

CHM 111 Acid/Base Quiz 2

Name: SOLUTIONS

Due: 4/4/08, 11:15 AM

Solve the following. Round pH values to the nearest 0.01 pH units (ex: pH = 2.34) [20]

1) Calculate the pH of a buffer solution that contains 0.25 M ammonium chloride mixed with 0.20 M ammonia. $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3$; $\text{p}K_{\text{a}, \text{NH}_4^+} = 9.26$

$$\text{pH} = 9.26 + \log\left(\frac{0.20 \text{ M}}{0.25 \text{ M}}\right) = \boxed{9.16}$$

Use the Henderson-Hasselbalch equation!

$$\begin{aligned} \text{p}K_{\text{a}} + \text{p}K_{\text{b}} &= 14; \quad K_{\text{b}, \text{NH}_3} = 1.8 \times 10^{-5}, \text{ so} \\ \text{p}K_{\text{b}} &= 4.74 \end{aligned}$$

2) Calculate the pH of the solution that results when you mix 150. mL of the buffer above (0.25 M ammonium chloride mixed with 0.20 M ammonia) with 5.0 mL of 0.25 M hydrochloric acid.



The HCl converts some ammonia to ammonium ion!

Initially,

$$\text{NH}_3: 0.20 \text{ M} \times 150 \text{ mL} = 30 \text{ mmol}$$

$$\text{NH}_4^+: 0.25 \text{ M} \times 150 \text{ mL} = 37.5 \text{ mmol}$$

$$\text{HCl}: 0.25 \text{ M} \times 5 \text{ mL} = 1.25 \text{ mmol}$$

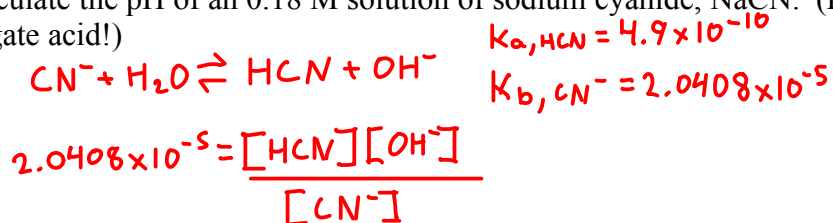
After HCl reaction:

$$\text{NH}_3: 30 \text{ mmol} - 1.25 \text{ mmol} = 28.75 \text{ mmol} \quad \frac{28.75 \text{ mmol}}{155 \text{ mL}} = 0.19 \text{ M}$$

$$\text{NH}_4^+: 37.5 \text{ mmol} + 1.25 \text{ mmol} = 38.75 \text{ mmol} \quad \frac{38.75 \text{ mmol}}{155 \text{ mL}} = 0.25 \text{ M}$$

$$\text{pH} = 9.26 + \log\left(\frac{0.19 \text{ M}}{0.25 \text{ M}}\right) = \boxed{9.13}$$

3) Calculate the pH of an 0.18 M solution of sodium cyanide, NaCN. (Hint: HCN is the conjugate acid!)



$$\frac{(x)(x)}{0.18 - x} = 2.0408 \times 10^{-5}$$

$$x = 1.9065 \times 10^{-3} = [\text{OH}^-]$$

$$\text{pOH} = 2.72, \quad \boxed{\text{pH} = 11.28}$$

4) Classify these salts as *acidic*, *basic*, or *neutral* in water.

• basic Na_2CO_3

• neutral NaNO_3

• acidic $\text{CH}_3\text{NH}_2\text{Cl}$ ↳ should be $\text{CH}_3\text{NH}_3\text{Cl}$!

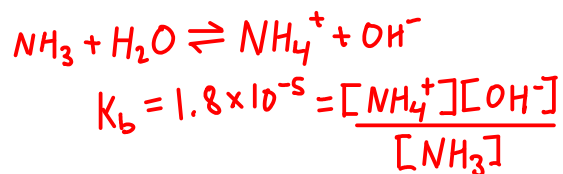
• neutral KCl

• acidic NH_4NO_3

5) Calculate the pH of the solution that results when you add 25.0 mL of water to 75.0 mL of 0.18 M NH_3 .

$$M_1 V_1 = M_2 V_2 \quad (75.0 \text{ mL})(0.18 \text{ M}) = (100.0 \text{ mL})M_2$$

$$M_2 = 0.135 \text{ M } \text{NH}_3$$



$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.135 - x}$$

$$x = 1.51 \times 10^{-3} = [\text{OH}^-]$$

$$\text{pOH} = 2.81, \quad \boxed{\text{pH} = 11.19}$$

First, figure out the ammonia concentration. Then, it's just another weak base problem!