17.1 Each box contains 4 molecules of HX and a different amount of X^-. The box with the greatest concentration of X^- will have the smallest concentration of free H^+, meaning box 2 will have the least free H^+ and therefore the highest pH.

17.2 a) Referencing Figure 16.7, the solution on the right has a pH in the range of 3. The solution on the left has a pH in the range of 4.5.

b) The beaker on the left, which contains a buffer created be a weak acid and a conjugate base salt, will be better able to maintain its pH when small amounts of NaOH are added.

17.3 In a buffer, if the concentrations of HX and X^- are equal, pH equals pK_a. If this were the case, given the pKa, the pH of the buffer would be 4.5. Since the pH of the buffer is 4.3 (more acidic) the concentration of conjugate base ion must be greater (along with the concentration of H^+) than the concentration of the weak acid. As such $[HX]<[X^-]$. If this doesn't make sense, remember the basis of the common ion effect.

17.4 a) Drawing 3 represents the buffer after the addition of strong acid. The additional H^+ from the strong acid have converted X^- ions into HX.

b) Drawing 1 represents the buffer after the addition of strong base. The additional OH^- from the strong base have converted HX into X^-.

c) Drawing 2 represents a situation that cannot arise from the addition of with an acid or a base. Both [HX] and [X^-] cannot both decrease with the addition of an acid or base.

17.14

a) The equilibrium constant for the reaction will not change

b) Adding a strong electrolyte salt of HB^+ will create HB^+ ion which will shift this equilibrium to the left, decreasing the concentration of B.

c) As the equilibrium shifts to the left, the concentration of OH- will also decrease, decreasing the pH of the solution.

17.15a)

	C_2H_5COOH	$\rightleftharpoons$	$C_2H_5COO^-$	+	H^+
I	0.085M		0.060 M		0 M
C	-x		+x		+x
E	0.085M - x		0.060 + x		x

$$K_a = \frac{[C_2H_5COO^-][H^+]}{[C_2H_5COOH]} \quad 1.3 \times 10^{-5} = \frac{(0.060 + x)x}{(0.085 - x)} \approx \frac{0.060x}{0.085}$$

$x = 1.84 \times 10^{-5}$

$pH = -\log[H^+]$ $\quad pH = -\log[1.84 \times 10^{-5}]$ $\quad pH = 4.73$

17.15b)

	$(CH_3)_3N$	+	H_2O	$\rightleftharpoons$	$(CH_3)_3NH^+$	+	OH^-
I	0.075M		———		0.10 M		0 M
C	-x		———		+x		+x
E	0.075 - x		———		0.10 + x		x

$$K_b = \frac{[C_5H_5NH^+][OH^-]}{[(CH_3)_3N]} \quad 6.4 \times 10^{-5} = \frac{(0.10+x)x}{0.075 - x} \approx \frac{0.10x}{0.075}$$

$x = 4.8 \times 10^{-5}$

$pOH = -\log[OH^-]$ $\quad pOH = -\log[4.8 \times 10^{-5}]$ $\quad pOH = 4.32$

$pH + pOH = 14$ $\quad pH = 14 - pOH$ $\quad pH = 14 - 4.32 = 9.68$

17.17a)

	buCOOH	$\rightleftharpoons$	$buCOO^-$	+	H^+
I	0.0075M		0 M		0 M
C	-x		+x		+x
E	0.0075M - x		x		x

$$K_a = \frac{[buCOO^-][H^+]}{[buCOOH]} \quad 1.5 \times 10^{-5} = \frac{x^2}{(0.0075 - x)} \approx \frac{x^2}{0.0075}$$

$$x = 3.35 \times 10^{-4}$$

$$\%\ \text{Ion} = \frac{[H^+]}{[HA_{initial}]} = \frac{[3.35 \times 10^{-4}]}{[0.0075M]} = 4.47\%$$

17.17b)

	$buCOOH$	$\rightleftharpoons$	$buCOO^-$	+	H^+
I	0.0075M		0.085 M		0 M
C	-x		+x		+x
E	0.0075M - x		0.085 + x		x

$$K_a = \frac{[buCOO^-][H^+]}{[buCOOH]} \quad 1.5 \text{ x } 10^{-5} = \frac{(0.085 + x)x}{(0.0075 - x)} \approx \frac{0.085x}{0.0075}$$

$$x = 1.32 \text{ x } 10^{-6}$$

$$\% \text{ Ion} = \frac{[H^+]}{[HA_{initial}]} = \frac{[1.32 \text{ x } 10^{-6}]}{[0.0075M]} = 0.018\%$$

17.19 A solution of CH_3COOH and $NaCH_3COO$ act as a buffer because the solution contains elevated concentrations of both weak acid and conjugate base ions. As a result, the solution can resist pH changes from the addition of both a base and an acid. In contrast, a solution of HCl (a strong acid) and NaCl will not contain an elevated concentration of unprotonated acid and the conjugate base Cl^- does not have the ability to react with acids.

17.21b

$$\frac{85\text{mL}}{} \Bigg| \frac{1\text{ liter}}{1000\text{ mL}} \Bigg| \frac{0.13\text{ mol lactic}}{1\text{ liter}} = 0.011\text{ mol}/.180\text{L} = 0.061\text{M Lactic}$$

$$\frac{95\text{mL}}{} \Bigg| \frac{1\text{ liter}}{1000\text{ mL}} \Bigg| \frac{0.15\text{ mol lac}^-}{1\text{ liter}} = 0.014\text{ mol}/.180\text{L} = 0.078\text{M Lac}^-$$

	Lactic	$\rightleftharpoons$	Lac^-	+	H^+
I	0.061M		0.078 M		0 M
C	-x		+x		+x
E	0.061M - x		0.078 + x		x

$$K_a = \frac{[Lac^-][H^+]}{[Lactic]} \quad 1.4 \times 10^{-4} = \frac{(0.078 + x)x}{(0.061 - x)} \approx \frac{0.078x}{0.061}$$

$x = 1.1 \times 10^{-4}$

$pH = -\log[H^+] \quad pH = -\log[1.1 \times 10^{-4}] \quad pH = 3.96$

17.22a

	HCO_3^-	$\rightleftharpoons$	CO_3^{2-}	+	H^+
I	0.105		0.125 M		0 M
C	-x		+x		+x
E	0.105M - x		0.125 + x		x

$$K_a = \frac{[CO_3^{2-}][H^+]}{[HCO_3^-]} \quad 5.6 \times 10^{-11} = \frac{(0.125 + x)x}{(0.105 - x)} \approx \frac{0.125x}{0.105}$$

$$x = 4.7 \times 10^{-11}$$

$$pH = -\log[H^+] \qquad pH = -\log[4.7 \times 10^{-11}] \qquad pH = 10.33$$