

## AP Chemistry - Packet #1 - Equilibrium

1) a) 5 points

$$K_p = (P_{\text{NH}_3}) (P_{\text{H}_2\text{S}}) \quad P_{\text{NH}_3} = P_{\text{H}_2\text{S}} = (0.659 \text{ atm} / 2) = 0.330 \text{ atm}$$

$$K_p = (0.330 \text{ atm}) (0.330 \text{ atm}) = 0.109 \text{ atm}^2$$

b) 5 points

$$P_{\text{NH}_3} = 2 P_{\text{H}_2\text{S}} \quad (2x) (x) = 0.109 \quad x = 0.233 \text{ atm} = P_{\text{H}_2\text{S}} \quad 2x = 0.466 \text{ atm} = P_{\text{NH}_3}$$

c) 5 points

equilibrium pressure of  $\text{NH}_3$  = equilibrium pressure of  $\text{H}_2\text{S}$  = 0.330 atm

$$P_{\text{NH}_3} \text{ that has reacted} = P_{\text{H}_2\text{S}} = 0.500 - 0.330 = 0.170 \text{ atm}$$

$$n = PV / RT = (0.170 \times 1.00) / (0.08205 \times 298) = 6.95 \times 10^{-3}$$

If a student calculated  $K_c$  rather than  $K_p$ , one point was deducted.

2) a) two points



$$(x) (x) = K_{\text{sp}} = 7.6 \times 10^{-7}$$

$$(x) = 8.7 \times 10^{-4} \text{ mol / liter} = \text{solubility of SrSO}_4$$

b) three points



$$K_{\text{sp}} = [\text{Sr}^{2+}] [\text{F}^-]^2 = (x) (2x)^2 = 7.9 \times 10^{-10}$$

$$x = 5.8 \times 10^{-4} \text{ mol / liter} = \text{solubility of SrF}_2$$

c) two points

Solve for  $[\text{Sr}^+]$  required for precipitation of each salt.

$$K_{\text{sp}} = [\text{Sr}^{2+}] [\text{F}^-]^2 = 7.9 \times 10^{-10} = (x) (0.020 \text{ mole} / 1.0 \text{ L})^2 = 7.9 \times 10^{-10}$$

$$x = 2.0 \times 10^{-6} \text{ M}$$

$$K_{\text{sp}} = [\text{Sr}^{2+}] [\text{SO}_4^{2-}] = 7.6 \times 10^{-7} = (y) (0.10 \text{ mole} / 1.0 \text{ liter}) = 7.6 \times 10^{-7}$$

$$y = 7.6 \times 10^{-6} \text{ M}$$

Since  $2.0 \times 10^{-6} \text{ M} < 7.6 \times 10^{-6} \text{ M}$ ,  $\text{SrF}_2$  must precipitate first.

When  $\text{SrF}_2$  precipitates,  $[\text{Sr}^{2+}] = 2.0 \times 10^{-6} \text{ M}$

d) two points

The second precipitate to form is  $\text{SrSO}_4$ , which appears when  $[\text{Sr}^{2+}] = 7.6 \times 10^{-6} \text{ M}$  (from Part c.)

When  $[\text{Sr}^{2+}] = 7.6 \times 10^{-6} \text{ M}$ ,  $[\text{F}^-]$  is determined as follows:

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$$K_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2 = 7.9 \times 10^{-10} = (7.6 \times 10^{-6})(z)^2 = 7.9 \times 10^{-10}$$

$$z = 1.0 \times 10^{-2} \text{M}$$

$$\% \text{F}^- \text{ still in solution} = 1.0 \times 10^{-2} / 2.0 \times 10^{-2} \times 100 = 50. \%$$

3a) two points

$$\text{mole fraction for CO} = \text{moles CO} \div \text{moles (CO + CO}_2 + \text{H}_2\text{O + H}_2)$$

$$= 0.55 \div (0.55 + 0.55 + 0.20 + 0.30) = 0.55 \div 1.60 = 0.34 \text{ (or 34\%)}$$

Either mole fraction for CO definition above or numerator or denominator correct: one pt.

Numerator, denominator and calculation correct or fraction 11/32: one pt.

One pt. deducted if units included with answer.

b) two points

$$K_c = ([\text{H}_2\text{O}][\text{CO}]) \div ([\text{H}_2][\text{CO}_2]) = [ (0.55)(0.55) ] \div [ (0.20)(0.30) ] = 5.0$$

Correct set up and/or correct substitution: one pt.

Calculated answer: one pt.

One pt. deducted if expression(s) inverted but otherwise correct.

One pt. deducted if H<sub>2</sub>O missing.

c) one point  $K_p = K_c (RT)^{[\text{delta}]n}$ ;  $[\text{delta}]n = 0$  therefore  $K_p = K_c$  (or,  $K_p = 5.0$ )

The above earns one pt. Note: consistent  $[\text{delta}]n$  calculation is allowed. Numerical value same as part (b) allowed.

d) two points

$$\text{moles CO reacting} = \text{moles H}_2\text{O reacting} = \text{moles CO}_2 \text{ formed} = 0.16 \text{ moles}$$

$$\text{equilibrium [ ]'s at lower T: } [\text{H}_2] = 0.36 \text{ mol/L} \quad [\text{CO}_2] = 0.46 \text{ mol/L} \quad [\text{H}_2\text{O}] = [\text{CO}] = 0.39 \text{ mol/L}$$

$$K_c = (0.39)^2 \div [ (0.46)(0.36) ] = 0.92$$

e) two points

x = number of moles that react

	[H <sub>2</sub> ]	[CO <sub>2</sub> ]	[H <sub>2</sub> O]	[CO]
Initial	0.50/3.0	0.50/3.0	0	0
Change	-x/3.0	-x/3.0	x/3.0	x/3.0
Equilibrium	(0.50-x)/3.0	(0.50-x)/3.0	x/3.0	x/3.0

$$[x^2 / 9] \div [(0.50-x) / 3.0]^2 = 5.0$$

$$x \div (0.50 - x) = 2.2$$

$$x = 0.34 \text{ mol}$$

$$[\text{CO}] = 0.34 \text{ mol} \div 3.0 \text{ L} = 0.11 \text{ mol/L}$$

Note: without final correct [CO], one point awarded if:

all 3.0's in denominators missing, or

all four [ ]'s substituted correctly, or

any two [ ]'s and proper K substituted correctly

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4a) three points

$$K_a = ([H^+][C_3H_5O_2^-]) \div [HC_3H_5O_2] \quad 1.3 \times 10^{-5} = x^2 \div 0.20 \quad x = [H^+] = 1.6 \times 10^{-3}$$

b) one point

$$\% \text{ dissociation} = [H^+] \div [HC_3H_5O_2] = 1.6 \times 10^{-3} \div 0.20 = 0.80\%$$

c) two points

$$[H^+] = \text{antilog}(-5.20) = 6.3 \times 10^{-6}$$

$$1.3 \times 10^{-5} = (6.3 \times 10^{-6}) \times ([C_3H_5O_2^-] \div [HC_3H_5O_2])$$

$$[C_3H_5O_2^-] \div [HC_3H_5O_2] = 1.3 \times 10^{-5} \div 6.3 \times 10^{-6} = 2.1$$

An alternate solution for (c) based on the Henderson-Hasselbalch equation.

$$pH = pK_a + \log([base] \div [acid]) \quad 5.20 = 4.89 + \log([C_3H_5O_2^-] \div [HC_3H_5O_2])$$

$$\log([C_3H_5O_2^-] \div [HC_3H_5O_2]) = 0.31$$

$$[C_3H_5O_2^-] \div [HC_3H_5O_2] = 2.0$$

d) six points

$$0.10 \text{ L} \times 0.35 \text{ mol/L} = 0.035 \text{ mol } HC_3H_5O_2$$

$$0.10 \text{ L} \times 0.50 \text{ mol/L} = 0.050 \text{ mol } C_3H_5O_2^-$$

$$0.035 \text{ mol} - 0.004 \text{ mol} = 0.031 \text{ mol } HC_3H_5O_2$$

$$0.050 \text{ mol} + 0.004 \text{ mol} = 0.054 \text{ mol } C_3H_5O_2^-$$

$$1.3 \times 10^{-5} = [H^+] \times [(0.054 \text{ mol}/0.1 \text{ L}) \div (0.031 \text{ mol}/0.1 \text{ L})]$$

Can use 0.54 and 0.31 instead.

$$[H^+] = 7.5 \times 10^{-6} \quad pH = 5.13$$

An alternate solution for (d) based on the Henderson-Hasselbalch equation.

use [ ]s or moles of  $HC_3H_5O_2$  and  $C_3H_5O_2^-$

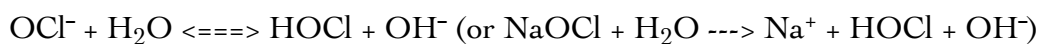
$$pH = pK_a + \log(0.054 / 0.031) = 4.89 + 0.24 = 5.13$$

5a) two points total ; one point for correct substitutions; one point for computation

$$[H^+] = [OCl^-] = \text{square root}(0.14 \times 3.2 \times 10^{-8}) = 6.7 \times 10^{-5} \text{ M}$$

$$\text{since } K_a = ([H^+][OCl^-]) / [HOCl] = [H^+]^2 / C_{HOCl}$$

(b) two points total: one point each



$$K_b = K_w / K_a = 1 \times 10^{-14} \div 3.2 \times 10^{-8} = 3.1 \times 10^{-7}$$

(c) two points total; one for concentrations and one for pH calc.

Concentrations before reaction:

$$[HOCl] = [(0.0400)(0.14)] / 0.050 = 0.11 \text{ M}$$

$$[OH^-] = [(0.0100)(0.56)] / 0.050 = 0.11 \text{ M}$$

Thus reaction is essentially complete and exactly equals a solution of NaOCl and  $[OCl^-] = 0.11 \text{ M}$  (or reaction is at equivalence point).

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Then  $[\text{OH}^-] = [\text{HOCl}]$   $K_b = [\text{OH}^-]^2 / 0.11 = 3.1 \times 10^{-7}$

$[\text{OH}^-] = \text{square root} [(0.11)(3.1 \times 10^{-7})] = 1.8 \times 10^{-4}$

$\text{pOH} = 3.73$   $\text{pH} = 14 - 3.73 = 10.27$

(d) two points; one for half-neutralized; one for mmol calcs.

$\text{pH} = 7.49$  therefore  $[\text{H}^+] = 3.2 \times 10^{-8}$

$\text{pH} = \text{p}K_a$ , or  $[\text{H}^+] = K_a$ . So  $[\text{OCl}^-] / [\text{HOCl}] = 1$ , or solution must be half neutralized.

initial mmol HOCl =  $50.0 \times 0.20 = 10.0$  mmol

mmol NaOH required =  $10.0 \div 2 = 5.0$  mmol

(e) one point

From equation, 1 mol  $\text{H}^+$  produced for each 1 mole of HOCl produced, thus  $[\text{H}^+] = [\text{HOCl}] = 0.065$  therefore  $\text{pH} = 1.19$

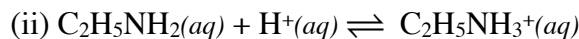
$$6. \text{ (a) molarity} = \frac{\text{moles}}{\text{liter}} = \frac{14.85\text{g} \times \frac{1 \text{ mol}}{45.0 \text{ g}}}{0.500 \text{ L}} = 0.660 \text{ M}$$

$$\text{(b) } K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]}$$

(c)  $\text{C}_2\text{H}_5\text{NH}_2$ ; weak base means  $K_b$  is very small so little of the ethylamine dissociates

(d) (i)  $\text{pOH} = 14 - \text{pH} = 14 - 10.93 = 3.07$

$\text{pOH} = -\log[\text{OH}^-] = 3.07$ ;  $[\text{OH}^-] = 8.51 \times 10^{-4} \text{ M}$



(iii)  $(0.500 \text{ L})(0.200 \text{ M}) = 0.100 \text{ mol H}^+ = \# \text{ moles of } \text{C}_2\text{H}_5\text{NH}_3^+$  since the stoichiometry is 1:1

$0.100 \text{ mol } \text{C}_2\text{H}_5\text{NH}_3^+ / 1 \text{ L of solution} = 0.100 \text{ M}$

(iv)  $[\text{C}_2\text{H}_5\text{NH}_3^+] = 0.100 \text{ M}$ ;  $[\text{OH}^-] = 8.5 \times 10^{-4}$

$$K_b = \frac{(0.100)(8.51 \times 10^{-4})}{(0.150)} = 5.67 \times 10^{-4}$$

7. (a)  $K_p = \frac{(P_{\text{CO}})^2}{P_{\text{CO}_2}}$

(b)  $n = \frac{PV}{RT} = \frac{(5.00 \text{ atm})(2.00 \text{ L})}{(0.0821 \frac{\text{L atm}}{\text{mol K}})(1160 \text{ K})} = 0.105 \text{ mol CO}_2$

(c) (i)  $P_{\text{CO}} + P_{\text{CO}_2} = 8.37 \text{ atm}$ ;  $P_{\text{CO}} = 8.37 \text{ atm} - 1.63 \text{ atm} = 6.74 \text{ atm}$

(ii)  $K_p = \frac{(6.74)^2}{1.63} = 27.9$

(d) same; catalyst increases the rate of the forward rate the same as the reverse rate.

(e) decrease; the trial  $K_p = (2.00)^2 / 2.00 = 2$ , since this is less than the calculated  $K_p$ , the reaction will produce more CO and consume  $\text{CO}_2$  to reach equilibrium.

8 (a)  $K_{sp} = [\text{Mg}^{2+}][\text{F}^-]^2 = (1.21 \times 10^{-3})(2.42 \times 10^{-3})^2$

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$$= 7.09 \times 10^{-9}$$

(b)  $X$  = concentration loss by  $\text{Mg}^{2+}$  ion

$2X$  = concentration loss by  $\text{F}^-$  ion

$$[\text{Mg}^{2+}] = (1.21 \times 10^{-3} - X) \text{ M}$$

$$[\text{F}^-] = (0.100 + 2.42 \times 10^{-3} - 2X) \text{ M}$$

since  $X$  is a small number then  $(0.100 + 2.42 \times 10^{-3} - 2X) \approx 0.100$

$$K_{sp} = 7.09 \times 10^{-9} = (1.21 \times 10^{-3} - X)(0.100)^2$$

$$X = 1.2092914 \times 10^{-3}$$

$$[\text{Mg}^{2+}] = 1.21 \times 10^{-3} - 1.20929 \times 10^{-3} = 7.09 \times 10^{-7} \text{ M}$$

(c)  $[\text{Mg}^{2+}] = 3.00 \times 10^{-3} \text{ M} \times 100.0 \text{ mL} / 300.0 \text{ mL} = 1.00 \times 10^{-3} \text{ M}$

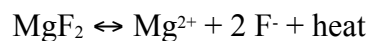
$$[\text{F}^-] = 2.00 \times 10^{-3} \text{ M} \times 200.0 \text{ mL} / 300.0 \text{ mL} = 1.33 \times 10^{-3} \text{ M}$$

$$\text{trial } K_{sp} = (1.00 \times 10^{-3})(1.33 \times 10^{-3})^2 = 1.78 \times 10^{-9}$$

trial  $K_{sp} < 7.09 \times 10^{-9}$ ,  $\therefore$  no ppt.

(d) @  $18^\circ\text{C}$ ,  $1.21 \times 10^{-3} \text{ M}$   $\text{MgF}_2$  dissolves

@  $27^\circ\text{C}$ ,  $1.17 \times 10^{-3} \text{ M}$   $\text{MgF}_2$  dissolves



dissolving is exothermic; if heat is increased it forces the equilibrium to shift left (according to LeChatelier's Principle) and less  $\text{MgF}_2$  will dissolve.