WEAK ACIDS AND THE EQUILIBRIUM CONSTANT

NAME SECTION

1. The chemical equation which describes how the weak acid $HC_2H_3O_2$ dissociates in aqueous solution is,

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

a. In the data you obtained earlier (Acids, Bases and pH, pg. 55), the initial concentration of $HC_2H_3O_2$ is 0.100 M. In the space provided below (**ICE** Table), enter the initial concentration of $HC_2H_3O_2$, H^+, and $C_2H_3O_2^-$. Based on the measured pH of this solution, calculate and enter the equilibrium concentration of H^+.

	$HC_2H_3O_2(aq)$	$\rightleftharpoons H^+(aq)$	$+ C_2H_3O_2^-(aq)$
Initial Concentrations	0.100	0	0
Change	−x	+x	+x
Equilibrium Concentrations	.1−x	+x	+x

b. Calculate the change in $[H^+]$.

pH = 2.878 = −log[H⁺] $10^{-2.878} = [H^+]$

$[H^+] = 0.00132M$

c. Using the balanced chemical equation and the calculated change in $[H^+]$, calculate the change in $[HC_2H_3O_2]$ and $[C_2H_3O_2^-]$.

	$HC_2H_3O_2$	H^+	$C_2H_3O_2^-$
Change	−.00132	+0.00132	+0.00132

d. Calculate the equilibrium concentration of $HC_2H_3O_2(aq)$ and $C_2H_3O_2^-(aq)$.

$[C_2H_3O_2^-] = [H^+] = 0.00132M$

$[HC_2H_3O_2] = 0.0987M$

e. Estimate the equilibrium constant for the dissociation of $HC_2H_3O_2$*(aq)*.

$$K_a(HC_2H_3O_2) = \frac{(0.00132)(.00132)}{.0982} = 1.8 \times 10^{-5}$$

f. Calculate the magnitude of the equilibrium constant for benzoic acid, $HC_7H_5O_2$, if a 0.100 M solution has a pH = 2.59.

$$HC_2H_5O_2 \rightleftharpoons H^+ + C_7H_5O_2^-$$

	$HC_2H_5O_2$	H^+	$C_7H_5O_2^-$
I	0.100	~0	0
C	−.00257	+.00257	+.00257
E	.0974	.00257	.00257

pH = 2.59

$10^{-2.59} = [H^+]$

$$K_a = \frac{(.00257)(.00257)}{.0974} = 6.78 \times 10^{-5}$$

g. Calculate the magnitude of the equilibrium constant for an aqueous solution of ammonia, if a 0.100 M solution has a pH = 11.13.

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + {}^-OH$$

	NH_3	H_2O	NH_4^+	^-OH
I	.1	—	0	~0
C	−.00135		+.00135	+.00135
E	.0987		.00135	.00135

pOH = 14 − 11.13 = 2.87

$[^-OH] = 10^{-2.87} =$

$$K_b = \frac{[.00135](.00135)}{.0987} = 1.8 \times 10^{-5}$$

h. Calculate the pH of a solution which is 0.53 M $HC_6H_4NO_2$ (nicotinic acid). ($K_a = 1.4 \times 10^{-5}$)

$$HC_6H_4NO_2 \rightleftharpoons H^+ + C_6H_4NO_2^-$$

	$HC_6H_4NO_2$	H^+	$C_6H_4NO_2^-$
I	0.53	~0	0
C	−x	+x	+x
E	.53−x	+x	+x

$$K_a = \frac{[H^+][C_6H_4NO_2^-]}{[HC_6H_4NO_2]}$$

$$1.4 \times 10^{-5} = \frac{(x)(x)}{.53 - x}$$

assume $.53 - x \cong .53$

$1.4 \times 10^{-5} \cdot .53 = x^2$

$2.7 \times 10^{-3}\,M = x = [H^+]$

$pH = -\log[H^+]$
$= -\log(2.7 \times 10^{-3})$
$= 2.56$

i. Calculate the pH of a solution which is 0.712 M CH_3NH_2 (methylamine). ($K_b = 4.4 \times 10^{-4}$)

$$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + {}^-OH$$

	CH_3NH_2	H_2O	$CH_3NH_3^+$	^-OH
I	.712	—	0	~0
C	−x	—	+x	+x
E	.712−x		+x	+x

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

$$4.4 \times 10^{-4} = \frac{(x)(x)}{.712 - x}$$

assume $.712 - x \cong .712$

$4.4 \times 10^{-4} \cdot .712 = x^2$

$[^-OH] = 5.34 \times 10^{-3}\,M$

pOH = 2.27

pH = 14 − 2.27
= 11.73