

AP CHEMISTRY

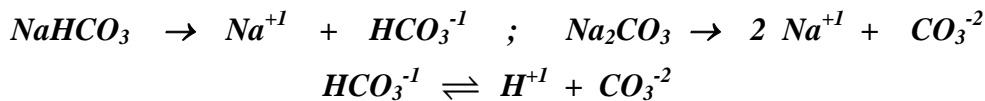
TOPIC 7: ACIDS & BASES, PART E,

Day 87:

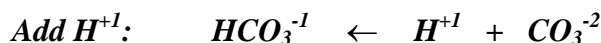
- Buffers

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^+ or OH^- is added to the solution (these take place in separate containers). You will not have both H^+ and OH^- in the same equation (except for water).



The above reaction is the buffered solution ... at equilibrium



when hydrogen ions are added to the solution the reaction forms more 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 ($\text{HC}_3\text{H}_5\text{O}_2$, $K_a = 1.3 \times 10^{-5}$)

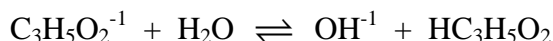


	[$\text{HC}_3\text{H}_5\text{O}_2$]	\rightleftharpoons	[H^+]	+	[$\text{C}_3\text{H}_5\text{O}_2^{-1}$]
I	0.100 M		0		0
C	- x		+ x		+ x
E	0.100 - x		x		x

$$K_a = \frac{[\text{H}^+] [\text{C}_3\text{H}_5\text{O}_2^{-1}]}{[\text{HC}_3\text{H}_5\text{O}_2]} = 1.3 \times 10^{-5} ; \quad x^2 = (1.3 \times 10^{-5}) (0.100)$$

$$x = \sqrt{ 1.3 \times 10^{-6} } = 1.14 \times 10^{-3} ; \quad \text{pH} = - \log (1.14 \times 10^{-3}) = \mathbf{2.94}$$

- b) 0.100 M sodium propanate, $\text{NaC}_3\text{H}_5\text{O}_2$ $\text{NaC}_3\text{H}_5\text{O}_2 \rightarrow \text{Na}^+ + \text{C}_3\text{H}_5\text{O}_2^{-1}$ (weak base)



	[$\text{C}_3\text{H}_5\text{O}_2^{-1}$]	+	[H_2O]	\rightleftharpoons	[OH^-]	+	[$\text{HC}_3\text{H}_5\text{O}_2$]
I	0.100 M		-		0		0
C	- x		-		+ x		+ x
E	0.100 - x		-		x		x

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

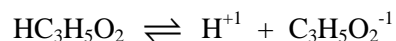
$$K_b = \frac{[\text{OH}^-] [\text{HC}_3\text{H}_5\text{O}_2]}{[\text{C}_3\text{H}_5\text{O}_2^{-1}]} ; \quad x^2 = (7.69 \times 10^{-10}) (0.100)$$

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

c) pure water, H₂O

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

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



	[HC ₃ H ₅ O ₂]	⇌	[H ⁺]	+	[C ₃ H ₅ O ₂ ⁻¹]
I	0.100 M		~ 0		0.100 M
C	- x		+ x		+ x
E	0.100 - x		x		0.100 + x

$$K_a = 1.3 \times 10^{-5} = \frac{[H^{+}][C_3H_5O_2^{-1}]}{[HC_3H_5O_2]} = \frac{(x)(0.100 + x)}{0.100 - x} = \frac{(x)(0.100)}{0.100}$$

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

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)

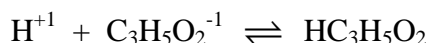


	[HC ₃ H ₅ O ₂]	⇌	[C ₃ H ₅ O ₂ ⁻¹]	+	[H ⁺]
I	0.100 M		1.14 x 10 ⁻³		0.020 M
C	+ 1.14 x 10 ⁻³		- 1.14 x 10 ⁻³		- 1.14 x 10 ⁻³
E	0.100 + 1.14 x 10 ⁻³		~ 0		0.01886

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

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

b)



	[HC ₃ H ₅ O ₂]	←	[C ₃ H ₅ O 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

$$pH = pK_a + \log \frac{[A^{-1}]}{[HA]}; \quad pK_a = -\log (1.3 \times 10^{-5}) = 4.89; \quad pH = 4.89 + \log \frac{[0.080]}{[0.020]} = 5.49$$

OR

$$K_a = 1.3 \times 10^{-5} = \frac{[H^{+1}][C_3H_5O_2^{-1}]}{[HC_3H_5O_2]} = \frac{[H^{+1}](0.080)}{0.020}$$

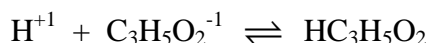
$$[H^{+1}] = \frac{(1.3 \times 10^{-5})(0.020)}{0.080} = 3.25 \times 10^{-6}$$

$$pH = -\log (3.25 \times 10^{-6}) = 5.49$$

c)

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

d)



	[HC ₃ H ₅ O ₂]	←	[C ₃ H ₅ O 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

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

$$pH = (-\log 1.3 \times 10^{-5}) + \log \frac{[0.080]}{[0.120]} = 4.71$$

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 (HC₃H₅O₂) and its conjugate ion (common ion) = (weak base, C₃H₅O₂⁻¹). 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.