General Chemistry w/ Doctor Mattson

Unit 4 material: 16.10, 16.11, 16.12*, 16.13, 16.14, Chp 17**, Chp 18

* not formation constants pgs 693-698

** skip 17.3

Today: 3/20 Finish K_{sp}
Start Thermodynamics (17)

Tuesday 3/21 Review 5-6:30 pm HTTC 108

Wednesday 3/22 Chp 17

Thursday 3/23 Expt 9

Friday 3/24 Class!

Monday 3/27 Class cancelled!

What is the molar solubility of calcium phosphate?

$$Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+} + 2PO_4^{3-}$$

M	$Ca_3(PO_4)_2(s)$	$3Ca^{2+}$	$2PO_4^{3-}$
I	lots	0	0
C	$-x$	$+3x$	$+2x$
E	lots	$3x$	$2x$

$$K_{sp} = [Ca^{+2}]^3 [PO_4^{3-}]^2 = 2.1 \times 10^{-33}$$
$$= (3x)^3 (2x)^2$$
$$= (27x^3)(4x^2)$$
$$= 108x^5 = 2.1 \times 10^{-33} \Rightarrow \boxed{x = 1.14 \times 10^{-7}\ M}$$

* What is the molar solubility of barium fluoride in a sol'n that is 0.040 M KF?

K+ ✓ ↳F−

~~Type I~~ Question Type II: Common ion effect

	$BaF_2(s)$	$\rightleftharpoons$	Ba^{+2}	$+2F^-$
I	lots		0	0.040
C	$-x$		$+x$	$+2x$
E	lots		x	$2x+0.040$

$$K_{sp} = [Ba^{+2}][F^-]^2 = 1.8\times10^{-7}$$

$$(x)(0.040+2x)^2 = 1.8\times10^{-7}$$

↳ x is super small

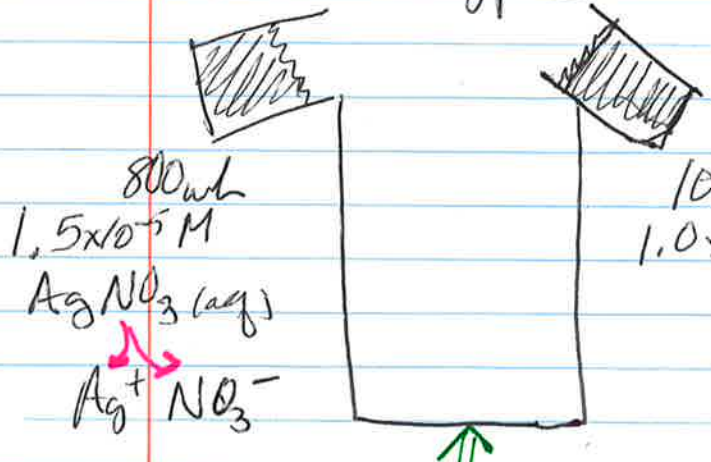
$$(x)(0.040)^2 = 1.8\times10^{-7}$$

$$x = 1.13\times10^{-4}\ M$$

$$[Ba^{+2}] = x = 1.13\times10^{-4}\ M$$

$$[F^-] = 0.040 + 2x = 0.040\ M$$

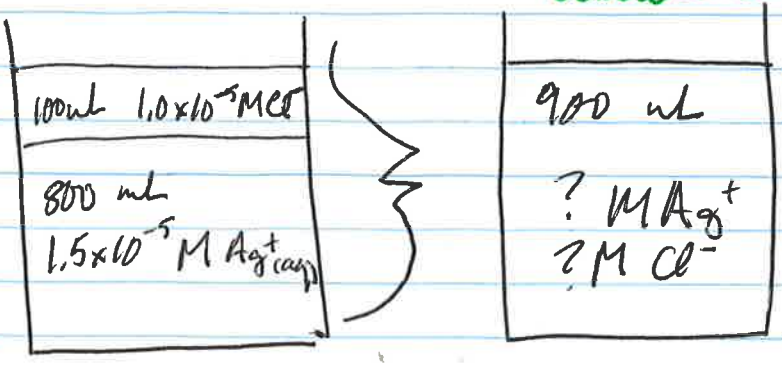
Question Type III: Does a precipitate form?



Review Solubility Rules if this isn't intuitive

If precipitate forms, its AgCl.

The 2 sol'ns mutually dilute each other



$Ag^+ \Rightarrow M_c V_c = M_d V_d$

$1.5\times10^{-5} \times 800\,mL = M_d \times 900\,mL$

$M_d = 1.3\times10^{-5}\,M$

$Cl^- \Rightarrow M_c V_c = M_d V_d$

$0.010\,M \times 100\,mL = M_d \times 900\,mL$

$M_d = 1.1\times10^{-3}\,M$

	$AgCl_{(s)}$	$\rightleftharpoons$	Ag^+	$+$	Cl^-
M					
I	0		1.3×10^{-5}		1.1×10^{-3}
C					
E					

$K_{sp} = [Ag^+][Cl^-] = 1.8\times10^{-10}$

$Q_{sp} = [Ag^+][Cl^-] = 1.3\times10^{-5} \times 1.1\times10^{-3} = 1.43\times10^{-8}$

If $Q_{sp} > K_{sp}$, shift left, forming a precipitate.

[True, we will have a ppt!]

If $Q_{sp} < K_{sp}$, no shift. no ppt, no equil.

Chapter 17

ΔH Enthalpy $\Delta H < 0$ exothermic (enthalpy favored)
$\Delta H > 0$ endothermic

→ Does it give off heat?
Does it get hot?

ΔS Entropy Is there a decrease in order? entropy favored
Is there an increase in order? (become more ordered) entropy not favored

$S_{gas} >> S_{liquid} >> S_{solid}$
high entropy — low entropy

ΔG spontaneous: $\Delta G < 0$ capable of happening on its own. Does it happen?
non spontaneous $\Delta G > 0$

Gibbs - Helmholtz Eqn

Dissolve sugar in hot tea

ΔG spontaneous. it happens $\Delta G < 0$

ΔH hard to answer

ΔS Entropy favored: solid sugar ⇒ aqueous
$\Delta S > 0$

$H_2O_{(s)} \longrightarrow H_2O_{(l)}$ at ~~15~~ 15°C

ΔG snow melts at 15°C. spontaneous
$\Delta G < 0$

ΔH to melt snow, you need to add heat.
endothermic $\Delta H > 0$

ΔS favored. S ⇒ ℓ $\Delta S > 0$

$CO_{2\,(g)} + NaOH_{(aq)} \longrightarrow NaHCO_{3\,(aq)}$

ΔG spontaneous. it happened
$\Delta G < 0$

ΔH must be exothermic to work $\Delta G = \Delta H - T\Delta S$
$\Delta H < 0$

ΔS not favored. 1 mol g ⇒ 0 mol gas
$\Delta S < 0$

$CO_{2(g)} + NaOH_{(aq)} \longrightarrow NaHCO_{3(aq)}$

ΔH_f° $-394\ \frac{kJ}{mol}$ $\quad -470\ kJ/mol \quad -951\ kJ/mol$

products − react ⇒ $x\ \frac{-1\ mol}{}$ $\quad x\ \frac{-1\ mol}{}$ $\quad x\ \frac{+1\ mol}{}$

$\Delta H? = 394\ kJ + 470\ kJ + -951\ kJ$

$\Delta H^\circ_{rxn} = 394\ kJ\ 470\ kJ - 951\ kJ$

$= -87\ kJ$ exothermic

S° $+214\ J/mol\,K$ $\quad +48\ J/mol\,K \quad 102\ J/mol\,K$

prod − react $\quad \frac{x\ -1\ mol}{-214\ J/K} \quad \frac{x\ -1\ mol}{-48\ J/K} \quad \frac{x\ 1\ mol}{102\ J/K}$

$\Delta S^\circ = -214\ J/K + (-48\ J/K) + 102\ J/K$

$= -160\ J/K$ not favored