1. Find the pH of a solution that is 0.195 M $HC_2H_3O_2$ and 0.125 M $NaC_2H_3O_2$, using both an ICE chart and the Henderson-Hasselbalch equation. (For $HC_2H_3O_2$, $K_a = 1.8 \times 10^{-5}$).

	$HC_2H_3O_2$	+	H_2O	$\rightleftharpoons$	H_3O^+	+	$C_2H_3O_2^-$
I	0.195 M		–		0		0.125 M
C	-x		–		+x		+x
E	0.195 - x		–		x		0.125 + x

$K_a = 1.8 \times 10^{-5} = \frac{x(0.125+x)}{0.195-x} \approx \frac{x(0.125)}{(0.195)}$

$x = 2.81 \times 10^{-5}\ M = [H_3O^+]$

$x/0.125 = 0.02\%$, or $x/0.195 = 0.01\%$, so clearly assumption was good.

$pH = 4.55$

Henderson-Hasselbalch: $pH = pK_a + log\left(\frac{[A^-]}{[HA]}\right) = 4.74 + log\left(\frac{0.125}{0.195}\right) = 4.55$

2. Find the pH of a solution that is 0.255 M CH_3NH_2 and 0.135 M CH_3NH_3Br, using both an ICE chart and the Henderson-Hasselbalch equation. (For CH_3NH_2, $K_b = 4.4 \times 10^{-4}$).

	CH_3NH_2	+	H_2O	$\rightleftharpoons$	$CH_3NH_3^+$	+	OH^-
I	0.225 M		–		0.135 M		0
C	-x		–		+x		+x
E	0.225 - x		–		0.135 + x		x

$K_b = \frac{x(0.135+x)}{0.225-x} \approx \frac{x(0.135)}{0.225} \Rightarrow x = 8.31 \times 10^{-4}$ M (or 7.27×10^{-4} M if solving quadratic equation)

$pOH = 3.08 \Rightarrow pH = 10.92$ (or 3.14 and 10.86, respectively, when using quadratic equation)

Could also set up as the acid equilibrium and you should end up with the same answer.

Or: $pH = pK_a + \log\left(\frac{[B]}{[BH^+]}\right) = 10.64 + \log\left(\frac{0.255}{0.135}\right) = 10.92$
Remember: $pK_a + pK_b = pK_w = 14$

3. For the acetic acid / acetate buffer in Problem 1, write the pertinent chemical reactions that would occur if a strong acid were added and if a strong base were added.

For strong acid addition: $C_2H_3O_2^-(aq) + H_3O^+(aq) \longrightarrow HC_2H_3O_2(aq) + H_2O(\ell)$

For strong base addition: $HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow C_2H_3O_2^-(aq) + H_2O(\ell)$

4. Calculate the molar ratio and the mass ratio of NaF to HF required to create a buffer with pH = 4.00. (For HF, $K_a = 3.4 \times 10^{-4}$)

   $pH = pK_a + log\left(\frac{[F^-]}{[HF]}\right) \Rightarrow 4.00 = 3.47 + log\left(\frac{[F^-]}{[HF]}\right)$

   $log\left(\frac{[F^-]}{[HF]}\right) = 0.53 \Rightarrow \frac{[F^-]}{[HF]} = 10^{0.53} = 3.39 = \frac{moles\ F^-}{moles\ HF} = \frac{moles\ NaF}{moles\ HF}$ (molar ratio).

   Molar mass of NaF = 41.99 g/mol, molar mass of HF = 20.01 g/mol.

   $\left(\frac{3.39\ mol\ NaF}{1\ mol\ HF}\right)\left(\frac{41.99\ g/molNaF}{20.01\ g/molHF}\right) = 7.11\frac{g\ NaF}{g\ HF}$

5. Determine whether or not the mixing of each pair of solutions results in a buffer.

   (a) 75.0 mL of 0.10 M HF; 55.0 mL of 0.15 M NaF $\Rightarrow$ Buffer

   (b) 150.0 mL of 0.10 M HF; 135.0 mL of 0.175 M HCl $\Rightarrow$ Not a buffer

   (c) 125.0 mL of 0.15 M CH_3NH_2; 120.0 mL of 0.25 M CH_3NH_3Cl $\Rightarrow$ Buffer

6. A 100.0 mL buffer solution is 0.175 M HClO ($K_a = 2.9 \times 10^{-8}$) and 0.150 M in NaClO.

   (a) What is the initial pH of this solution?
   (b) What is the pH after addition of 0.00185 moles of HBr?
   (c) What is the pH after addition of 0.00213 moles of NaOH?

   Initial pH: $pH = pK_a + log\left(\frac{[ClO^-]}{[HClO]}\right) = 7.54 + log\left(\frac{0.150}{0.175}\right) = 7.47$

   After addtion of 0.00185 moles of strong acid (HBr): Start with 0.0150 moles of ClO^-, 0.00185 moles will react with H_3O^+, leaving 0.01315 mol ClO^-. Start with 0.0175 moles of HClO, add 0.00185 moles, leaving 0.01935 mol HClO.

   Then: $pH = 7.54 + log\left(\frac{0.01315\ mol\ ClO^-/0.100\ L}{0.01935\ mol\ HClO/0.100\ L}\right) = 7.37$

   After addition of 0.00213 moles of strong base (NaOH): Start with 0.0175 mol HClO. 0.00213 mol will react with OH^-, leaving 0.01537 mol HClO. Start with 0.0150 mol ClO^- and produce an additional 0.00213 moles during the reaction, resulting in 0.01713 mol ClO^-.

   Then: $pH = 7.54 + log\left(\frac{0.01713\ mol\ ClO^-/0.100\ L}{0.01537\ mol\ HClO/0.100\ L}\right) = 7.59$