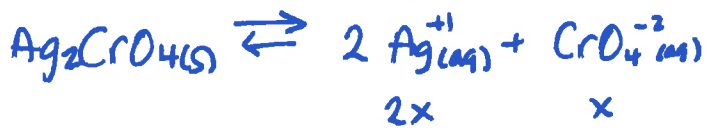


1. Fill out the chart: (the first compound is done as an example)

Compound	Ionization Equation	K_{sp} Expression	K_{sp} in terms of x
Zn(OH) ₂	$Zn(OH)_2 \rightleftharpoons Zn^{+2} + 2 OH^{-1}$	$K_{sp} = [Zn^{+2}][OH^{-1}]^2$	$K_{sp} = x(2x)^2 = 4x^3$
AgCl	$AgCl_{(s)} \rightleftharpoons Ag^{+1}_{(aq)} + Cl^{-1}_{(aq)}$	$[Ag^{+1}][Cl^{-}]$	$x \cdot x = x^2$
Al(OH) ₃	$Al(OH)_3 \rightleftharpoons Al^{+3} + 3OH^{-1}$	$[Al^{+3}][OH^{-}]^3$	$x(3x)^3 = 27x^4$
CaF ₂	$CaF_2 \rightleftharpoons Ca^{+2} + 2F^{-1}$	$[Ca^{+2}][F^{-}]^2$	$x(2x)^2 = 4x^3$
Ag ₂ CO ₃	$Ag_2CO_3 \rightleftharpoons 2Ag^{+1} + CO_3^{-2}$	$[Ag^{+1}]^2[CO_3^{-2}]$	$(2x)^2(x) = 4x^3$
Pb(OH) ₄	$Pb(OH)_4 \rightleftharpoons Pb^{+4} + 4OH^{-1}$	$[Pb^{+4}][OH^{-}]^4$	$x(4x)^4 = 256x^5$

2. Silver chromate (Ag₂CrO₄), has a K_{sp} of 3.3×10^{-12} .

a. Calculate the molar solubility of silver chromate into water.



The molar solubility is .000094 M;

9.4×10^{-5} moles Ag₂CrO₄ can dissolve per liter of solution.

$$K_{sp} = [Ag^+]^2 [CrO_4^{2-}] = (2x)^2 x = 4x^3$$

$$K_{sp} = 4x^3 = 3.3 \times 10^{-12} \quad x = 0.00009379 \text{ M} \approx 9.4 \times 10^{-5} \text{ M}$$

b. What are the concentrations of silver ion, and of chromate ion, in a saturated solution of silver chromate?

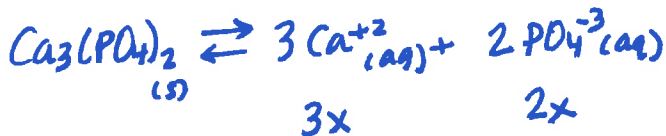
$[CrO_4^{2-}] = x = .000094 \text{ M}$

$$[Ag^+] = 2x = .0001876$$

$[Ag^+] = .00019 \text{ M}$

3. Calcium phosphate (Ca₃(PO₄)₂) has a K_{sp} of 2.0×10^{-29} .

a. Calculate the molar solubility of calcium phosphate into water.



$$K_{sp} = [Ca^{+2}]^3 [PO_4^{-3}]^2 = 2.0 \times 10^{-29}$$

$$(3x)^3 (2x)^2 = 108 x^5 = 2.0 \times 10^{-29}$$

$$x = 7.137 \times 10^{-7} \text{ M}$$

the molar solubility of Ca₃(PO₄)₂ is $7.1 \times 10^{-7} \text{ M}$

b. What are the concentrations of calcium ion, and of phosphate ion, in a saturated solution of calcium phosphate?

$[Ca^{+2}] = 3x = 3(7.137 \times 10^{-7} \text{ M}) = 2.1 \times 10^{-6} \text{ M } Ca^{+2}$
 $[PO_4^{-3}] = 2x = 2(7.137 \times 10^{-7} \text{ M}) = 1.4 \times 10^{-6} \text{ M } PO_4^{-3}$

4. The solubility of magnesium fluoride in water is 0.13 grams per liter.

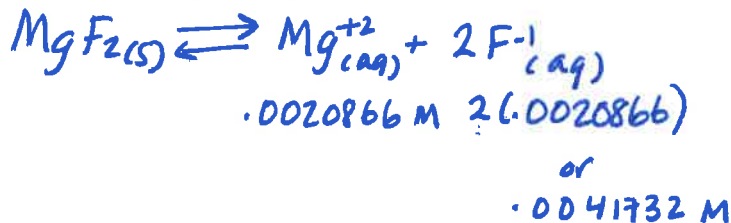
a. What is the formula for magnesium fluoride? MgF₂

b. Calculate the molar solubility of magnesium fluoride into water.

$$\left(\frac{0.13 \text{ g MgF}_2}{\text{Liter}} \right) \left(\frac{1 \text{ mole}}{62.3018 \text{ g}} \right) = 0.0020866 \frac{\text{moles MgF}_2}{\text{Liter}}$$

So the "molar solubility" of MgF₂ is 0.0021 M

c. Calculate K_{sp} of magnesium fluoride.



$$K_{sp} = [\text{Mg}^{2+}][\text{F}^{-}]^2 = (.0020866)(.0041732)^2 = 3.63 \times 10^{-8}$$

$$\boxed{K_{sp} = 3.6 \times 10^{-8}}$$

5. The maximum mass of magnesium phosphate able to dissolve into 4.00 liters of solution is 4.12 milligrams.

a. What is the formula for magnesium phosphate? Mg₃(PO₄)₂

b. Calculate the molar solubility of magnesium phosphate into water.

$$(4.12 \text{ mg}) \left(\frac{1 \text{ g}}{1000 \text{ mg}} \right) \left(\frac{1 \text{ mole}}{262.85779} \right) = 0.000015674 \text{ moles}$$

$$\frac{.000015674 \text{ moles}}{4.00 \text{ L}} = 0.0000039185 \text{ M} \rightarrow \boxed{3.92 \times 10^{-6} \text{ M}}$$

c. Calculate K_{sp} of magnesium phosphate.



$$\left(3.9185 \times 10^{-6} \frac{\text{mole Mg}_3(\text{PO}_4)_2}{\text{L}} \right) \left(\frac{3 \text{ mole Mg}^{2+}}{1 \text{ mole Mg}_3(\text{PO}_4)_2} \right) = 1.1755 \times 10^{-5} \text{ M Mg}^{2+}$$

$$\left(3.9185 \times 10^{-6} \frac{\text{mole Mg}_3(\text{PO}_4)_2}{\text{L}} \right) \left(\frac{2 \text{ mole PO}_4^{-3}}{1 \text{ mole Mg}_3(\text{PO}_4)_2} \right) = 7.8369 \times 10^{-6} \text{ M PO}_4^{-3}$$

$$K_{sp} = [\text{Mg}^{2+}]^3 [\text{PO}_4^{-3}]^2 = (1.1755 \times 10^{-5})^3 (7.8369 \times 10^{-6})^2$$

$$K_{sp} = 9.976 \times 10^{-26} \rightarrow \boxed{9.98 \times 10^{-26}}$$

6. Two saturated solutions of magnesium phosphate are prepared, both at room temperature. One has a volume of 10 mL, and one has a volume of 200 mL. Circle the correct answer for each question:

Both solutions are Saturated so the molarities should be equal. ($Q_{sp} = K_{sp}$)

a. Which solution has a higher concentration (molarity)?

the 10 mL one the 200 mL one they have equal molarities

b. Which solution contains more mass of solute?

the 10 mL one the 200 mL one they contain equal solute masses

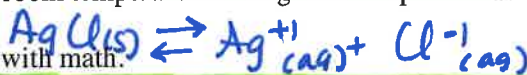
$$\# \text{ moles solute} = (\text{Volume in Liters}) \left(\text{Molarity} \left(\frac{\text{mole}}{\text{L}} \right) \right)$$

the molarities are equal, but the 200 mL solution has 20x the volume, so 20x as many moles of solute, and 20x the mass of solute.

7. AgCl has a K_{sp} of 1.8×10^{-10} .

A solution with a volume of 75 mL contains silver chloride dissolved at a concentration of 0.0000134 M. No solid is present. Assume that everything is at room temperature throughout this problem.

a. Is this solution saturated? Justify your answer with math.



yes no

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.8 \times 10^{-10} = x^2$$

$x = .000134 \text{ M}$, which matches the given concentration, so yes.

$$Q_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$\leftarrow \text{OR} \rightarrow = (.0000134)(.0000134)$$

$$= 1.80 \times 10^{-10}$$

$Q_{sp} = K_{sp}$, so yes, it is saturated.

b. Suppose that distilled water is added to the original solution so that the total volume of solution is 150 mL. The new solution is stirred so that it is uniform throughout.

Which solution is/was more concentrated?

the 75 mL one the 150 mL one they are equally concentrated

Which solution is/was more dilute?

the 75 mL one the 150 mL one they are equally dilute

Which solution contained/contains more mass of solute?

the 75 mL one the 150 mL one they contain equal masses of solute

The same # of moles of AgCl are now spread out into twice the volume, so the new molarity (in the 150 mL) is half as large as the original molarity.

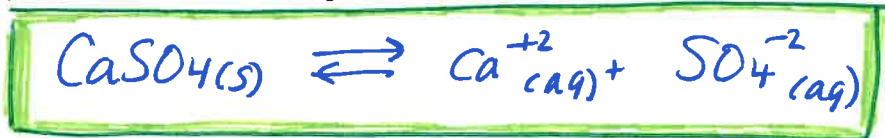
In the 75 mL solution, $[\text{Ag}^+] = [\text{Cl}^-] = 0.0000134 \text{ M}$.

In the 150 mL soln, $[\text{Ag}^+] = [\text{Cl}^-] = \frac{1}{2}(0.0000134) = .0000067 \text{ M}$

(Since no solid was present, the mixture is unable to reestablish equilibrium. If solid had been present, the rxn would "shift right," meaning that some AgCl(s) would dissolve until $Q_{sp} = K_{sp}$ again.)

8. Calcium sulfate has the formula CaSO_4 .

a. Write the chemical equation for this compound dissolving into water, with phase subscripts. (similar to the "ionization equation" format in problem 1.)



A large amount of solid CaSO_4 is added to water, and stirred in so that the calcium sulfate dissolves as much as possible. At least 90% of the CaSO_4 remains as undissolved solid, and settles to the bottom of the container. The aqueous part of the solution has a volume of 500. mL.

b. Is the solution saturated?

yes

no

Explain how you can tell based on the info given so far.

The CaSO_4 "dissolves as much as possible."

c. In the solution, the calcium ion and sulfate ion are each dissolved at a concentration of 0.0030 M. Calculate K_{sp} of calcium sulfate.

$$K_{sp} = [\text{Ca}^{+2}][\text{SO}_4^{-2}] = (0.0030)(0.0030) = 9.0 \times 10^{-6}$$

For d-f, Suppose you add another 500. mL of water, so that the total aqueous solution volume is 1000. mL.

d. What would be the immediate effect on the concentrations of calcium and sulfate ion, before the K_{sp} reaction has had a chance to reestablish equilibrium? (what will the new concentrations be?)

Imagine that the 1000 mL mixes instantly throughout before the reaction has any chance to "respond."

$$M_1 V_1 = M_2 V_2 \quad (0.0030 \text{ M})(500. \text{ mL}) = M_2 (1000. \text{ mL})$$
$$M_2 = .0015 \text{ M}$$

Since the volume doubles, both molarities will cut in half, to 0.0015 M.

e. What will be the concentration of calcium and sulfate ion once equilibrium has been reestablished?

More $\text{CaSO}_4(s)$ can dissolve until the Ca^{+2} and SO_4^{-2} return to their equilibrium values of 0.0030 M.

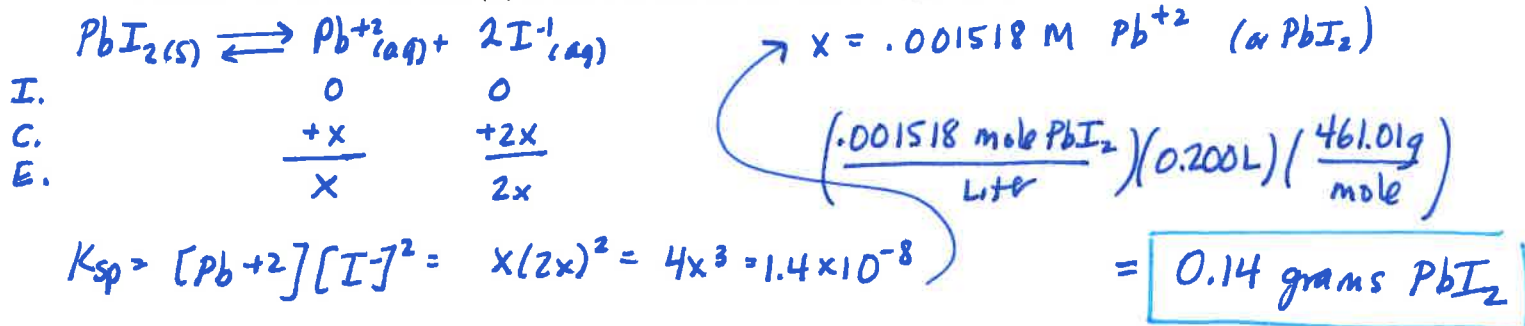
$$[\text{Ca}^{+2}] = [\text{SO}_4^{-2}] = 0.0030 \text{ M.}$$

f. How did the mass of solid calcium sulfate change, as equilibrium was reestablished?

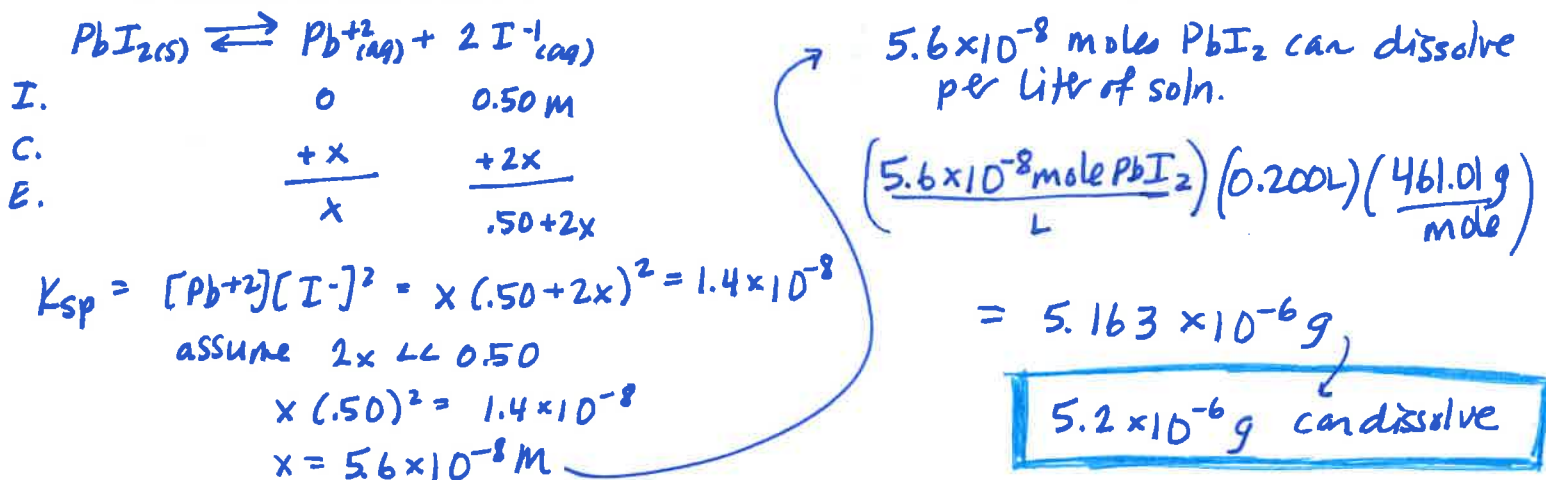
It decreased, since some of it dissolved.

9. At 15°C, the K_{sp} of Lead(II)iodide is 7.47×10^{-9} .
At 25°C, the K_{sp} of Lead(II) iodide is 1.4×10^{-8} .

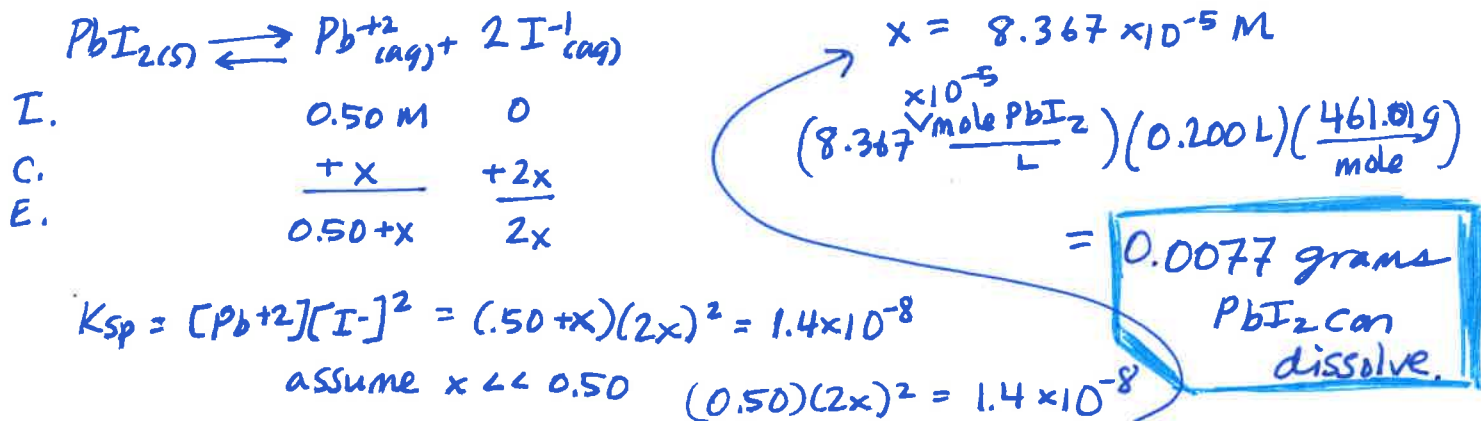
a. Determine the mass of lead (II) iodide that can dissolve into 200. ^{ml} of water, at 25°C.



b. Determine the mass of lead iodide that can dissolve into 200. mL of 0.50 Molar KI, at 25°C.



c. Determine the mass of lead iodide that can dissolve into 200. mL of 0.50 M lead (II) nitrate, at 25°C.



d. Based on the K_{sp} values at 15°C and 25°C,

does lead iodide have a positive or negative heat of solution, $\Delta H_{\text{solution}}$? positive
When lead iodide dissolves into water, does this process absorb or release heat? absorb

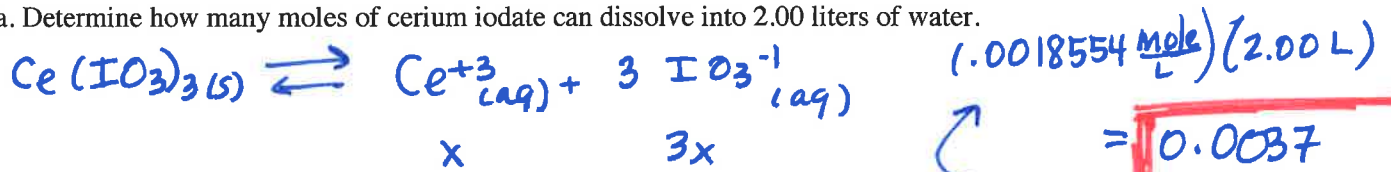
(Since K_{eq} increases as temp increases, ΔH_{soln} is + (endo)

e. Discuss the results to 9a vs 9b vs 9c. Include an explanation of the "common ion effect."

In a, a larger amt of solid can dissolve because no "common ions" are already present in solution. But in (b) and (c), less solid can dissolve because I^{-} or Pb^{+2} ions are already present in solution. (Pb^{+2} and I^{-} are "common ions" with PbI_2 ...) Since $[Pb^{+2}]$ and $[I^{-}]$ multiply to a constant value (K_{sp}) at equilibrium, if the solution already contains one of these ions, the other ion will have to decrease in concentration for Q_{sp} to equal K_{sp} .

10. Cerium iodate, $\text{Ce}(\text{IO}_3)_3$, has a K_{sp} of 3.2×10^{-10}

a. Determine how many moles of cerium iodate can dissolve into 2.00 liters of water.

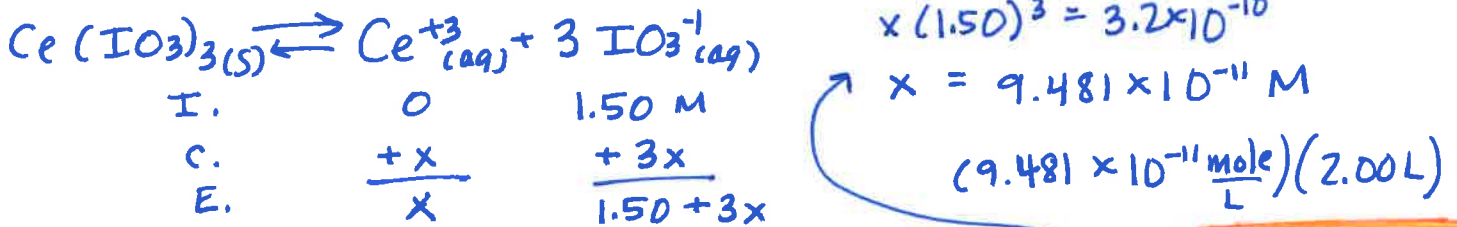


$$K_{sp} = [\text{Ce}^{+3}][\text{IO}_3^{-1}]^3 = x(3x)^3 = 3.2 \times 10^{-10}$$

$$27x^4 = 3.2 \times 10^{-10}$$

$$x = .0018554 \text{ M}$$

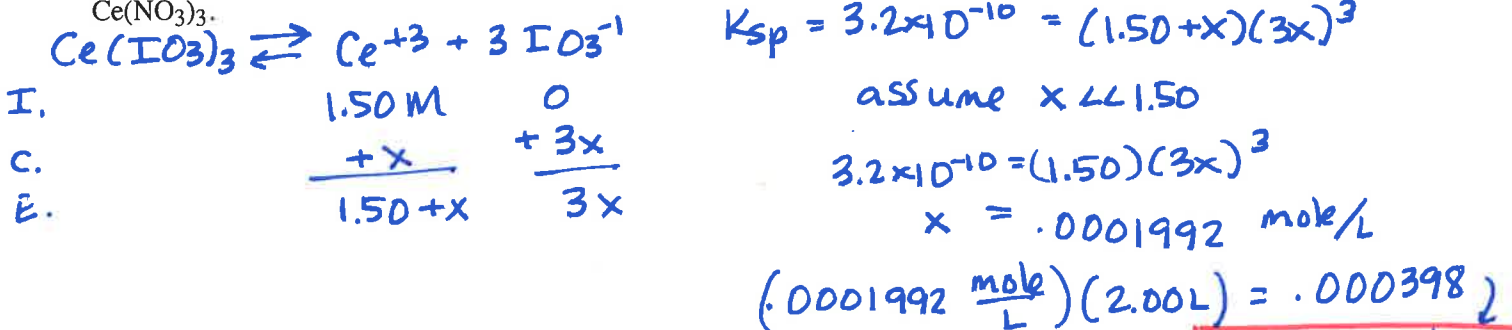
b. Determine how many moles of cerium iodate can dissolve into 2.00 liters of 1.50 Molar NaIO_3 .



$$K_{sp} = [\text{Ce}^{+3}][\text{IO}_3^{-1}]^3 = x(1.50 + 3x)^3 = 3.2 \times 10^{-10}$$

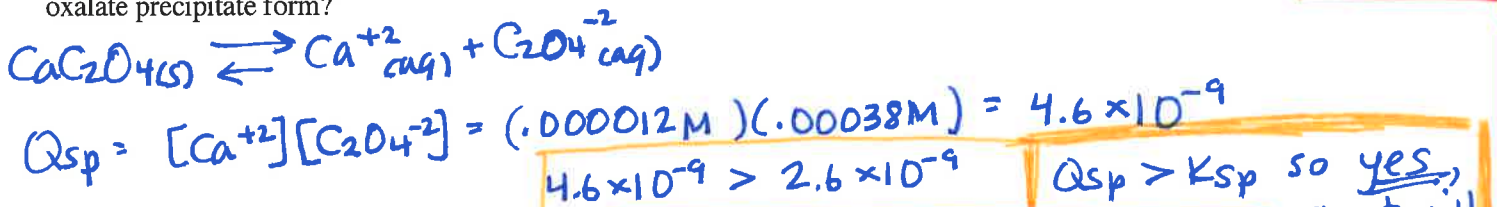
assume $3x \ll 1.50$
 $= 1.9 \times 10^{-10}$
moles $\text{Ce}(\text{IO}_3)_3$ can dissolve

c. Determine how many moles of cerium iodate can dissolve into 2.00 liters of 1.50 Molar Cerium nitrate, $\text{Ce}(\text{NO}_3)_3$.

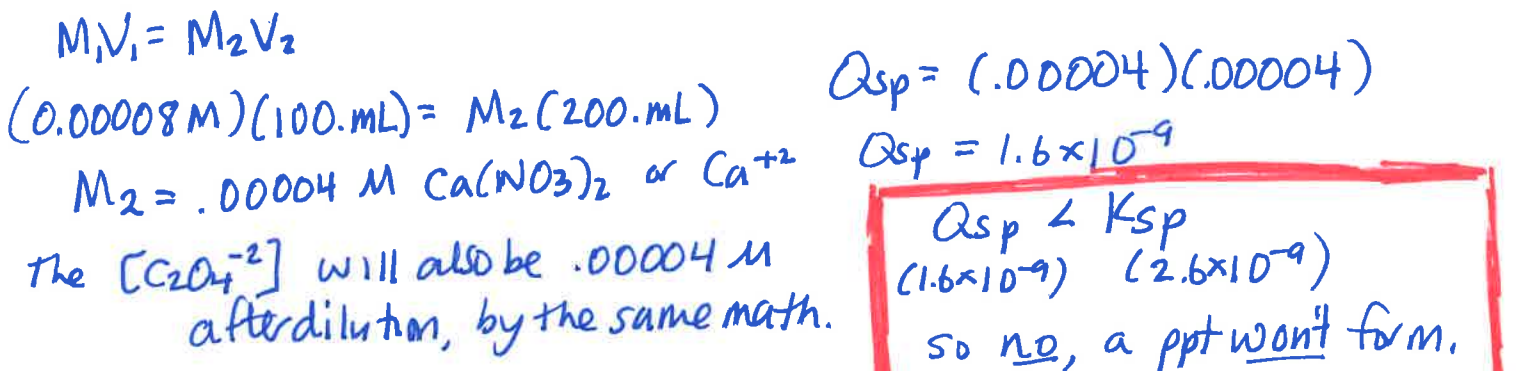


11. Calcium oxalate, CaC_2O_4 , has a K_{sp} of 2.6×10^{-9} .

a. If a solution contains 0.000012 moles Ca^{+2} per liter, and 0.00038 moles $\text{C}_2\text{O}_4^{-2}$ per liter, will a calcium oxalate precipitate form?



b. 100. mL of 0.00008 M $\text{Ca}(\text{NO}_3)_2$ are mixed with 100. mL of 0.00008 M $\text{Na}_2\text{C}_2\text{O}_4$. Will a calcium oxalate precipitate form? (Assume volumes are additive, so total $V = 200. \text{ mL}$)



12. Silver Chromate, Ag_2CrO_4 , has a K_{sp} of 1.7×10^{-12} .

If 0.000010 moles of silver chromate are stirred into 100. mL water, will all of the silver chromate be able to dissolve?

$$\frac{0.000010 \text{ mole}}{0.100 \text{ L}} = 0.00010 \text{ M } \text{Ag}_2\text{CrO}_4$$

$$\text{So } [\text{Ag}^+] = 2(0.00010 \text{ M}) = 0.00020 \text{ M}$$

$$[\text{CrO}_4^{2-}] = 1(0.00010 \text{ M}) = 0.00010 \text{ M}$$

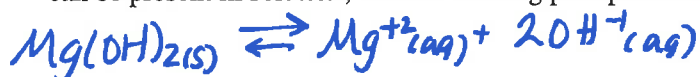
$$Q_{sp} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}] = (0.00020)^2(0.00010) = 4.0 \times 10^{-12}$$

$$Q_{sp} (4.0 \times 10^{-12}) > K_{sp} (1.7 \times 10^{-12})$$

so no, it won't all dissolve.

13. Magnesium hydroxide has a K_{sp} of 1.6×10^{-12} .

a. If the concentration of magnesium ion is 0.10 Molar, What is the maximum concentration of hydroxide that can be present in solution, without causing precipitation?



$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2 \quad 1.6 \times 10^{-12} = (0.10)[\text{OH}^-]^2$$

$$[\text{OH}^-]_{\text{max}} = 4.0 \times 10^{-6} \text{ M}$$

b. If it essential to keep magnesium ion dissolved in a solution, should the solution be kept at high pH or low pH, and why?

To keep Mg^{2+} dissolved, we need to keep $[\text{OH}^-]$ low, so that $Q_{sp} < K_{sp}$. $[\text{OH}^-]$ is lower in acidic pH, so we need to keep pH low.

14. Solutions containing silver nitrate or gold nitrate must be made with distilled or deionized water, because silver ion and gold ion can both precipitate with the chloride ions in tapwater. Given the following data:

$$\text{AuCl} \quad K_{sp} = 2.0 \times 10^{-13}$$

$$\text{AuCl}_3 \quad K_{sp} = 3.2 \times 10^{-25}$$

$$\text{AgCl} \quad K_{sp} = 1.8 \times 10^{-10}$$

a. Determine the minimum chloride concentration that would cause a precipitate to form in a solution of 0.010 Molar Gold (I) nitrate.

$$K_{sp} = [\text{Au}^+][\text{Cl}^-] \quad 2.0 \times 10^{-13} = (0.010)[\text{Cl}^-]$$

$$[\text{Cl}^-] = 2.0 \times 10^{-11} \text{ M}$$

b. Determine the minimum chloride concentration that would cause a precipitate to form in a solution of 0.010 Molar Gold (III) nitrate.

$$K_{sp} = [\text{Au}^{3+}][\text{Cl}^-]^3 \quad 3.2 \times 10^{-25} = (0.010)[\text{Cl}^-]^3$$

$$[\text{Cl}^-] = 3.2 \times 10^{-8} \text{ M}$$

c. Suppose that a 0.010 M silver nitrate solution is slowly added to a solution 0.00020 M solution of NaCl.

The silver nitrate solution is added 1 drop at a time into 1250 mL of the 0.00020M NaCl solution.

Assume that the solution mixes instantly after each drop is added.

What volume of silver nitrate, in mL, must be added for precipitation to begin? if 1 mL is equivalent to about 20 drops of solution, how many drops of solution must be added for precipitation to begin?

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.8 \times 10^{-10} = [\text{Ag}^+](0.00020)$$

$$[\text{Ag}^+] = 9.0 \times 10^{-7} \text{ M}$$

$$M_1V_1 = M_2V_2$$

$$(0.010 \text{ M})V_1 = (9.0 \times 10^{-7} \text{ M})(1250 \text{ mL})$$

$$V_1 = 0.1125 \text{ mL}$$

if $[\text{Ag}^+]$ is larger than $9.0 \times 10^{-7} \text{ M}$ after dilution, a ppt will form.

0.11 mL are necessary (or slightly more than that)

$$(0.1125 \text{ mL}) \left(\frac{20 \text{ drops}}{1 \text{ mL}} \right) = 2.25 \rightarrow 2 \text{ drops (or really, 3!)}$$

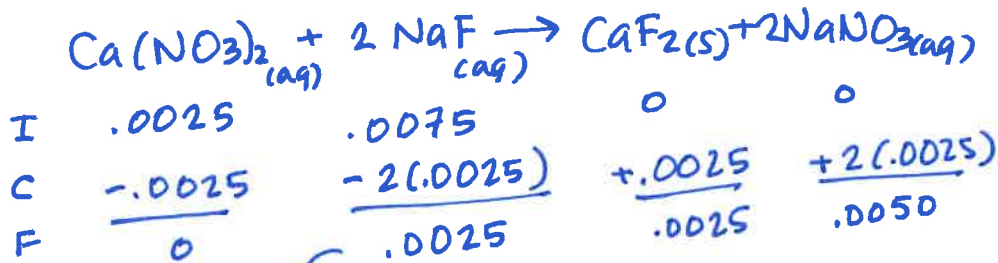
15. Calculate the concentration of calcium ion and of fluoride ion, if 50.0 mL of 0.050 M $\text{Ca}(\text{NO}_3)_2$ are mixed with 50.0 mL of 0.15 M NaF. K_{sp} of CaF_2 is 3.4×10^{-11} .

$$(0.0500 \text{ L})(0.050 \text{ M}) = 0.0025 \text{ moles}$$

$$(0.0500 \text{ L})(0.15 \text{ M}) = 0.0075 \text{ moles}$$

$$50.0 \text{ mL} + 50.0 \text{ mL} = 100.0 \text{ mL}$$

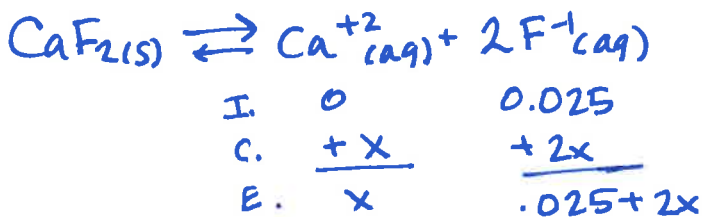
(assuming volumes are "additive")



$$[\text{Ca}^{+2}] = 5.4 \times 10^{-8} \text{ M}$$

$$[\text{F}^-] = 0.025 \text{ M}$$

$\frac{.0025 \text{ moles}}{0.1000 \text{ L}} = 0.025 \text{ M NaF}$, so $[\text{F}^-] = 0.025 \text{ M}$



$$K_{sp} = [\text{Ca}^{+2}][\text{F}^-]^2 = 3.4 \times 10^{-11}$$

$$x(0.025 + 2x)^2 = 3.4 \times 10^{-11}$$

assume $2x \ll 0.025$

$$x(0.025)^2 = 3.4 \times 10^{-11}$$

$$x = 5.4 \times 10^{-8} \text{ M}$$

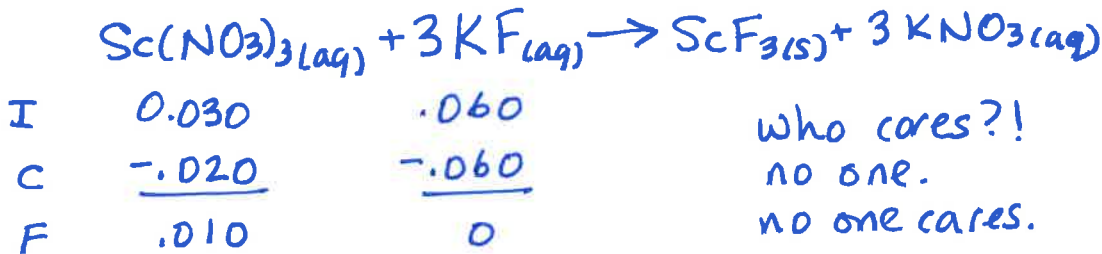
16. Scandium fluoride, ScF_3 , has a K_{sp} of 4.2×10^{-18} .

a. Calculate the concentration of scandium ion and of fluoride ion if 100.0 mL 0.30 M scandium nitrate react with 200. mL of 0.30 Molar potassium fluoride.

$$200. \text{ mL} + 100.0 \text{ mL} = 300. \text{ mL}$$

$$(0.1000 \text{ L})(0.30 \text{ M}) = 0.030 \text{ moles}$$

$$(0.200 \text{ L})(0.30 \text{ M}) = 0.060 \text{ moles}$$

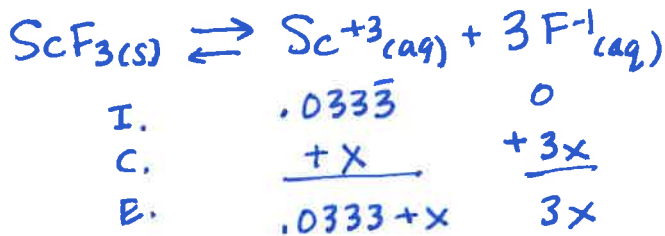


who cares?!
no one.
no one cares.

$$[\text{Sc}^{+3}] = 0.033 \text{ M}$$

$$[\text{F}^-] = 5.0 \times 10^{-6} \text{ M}$$

$\frac{.010 \text{ mole}}{0.300 \text{ L}} = 0.0333 \text{ M}$, so $[\text{Sc}^{+3}] = 0.033 \text{ M}$



$$K_{sp} = [\text{Sc}^{+3}][\text{F}^-]^3 = 4.2 \times 10^{-18}$$

$$(0.0333 + x)(3x)^3 = 4.2 \times 10^{-18}$$

assume $x \ll 0.0333$

$$(0.0333)(3x)^3 = 4.2 \times 10^{-18}$$

$$x = 1.671 \times 10^{-6} \text{ M}$$

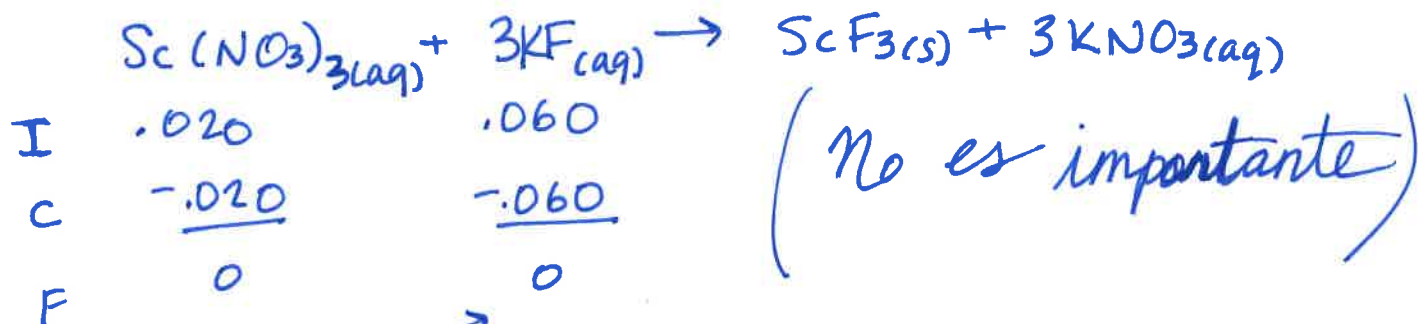
$$[\text{F}^-] = 3x = 5.0 \times 10^{-6} \text{ M}$$

16, cont'd Scandium fluoride, ScF_3 , has a K_{sp} of 4.2×10^{-18} .

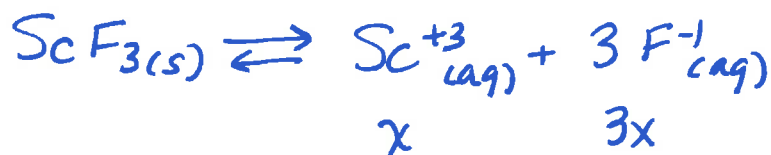
b. Calculate the concentration of scandium ion and of fluoride ion if 100.0 mL 0.20 M scandium nitrate react with 200. mL of 0.30 Molar potassium fluoride.

$$(0.1000 \text{ L})(0.20 \text{ M}) = 0.020 \text{ moles } \text{Sc}(\text{NO}_3)_3$$

$$(0.200 \text{ L})(0.30 \text{ M}) = 0.060 \text{ moles } \text{KF}$$



nothing left, so no common ion effect will occur.



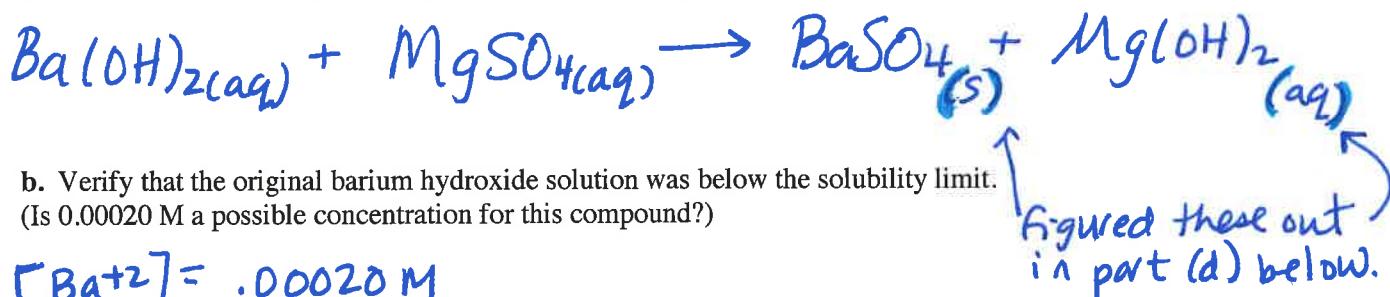
$$K_{sp} = [\text{Sc}^{+3}][\text{F}^{-}]^3 = \text{x}(3\text{x})^3 = 4.2 \times 10^{-18}$$
$$27\text{x}^4 = 4.2 \times 10^{-18}$$
$$\text{x} = 1.98596 \times 10^{-5} \text{ M}$$

$$[\text{Sc}^{+3}] = \text{x} = 2.0 \times 10^{-5} \text{ M}$$
$$[\text{F}^{-}] = 3\text{x} = 6.0 \times 10^{-5} \text{ M}$$

Compound	K _{sp}
Ba(OH) ₂	5 x 10 ⁻³
BaSO ₄	1.1 x 10 ⁻¹⁰
Mg(OH) ₂	1.8 x 10 ⁻¹¹
MgSO ₄	very soluble

17. Dilute solutions of magnesium sulfate and barium hydroxide are mixed:
300. mL of 0.00020 M MgSO₄ were mixed with 100. mL of 0.00020 M Ba(OH)₂.

a. Write the double replacement reaction between barium hydroxide and magnesium sulfate.
(you'll add phase subscripts to the products after you do part (d).)



b. Verify that the original barium hydroxide solution was below the solubility limit.
(Is 0.00020 M a possible concentration for this compound?)

$$[\text{Ba}^{+2}] = 0.00020 \text{ M}$$

$$[\text{OH}^-] = 2(0.00020 \text{ M}) = 0.00040 \text{ M}$$

$$Q_{\text{sp}} = [\text{Ba}^{+2}][\text{OH}^-]^2 = (0.00020)(0.00040)^2 = 3.2 \times 10^{-11}$$

$Q_{\text{sp}} (3.2 \times 10^{-11}) < K_{\text{sp}} (5 \times 10^{-3})$ so yes, it can be dissolved at this concentration.

c. Find the concentration of all four ions in solution: after diluting to the total volume of 400. mL, but before any precipitation occurs. (so, find the concentration of magnesium ion, sulfate ion, barium ion, and hydroxide ion.) 300. mL + 100. mL = 400. mL total solution volume

$$M_1 V_1 = M_2 V_2$$

$$\text{MgSO}_4: (0.00020 \text{ M})(300. \text{ mL}) = M_2(400. \text{ mL})$$

$$M_2 = 0.00015 \text{ M} \quad \text{so} \quad [\text{Mg}^{+2}] = [\text{SO}_4^{-2}] = 0.00015 \text{ M}$$

$$2(0.000050) = 0.00010$$

$$\text{Ba(OH)}_2: (0.00020 \text{ M})(100. \text{ mL}) = M_2(400. \text{ mL})$$

$$M_2 = 0.000050 \text{ M} \quad \text{so} \quad [\text{Ba}^{+2}] = 0.000050 \text{ M}, \quad [\text{OH}^-] = 0.00010 \text{ M}$$

d. Which precipitate, if any, will form? Once you figure this out, add subscripts to the rxn in (a).

Need to compare Q_{sp} to K_{sp} for BaSO₄ and Mg(OH)₂.

$$\text{BaSO}_4: Q_{\text{sp}} = [\text{Ba}^{+2}][\text{SO}_4^{-2}] = (0.000050)(0.00015) = 7.5 \times 10^{-9}$$

$Q_{\text{sp}} (7.5 \times 10^{-9}) > K_{\text{sp}} (1.1 \times 10^{-10})$ so BaSO₄ ppt will form

$$\text{Mg(OH)}_2: Q_{\text{sp}} = [\text{Mg}^{+2}][\text{OH}^-]^2 = (0.00015)(0.00010)^2 = 1.5 \times 10^{-12}$$

$Q_{\text{sp}} (1.5 \times 10^{-12}) < K_{\text{sp}} (1.8 \times 10^{-11})$ so Mg(OH)₂ will stay dissolved.

18. Lead II carbonate (PbCO_3) has a K_{sp} of 7.4×10^{-14} . It has a positive $\Delta H_{\text{solution}}$. Suppose that 100 grams of lead II carbonate are added to water in a beaker, and stirred so that it can dissolve as much as possible (to form a "saturated" solution). Most of the solid remains undissolved, at the bottom of the beaker. The total solution volume (not including the volume of the solid at the bottom) is 400. mL. If the following changes are made, how will the aqueous concentration of lead II ion be affected, once equilibrium has been reestablished? Assume that temperature is constant unless otherwise noted. (increase, decrease, or no change)



no change 100 mL of distilled water are added.

since eqm is reestablished, $[\text{Pb}^{+2}]$ and $[\text{CO}_3^{-2}]$ will be same as before adding H_2O .

increase some solid lead II nitrate is added, and it dissolves

The rxn will shift left to use up some Pb^{+2} and some CO_3^{-2} , but $[\text{Pb}^{+2}]$ will still be higher than before adding it.

decrease some solid potassium carbonate is added, and it dissolves

Adding CO_3^{-2} means the rxn shifts left, using up some Pb^{+2} .

decrease The solution is cooled down; temp decreases by about 10°C

since ΔH_{soln} is positive (given above), PbCO_3 is less soluble at lower temp. so $[\text{Pb}^{+2}]$ and $[\text{CO}_3^{-2}]$ will both decrease.

no change 20 grams of the solid lead II carbonate are removed.

still @ Equilibrium with 1:1 Pb^{+2} to CO_3^{-2} molarity

increase nitric acid ($\text{HNO}_3(aq)$) is added.

The nitric acid reacts with some of the carbonate according to the following:



as CO_3^{-2} is consumed, the rxn shifts right, making more Pb^{+2}

no change some of the water from the solution evaporates while sitting out over several days.

since it establishes equilibrium again, $[\text{Pb}^{+2}]$ and $[\text{CO}_3^{-2}]$ will be same as before

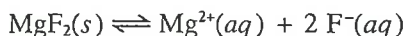
no change More solid lead II carbonate are added to the beaker

same explanation as \uparrow

Increase some solid cadmium nitrate ($\text{Cd}(\text{NO}_3)_2$) is added. The cadmium nitrate dissolves, and then the Cd^{+2} ions react with some of the carbonate ions to form a CdCO_3 precipitate.

since $\text{CO}_3^{-2}(aq)$ is consumed, the rxn will shift right, so $[\text{Pb}^{+2}]$ will increase

19



(AP 1994)

In a saturated solution of MgF_2 at 18°C , the concentration of Mg^{2+} is 1.21×10^{-3} molar. The equilibrium is represented by the equation above.

- Write the expression for the solubility-product constant, K_{sp} , and calculate its value at 18°C .
- Calculate the equilibrium concentration of Mg^{2+} in 1.000 liter of saturated MgF_2 solution at 18°C to which 0.100 mole of solid KF has been added. The KF dissolves completely. Assume the volume change is negligible.

(a) and (b) are optional

$$(a) K_{sp} = [\text{Mg}^{+2}][\text{F}^{-}]^2$$

given: $[\text{Mg}^{+2}] = .00121\text{ M}$ in a sat'd soln

$$[\text{F}^{-}] = 2(.00121\text{ M}) = .00242\text{ M}$$

$$K_{sp} = (.00121)(.00242)^2 = \boxed{7.1 \times 10^{-9}} \quad (7.086 \times 10^{-9} \text{ before rounding})$$

$$(b) [\text{F}^{-}] = \frac{.100 \text{ mole}}{1.000\text{ L}} = 0.100\text{ M}$$



I.	0	0.100
C.	$\frac{+x}{\quad}$	$\frac{+2x}{\quad}$
E.	x	.100 + 2x

$$K_{sp} = x(.100 + 2x)^2 = 7.086 \times 10^{-9}$$

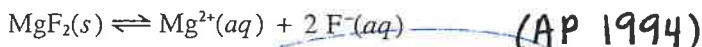
assume $2x \ll .100$

$$x(.100)^2 = 7.086 \times 10^{-9}$$

$$x = 7.086 \times 10^{-7}$$

$$\boxed{[\text{Mg}^{+2}] = 7.1 \times 10^{-7}\text{ M}}$$

19



In a saturated solution of MgF_2 at $18^{\circ}C$, the concentration of Mg^{2+} is 1.21×10^{-3} molar. The equilibrium is represented by the equation above.

Write the expression for the solubility-product constant, K_{sp} , and calculate its value at $18^{\circ}C$.

Calculate the equilibrium concentration of Mg^{2+} in 1.000 liter of saturated MgF_2 solution at $18^{\circ}C$ to which 0.100 mole of solid KF has been added. The KF dissolves completely. Assume the volume change is negligible.

(a) and (b) are optional

Predict whether a precipitate of MgF_2 will form when 100.0 milliliters of a 3.00×10^{-3} -molar $Mg(NO_3)_2$ solution is mixed with 200.0 milliliters of a 2.00×10^{-3} -molar NaF solution at $18^{\circ}C$. Calculations to support your prediction must be shown.

$100.0 + 200.0 = 300.0 \text{ mL total volume}$

$M_1V_1 = M_2V_2$ $(.00300 \text{ M})(100.0 \text{ mL}) = M_2(300.0 \text{ mL})$ $Mg(NO_3)_2$
 $M_2 = 0.00100 \text{ M} = [Mg^{2+}]$

$(.00200 \text{ M})(200.0 \text{ mL}) = M_2(300.0 \text{ mL})$ NaF
 $M_2 = .00133 \text{ M} = [F^{-}]$

$Q_{sp} = [Mg^{2+}][F^{-}]^2 = (.00100)(.00133)^2 = 1.8 \times 10^{-9}$

in part (a) we got a K_{sp} of 7.1×10^{-9} .

$Q_{sp} < K_{sp}$ so no ppt forms.

At $27^{\circ}C$ the concentration of Mg^{2+} in a saturated solution of MgF_2 is 1.17×10^{-3} molar. Is the dissolving of MgF_2 in water an endothermic or an exothermic process? Give an explanation to support your conclusion.

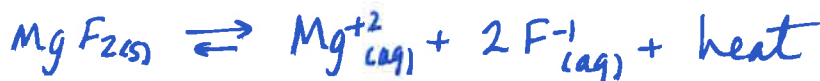
given: at $18^{\circ}C$, $[Mg^{2+}] = .00121 \text{ M}$ in a sat'd solution

at $27^{\circ}C$, $[Mg^{2+}] = .00117 \text{ M}$ in a sat'd solution

$.00117 < .00121$

since less Mg^{2+} can dissolve at the higher temp (solubility decreases as temp increases)

$\Delta H_{\text{solution}}$ must be negative; dissolving is exothermic



20. Solve the following problem related to the solubility equilibria of some metal hydroxides in aqueous solution. (AP 1998)

- (a) The solubility of $\text{Cu}(\text{OH})_2(s)$ is 1.72×10^{-6} gram per 100. milliliters of solution at 25°C .
- optional $\left\{ \begin{array}{l} \text{A) Write the balanced chemical equation for the dissociation of } \text{Cu}(\text{OH})_2(s) \text{ in aqueous solution.} \\ \text{B) Calculate the solubility (in moles per liter) of } \text{Cu}(\text{OH})_2 \text{ at } 25^\circ\text{C}. \\ \text{C) Calculate the value of the solubility-product constant, } K_{sp}, \text{ for } \text{Cu}(\text{OH})_2 \text{ at } 25^\circ\text{C}. \end{array} \right. \text{ optional}$

(b) The value of the solubility-product constant, K_{sp} , for $\text{Zn}(\text{OH})_2$ is 7.7×10^{-17} at 25°C .

optional $\left\{ \begin{array}{l} \text{A) Calculate the solubility (in moles per liter) of } \text{Zn}(\text{OH})_2 \text{ at } 25^\circ\text{C} \text{ in a solution with a pH of } 9.35. \leftarrow \text{optional} \end{array} \right.$

(ii) At 25°C , 50.0 milliliters of 0.100-molar $\text{Zn}(\text{NO}_3)_2$ is mixed with 50.0 milliliters of 0.300-molar NaOH . Calculate the molar concentration of $\text{Zn}^{2+}(aq)$ in the resulting solution once equilibrium has been established. Assume that volumes are additive.

only this one is assigned.



(ii)
$$\frac{(1.72 \times 10^{-6} \text{ g}) \left(\frac{1 \text{ mole}}{97.5606 \text{ g}} \right)}{0.100 \text{ L}} = 1.763 \times 10^{-7} \text{ M}$$

molar solubility = $1.76 \times 10^{-7} \text{ M}$

(iii) $[\text{Cu}^{2+}] = 1.763 \times 10^{-7} \text{ M}$

$[\text{OH}^{-}] = 2(1.763 \times 10^{-7} \text{ M}) = 3.526 \times 10^{-7} \text{ M}$

$K_{sp} = [\text{Cu}^{2+}][\text{OH}^{-}]^2 = (1.763 \times 10^{-7})(3.526 \times 10^{-7})^2 = 2.192 \times 10^{-20}$

$K_{sp} = 2.19 \times 10^{-20}$

(b) (i) if $\text{pH} = 9.35$, $[\text{OH}^{-}] = 2.24 \times 10^{-5} \text{ M}$ (you'll learn this when we do acids/bases/pH)

	$\text{Zn}(\text{OH})_2$	\rightleftharpoons	$\text{Zn}^{2+}(aq)$	+	$2\text{OH}^{-}(aq)$
I.			0		2.24×10^{-5}
C.			+x		+2x
E.			x		$2.24 \times 10^{-5} + 2x$

$K_{sp} = [\text{Zn}^{2+}][\text{OH}^{-}]^2$
 $7.7 \times 10^{-17} = x(2.24 \times 10^{-5} + 2x)^2$
 assume $2x \ll 2.24 \times 10^{-5}$

$7.7 \times 10^{-17} = x(2.24 \times 10^{-5})^2$
 $x = [\text{Zn}^{2+}] = 1.5 \times 10^{-7} \text{ M}$
 this also equals the molar solubility :)

20. Solve the following problem related to the solubility equilibria of some metal hydroxides in aqueous solution. (AP 1998)

- (a) The solubility of $\text{Cu}(\text{OH})_2(s)$ is 1.72×10^{-6} gram per 100. milliliters of solution at 25°C .
- optional ~~X~~ Write the balanced chemical equation for the dissociation of $\text{Cu}(\text{OH})_2(s)$ in aqueous solution. optional
- ~~OX~~ Calculate the solubility (in moles per liter) of $\text{Cu}(\text{OH})_2$ at 25°C .
- ~~X~~ Calculate the value of the solubility-product constant, K_{sp} , for $\text{Cu}(\text{OH})_2$ at 25°C .
- (b) The value of the solubility-product constant, K_{sp} , for $\text{Zn}(\text{OH})_2$ is 7.7×10^{-17} at 25°C .

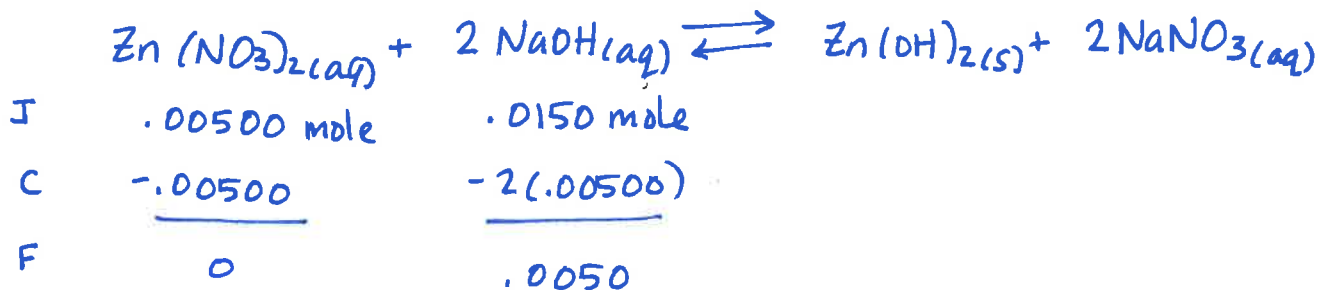
optional ~~X~~ Calculate the solubility (in moles per liter) of $\text{Zn}(\text{OH})_2$ at 25°C in a solution with a pH of 9.35. ← optional

only this one is assigned. (ii) At 25°C , 50.0 milliliters of 0.100-molar $\text{Zn}(\text{NO}_3)_2$ is mixed with 50.0 milliliters of 0.300-molar NaOH . Calculate the molar concentration of $\text{Zn}^{2+}(aq)$ in the resulting solution once equilibrium has been established. Assume that volumes are additive.

(b. ii.)

$$(.0500 \text{ L})(.100 \text{ M}) = .00500 \text{ moles Zn}(\text{NO}_3)_2$$

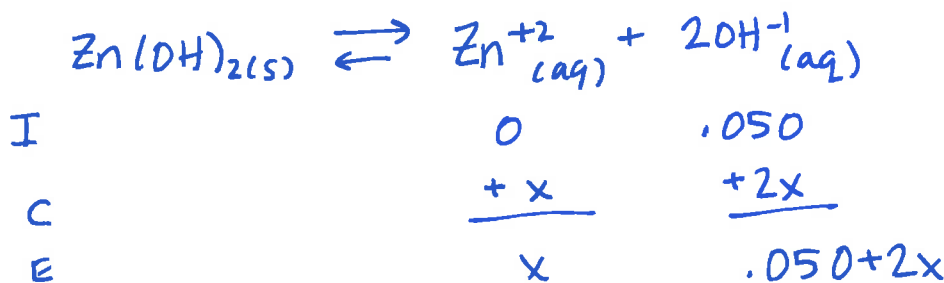
$$(.0500 \text{ L})(.300 \text{ M}) = .0150 \text{ moles NaOH}$$



$$[\text{OH}^-] = \frac{.0050 \text{ moles}}{0.100 \text{ L}} = 0.050 \text{ M}$$

↑

$$50.0 \text{ mL} + 50.0 \text{ mL} = 100.0 \text{ mL} \approx .1000 \text{ L}$$



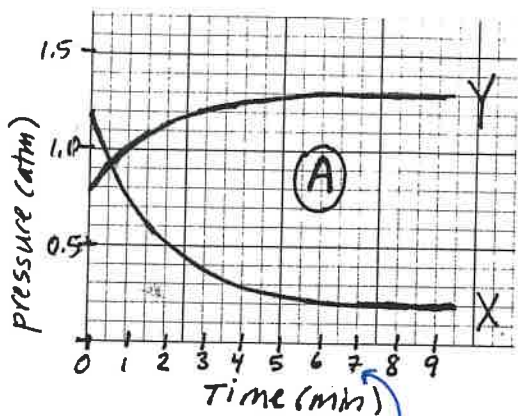
$$K_{sp} = [\text{Zn}^{2+}][\text{OH}^-]^2$$

$$7.7 \times 10^{-17} = x(.050 + 2x)^2$$

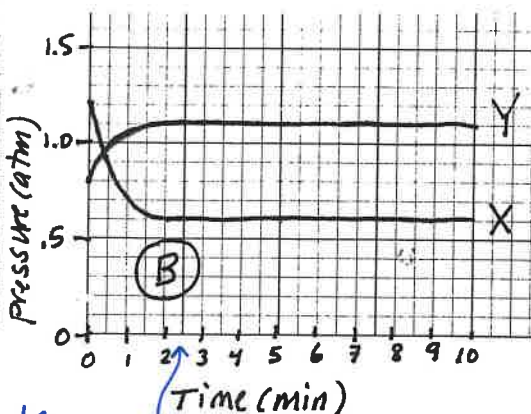
assume $2x \ll .050$

$$7.7 \times 10^{-17} = x(.050)^2$$

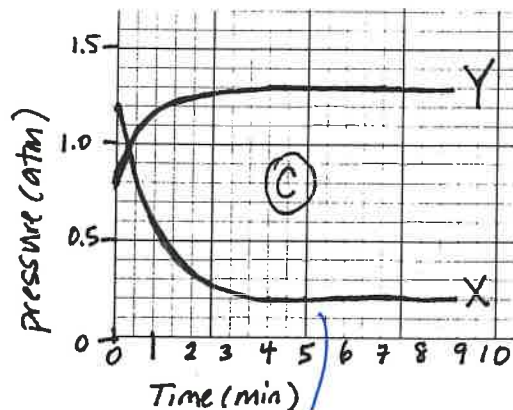
$$x = [\text{Zn}^{2+}] = 3.1 \times 10^{-14}$$



takes ≈ 7 minutes to reach eqm



takes 2-3 min to reach eqm



takes ≈ 5 min to reach eqm

21.

Suppose that two gases are reacting according to: $2X \rightleftharpoons Y$

Graphs A-C, above, show the partial pressures of gases X and Y as a function of time.

Graph A shows the pressures when the reaction is done at 200°C , with no catalyst present.

- a. Write an expression for K_p in terms of the appropriate partial pressures:

$$K_p = \frac{P_y}{(P_x)^2}$$

- b. Calculate K_p for the reaction, based on graph A (report extra S.F.):

$$K_p = \frac{(1.3)}{(0.2)^2} = 32.5$$

- c. In a separate experiment, the reaction was done at 200°C , with a catalyst present.

In yet another experiment, the reaction was done at 250°C , without a catalyst.

Which graph (B or C), could correspond to the reaction at 250°C , with no catalyst? **B**

Which graph (B or C), could correspond to the reaction at 200°C , with a catalyst? **C**

Explain your choices:

graphs A and C have the same K_{eq} of 32.5. graph B has a K_{eq} of 3.06. Both B and C show a faster rate of rxn than graph A; both B and C reach equilibrium in less time than A, and have steeper slopes, but this could be due to increasing T or adding a catalyst. Adding a catalyst does not change K_{eq} . changing temp does change K_{eq} . so B must be at a different temp than A.

- e. Calculate K_p for graph B. $\frac{1.1}{(0.6)^2} = 3.06$ f. Calculate K_p for graph C. $\frac{1.3}{(0.2)^2} = 32.5$

- g. Is this reaction exothermic or endothermic? **exothermic**

Justify your answer:

As Temperature increases from 200°C to 250°C , K_p/K_{eq} decreases from 32.5 to 3.06 since it is less favorable at higher temps, Rxn must be exothermic.

h. Fill out this chart:	Effect on K_{eq}	Effect on rate of reaction
Raising the temperature (if the rxn is endothermic)	K_{eq} increases	rate increases (reaches eqm faster)
Raising the temperature (if the rxn is exothermic)	K_{eq} decreases	rate increases (reaches eqm faster)
Adding a catalyst	no effect	rate increases

- i. Make a graph (use graph D) showing the partial pressures vs. time, if this reaction is done at 250°C , with a catalyst. Assume the same initial pressures as on graphs A-C.

