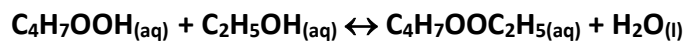


Chem1403 Worksheet
RICE table

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



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

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

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

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

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

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

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

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

$$x = 0.98 \text{ and } 0.27$$

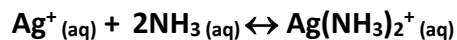
*0.27 is chosen because 0.98 is larger any initial concentration. It is absurd to consumed more than initial concentration.

$$x = 0.98 \text{ and } 0.27$$

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

2. Calculate the initial concentration of $[\text{NH}_3]$.

Given: $[\text{Ag}^+]_{\text{initial}} = 0.0100\text{M}$, $[\text{Ag}(\text{NH}_3)_2^+]_{\text{equilibrium}} = 0.00278\text{M}$, and $K_{\text{eq}} = 7.99$.



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

$$K = \frac{[\text{Ag}(\text{NH}_3)_2^+]}{[\text{Ag}^+][\text{NH}_3]^2}$$

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

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

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

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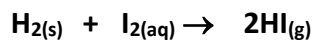
$$(x - 0.00556)^2 = 2.51 \times 10^{-6}$$

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

$$x = .00714\text{M}$$

3. Calculate the equilibrium concentration for I_2 .

Given: $[\text{H}_2]_{\text{initial}} = 0.25\text{M}$, $[\text{I}_2]_{\text{initial}} = 0.25\text{M}$, $[\text{HI}]_{\text{initial}} = 1.2\text{M}$ and $K = 54$



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

$$K = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

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

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

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

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

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

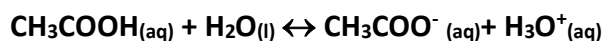
$$1.2 + 2x = \sqrt{54}(0.25) - \sqrt{54}x$$

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

$$9.35x = .637$$

$$x = 0.0681M$$

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



I	1.6×10^{-4}	X	0	0
C	-x	X	+x	+x
E	$1.6 \times 10^{-4} - x$	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: $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

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

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