Chem1403 Worksheet
RICE table

1. **With the following condensation reaction, calculate the equilibrium concentration for C_2H_5OH. Given $[C_4H_7OOH]_{initial}$ = 0.590M, $[C_2H_5OH]_{initial}$ = 0.450M, and K_{eq} = 4.69**

$C_4H_7OOH_{(aq)} + C_2H_5OH_{(aq)} \leftrightarrow C_4H_7OOC_2H_{5(aq)} + H_2O_{(l)}$

I	0.590M	0.450M	0	X
C	-x	-x	+x	X
E	0.590 - x	0.450 - x	x	X

$$K = \frac{[C_4H_7OOC_2H_5]}{[C_4H_7OOH][C_2H_5OH]}$$

$$4.69 = \frac{x}{(0.590 - x)(0.450 - x)}$$

$$4.69 = \frac{x}{(.2655 - 1.04x + x^2)}$$

$$4.69(.2655 - 1.04x + x^2) = x$$

$$1.25 - 4.88x + 4.69x^2 = x$$

$$1.25 - 5.88x + 4.69x^2 = 0 \rightarrow 4.69x^2 - 5.88x + 1.25 = 0$$

$$*\ Using\ quadratic\ formula: x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = 0.98\ and\ 0.27$$

*0.27 is chosen because 0.98 is larger any initial concentration. It is absurd to consumed more than initial concentration.

$$x = 0.98\ and\ 0.27$$

$$[C_2H_5OH]_{equilibrium} = 0.450 - 0.270 = 0.180M$$

2. **Calculate the initial concentration of $[NH_3]$.**
   Given: $[Ag^+]_{initial}$ = 0.0100M, $[Ag(NH_3)_2^+]_{equilibrium}$ = 0.00278M, and K_{eq} = 7.99.

$Ag^+{}_{(aq)}$ + $2NH_{3\,(aq)} \leftrightarrow Ag(NH_3)_2^+{}_{(aq)}$

I	0.0100M	x	0
C	-0.00278M	-2(0.00278M)	+0.00278
E	0.00722M	x - 0.00556M	0.00278

$$K = \frac{[Ag(NH_3)_2{}^+]}{[Ag^+][NH_3]^2}$$

$$7.99 = \frac{0.00278}{(0.00722)(x - 0.00556)^2}$$

$$(0.00722)(x - 0.00556)^2 = \frac{0.00278}{7.99}$$

$$(x - 0.00556)^2 = \frac{0.00278}{7.99}(0.00722)$$

$$(x - 0.00556)^2 = \frac{0.00278}{7.99}(0.00722)$$

$$(x - 0.00556)^2 = 2.51 \times 10^{-6}$$

$$x - 0.00556 = \sqrt{(2.51 \times 10^{-6})}$$

$$x = .00714M$$

3. **Calculate the equilibrium concentration for I_2.**
   Given: $[H_2]_{initial}$ = 0.25M, $[I_2]_{initial}$ = 0.25M, $[HI]_{intial}$ = 1.2M and K = 54

$H_{2(s)}$ + $I_{2(aq)} \rightarrow$ $2HI_{(g)}$

I	0.25M	0.25M	1.2
C	-x	-x	+2x
E	0.25 -x	0.25 -x	1.2 + 2x

$$K = \frac{[HI]^2}{[H_2][I_2]}$$

$$\frac{(1.2 + 2x)^2}{(0.25 - x)(0.25 - x)} = 54$$

$$\frac{(1.2 + 2x)^2}{(0.25 - x)^2} = 54$$

$$\left[\frac{(1.2 + 2x)}{(0.25 - x)}\right]^2 = 54$$

$$\frac{(1.2 + 2x)}{(0.25 - x)} = \sqrt{54}$$

$$1.2 + 2x = \sqrt{54}\,(0.25 - x)$$

$$1.2 + 2x = \sqrt{54}(0.25) - \sqrt{54}x$$

$$2x + \sqrt{54}x = \sqrt{54}(0.25) - 1.2$$

$$9.35x = .637$$

$$x = 0.0681M$$

4. **This is a dissociation reaction of acetic acid (a weak acid). With K = 1.76 x 10^{-5} and [CH_3COOH] = 1.6 x 10^{-4}, determine the equilibrium concentration for [H_3O^+].**

$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightarrow CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

I	1.6 x 10^{-4}	X	0	0
C	-x	X	+x	+x
E	1.6 x 10^{-4} - x	X	x	x

$$K = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$1.76 \times 10^{-5} = \frac{x^2}{(1.6 \times 10^{-4}) - x}$$

$$(1.76 \times 10^{-5})(1.6 \times 10^{-4} - x) = x^2$$

$$(2.82 \times 10^{-9}) - (1.76 \times 10^{-5})x = x^2$$

$$0 = x^2 + (1.76 \times 10^{-5})x - (2.82 \times 10^{-9})$$

$$* \, Using\ quadratic\ formula: x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = 4.50 \times 10^{-5}\ and - 6.26 \times 10^{-5}$$

*4.50×10^{-5} is chosen because -6.26×10^{-5} is negative. Cannot have negative concentration from nothing.