- Ch. 16.1-16.2: Solution Equilibria; Common Ion Effect
- **Homework 16-1:** Problems pg. 722 #2, 4, 5, 6

2 **Describe the effect of pH (increase, decrease, or no change) chat results from each of the following additions:**

(a) Adding CH_3COO^- to CH_3COOH will reduce the acidity and **increase** the pH.
(b) Adding NH_4^+ to NH_3 will reduce the alkalinity and **decrease** the pH.
(c) Adding $HCOO^-$ to HCOOH will reduce the acidity and **increase** the pH.
(d) Adding KCl to HCl will have **no effect** on the pH (HCl is a strong acid).
(e) Adding BaI_2 to HI will have **no effect** on the pH (HI is a strong acid).

4 **The pK_as of two monoprotic acids HA and HB are 5.9 and 8.1, respectively. Which of the two is the stronger acid?**

Since $pK_a = -\log(K_a)$, the lower the pK_a the stronger the acid. HA has a lower pK_a than HB so it is stronger.

5 **(a) Determine the pH of a 0.40 *M* CH_3COOH solution.**

This is a weak acid problem. Setting up the standard equilibrium table:

	$CH_3COOH(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$CH_3COO^-(aq)$
Initial (*M*):	0.40		0.00		0.00
Change (*M*):	$-x$		$+x$		$+x$
Equilibrium (*M*):	$(0.40 - x)$		x		x

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x^2}{(0.40 - x)} \approx \frac{x^2}{0.40} = 1.8 \times 10^{-5}$$

$$x = [H^+] = 2.7 \times 10^{-3}\ M; \quad \mathbf{pH} = \mathbf{2.57}$$

(b) Determine the pH of a solution that is 0.40 *M* CH_3COOH and 0.20 *M* CH_3COONa.

In addition to the acetate ion formed from the ionization of acetic acid, we also have acetate ion formed from the sodium acetate dissolving.

$$CH_3COONa(aq) \rightarrow CH_3COO^-(aq) + Na^+(aq)$$

Dissolving 0.20 *M* sodium acetate initially produces 0.20 *M* CH_3COO^- and 0.20 *M* Na^+. The sodium ions are not involved in any further equilibrium (why?), but the acetate ions must be added to the equilibrium in part (a).

	$CH_3COOH(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$CH_3COO^-(aq)$
Initial (*M*):	0.40		0.00		0.20
Change (*M*):	$-x$		$+x$		$+x$
Equilibrium (*M*):	$(0.40 - x)$		x		$(0.20 + x)$

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \frac{(x)(0.20 + x)}{(0.40 - x)} \approx \frac{x(0.20)}{0.40} = 1.8 \times 10^{-5}$$

$$x = [H^+] = 3.6 \times 10^{-5}\ M; \ \mathbf{pH} = \mathbf{4.44}$$

Could you have predicted whether the pH should have increased or decreased after the addition of the sodium acetate to the pure 0.40 *M* acetic acid in part (a)?
An alternate way to work part (b) of this problem is to use the Henderson-Hasselbalch equation.

$$\mathbf{pH} = p\boldsymbol{K}_a + \log\frac{[\mathrm{A}^-]}{[\mathrm{HA}]} = -\log(1.8\times10^{-5}) + \log\frac{0.20\ M}{0.40\ M} = 4.74 - 0.30 = \mathbf{4.44}$$

6 (a) Determine the pH of a 0.20 *M* NH_3 solution:

This is a weak base calculation.

	$NH_3(aq) + H_2O(l)$	$\rightleftharpoons$ $NH_4^+(aq)$	$+\ OH^-(aq)$
Initial (*M*):	0.20	0	0
Change (*M*):	$-x$	$+x$	$+x$
Equilibrium (*M*):	$0.20 - x$	x	x

$$\boldsymbol{K}_b = \frac{[\mathrm{NH_4^+}][\mathrm{OH^-}]}{[\mathrm{NH_3}]} = \frac{(x)(x)}{0.20 - x} \approx \frac{\boldsymbol{x}^2}{0.20} = 1.8\times10^{-5}$$

$$x = 1.9\times10^{-3}\ M = [\mathrm{OH^-}];\ \ \mathrm{pOH} = 2.72;\ \ \mathbf{pH = 11.28}$$

(b) Determine the pH of a solution that is 0.20 *M* in NH_3 and 0.30 *M* NH_4Cl.

The initial concentration of NH_4^+ is 0.30 *M* from the salt NH_4Cl. We set up a table as in part (a).

	$NH_3(aq) + H_2O(l)$	$\rightleftharpoons$ $NH_4^+(aq)$	$+\ OH^-(aq)$
Initial (*M*):	0.20	0.30	0
Change (*M*):	$-x$	$+x$	$+x$
Equilibrium (*M*):	$0.20 - x$	$0.30 + x$	x

$$\boldsymbol{K}_b = \frac{[\mathrm{NH_4^+}][\mathrm{OH^-}]}{[\mathrm{NH_3}]} = \frac{(x)(0.30 + x)}{0.20 - x} \approx \frac{\boldsymbol{x}(0.30)}{0.20} = 1.8\times10^{-5}$$

$$x = 1.2\times10^{-5}\ M = [\mathrm{OH^-}];\ \mathrm{pOH} = 4.92;\ \mathbf{pH = 9.08}$$

Alternatively, we could use the Henderson-Hasselbalch equation to solve this problem. Table 15.4 gives the value of K_b for ammonia. Substituting into the Henderson-Hasselbalch equation gives:

$$K_a = \frac{K_w}{K_b} = \frac{1\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$$

$$\mathbf{pH} = p\boldsymbol{K}_a + \log\frac{[\mathrm{NH_3}]}{[\mathrm{NH_4^+}]} = -\log(5.6\times10^{-10}) + \log\frac{(0.20)}{(0.30)} = 9.25 + (-0.18) = \mathbf{9.07}$$

Notice that a higher concentration of acid will reduce the pH, as expected for a more acidic solution.