AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART E,

Day 87:

- Buffers

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1. A certain buffer is made by dissolving sodium bicarbonate, $NaHCO_3$, and sodium carbonate, Na_2CO_3 , in water. Write the necessary **EQUATIONS** to show how this buffer behaves when H^{+1} or OH^{-1} is added to the solution (these take place in separate containers). You will not have both H^{+1} and OH^{-1} in the same equation (except for water).

$$NaHCO_3 \rightarrow Na^{+1} + HCO_3^{-1} \quad ; \quad Na_2CO_3 \rightarrow 2\ Na^{+1} + CO_3^{-2}$$

$$HCO_3^{-1} \rightleftharpoons H^{+1} + CO_3^{-2}$$

The above reaction is the buffered solution … at equilibrium

$$\textbf{Add } H^{+1}\textbf{:} \qquad HCO_3^{-1} \leftarrow H^{+1} + CO_3^{-2}$$

when hydrogen ions are added to the solution the reaction forms more HCO_3^{-1}.

$$\textbf{Add } OH^{-1}\textbf{:} \qquad CO_3^{-2} + H_2O \leftarrow OH^{-1} + HCO_3^{-1}$$

when hydroxide ions are added to the solution the reaction forms water and the conjugate base CO_3^{-2}.

2. Calculate the pH of each of the following solutions.

a) 0.100 *M* propanoic acid ($HC_3H_5O_2$, $K_a = 1.3 \times 10^{-5}$)

$$HC_3H_5O_2 \rightleftharpoons H^{+1} + C_3H_5O_2^{-1}$$

	[$HC_3H_5O_2$]	$\rightleftharpoons$	[H^{+1}]	+	[$C_3H_5O_2^{-1}$]
I	0.100 *M*		0		0
C	- *x*		+ *x*		+ *x*
E	0.100 - *x*		*x*		*x*

$$K_a = \frac{\left[\ H^{+1}\ \right]\left[\ C_3H_5O_2^{-1}\ \right]}{\left[\ HC_3H_5O_2\ \right]} = 1.3\times10^{-5} \ \ ; \ \ x^2 = \left(\ 1.3\times10^{-5}\ \right)\left(\ 0.100\ \right)$$

$$x = \sqrt{1.3\times10^{-6}} = 1.14\times10^{-3} \ \ ; \ \ \textbf{pH} = -\log\left(1.14 \times 10^{-3}\right) = \mathbf{2.94}$$

b) 0.100 *M* sodium propanate, $NaC_3H_5O_2$ ***$NaC_3H_5O_2 \rightarrow Na^{+1} + C_3H_5O_2^{-1}$ (weak base)***

$$C_3H_5O_2^{-1} + H_2O \rightleftharpoons OH^{-1} + HC_3H_5O_2$$

	[$C_3H_5O_2^{-1}$]	+	[H_2O]	$\rightleftharpoons$	[OH^{-1}]	+	[$HC_3H_5O_2$]
I	0.100 *M*		-		0		0
C	- *x*		-		+ *x*		+ *x*
E	0.100 - *x*		-		*x*		*x*

$$K_b = \frac{K_w}{K_a} = \frac{1.00\times10^{-14}}{1.3\times10^{-5}} = 7.69\times10^{-10}$$

$$K_b = \frac{\left[\ OH^{-1}\ \right]\left[\ HC_3H_5O_2\ \right]}{\left[\ C_3H_5O_2^{-1}\ \right]} \ \ ; \ \ x^2 = \left(\ 7.69\times10^{-10}\ \right)\left(\ 0.100\ \right)$$

$$x = \sqrt{7.69\times10^{-11}} = 8.77\times10^{-6}M \ \ ; \ \ \text{pOH} = -\log\left(8.77 \times 10^{-6}\right) = 5.06 \ ; \ \ \textbf{pH} = 14 - 5.06 = \mathbf{8.94}$$

c) pure water, H_2O

$$[\ H^{+1}\ ] = [\ OH^{-1}\ ] = 1.0 \times 10^{-7};\quad \mathbf{pH} = -\log(\ 1.0 \times 10^{-7}\ ) = \mathbf{7.00}$$

d) a mixture containing 0.100 *M* propanoic acid and 0.100 *M* sodium propanate.

$$HC_3H_5O_2 \rightleftharpoons H^{+1} + C_3H_5O_2^{-1}$$

	[$HC_3H_5O_2$]	$\rightleftharpoons$	[H^{+1}]	+	[$C_3H_5O_2^{-1}$]
I	0.100 *M*		~ 0		0.100 *M*
C	- *x*		+ *x*		+ *x*
E	0.100 - *x*		*x*		0.100 + *x*

$$K_a = 1.3\times10^{-5} = \frac{[\ H^{+1}\ ][\ C_3H_5O_2^{-1}\ ]}{[\ HC_3H_5O_2\ ]} = \frac{(x)(0.100+x)}{0.100-x} = \frac{(x)(0.100)}{0.100}$$

$$x = \frac{(\ 1.3\times10^{-5}\ )(0.100)}{0.100} = 1.3\times10^{-5} = [\ H^{+1}\ ] \qquad ; \qquad \mathbf{pH} = -\log(\ 1.3 \times 10^{-5}\ ) = \mathbf{4.89}$$

OR

$$pH = pK_a + \log\frac{[A^{-1}]}{[HA]};\quad pK_a = -\log(1.3 \times 10^{-5}) = 4.89\ ;\quad pH = 4.89 + \log\frac{[0.100]}{[0.100]} = 4.89$$

3. Calculate the pH after 0.020 mol HCl is added to 1.00 liter of each of the items in question #2.

a)

$$HC_3H_5O_2 \rightleftharpoons C_3H_5O_2^{-1} + H^{+1}$$

	[$HC_3H_5O_2$]	$\rightleftharpoons$	[$C_3H_5O_2^{-1}$]	+	[H^{+1}]
I	0.100 *M*		1.14 x 10^{-3}		0.020 *M*
C	+ 1.14 x 10^{-3}		- 1.14 x 10^{-3}		- 1.14 x 10^{-3}
E	0.100 + 1.14 x 10^{-3}		~ 0		0.01886

$$\mathbf{pH} = -\log(\ 0.01886\ ) = \mathbf{1.72}$$

Either answer is correct – AS LONG AS YOU SHOW ALL YOUR WORK !!!

b)

$$H^{+1} + C_3H_5O_2^{-1} \rightleftharpoons HC_3H_5O_2$$

	[$HC_3H_5O_2$]	←	[$C_3H_5O_2^{-1}$]	+	[H^{+1}]
I	~ 0		0.100 *M*		0.020 *M*
C	+ 0.020		- 0.020		- 0.020
E	0.020 *M*		0.080 *M*		~ 0

$$pH = pK_a + \log\frac{[A^{-1}]}{[HA]};\ \ pK_a = -\log(1.3\times10^{-5}) = 4.89;\ \ pH = 4.89 + \log\frac{[0.080]}{[0.020]} = 5.49$$

OR

$$K_a = 1.3\times10^{-5} = \frac{[\ H^{+1}\ ][\ C_3H_5O_2^{-1}\ ]}{[\ HC_3H_5O_2\ ]} = \frac{[H^{+1}](0.080)}{0.020}$$

$$[\ H^{+1}\ ] = \frac{(\ 1.3\times10^{-5}\ )(\ 0.020\ )}{0.080} = 3.25\times10^{-6}$$

pH = - log (3.25 x 10^{-6}) = **5.49**

c)

This a strong acid problem, [H^{+1}] = 0.020 *M* ; **pH** = - log (0.020) = **1.70**

d)

$$H^{+1} + C_3H_5O_2^{-1} \rightleftharpoons HC_3H_5O_2$$

	[$HC_3H_5O_2$]	←	[$C_3H_5O_2^{-1}$]	+	[H^{+1}]
I	0.100 *M*		0.100 *M*		0.020 *M*
C	+ 0.020		- 0.020		- 0.020
E	0.120		0.080		~ 0

$$pH = pK_a + \log\frac{[A^{-1}]}{[HA]}$$

$$pH = \left(-\log\ 1.3\times10^{-5}\right) + \log\frac{[\ 0.080\ ]}{[\ 0.120\ ]} = 4.71$$

4. Which of the solutions in question #3 shows the least change in pH upon the addition of the acid? Explain.

The solution in part "d" is a buffer solution, it contains both a weak acid ($HC_3H_5O_2$) and a its conjugate ion (common ion) = (weak base, $C_3H_5O_2^{-1}$). Solution "d" shows the greatest resistance to changes in pH when either a strong acid or strong base is added, which is the primary property of buffers.