Chem 116 POGIL Worksheet - Week 10 - Solutions
Common Ion Effect and Buffers

Key Questions

1. What is the pH of a solution prepared by adding 0.20 mole of formic acid, HCO_2H, and 0.25 mole of sodium formate, $NaHCO_2$, in enough water to make a liter of solution? The K_a of formic acid is 1.8×10^{-4}.

$$K_a = \frac{[H_3O^+][HCO_2^-]}{[HCO_2H]} = \frac{[H_3O^+](0.25)}{0.20} = 1.8 \times 10^{-4}$$

$$[H_3O^+] = 1.44 \times 10^{-4}$$

$$pH = 3.84$$

2. What is the pH of a solution prepared by adding 0.60 mole of formic acid and 0.75 mole of sodium formate in enough water to make a liter of solution? How does your answer compare to the answer you obtained in Key Question 8? Explain.

$$K_a = \frac{[H_3O^+][HCO_2^-]}{[HCO_2H]} = \frac{[H_3O^+](0.75)}{0.60} = 1.8 \times 10^{-4}$$

$$[H_3O^+] = 1.44 \times 10^{-4}$$

$$pH = 3.84$$

The pH is the same in both cases because the ratio $[HCO_2^-]/[HCO_2H]$ is the same in both cases.

3. The solution prepared in Key Question 9 was diluted with enough water to make ten liters of solution. What is the pH of the dilute solution? How does it compare to the previous two solutions? Explain.

Diluting to ten liters makes the analytical concentrations of both the acid and its conjugate base become one tenth what they were; i.e.,

$$C_{HCO_2H} = 0.060\,M \qquad C_{HCO_2^-} = 0.075\,M$$

But the ratio $[HCO_2^-]/[HCO_2H]$ remains the same as it was in the solutions in Key Questions 8 and 9, so the pH is still 3.84.

$$K_a = \frac{[H_3O^+][HCO_2^-]}{[HCO_2H]} = \frac{[H_3O^+](0.075)}{0.060} = 1.8 \times 10^{-4}$$

$$[H_3O^+] = 1.44 \times 10^{-4}$$

$$pH = 3.84$$

It is the ratio between the concentrations of the conjugate pair that fixes the pH in solutions of this type, not the actual concentrations themselves.

4. A student wishes to prepare a buffer with a pH of 4.52. She starts by preparing one liter of 0.10 M formic acid, HCO_2H. How many grams of sodium formate, $NaHCO_2$, must she add to the solution to achieve the desired pH? $K_a = 1.8 \times 10^{-4}$ for formic acid. The formula weight of sodium formate is 68.00 u.

   First we need to determine what concentration of hydronium ion corresponds to pH = 4.54.

$$[H_3O^+] = 10^{-4.52} = 3.0 \times 10^{-5}\ M$$

   Next, by substituting into the K_a expression, we determine the ratio $[HCO_2^-]/[HCO_2H]$ that will give $[H_3O^+] = 3.0 \times 10^{-5}$ M..

$$\frac{K_a}{[H_3O^+]} = \frac{[HCO_2^-]}{[HCO_2H]} = \frac{1.8\times10^{-4}}{3.0\times10^{-5}} = 6.0$$

   One liter of 0.10 M HCO_2H contains 0.10 mol of the acid. Therefore,

$$\text{g NaHCO}_2 = (0.10\ \text{mol HCO}_2\text{H})\left(\frac{6\ \text{mol NaHCO}_2}{1\ \text{mol HCO}_2\text{H}}\right)\left(\frac{68.00\,\text{g NaHCO}_2}{\text{mol NaHCO}_2}\right) = 40.8\,\text{g NaHCO}_2$$

5. A buffer solution is made by adding an equal number of moles of the base pyradine, C_5H_5N, and the chloride salt of it conjugate base, C_5H_5NHCl. What is the pH of the buffer solution? The K_b of C_5H_5N is 1.7×10^{-9}.

   If $[C_5H_5N] = [C_5H_5NH^+]$, then substituting into the K_b expression gives

$$K_b = \frac{[C_5H_5NH^+][OH^-]}{[C_5H_5N]} = [OH^-]$$

   For any equimolar mixture of a base and its conjugate acid, $[OH^-] = K_b$, from which it follows $pOH = pK_b$.

   $pOH = -\log(1.7 \times 10^{-9}) = 8.77$

   $pH = 14.00 - pH = 5.23$