AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART E, EXAMPLES Day 87:

- Buffers

Answers:

Example #1, part 1: A buffered solution (at equilibrium) contains 0.75 *M* acetic acid ($HC_2H_3O_2$, K_a = 1.8 x 10^{-5}) and 0.75 *M* sodium acetate, $NaC_2H_3O_2$. Calculate the pH of this solution.

	[$HC_2H_3O_2$]	$\rightleftharpoons$	[$C_2H_3O_2^{-1}$]	+	[H^{+1}]
I	0.75 *M*		0.75 *M*		~ 0
C	- *x*		+ *x*		+ *x*
E	0.75 - *x*		0.75 + *x*		*x*

$$\frac{\left[\ C_2H_3O_2^{-1}\right]\left[H^{+1}\right]}{\left[\ HC_2H_3O_2\ \right]} = \frac{(\ 0.75\ -\ x\ )(\ x\ )}{0.75\ +\ x} = \frac{(\ 0.75\ )(\ x\ )}{0.75} = 1.8\times10^{-5}$$

$$x = \frac{(\ 0.75\ )\left(\ 1.8\times10^{-5}\ \right)}{0.75} = \left[\ H^{+}\ \right]$$

$$pH = -\log\ 1.8\times10^{-5} = 4.74$$

Example #1, part 2: Calculate the change in pH that occurs when 0.010 mol of gaseous hydrochloric acid, HCl, is added to 1.0 liter of the buffered solution (from part 1). Compare the above pH with the pH when the hydrochloric acid is added.

$$0.010\ M\ \text{HCl} \rightarrow 0.010\ M\ \text{H}^{+1} + 0.010\ M\ \text{Cl}^{-1}$$

	[$HC_2H_3O_2$]	$\leftarrow$	[$C_2H_3O_2^{-1}$]	+	[H^{+1}]
I	0.75 *M*		0.75 *M*		0.010
C	+ 0.010		- 0.010		- 0.010
E	0.75 + 0.010 = 0.760		0.75 - 0.010 = 0.740		~ 0

$$\frac{\left[\ C_2H_3O_2^{-1}\right]\left[H^{+1}\right]}{\left[\ HC_2H_3O_2\ \right]} = \frac{(\ 0.740\ )\left[H^{+1}\right]}{0.760} = 1.8\times10^{-5}$$

$$\frac{(\ 0.760\ )\left(\ 1.8\times10^{-5}\right)}{0.740} = \left[H^{+1}\right] = 1.85\times10^{-5}$$

$$pH = -\log\ 1.85\times10^{-5} = 4.73$$

Change is – 0.01 in pH

The pH of 0.010 mol of HCl (in this solution) would equal 2.00 had the buffer not been in place.

The pH would have changed MUCH more than 0.01.