AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART E,   EXAMPLES, PART II

Day 82:

- Common-Ion Effect
- Percent Dissociation (weak acids & bases)

Example:

$$HC_3H_4O_{2\,(aq)} \rightleftharpoons H^+_{(aq)} + C_3H_4O_2{}^{-1}{}_{(aq)}$$

Acrylic acid, $HC_3H_4O_2$, is a monoprotic acid that dissociates in an aqueous solution, as represented in the equation above. Acrylic acid is 0.6433 percent dissociated in 1.35 *M* $HC_3H_4O_{2\,(aq)}$ at 298 K. For parts (a) through (c) below, assume the temperature remains at 298 K.

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

$$K_a = \frac{\left[\ H^{+1}\ \right]\left[\ C_3H_4O_2{}^{-1}\ \right]}{\left[\ HC_3H_4O_2\ \right]}$$

$$(\ 1.35\ M\ )(\ 0.006433\ )\ =\ 0.008684\ M\ =\ x$$

	[$HC_3H_4O_2$]	$\rightleftharpoons$	[H^{+1}]	+	[$C_3H_4O_2{}^{-1}$]
I	1.35 *M*		0		0
C	- x		+ x		+ x
E	1.35 *M* - 0.008684 *M*		0.008684 *M*		0.008684 *M*

$$K_a = \frac{x^2}{1.35 - x} = \frac{(\ 0.008684\ )^2}{1.341316} = 5.62 \times 10^{-5}$$

(b) Calculate the pH of 1.35 *M* $HC_3H_4O_2$

$$\mathbf{pH} = -\log\ [\ H^{+1}\ ]\ =\ -\log\ (\ 0.008684\ )\ =\ \mathbf{2.06}$$

(c) Calculate the pH of a solution formed by dissolving 1.35 mol of solid potassium , $KC_3H_4O_2$. in 2550 mL of 1.35 *M* $HC_3H_4O_2$. Assume the volume change is negligible.

$$KC_3H_4O_2 \longrightarrow K^{+1} + C_3H_4O_2{}^{-1}$$

(this is a SALT – that produces a “common ion”)

$$M = \frac{1.35\ mol\ KC_3H_4O_2}{2.550\ L} = 0.529\ M$$

	[$C_3H_5O_3{}^{-1}$]	+	HOH	$\rightleftharpoons$	[$HC_3H_5O_3$]	+	[OH^{-1}]
I	0.529 *M*		-		1.35 *M*		~ 0
C	- x		-		+ x		+ x
E	0.529 - x		-		1.35 + x		[OH^{-1}]

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{5.62 \times 10^{-5}} = 1.78 \times 10^{-10}$$

$$K_b = \frac{[\ HC_3H_5O_3\ ][\ OH^{-1}\ ]}{[\ C_3H_5O_3^{-1}\ ]} = \frac{(\ 1.35 + x\ )[\ OH^{-1}\ ]}{(\ 0.529 - x\ )} = \frac{(\ 1.35\ )[\ OH^{-1}\ ]}{(\ 0.529\ )} = 1.78 \times 10^{-10}$$

$$[\ OH^{-1}\ ] = \frac{(\ 1.78 \times 10^{-10}\ )(\ 0.529\ )}{(\ 1.35\ )} = 6.97 \times 10^{-11}$$

$$\mathbf{pOH} = -\log\ [\ OH^{-1}\ ] = -\log\ (\ 6.97 \times 10^{-11}\ ) = \mathbf{10.16}$$

$$\mathbf{pH} = 14 - 10.16 = \mathbf{3.84}$$

OR

$$\mathbf{pH} = (\ 14 - (\ -\log\ [\ OH^{-1}\ ]\ ) = 14 - (\ -\log\ (\ 6.97 \times 10^{-11}\ )) = \mathbf{3.84}$$

2. Calculate the pH of a 1.00 liter 2.22 *M* nitrous acid, HNO_2 , solution (***at equilibrium***). THEN calculate the pH of the solution when 15.3 grams of solid lithium nitrite, $LiNO_2$, is added to the solution at equilibrium. Assume that the volume change is negligible. The K_a value for HNO_2 is 4.0×10^{-4}.

$$HNO_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^{+1}{}_{(aq)} + NO_2^{-1}{}_{(aq)}$$

	$[\ HNO_2\ ]$	+	$[\ H_2O\ ]$	$\rightleftharpoons$	$[\ H_3O^{+1}\ ]$	+	$[\ NO_2^{-1}\ ]$
I	2.22 *M*		-		0		0
C	$-x$		-		$+x$		$+x$
E	$2.22 - x$		-		x		x

$$K_a = 4.0 \times 10^{-4} = \frac{[\ H_3O^{+1}\ ][\ NO_2^{-1}\ ]}{[\ HNO_2\ ]} = \frac{x^2}{2.22 - x} = \frac{x^2}{2.22}$$

$$x^2 = (\ 2.22\ )(\ 4.0 \times 10^{-4}\ ) = 8.88 \times 10^{-4}$$

$$x = \sqrt{8.88 \times 10^{-4}} = 0.0298$$

$$[\ H^{+1}\ ] = 0.0298\ M$$

$$\mathbf{pH} = -\log\ (\ 0.0298\ ) = \mathbf{1.53}$$

At equilibrium, 15.3 grams of solid lithium nitrite, $LiNO_2$, is added to the equilibrium system. Lithium nitrite dissociates to 100%. The lithium does nothing to affect the solution (in regards to altering the pH because it is a conjugate of a STRONG BASE) Lithium ion is a spectator ion in the solution.

$$LiNO_{2\,(aq)} \rightleftharpoons Li^{+1}{}_{(aq)} + NO_2^{-1}{}_{(aq)}$$

(this is a SALT – that produces a "common ion")

Calculate the concentration of the nitrite ion once it has dissociated in the nitrous acid solution at equilibrium.

$$\frac{15.3\ g\ LiNO_2}{} \times \frac{1\ mol\ LiNO_2}{52.95\ g} = 0.289\ mol$$

$$M = \frac{mol}{L} = \frac{0.289\ mol\ LiNO_2}{1.00\ L} = 0.289\ M\ LiNO_2$$

$$LiNO_{2\,(aq)} \rightleftharpoons Li^{+1}{}_{(aq)} + NO_2^{-1}{}_{(aq)}$$

$$[\ NO_2^{-1}\ ] = 0.289\ M$$

	$[\ NO_2^{-1}\ ]$	+	HOH	$\rightleftharpoons$	$[\ HNO_2\ ]$	+	$[\ OH^{-1}\ ]$
I	0.289 M		-		2.22 M		~ 0
C	$-x$		-		$+x$		$+x$
E	$0.289 - x$		-		$2.22 + x$		$[\ OH^{-1}\ ]$

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{4.00 \times 10^{-4}} = 2.50 \times 10^{-11}$$

$$K_b = \frac{[\ HNO_2\ ][\ OH^{-1}\ ]}{[\ NO_2^{-1}\ ]} = \frac{(\ 2.22 + x\ )[\ OH^{-1}\ ]}{(\ 0.289 - x\ )} = \frac{(\ 2.22\ )[\ OH^{-1}\ ]}{(\ 0.289\ )} = 2.50 \times 10^{-11}$$

$$[\ OH^{-1}\ ] = \frac{(\ 2.50 \times 10^{-11}\ )(\ 0.289\ )}{(\ 2.22\ )} = 3.25 \times 10^{-12}$$

pOH = - log (3.25×10^{-12}) = **11.49**

pH = 14 - pOH = 14 - **11.49 = 2.51**

OR

pH = (14 - (- log $[\ OH^{-1}\ ]$) = 14 – (- log (3.25×10^{-12})) = **2.51**