AP CHEMISTRY

TOPIC 7: ACIDS & BASES, REVIEW, PART II

Day 83:

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

$$HC_3H_5O_{2\,(aq)} \rightleftharpoons H^+_{(aq)} + C_3H_5O_2^{-1}{}_{(aq)}$$

1. Propanoic acid, $HC_3H_5O_{2\,(aq)}$, is a monoprotic acid that dissociates in an aqueous solution, as represented by the equation above. Propanoic acid is 2.03 percent dissociated in 0.85 *M* $HC_3H_5O_{2\,(aq)}$ 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 propanoic acid and calculate its value.

$$K_a = \frac{\left[\ H^{+1}\ \right]\left[\ C_3H_5O_2^{-1}\ \right]}{\left[\ HC_3H_5O_2\ \right]}$$

Propanoic acid is a weak acid, Strong acids (and strong bases) always dissociate to 100%. Since the question gives the percent dissociation for this weak acid, we can determine the concentration for the products (x) AT EQUILIBRIUM.

$$0.85\ M \text{ x } 0.0203 = 0.017255\ M$$

	$[\ HC_3H_5O_2\ ]$	$\rightleftharpoons$	$[\ H^{+1}\ ]$	+	$[\ C_3H_5O_2^{-1}\ ]$
I	0.85 *M*		0		0
C	$- x$		$+ x =$ 0.017255 *M*		$+ x$
E	0.85 - 0.017255		0.017255		0.017255

$$K_a = \frac{\left[\ H^{+1}\ \right]\left[\ C_3H_5O_2^{-1}\ \right]}{\left[\ HC_3H_5O_2\ \right]} = \frac{(\ 0.017255\ )^2}{(0.85 - 0.017255)} = 3.58 \times 10^{-4}$$

Since we KNOW "x", do not drop the "x" in the equation to solve for K_a.

   (b) Calculate the pH of 0.85 *M* $HC_3H_5O_{2\,(aq)}$.

$$[\ H^{+1}\ ] = 0.017255\ \boldsymbol{M}$$

$$\mathbf{pH} = -\log(\ 0.017255\ ) = \mathbf{1.76}$$

(c) Calculate the pH of a solution formed by dissolving 0.0864 mole of solid potassium propanate, $KC_3H_5O_2$, in 250 mL of 0.85 *M* $HC_3H_5O_{2\,(aq)}$. Assume that volume change is negligible.

$$\mathbf{KC_3H_5O_2 \rightarrow K^{+1} + C_3H_5O_2^{-1}}$$

$$\frac{0.0864\ mol\ C_3H_5O_2^{-1}}{0.250\ L} = 0.3456\ M$$

	$[C_3H_5O_2^{-1}]$	+	HOH	$\rightleftharpoons$	$[HC_3H_5O_2]$	+	$[OH^{-1}]$
I	0.3456 *M*		-		0.85 *M*		~ 0
C	- *x*		-		+ *x*		+ *x*
E	0.3456 - *x*		-		0.85 + *x*		$[OH^{-1}]$

$$K_b = \frac{K_w}{K_a} = \frac{1.00\times10^{-14}}{3.58\times10^{-4}} = 2.79\times10^{-11}$$

$$K_b = \frac{\left[HC_3H_5O_2\right]\left[OH^{-1}\right]}{\left[C_3H_5O_2^{-1}\right]} = \frac{(0.85 + x)\left[OH^{-1}\right]}{(0.3456 - x)} = \frac{(0.85)\left[OH^{-1}\right]}{(0.3456)} = 2.79\times10^{-11}$$

$$\left[OH^{-1}\right] = \frac{(2.79\times10^{-11})(0.3456)}{(0.85)} = 1.14\times10^{-11}$$

$$\text{pOH} = -\log(1.14 \text{ x } 10^{-11}) = 10.94$$

$$\mathbf{pH} = 14 - \text{pOH} = 14 - 10.94 = \mathbf{3.06}$$

2. Calculate the pOH for a 1.44 *M* solution of a hydroxylamine , $HONH_2$, solution. $K_b = 1.1 \times 10^{-8}$.

Answers: ***WEAK BASE (you are given a K_b value - NOT a conjugate base - hydrogen is added to the end !***

Remember, add H^{+1} to the "BACK" of Bases

	[$HONH_2$]	+	[H_2O]	$\rightleftharpoons$	[$HONH_3^{+1}$]	+	[OH^{-1}]
I	1.44 *M*		-		0		0
C	- *x*		-		+ *x*		+ *x*
E	1.44 - *x*		-		*x*		*x*

$$K_b = 1.1 \times 10^{-8} = \frac{\left[HONH_3^{+1} \right]\left[OH^{-1} \right]}{\left[HONH_2 \right]} = \frac{(x)(x)}{1.44 - x} = \frac{x^2}{1.44}$$

$$K_b = 1.1 \times 10^{-8} = \frac{x^2}{1.44}$$

$$\left(1.1 \times 10^{-8} \right)\left(1.44 \right) = x^2$$

$$x = \sqrt{1.58 \times 10^{-8}} = 1.26 \times 10^{-4}\ M = \left[OH^{-1} \right]$$

pOH = - log (1.26 x 10^{-4}) = **3.90**

3. Calculate the pH of a 0.95 M LiF solution. K_a value for HF is 7.2×10^{-4}.

$$LiF \rightarrow Li^{+1} + F^{-1}$$

(Li^{+1}, conjugate acid of a strong base, F^{-1}, conjugate base of a weak acid)

The sodium ion, Li^{+1}, has no effect on the solution to make it basic or acidic.

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{7.2 \times 10^{-4}} = 1.39 \times 10^{-11}$$

Pure water (pH or pure water equals 7.00) and then the salt is added !

	$[F^{-1}]$	+	$[H_2O]$	$\rightleftharpoons$	[HF]	+	$[OH^{-1}]$
I	0.95 M		-		0		0
C	$-x$		-		$+x$		$+x$
E	0.95 M - x		-		x		x

$$K_b = 1.39 \times 10^{-11} = \frac{\left[\ HF\ \right]\left[\ OH^{-1}\ \right]}{\left[\ F^{-1}\ \right]} = \frac{x^2}{0.95 - x} = \frac{x^2}{0.95}$$

$$x^2 = (0.95)\left(\ 1.39 \times 10^{-11}\ \right) = 1.32 \times 10^{-11}$$

$$x = \sqrt{1.32 \times 10^{-11}} = 3.63 \times 10^{-6}$$

$$[\ OH^{-1}\ ] = 3.63 \times 10^{-6}\ M = x$$

$$pOH = -\log(\ 3.63 \times 10^{-6}\ ) = 5.44$$

pH = 14 - 5.44 = **8.56**