac{(0.10 \text{ M}) V_a}{1.00 \text{ L}}} = \frac{V_b}{V_a} = \frac{1.00 \text{ L} - V_a}{V_a} = 0.452$$

$$1.00 \text{ L} - V_a = 0.452 V_a \Rightarrow 1.452 V_a = 1.00 \text{ L} \Rightarrow \boxed{V_a = 0.69 \text{ L}}$$

\Downarrow

$$\boxed{V_b = 0.31 \text{ L}}$$