

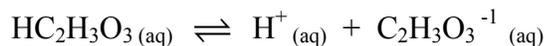
AP CHEMISTRY

TOPIC 7: ACIDS & BASES, EXAMPLES FOR THE REVIEW

Day 88:

- Buffers

CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS.



1. Glycolic acid, $\text{HC}_2\text{H}_3\text{O}_3$, is a monoprotic acid that dissociates in an aqueous solution, as represented in the equation above. Glycolic acid is 0.6179 percent dissociated in 3.85 M $\text{HC}_2\text{H}_3\text{O}_3(\text{aq})$ at 298 K. For parts (a) through (d) below, assume the temperature remains at 298 K.

- (a) Write the expression for the acid-dissociation constant, K_a , for glycolic acid and calculate its value.

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_3^{-1}]}{[\text{HC}_2\text{H}_3\text{O}_3]}$$

$$3.85 \text{ M} \times 0.006179 = 0.0237893 \text{ M} = x$$

	$[\text{HC}_2\text{H}_3\text{O}_3]$	\rightleftharpoons	$[\text{H}^+]$	+	$[\text{C}_2\text{H}_3\text{O}_3^{-1}]$
I	3.85 M		0		0
C	3.85 - 0.0237893		+ 0.0237893		+ 0.0237893
E	3.82621		0.0237893		0.0237893

$$K_a = \frac{(0.0237893)^2}{3.82621} = 1.48 \times 10^{-4}$$

- (b) Calculate the pH of 3.85 M $\text{HC}_2\text{H}_3\text{O}_3$.

$$\text{pH} = -\log [\text{H}^+] = -\log (0.0237893) = 1.63$$

- (c) Calculate the pH of a solution formed by dissolving 1.72 mol of solid sodium glycolate, $\text{NaC}_2\text{H}_3\text{O}_3$. in 850 mL of 3.85 M $\text{HC}_2\text{H}_3\text{O}_3$. Assume the volume changes is negligible.

$$M = \frac{1.72 \text{ mol NaC}_2\text{H}_3\text{O}_3}{0.850 \text{ L}} = 2.02 \text{ M}$$



	$[\text{HC}_2\text{H}_3\text{O}_3]$	\rightleftharpoons	$[\text{H}^{+1}]$	+	$[\text{C}_2\text{H}_3\text{O}_3^{-1}]$
I	3.85 M		~ 0		2.02 M
C	-x		+x		+x
E	3.85 - x		x		2.02 + x

$$K_a = \frac{[\text{H}^{+1}][\text{C}_2\text{H}_3\text{O}_3^{-1}]}{[\text{HC}_2\text{H}_3\text{O}_3]} = \frac{(x)(2.02+x)}{(3.85-x)} = \frac{(x)(2.02)}{(3.85)} = 1.48 \times 10^{-4}$$

$$x = \frac{(1.48 \times 10^{-4})(3.85)}{(2.02)} = 2.82 \times 10^{-4} = [\text{H}^{+1}]$$

$$\text{pH} = -\log [\text{H}^{+1}] = -\log (2.82 \times 10^{-4}) = \mathbf{3.55}$$

OR

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^{-1}]}{[\text{HA}]} = -\log(1.48 \times 10^{-4}) + \log\left(\frac{2.02}{3.85}\right)$$

$$\text{pH} = 3.83 + (-0.280) = 3.55$$

(d) 0.75 mol HCl is added to 850 mL of the solution created in part (c). Calculate the pH of the solution.

$$\frac{0.75 \text{ mol HCl}}{850 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.882 \text{ M HCl}$$

	$[\text{HC}_2\text{H}_3\text{O}_3]$	\rightleftharpoons	$[\text{H}^+]$	+	$[\text{C}_2\text{H}_3\text{O}_3^{-1}]$
I	3.85 M		0.882 M		2.02 M
C	+x = +0.882		-x = -0.882		-x = -0.882
E	3.85 + 0.882 = 4.73		~ 0		2.02 - 0.882 = 1.14

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_3^{-1}]}{[\text{HC}_2\text{H}_3\text{O}_3]} = \frac{[\text{H}^+](1.14)}{(4.73)} = 1.48 \times 10^{-4}$$

$$[\text{H}^+] = \frac{(1.48 \times 10^{-4})(4.73)}{(1.14)} = 6.14 \times 10^{-4}$$

$$\text{pH} = -\log [\text{H}^+] = -\log 6.14 \times 10^{-4} = 3.21$$

OR

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^{-1}]}{[\text{HA}]} ; \text{pH} = \text{p}K_a + \log \frac{[\text{C}_2\text{H}_3\text{O}_3^{-1}]}{[\text{HC}_2\text{H}_3\text{O}_3]}$$

$$\text{pH} = -\log(1.48 \times 10^{-4}) + \log\left(\frac{1.14}{4.73}\right) = 3.21$$

2. (**NEW QUESTION**) An aqueous solution contains dissolved $C_6H_5NH_3Cl$ and $C_6H_5NH_2$. The concentration of $C_6H_5NH_2$ is $0.50 M$ and the pH is 4.20. For $C_6H_5NH_2$, $K_b = 3.8 \times 10^{-10}$.

a) Calculate the concentration of $C_6H_5NH_3^{+1}$ in this buffer solution.



$$pOH = 14 - pH = 14 - 4.20 = \mathbf{9.80} ; \text{anti log}(-9.80) = 1.58 \times 10^{-10}$$

	$[C_6H_5NH_2]$	+	$[H_2O]$	\rightleftharpoons	$[C_6H_5NH_3^{+1}]$	+	$[OH^{-1}]$
I	0.500		-		?		0
C	- x		-		+ x		+ x
E	0.500 - x		-		? + x		1.58×10^{-10}

$$K_b = \frac{[C_6H_5NH_3^{+1}] [OH^{-1}]}{[C_6H_5NH_2]} = 3.8 \times 10^{-10} = \frac{[C_6H_5NH_3^{+1}] (1.58 \times 10^{-10})}{(0.500 - 1.58 \times 10^{-10})} :$$

$$[C_6H_5NH_3^{+1}] = \frac{(3.8 \times 10^{-10})(0.500 - 1.58 \times 10^{-10})}{(1.58 \times 10^{-10})} = 1.20$$

--- OR ---

$$pOH = pK_b + \log \frac{[HB^{+1}]}{[B]} ; 9.80 = pK_a + \log \frac{[HB^{+1}]}{[0.50]}$$

$$9.80 = -\log(3.8 \times 10^{-10}) + \log \frac{[HB^{+1}]}{[0.50]}$$

$$9.80 = 9.42 + \log \frac{[HB^{+1}]}{[0.50]} ; 0.3798 = \log \frac{[HB^{+1}]}{[0.50]}$$

We need to “get rid” of the log function and work with only “numbers” ... Therefore, take the antilog of both sides ...

$$\text{anti log}(-0.380) = \text{anti log} \left(\log \frac{[HB^{+1}]}{[0.50]} \right) ; 2.398 = \frac{[HB^{+1}]}{[0.50]}$$

Now we can solve for HB^{+1} ...

$$[HB^{+1}] = (2.398) \times [0.50] = [1.20]$$

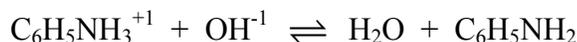
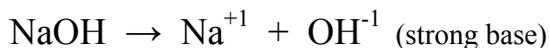
$$[HB^{+1}] = [1.20] = [C_6H_5NH_3^{+1}] = [C_6H_5NH_3Cl]$$

b) Calculate the pH after 4.0 grams of NaOH (s) is added to 1.0 L of this solution (volume change is negligible).



(K_b , from previous problem)

$$\frac{4.0 \text{ g NaOH}}{39.9979 \text{ g}} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol NaOH}} = 0.100 \text{ mol NaOH} \quad ; \quad \frac{0.100 \text{ mol NaOH}}{1.0 \text{ L}} = 0.100 \text{ M NaOH}$$



	[C ₆ H ₅ NH ₂]	+	[H ₂ O]	←	[C ₆ H ₅ NH ₃ ⁺¹]	+	[OH ⁻¹]
I	0.500 M		-		1.20 M		0.100 M
C	+ 0.100		-		1.20 - 0.100		- 0.100
E	0.600		-		1.10		~ 0

$$pOH = pK_b + \log \frac{[HB^{+1}]}{[B]} \quad ; \quad pOH = (-\log 3.8 \times 10^{-10}) + \log \left(\frac{1.10}{0.600} \right) = 9.68$$

$$\mathbf{pH} = 14 - pOH = 14 - 9.68 = \mathbf{4.32}$$

OR

$$K_b = \frac{[\text{C}_6\text{H}_5\text{NH}_3^{+1}] [\text{OH}^{-1}]}{[\text{C}_6\text{H}_5\text{NH}_2]} = 3.8 \times 10^{-10}$$

$$3.8 \times 10^{-10} = \frac{(1.10) [\text{OH}^{-1}]}{(0.600)}$$

$$[\text{OH}^{-1}] = \frac{(3.8 \times 10^{-10})(0.600)}{(1.10)} = 2.07 \times 10^{-10}$$

$$pOH = -\log (2.07 \times 10^{-10}) = 9.68$$

$$\mathbf{pH} = 14 - pOH = 14 - 9.68 = \mathbf{4.32}$$