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. $NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH_3$; $pK_{a, NH_4^+} = 9.26$

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

Use the Henderson-Hasselbalch equation!

$pK_a + pK_b = 14$; $K_{b, NH_3} = 1.8 \times 10^{-5}$, so
$pK_b = 4.74$

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.

buffer: $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$; $pK_a = 9.26$

on addition of HCl, $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$

The HCl converts some ammonia to ammonium ion!

Initially,

NH_3: 0.20 M × 150 mL = 30 mmol

NH_4^+: 0.25 M × 150 mL = 37.5 mmol

HCl: 0.25 M × 5 mL = 1.25 mmol

After HCl reaction:

NH_3: 30 mmol − 1.25 mmol = 28.75 mmol $\quad \frac{28.75\text{ mmol}}{155\text{ mL}} = 0.19\,M$

NH_4^+: 37.5 mmol + 1.25 mmol = 38.75 mmol $\quad \frac{38.75\text{ mmol}}{155\text{ mL}} = 0.25\,M$

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

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

$K_{a,HCN} = 4.9\times10^{-10}$

$CN^- + H_2O \rightleftharpoons HCN + OH^-$

$K_{b,CN^-} = 2.0408\times10^{-5}$

$$2.0408\times10^{-5} = \frac{[HCN][OH^-]}{[CN^-]}$$

$$\frac{(x)(x)}{0.18-x} = 2.0408\times10^{-5}$$

$$x = 1.9065\times10^{-3} = [OH^-]$$

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

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

- basic Na_2CO_3
- neutral $NaNO_3$
- acidic CH_3NH_2Cl ← should be CH_3NH_3Cl!
- 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.

$M_1V_1 = M_2V_2$ $(75.0\ mL)(0.18\ M) = (100.0\ mL)M_2$

$M_2 = 0.135\ M\ NH_3$

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

$$K_b = 1.8\times10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8\times10^{-5} = \frac{(x)(x)}{0.135-x}$$

$$x = 1.51\times10^{-3} = [OH^-]$$

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

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