

## General Chemistry II

RR # 8 Answer Key  
Summer 2022

1. A student prepares a 0.45 M solution of a monoprotic weak acid and determines the pH to be 3.68. What is the  $K_a$  of the weak acid?

①  $pH = -\log[H_3O^+] \text{ so } [H_3O^+] = 10^{-3.68} = 2.09 \times 10^{-4} M$



I	.45 M	-	0	
C	-X	-	+X	+X
E	.45 - X	-	X	X

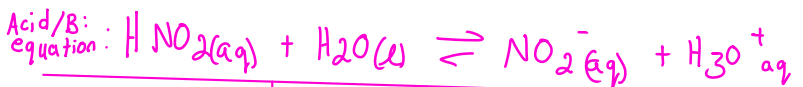
$X = 2.09 \times 10^{-4}$  so  $[HA] = .45 - 2.09 \times 10^{-4}$   
 $[HA] = .4498 M$

③  $K_a = \frac{[A^-][H_3O^+]}{[HA]} = \frac{(2.09 \times 10^{-4})^2}{.4498} = 9.7 \times 10^{-8}$

2. In the lab you make a solution that is 0.0048 M  $HNO_2$  and 0.00056 M  $LiNO_2$ .

What is the pH of your solution?  $K_a = 4.5 \times 10^{-4}$

$Li^+ NO_2^-$   
 $LiNO_2$  in  $H_2O$ , so initial  $[NO_2^-] = [LiNO_2] = .00056 M$ .  $HNO_2$  is a weak acid and will ionize in  $H_2O$ .



I	.0048	-	.00056	0
C	-X	-	+X	+X
E	.0048 - X	-	.00056 + X	X

$K_a = \frac{[NO_2^-][H_3O^+]}{[HNO_2]} = \frac{(.00056 + X)(X)}{.0048 - X} = 4.5 \times 10^{-4}$

$X^2 + .00056X = 2.16 \times 10^{-6} - 4.5 \times 10^{-4}X$

$X^2 + .00101X - 2.16 \times 10^{-6} = 0$

$X = [H_3O^+] = .001049 M$

$pH = -\log[.001049] = 2.98 = pH$

3. Lactate,  $CH_3CH(OH)CO_2^-$ , is constantly produced from pyruvate during normal metabolism. When the citric acid cycle backs up due to insufficient oxygen supply, lactate builds up in your exercising muscles and you feel that painful burning sensation. Lactate has  $K_b = 7.24 \times 10^{-11}$  at 25 °C.

a. Write the reaction equation described by this  $K_b$ .



b. What is the  $K_a$  of lactic acid at 25 °C?

$$K_w = K_a \cdot K_b$$

$$1.00 \times 10^{-14} = K_a \cdot 7.24 \times 10^{-11}$$

$$K_a = 1.38 \times 10^{-4}$$

c. If a solution is initially 0.210 M lactic acid, what is the pH at 25 °C?

You can assume that the  $K_a$  is considered very small.

	$[CH_3CH(OH)CO_2H]$	$[H_2O]$	$[CH_3CH(OH)CO_2^-]$	$[H_3O^+]$
I	0.210 M	—	0	0
C	-X	—	+X	+X
E	0.210 - X	—	X	X

we are told  
 $K_a$  is small  
so we can  
ignore

$$K_a = \frac{X^2}{0.210}$$

$$1.38 \times 10^{-4} = \frac{X^2}{0.210}$$

$$X = 5.39 \times 10^{-3} M = [H_3O^+]$$

$$pH = -\log 5.39 \times 10^{-3} = 2.27$$

4. Determine whether an aqueous solution of each of the following salts will be acidic, basic, or neutral. You may have to look up and compare  $K_a$ s and  $K_b$ s for a few of these.

KClO <sub>4</sub> neutral	NaCN basic	NH <sub>4</sub> CH <sub>3</sub> COO neutral
NaBr neutral	NH <sub>4</sub> ClO basic	K <sub>2</sub> CO <sub>3</sub> basic
CaBr <sub>2</sub> neutral	NaF basic	LiClO <sub>4</sub> neutral
NH <sub>4</sub> Br acidic	NaHCO <sub>3</sub> basic	Al <sub>2</sub> (SO <sub>4</sub> ) <sub>3</sub> acidic

5. A 0.150 M solution of morphine ( $C_{17}H_{19}NO_3$ ) has a pH of 10.5 at 25°C. What is morphine's?

$$pH: 10.5 \quad pOH: 14 - 10.5 = 3.5$$

$$[OH^-]: 10^{-3.5} = 3.6 \times 10^{-4} M$$

Now set up an equation for hydrolysis of morphine:



	$C_{17}H_{19}NO_3(aq)$	$H_2O(l)$	$C_{17}H_{19}NO_3H^+$	$OH^-$
I	.150	—	0	0
C	-X	—	+X	+X
E	.150 - X	—	X	X

We know  $[\text{OH}^-]$  from pOH so  $x = 3.16 \times 10^{-4} \text{ M}$

$$K_b = \frac{[\text{C}_{17}\text{H}_{19}\text{NO}_3\text{H}^+][\text{OH}^-]}{[\text{C}_{17}\text{H}_{19}\text{NO}_3]} = \frac{x^2}{.150 - x} =$$

$$\frac{(3.16 \times 10^{-4} \text{ M})^2}{.150 - 3.16 \times 10^{-4}} = 6.67 \times 10^{-7}$$