AP CHEMISTRY
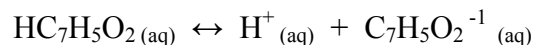
TOPIC 7: ACIDS & BASES, PART E,   PART II

- Buffers

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

$$HC_7H_5O_{2\,(aq)} \leftrightarrow H^+_{\,(aq)} + C_7H_5O_2^{\,-1}_{\,(aq)}$$

1. Benzoic acid, $HC_7H_5O_2$, is a monoprotic acid that dissociates in an aqueous solution, as represented in the equation above. Benzoic acid is 3.29 percent dissociated in 0.85 M $HC_7H_5O_{2\,(aq)}$ 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 benzoic acid and calculate its value.

$$K_a = \frac{\left[H^{+1} \right]\left[C_7H_5O_2^{\,-1} \right]}{\left[HC_7H_5O_2 \right]}$$

0.85 M x 0.0329 = 0.027965 M = x

	[$HC_2H_3O_3$]	$\leftrightarrow$	[H^{+1}]	+	[$C_2H_3O_3^{\,-1}$]
I	0.85 M		0		0
C	0.85 - 0.027965		+ 0.027965		+ 0.027965
E	0.822035		0.027965		0.027965

$$K_a = \frac{\left(0.027965 \right)^2}{0.822035} = 9.51\times10^{-4}$$

(b) Calculate the pH of 0.85 M $HC_7H_5O_2$.

$$\mathbf{pH} = -\log \left[H^{+1} \right] = -\log \left(0.027965 \right) = \mathbf{1.55}$$

(c) Calculate the pH of a solution formed by dissolving 0.067 mol of solid sodium benzonate, $NaC_7H_5O_2$. in 250 mL of 0.85 M $HC_7H_5O_2$. Assume the volume changes is negligible.

$$M = \frac{0.067 \ mol \ NaC_7H_5O_2}{0.250 \ L} = 0.268 \ M \ NaC_7H_5O_2$$

$$NaC_7H_5O_2 \rightarrow Na^{+1} + C_7H_5O_2^{-1} \ (\text{strong electrolyte})$$

	$[\,HC_7H_5O_2\,]$	$\leftrightarrow$	$[\,H^{+1}\,]$	+	$[\,C_7H_5O_2^{-1}\,]$
I	0.85 M		~ 0		0.268 M
C	$+\,x$		$-\,x$		$-\,x$
E	$0.85 + x$		$[\,H^{+1}\,]$		$0.268 - x$

$$K_a = \frac{[\,H^{+1}\,][\,C_7H_5O_2^{-1}\,]}{[\,HC_7H_5O_2\,]} = \frac{[\,H^{+1}\,](\,0.268 - x\,)}{(\,0.85 + x\,)} = \frac{[\,H^{+1}\,](\,0.268\,)}{(\,0.85\,)} = 9.51 \times 10^{-4}$$

$$[\,H^{+1}\,] = \frac{(\,9.51 \times 10^{-4}\,)(\,0.85\,)}{(\,0.268\,)} = 3.02 \times 10^{-3}$$

$$\mathbf{pH} = -\log\,[\,H^{+1}\,] = -\log\,(\,3.02 \times 10^{-3}\,) = \mathbf{2.52}$$

$$\boldsymbol{OR}$$

$$\mathbf{pH} = pK_a + \log\frac{[\,A^{-1}\,]}{[\,HA\,]} = -\log(\,9.51 \times 10^{-4}\,) + \log\left(\frac{0.268}{0.85}\right) = 2.52$$

(d) 5.33 grams of solid HCl is added to 250 mL of the solution created in part (c). Calculate the pH, in the resulting solution.

$$\frac{5.33 \ g \ HCl}{} \times \frac{1 \ mol \ HCl}{36.458 \ g} \ = \ 0.146 \ mol \ HCl$$

$$\frac{0.146 \ mol \ HCl}{0.250 \ L} \ = \ 0.585 \ M \ \ HCl$$

	[$HC_7H_5O_2$]	←	[H^{+1}]	+	[$C_7H_5O_2{}^{-1}$]
I	0.850 *M*		0.585 *M*		0.268 *M*
C	+ 0.268		- 0.268		- 0.268
E	0.850 + 0.268 = 1.118		0.585 - 0.268 = 0.317		~ 0

This is not a "good" buffer solution.

The solution could not handle the additional H^{+1} ions (from the strong acid). Since this solution is unable to handle the additional strong acid, we KNOW the concentration of the H^{+1} at equilibrium. Therefore:

$$\textbf{pH} = \ - \log \ [\, H^{+1} \,] \ = \ - \log \ (\, 0.317 \,) \ = \ \textbf{0.50}$$

2. A buffered solution (at equilibrium) contains 3.75 M acetic acid, ($HC_2H_3O_2$, $K_a = 1.80$ x 10^{-5}) and 2.43 M potassium acetate, $KC_2H_3O_2$.

a) Calculate the pH of this solution.

$$KC_2H_3O_2 \rightarrow K^{+1} + C_2H_3O_2^{-1} \text{ (strong electrolyte)}$$

	[$HC_2H_3O_2$]	$\leftrightarrow$	[H^{+1}]	+	[$C_2H_3O_2^{-1}$]
E	3.75 M		~ 0		2.43 M

$$K_a = \frac{\left[H^{+1} \right]\left[C_2H_3O_2^{-1} \right]}{\left[HC_2H_3O_2 \right]} = \frac{\left[H^{+1} \right](2.43)}{(3.75)} = 1.80 \times 10^{-5}$$

$$\left[H^{+1} \right] = \frac{(1.80 \times 10^{-5})(3.75)}{(2.43)} = 2.78 \times 10^{-5}$$

$$\mathbf{pH} = - \log \left[H^{+1} \right] = - \log (2.78 \text{ x } 10^{-5}) = \mathbf{4.56}$$

OR

$$\mathbf{pH} = pK_a + \log \frac{\left[A^{-1} \right]}{\left[HA \right]} = -\log (1.80 \times 10^{-5}) + \log \left(\frac{2.43}{3.75} \right) = 4.56$$

b) Calculate the pH after 34.3 grams of NaOH $_{(s)}$ is added to 1.0 L of this solution (volume change is negligible).

$$\frac{34.3 \text{ } g \text{ } NaOH}{} \times \frac{1 \text{ } mol \text{ } NaOH}{39.998 \text{ } g} = 0.858 \text{ } mol \text{ } NaOH$$

$$\frac{0.858 \text{ } mol \text{ } NaOH}{1 \text{ } L} = 0.858 \text{ } M \text{ } NaOH$$

	[$C_2H_3O_2^{-1}$]	+	[H_2O]	$\leftarrow$	[OH^{-1}]	+	[$HC_2H_3O_2$]
I	2.43 M		-		0.858 M		3.75 M
C	+ 0.858		-		- 0.858		- 0.858
E	3.288		-		~ 0		2.892

$$\mathbf{pH} = pK_a + \log \frac{\left[A^{-1} \right]}{\left[HA \right]} = -\log (1.80 \times 10^{-5}) + \log \left(\frac{3.288}{2.892} \right) = 4.80$$

<u>OR</u>

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{1.80 \times 10^{-5}} = 5.56 \times 10^{-10}$$

$$K_b = \frac{\left[OH^{-1} \right]\left[HC_2H_3O_2 \right]}{\left[C_2H_3O_2^{-1} \right]} = \frac{\left[OH^{-1} \right](2.892)}{(3.288)} = 5.56 \times 10^{-10}$$

$$\left[OH^{-1} \right] = \frac{(5.56 \times 10^{-10})(3.288)}{(2.892)} = 6.32 \times 10^{-10}$$

$$pOH = -\log \left[OH^{-1} \right] = -\log (6.32 \times 10^{-10}) = 9.20$$

$$\mathbf{pH} = 14 - pOH = 14 - 9.20 = \mathbf{4.80}$$