

AP CHEMISTRY

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

Day 82:

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

Example:



Acrylic acid, $\text{HC}_3\text{H}_4\text{O}_2$, 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 $\text{HC}_3\text{H}_4\text{O}_2(\text{aq})$ 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.

$$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_4\text{O}_2^{-1}]}{[\text{HC}_3\text{H}_4\text{O}_2]}$$

$$(1.35 \text{ M})(0.006433) = 0.008684 \text{ M} = x$$

	$[\text{HC}_3\text{H}_4\text{O}_2]$	\rightleftharpoons	$[\text{H}^+]$	+	$[\text{C}_3\text{H}_4\text{O}_2^{-1}]$
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 $\text{HC}_3\text{H}_4\text{O}_2$

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

- (c) Calculate the pH of a solution formed by dissolving 1.35 mol of solid potassium, $\text{KC}_3\text{H}_4\text{O}_2$, in 2550 mL of 1.35 M $\text{HC}_3\text{H}_4\text{O}_2$. Assume the volume change is negligible.



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

$$M = \frac{1.35 \text{ mol } \text{KC}_3\text{H}_4\text{O}_2}{2.550 \text{ L}} = 0.529 \text{ M}$$

	$[\text{C}_3\text{H}_5\text{O}_3^{-1}]$	+	HOH	\rightleftharpoons	$[\text{HC}_3\text{H}_5\text{O}_3]$	+	$[\text{OH}^{-1}]$
I	0.529 M		-		1.35 M		~ 0
C	-x		-		+x		+x
E	0.529 - x		-		1.35 + x		$[\text{OH}^{-1}]$

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

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

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

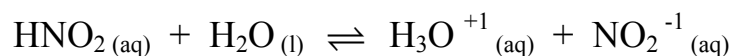
$$\text{pOH} = -\log [OH^{-1}] = -\log (6.97 \times 10^{-11}) = \mathbf{10.16}$$

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

OR

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

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]	+	[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

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

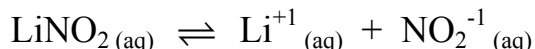
$$x^2 = (2.22)(4.0 \times 10^{-4}) = 8.88 \times 10^{-4}$$

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

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

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

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.

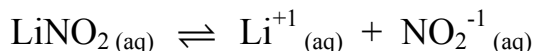


(this is a SALT – that produces a “common ion”)

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

$$\frac{15.3 \text{ g LiNO}_2}{52.95 \text{ g}} \times \frac{1 \text{ mol LiNO}_2}{1 \text{ mol LiNO}_2} = 0.289 \text{ mol}$$

$$M = \frac{\text{mol}}{L} = \frac{0.289 \text{ mol LiNO}_2}{1.00 \text{ L}} = 0.289 \text{ M LiNO}_2$$



$$[\text{NO}_2^-] = 0.289 \text{ M}$$

	$[\text{NO}_2^-]$	+	HOH	\rightleftharpoons	$[\text{HNO}_2]$	+	$[\text{OH}^-]$
I	0.289 M		-		2.22 M		~ 0
C	-x		-		+x		+x
E	0.289 - x		-		2.22 + x		$[\text{OH}^-]$

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

$$K_b = \frac{[\text{HNO}_2][\text{OH}^-]}{[\text{NO}_2^-]} = \frac{(2.22 + x)[\text{OH}^-]}{(0.289 - x)} = \frac{(2.22)[\text{OH}^-]}{(0.289)} = 2.50 \times 10^{-11}$$

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

$$\text{pOH} = -\log(3.25 \times 10^{-12}) = 11.49$$

$$\text{pH} = 14 - \text{pOH} = 14 - 11.49 = 2.51$$

OR

$$\text{pH} = (14 - (-\log[\text{OH}^-])) = 14 - (-\log(3.25 \times 10^{-12})) = 2.51$$