

# AP CHEMISTRY

## TOPIC 7: ACIDS & BASES, PART E,

## EXAMPLES, PART III

Day 87:

- Buffers

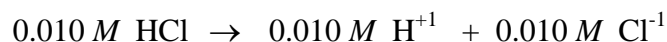
Let's try one from the examples from part I

Example #1, part 1: A buffered solution (at equilibrium) contains 0.75 M acetic acid ( $\text{HC}_2\text{H}_3\text{O}_2$ ,  $K_a = 1.8 \times 10^{-5}$ ) and 0.75 M sodium acetate,  $\text{NaC}_2\text{H}_3\text{O}_2$ . Calculate the pH of this solution.

	$[\text{HC}_2\text{H}_3\text{O}_2]$	$\rightleftharpoons$	$[\text{C}_2\text{H}_3\text{O}_2^{-1}]$	+	$[\text{H}^{+1}]$
<b>I</b>	0.75 M		0.75 M		$\sim 0$
<b>C</b>	-x		+x		+x
<b>E</b>	0.75 M - x		0.75 M + x		x

$$pH = pK_a + \log \frac{[A^{-1}]}{[HA]} = -\log(1.8 \times 10^{-5}) + \log \frac{[0.75]}{[0.75]} = 4.74 + \log(1) = 4.74$$

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.



	$[\text{HC}_2\text{H}_3\text{O}_2]$	$\leftarrow$	$[\text{C}_2\text{H}_3\text{O}_2^{-1}]$	+	$[\text{H}^{+1}]$
<b>I</b>	0.75 M		0.75 M		0.010
<b>C</b>	+0.010		-0.010		-0.010
<b>E</b>	0.75 + 0.010 = 0.760		0.75 - 0.010 = 0.740		$\sim 0$

$$pH = pK_a + \log \frac{[A^{-1}]}{[HA]} = -\log(1.8 \times 10^{-5}) + \log \frac{[0.74]}{[0.76]} = 4.74 + \log(0.974) = 4.73$$

Keep in mind, you may use the Henderson-Hasselbalch method or the "old fashion" way ...