

Answers to Chapter 17

1) Both Figure I & II depicted separate ions as they combine to form the precipitate. Figure III already showed cations bonded with anions.



$$K_{sp} = [\text{Th}^{4+}][\text{IO}_3^-]^4$$

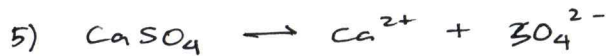
3) This is answered by process of elimination. Both Figure I & II will give  $K_{sp} = 4s^3$  while Figure III will give  $K_{sp} = s^2$  hence Figure IV is the only answer.



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

$$K_{sp} = 4s^3 \rightarrow \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4s^3}{4}}$$

$$s = \left(\frac{K_{sp}}{4}\right)^{1/3} \leftarrow \quad s = \sqrt[3]{\frac{K_{sp}}{4}} \leftarrow$$



$$K_{sp} = s \cdot s = s^2$$

$$K_{sp} = s^2 = (7.93 \times 10^{-3})^2$$

$$s = \frac{1.08 \text{ g/L}}{136.14 \text{ g/mol}} = 7.93 \times 10^{-3} = 6.30 \times 10^{-5}$$

6) All salts are equivalent (same charges magnitude i.e. +1 & -1 and +2 & -2) lowest  $K_{sp}$  to highest



$$K_{sp} = s^2 \quad 8.3 \times 10^{-17} = s^2$$

$$s = \sqrt{8.3 \times 10^{-17}} = 9.1 \times 10^{-9} \text{ mol/L}$$

what being asked is g/L  $(9.1 \times 10^{-9} \text{ mol/L})(234.77 \text{ g/mol}) = 2.1 \times 10^{-6}$



$$K_{sp} = [\text{Ba}^{2+}][\text{F}^-]^2 = (s)(2s)^2 = 4s^3$$

$$\sqrt[3]{\frac{1.0 \times 10^{-6}}{4}} = \sqrt[3]{\frac{4s^3}{4}}$$

$$6.3 \times 10^{-3} = s$$

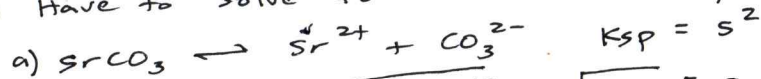
9) Similar to #8 instead of  $\text{Ba}^{2+}$ , its  $\text{Ca}^{2+}$



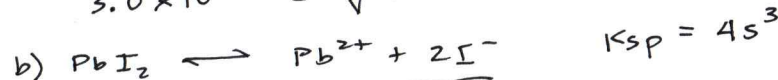
$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = (s)(2s)^2 = 4s^3$$

$$\sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4s^3}{4}} \quad s = \sqrt[3]{\frac{K_{sp}}{4}} = [\text{Ca}^{2+}]$$

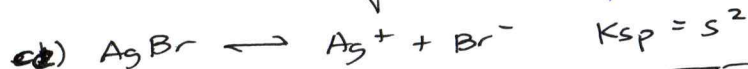
10) Have to solve for molar solubility



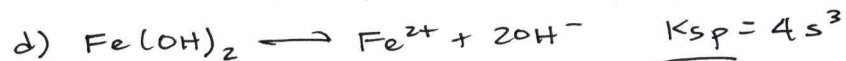
$$3.0 \times 10^{-5} = \sqrt{9.3 \times 10^{-10}} = \sqrt{K_{sp}} = s$$



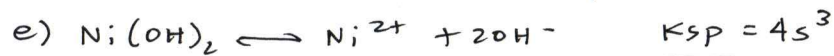
$$1.2 \times 10^{-3} = \sqrt[3]{\frac{6.5 \times 10^{-9}}{4}} = \sqrt[3]{\frac{K_{sp}}{4}} = s$$



$$7.1 \times 10^{-7} = \sqrt{5.0 \times 10^{-13}} = s$$

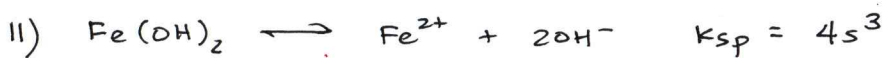


$$5.8 \times 10^{-6} = \sqrt[3]{\frac{8. \times 10^{-16}}{4}} = s$$



$$7.9 \times 10^{-6} = \sqrt[3]{\frac{2.0 \times 10^{-15}}{4}} = s$$

$\text{AgBr}$  has the lowest molar solubility



$$5.8 \times 10^{-6} = 3 \sqrt{\frac{8.0 \times 10^{-16}}{4}} \leftarrow s = \sqrt[3]{\frac{K_{sp}}{4}}$$

$$[\text{OH}^-] = 2s = 2(5.8 \times 10^{-6}) = 1.2 \times 10^{-5} = [\text{OH}^-]$$

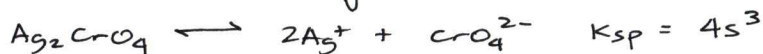
$$\text{pOH} = -\log [\text{OH}^-] = -\log 1.2 \times 10^{-5} = 4.93$$

$$\text{pH} = 14 - 4.93 = 9.07$$

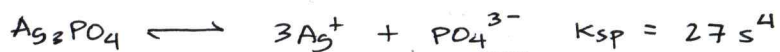
12) Have to calculate each.



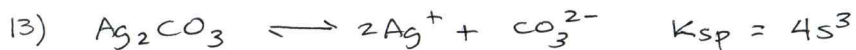
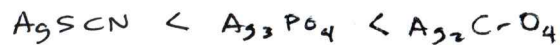
$$1.0 \times 10^{-6} = \sqrt{1.0 \times 10^{-12}} = s = \sqrt{K_{sp}}$$



$$6.5 \times 10^{-5} = \sqrt[3]{\frac{1.1 \times 10^{-12}}{4}} = s = \sqrt[3]{\frac{K_{sp}}{4}}$$

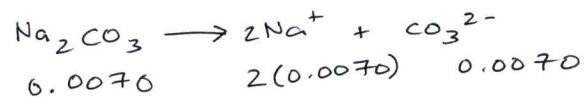


$$4.4 \times 10^{-5} = \sqrt[4]{\frac{1.0 \times 10^{-16}}{27}} = s = \sqrt[4]{\frac{K_{sp}}{27}}$$



The higher the concentration of the common ion, the more it will force the reaction to shift to the right decreasing the solubility of  $\text{Ag}_2\text{CO}_3$  hence  $0.10 \text{ M AgNO}_3$  is the answer

14) From the figure, a filtered saturated solution is shown which precipitate upon addition of some reagents. This could only happen if you add a common ion to  $\text{AgCl}$  hence both  $\text{HCl}$  &  $\text{NaCl}$  are possible reagents (Recall cation analysis).



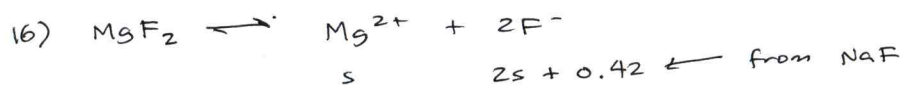
$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]$$

$$K_{sp} = (2s)^2 (0.0070 + s) \rightarrow s \text{ is too small}$$

$$8.6 \times 10^{-12} = (4s^2)(0.0070)$$

$$\sqrt{\frac{8.6 \times 10^{-12}}{0.028}} = \sqrt{\frac{0.028s^2}{0.028}}$$

$$1.75 \times 10^{-5} = s \quad [\text{Ag}^+] = 2s = 2(1.75 \times 10^{-5}) = 3.5 \times 10^{-5}$$

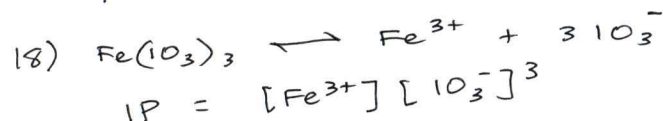


$$K_{sp} = (s)(2s + 0.42)^2$$

$$\frac{8.4 \times 10^{-8}}{0.176} = \frac{0.176s}{0.176}$$

$$4.8 \times 10^{-7} = s$$

17) a saturated solution means an equilibrium system exist so  $Q_c = K_{sp}$



$$IP = [\text{Fe}^{3+}][\text{IO}_3^-]^3$$

$$[\text{Fe}^{3+}] = 2.0 \times 10^{-5} \text{ M} \left( \frac{50 \text{ mL}}{100 \text{ mL}} \right) = 1.0 \times 10^{-5}$$

$$[\text{IO}_3^-] = 2.0 \times 10^{-4} \text{ M} \left( \frac{50 \text{ mL}}{100 \text{ mL}} \right) = 1.0 \times 10^{-4}$$

$$Q_c = IP = (1.0 \times 10^{-5})(1.0 \times 10^{-4})^3 = 1.0 \times 10^{-17}$$

$$K_{sp} > Q_c \text{ no ppt forms}$$

19) some  $\text{Ag}_2\text{CrO}_4$  will precipitate as the addition of a common ion ( $\text{Ag}^+$  from  $\text{AgNO}_3$ ) will lower the solubility of  $\text{Ag}_2\text{CrO}_4$ .



$$IP = Q_c = [\text{Pb}^{2+}]^3 [\text{PO}_4^{3-}]^2$$

$$[\text{Pb}^{2+}] = 1 \times 10^{-8} \text{ M} \left( \frac{50 \text{ mL}}{100 \text{ mL}} \right) = 5 \times 10^{-9}$$

$$[\text{PO}_4^{3-}] = 1 \times 10^{-6} \text{ M} \left( \frac{50 \text{ mL}}{100 \text{ mL}} \right) = 5 \times 10^{-7}$$

$$Q_c = (5 \times 10^{-9})^3 (5 \times 10^{-7})^2 = 3 \times 10^{-38}$$

$Q_c > K_{sp}$  precipitate form

$$21) [\text{Ag}^+] = 1 \times 10^{-7} \text{ M} \left( \frac{450}{900} \right) = 5 \times 10^{-8}$$

$$[\text{I}^-] = 1 \times 10^{-8} \text{ M} \left( \frac{450}{900} \right) = 5 \times 10^{-9}$$



$$Q_c = [\text{Ag}^+][\text{I}^-] = (5 \times 10^{-8})(5 \times 10^{-9}) = 2.5 \times 10^{-16}$$

$Q_c > K_{sp}$  AgI will ppt.

22) The one that precipitate first is the one with lowest  $K_{sp}$  which in this case it AuCl.

23) HCl will generate  $\text{H}^+$  and  $\text{Cl}^-$ . The  $\text{H}^+$  can react with  $\text{S}^{2-}$  from ZnS to form  $\text{H}_2\text{S}$ . The decrease in  $[\text{S}^{2-}]$  resulted in dissolution of ZnS.



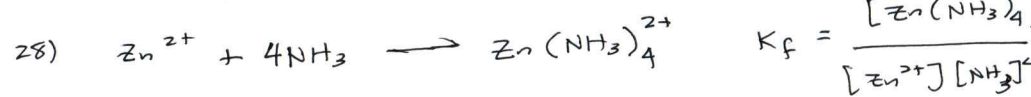
Since  $K_a$  of HX  $>$  HY, that means you have more  $\text{X}^-$  than  $\text{Y}^-$ . The  $\text{X}^-$  will cause AgX to have lower solubility since backward reaction is favored.

25) Choices A to C will results in lower molar solubility because of common ion effect. Choice D will have no effect but Choice E will cause  $\text{C}_2\text{O}_4^{2-}$  to react with  $\text{H}^+$  from HCl causing increase in molar

solubility.

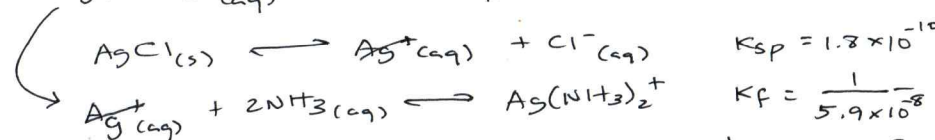
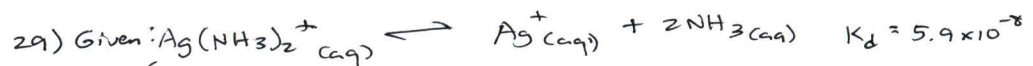
26) The initial figure shows presence of precipitate which dissolve upon addition of a strong acid. All choices has conjugate of strong acid except A where the  $\text{F}^-$  will react with  $\text{H}^+$  to form HF causing AgF to dissolve.

$$27) K_d = \frac{1}{K_f} = \frac{1}{1.0 \times 10^7} = 1.0 \times 10^{-7}$$



$$3.8 \times 10^9 = \frac{0.20}{[\text{Zn}^{2+}](0.0116)^4}$$

$$[\text{Zn}^{2+}] = \frac{0.20}{(3.8 \times 10^9)(0.0116)^4} = 2.9 \times 10^{-3}$$



What is  $K_c$  for  $\text{AgCl}(s) + 2\text{NH}_3 \rightleftharpoons \text{Ag}(\text{NH}_3)_2^+ + \text{Cl}^-$ ?

if you notice, you can reverse the first reaction and add with the 2nd reaction to get the reaction whose  $K_c$  is being asked.

$$K_c = K_{sp} \left( \frac{1}{K_d} \right) = (1.8 \times 10^{-10}) \left( \frac{1}{5.9 \times 10^{-8}} \right) = 3.1 \times 10^{-3}$$

30) Group I cations are known as chloride group since HCl precipitate but  $\text{Ag}^+$ ,  $\text{Hg}_2^{2+}$  and  $\text{Pb}^{2+}$ .