AP Chemistry - Packet #1 - Equilibrium

1) a) 5 points

$K_p = (P_{NH3})\ (P_{H2S})$ $P_{NH3} = P_{H2S} = (0.659\ atm\ /\ 2) = 0.330\ atm$

$K_p = (0.330\ atm)\ (0.330\ atm) = 0.109\ atm^2$

b) 5 points

$P_{NH3} = 2\ P_{H2S}$ $(2x)\ (x) = 0.109$ $x = 0.233\ atm = P_{H2S}$ $2x = 0.466\ atm = P_{NH3}$

c) 5 points

equilibrium pressure of NH_3 = equilibrium pressure of H_2S = 0.330 atm

P_{NH3} that has reacted = P_{H2S} = 0.500 - 0.330 = 0.170 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

$SrSO_4(s) \Longleftrightarrow Sr(aq)^{2+} + SO_4^{2-}(aq)$ at equilibrium: $[SO_4^{2-}] = x = [Sr^{2+}]$

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

$(x) = 8.7 \times 10^{-4}$ mol / liter = solubility of $SrSO_4$

b) three ponts

$SrF_2(s) \Longleftrightarrow Sr(aq)^{2+} + 2\ F(aq)^-$ at equilibrium: $[Sr^{2+}] = x$, $[F^-] = 2x$

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

$x = 5.8 \times 10^{-4}$ mol / liter = solubility of SrF_2

c) two points

Solve for $[Sr^+$ required for precipitation of each salt.

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

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

$K_{sp} = [Sr^{2+}][SO_4^{2-}] = 7.6 \times 10^{-7} = (y)\ (0.10\ mole/1.0\ liter) = 7.6 \times 10^{-7}$

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

Since 2.0×10^{-6} M < 7.6×10^{-6} M, SrF_2 must precipitate first.

When SrF_2 precipitates, $[Sr^{2+}] = 2.0 \times 10^{-6}$ M

d) two points

The second precipitate to form is $SrSO_4$, which appears when $[Sr^{2+}] = 7.6 \times 10^{-6}$ M (from Part c.)

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

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

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

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

3a) two points

mole fraction for CO = moles CO ÷ moles (CO + CO_2 + H_2O + H_2)

= 0.55 ÷ (0.55 + 0.55 + 0.20 + 0.30) = 0.55 ÷ 1.60 = 0.34 (or 34%)

Either mole fraction for CO definition above or numerator or denominator correct: one pt.
Numerator, denomiator and calculation correct or fraction 11/32: one pt.
One pt. deducted if units included with answer.

b) two points

K_c = ([H_2O] [CO]) ÷ ([H_2] [CO_2]) = [(0.55) (0.55)] ÷ [(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_2O missing.

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

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

d) two points

moles CO reacting = moles H_2O reacting = moles CO_2 formed = 0.16 moles

equilibrium [ ]'s at lower T: [H_2] = 0.36 mol/L [CO_2] = 0.46 mol/L [H_2O] = [CO] = 0.39 mol/L

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

e) two points

x = number of moles that react

	[H_2]	[CO_2]	[H_2O]	[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

$[x^2 / 9] \div [(0.50\text{-}x) / 3.0]^2 = 5.0$ x ÷ (0.50 - x) = 2.2 x = 0.34 mol

[CO] = 0.34 mol ÷ 3.0 L = 0.11 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

4a) three points

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

b) one point

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

c) two points

$[H^+]$ = antilog (- 5.20) = 6.3×10^{-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.

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

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

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

d) six points

0.10 L x 0.35 mol/L = 0.035 mol $HC_3H_5O_2$ 0.10 L x 0.50 mol/L = 0.050 mol $C_3H_5O_2^-$

0.035 mol - 0.004 mol = 0.031 mol $HC_3H_5O_2$ 0.050 mol + 0.004 mol = 0.054 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.

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

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

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

(b) two points total: one point each

$OCl^- + H_2O$ <===> $HOCl + OH^-$ (or $NaOCl + H_2O$ ---> $Na^+ + HOCl + OH^-$)

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

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

Concentrations before reaction:

$[HOCl] = [(0.0400)\ (0.14)] / 0.050 = 0.11$ M
$[OH^-] = [(0.0100)\ (0.56)] / 0.050 = 0.11$ M

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

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

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

pOH = 3.73 pH = 14 - 3.73 = 10.27

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

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

pH = pK_a , or $[H^+] = K_a$. So $[OCl^-] / [HOCl] = 1$, or solution must be half neutralized.

initial mmol HOCl = 50.0 x 0.20 = 10.0 mmol

mmol NaOH required = 10.0 ÷ 2 = 5.0 mmol

(e) one point

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

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

(b) $K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$

(c) $C_2H_5NH_2$; weak base means K_b is very small so little of the ethylamine dissociates

(d) (i) pOH = 14 – pH = 14 – 10.93 = 3.07

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

(ii) $C_2H_5NH_{2(aq)} + H^+_{(aq)} \rightleftharpoons C_2H_5NH_3^+{}_{(aq)}$

(iii) (0.500 L)(0.200 M) = 0.100 mol H^+ = # moles of $C_2H_5NH_3^+$ since the stoichiometry is 1:1

0.100 mol $C_2H_5NH_3^+$ /1 L of solution = 0.100 M

(iv) $[C_2H_5NH_3^+] = 0.100\ M$; $[OH^-] = 8.5 \times 10^{-4}$

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

7. (a) $K_p = \frac{(P_{CO})^2}{P_{CO_2}}$

(b) $n = \frac{PV}{RT} = \frac{(5.00\text{ atm})(2.00\text{ L})}{\left(0.0821 \frac{\text{L atm}}{\text{mol K}}\right)(1160\text{K})} = 0.105\text{ mol } CO_2$

(c) (i) $P_{CO} + P_{CO_2} = 8.37$ atm; $P_{CO} = 8.37$ atm – 1.63 atm = 6.74 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 CO_2 to reach equilibrium.

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

$= 7.09\times10^{-9}$

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

$2X$ = concentration loss by F^- ion

$[Mg^{2+}] = (1.21\times10^{-3} - X)$ M

$[F^-] = (0.100 + 2.42\times10^{-3} - 2X)$ M

since X is a small number then $(0.100 + 2.42\times10^{-3} - 2X) \approx 0.100$

$K_{sp} = 7.09\times10^{-9} = (1.21\times10^{-3} - X)(0.100)^2$

$X = 1.2092914\times10^{-3}$

$[Mg^{2+}] = 1.21\times10^{-3} - 1.20929\times10^{-3} = 7.09\times10^{-7}$M

(c) $[Mg^{2+}] = 3.00\times10^{-3}$M $\times$ 100.0 mL/300.0 mL = 1.00×10^{-3}M

$[F^-] = 2.00\times10^{-3}$ M $\times$ 200.0 mL/300.0 mL = 1.33×10^{-3} M

trial $K_{sp} = (1.00\times10^{-3})(1.33\times10^{-3})^2 = 1.78\times10^{-9}$

trial $K_{sp} < = 7.09\times10^{-9}$, $\therefore$ no ppt.

(d) @ 18ºC, 1.21×10^{-3} M MgF_2 dissolves

@ 27ºC, 1.17×10^{-3} M MgF_2 dissolves

$MgF_2 \leftrightarrow Mg^{2+} + 2\ F^- +$ heat

dissolving is exothermic; if heat is increased it forces the equilibrium to shift left (according to LeChatelier's Principle) and less MgF_2 will dissolve.