PSI AP Chemistry
Aqueous Equilibria part II –Solubility Product Name ____________________
Practice Problems

Introduction to Solubility Equilibria

Classwork

1. Write the chemical reaction for and the K_{sp} expression for the equilibrium that exists in a saturated solution of silver iodide.

2. Write the chemical reaction for and the K_{sp} expression for the equilibrium that exists in a saturated solution of cadmium (II) carbonate.

3. Write the chemical reaction for and the expression for the solubility product constant, K_{sp} that exists in a saturated solution of silver chromate?

Homework

4. What is the expression for the solubility product constant, K_{sp}, for iron (II) phosphate?

5. Write the K_{sp} expression for the equilibrium that exists in a saturated solution of iron (III) hydroxide, $Fe(OH)_3$.

6. Which of the following compounds is the least soluble?

Compound	K_{sp}
$MnCO_3$	1.8×10^{-11}
CuS	6.3×10^{-36}
CdS	8.0×10^{-27}
PbS	8.0×10^{-28}

A) $MnCO_3$ B) CuS C) CdS D) PbS E) both CdS and PbS

Calculating Ksp from Solubility

Classwork

7. The solubility of silver chloride is 1.34×10^{-5} at 25°C. At this temperature, what is the K_{sp} of AgCl?
A) 3.7×10^{-3} B) 1.1×10^{-6} C) 1.8×10^{-9} D) 3.2×10^{-20} E) 1.8×10^{-10}

8. The solubility of lead (II) chloride ($PbCl_2$) is 1.6×10^{-2} at 25°C. At this temperature, what is the K_{sp} of $PbCl_2$?
A) 5.0×10^{-4} B) 4.1×10^{-6} C) 3.1×10^{-7} D) 1.6×10^{-5} E) 1.6×10^{-2}

9. The solubility of silver sulfate is 2.2×10^{-5}. What is the K_{sp} of silver sulfate?
A) 1.1×10^{-14} B) 4.3×10^{-14} C) 2.1×10^{-14} D) 4.8×10^{-10} E) 2.2×10^{-5}

10. What is the solubility product for $AuCl_3$ if the molar solubility in a saturated solution is 3.3×10^{-7}?
A) 3.3×10^{-27} B) 1.2×10^{-26} C) 3.2×10^{-25} D) 3.3×10^{-13} E) 1.3×10^{-6}

Homework

11. The solubility of $CaSO_4$ in water is 0 .67 gram per liter of solution. Calculate the K_{sp}.

12. Calculate the K_{sp} of silver chromate, Ag_2CrO_4, if the solubility is 0.15 grams/L.

13. Calculate the K_{sp} of lithium carbonate, Li_2CO_3, if the solubility is 0.15 mol/L.

14. Given the following table of K_{sp} values, determine which compound listed has the greatest solubility.

Compound	K_{sp}
$CdCO_3$	5.2×10^{-12}
$Cd(OH)_2$	2.5×10^{-14}
AgI	8.3×10^{-17}
$Fe(OH)_3$	4.0×10^{-38}
$ZnCO_3$	1.4×10^{-11}

A) $CdCO_3$ B) $Cd(OH)_2$ C) AgI D) $Fe(OH)_3$ E) $ZnCO_3$

Calculating Solubility from Ksp

Classwork

15. What is the solubility, in mol/L, of AgBr if the $K_{sp} = 5.0 \times 10^{-13}$?
A) 2.5×10^{-25} B) 1.0×10^{-12} C) 7.1×10^{-7} D) 7.1×10^{-6}

16. Calculate the molar solubility of a saturated solution of lead (II) iodide, PbI_2 (K_{sp} is 1.4×10^{-8}).
A) 1.5×10^{-3} B) 3.1×10^{-3} C) 4.2×10^{-4} D) 3.5×10^{-9} E) 1.4×10^{-8}

17. Calculate the molar solubility of a saturated solution of Silver sulfide (K_{sp} is 6×10^{-51}).

18. Calculate the molar solubility of a saturated solution of aluminum chloride if the solubility product constant of $AlCl_3$ is 27×10^{-64}.

Homework

19. What is the solubility, in mol/L, of $BaCO_3$ if the $K_{sp} = 5.0 \times 10^{-9}$?
A) 7.1×10^{-5} B) 3.1×10^{-3} C) 2.4×10^{-10} D) 7.1×10^{-6}
E) 2.5×10^{-17}

20. Calculate the molar solubility of a saturated solution of silver carbonate, Ag_2CO_3 (K_{sp} is 8.1×10^{-12}
A) 1.3×10^{-4} B) 2.8×10^{-6} C) 1.5×10^{-3} D) 4.8×10^{-6} E) 1.4×10^{-8}

21. Calculate the molar solubility of a saturated solution of zinc hydroxide if the solubility product constant of $Zn(OH)_2$ is 3.0×10^{-16}.

22. Calculate the molar solubility of a saturated solution of lanthanum fluoride, LaF_3 (K_{sp} is 2×10^{-19}).

Factors Affecting Solubility: Common Ion Effect

Classwork

23. In which of the following aqueous solutions would you expect AgCl to have the lowest solubility?

A) pure water B) 0.020 M $BaCl_2$ C) 0.015 NaCl
D) 0.020 $AgNO_3$ E) 0.020 KCl

24. In which of the following aqueous solutions would you expect AgCl to have the highest solubility?

A) pure water B) 0.020 M $BaCl_2$ C) 0.015 M NaCl
D) 0.020 M $AgNO_3$ E) 0.020 M KCl

25. In which of the following aqueous solutions would you expect AgBr to have the lowest solubility?

A) pure water B) 0.20 M NaBr C) 0.10 M $AgNO_3$
D) 0.15 M KBr E) 0.10 M LiBr

26. Calculate the maximum concentration in (M) of chloride ions (Cl^-) in a solution that contains 0.100 M of Pb^{2+} if at 323 K the K_{sp} of $PbCl_2$ is 1.0 x10^{-4}.

A) 1.0 x10^{-4} B) 1.0 x10^{-3} C) 0.029 D) 0.032 E) 0.058

Homework

27. In which of the following aqueous solutions would you expect AgBr to have the highest solubility?

A) 0.10 M LiBr B) 0.10 M $AgNO_3$ C) 0.20 M NaBr
D) 0.15 M KBr E) pure water

28. In which of the following aqueous solutions would you expect $PbCl_2$ to have the lowest solubility?

A) 0.020 M KCl B) 0.015 M $BaCl_2$ C) 0.015 M $PbNO_3$
D) pure water E) 0.015 M NaCl

29. Of the following substances, which one would decrease the solubility of $CaCO_3$ in a saturated solution?

A) NaCl B) HCl C) HNO_3 D) $CaCl_2$ E) KNO_3

30. Calculate the maximum concentration in (M) of silver ions (Ag+) in a solution that contains 0.025 M of CO_3^{2-} .The K_{sp} of Ag_2CO_3 is 8.1 x10^{-12} .
A) 1.8 x10^{-5} B) 1.4 x10^{-6} C) 2.8 x10^{-6} D) 3.2 x10^{-10} E) 8.1 x10^{-12}

Factors Affecting Solubility: Changes in pH

Classwork

31. What is the solubility (M) of Nickel (II) hydroxide? The K_{sp} of $Ni(OH)_2$ at 25^{o}C is 6.0 x 10^{-16}.
A) 8.4 x10^{-6} B) 1.22 x10^{-8} C) 5.3 x10^{-6} D) 7.1 x 10^{-4} E) 1.5 x10^{-6}

32. What is the solubility (M) of $Ni(OH)_2$ in a solution buffered at pH = 10.0? The K_{sp} of $Ni(OH)_2$ at 25^{o}C is 6.0 x10^{-16}.
A) 2.0 x10^{-3} B) 6.0 x10^{-8} C) 1.8 x10^{-4} D) 6.0 x10^{-12} E) 1.5 x10^{--16}

33. What is the solubility (M) of $Ni(OH)_2$ in a solution buffered at pH = 11.5? The K_{sp} of $Ni(OH)_2$ at 25^{o}C is 6.0 x10^{-16}.
A) 6.0 x10^{-8} B) 1.1 x10^{-4} C) 1.9 x10^{-13} D) 6.0 x10^{-16} E) 5.9 x10^{-11}

34. Which of the following compounds will be more soluble if the pH of a saturated solution is lowered?
    A) AgCl B) AgI C) $PbCl_2$ D) NaCl E) $Cr(OH)_3$

Homework

35. What is the solubility (M) of calcium hydroxide? The K_{sp} of $Cu(OH)_2$ at 25°C is 4.8×10^{-20}.
    A) 1.9×10^{-2} B) 1.2×10^{-2} C) 2.5×10^{-3} D) 5.9×10^{-11} E) 1.6×10^{-5}

36. What is the solubility (M) of $Cu(OH)_2$ in a solution buffered at pH = 9.5? The K_{sp} of $Cu(OH)_2$ at 25°C is 4.8×10^{-20}.
    A) 2.0×10^{-3} B) 1.1×10^{-4} C) 1.8×10^{-4} D) 7.1×10^{-4} E) 1.6×10^{-5}

37. What is the solubility (M) of $Cu(OH)_2$ in a solution buffered at pH = 12.2? The K_{sp} of $Cu(OH)_2$ at 25°C is 4.8×10^{-20}.
    A) 2.0×10^{-3} B) 1.1×10^{-4} C) 1.8×10^{-4} D) 7.1×10^{-4} E) 1.6×10^{-5}

38. Which of the following compounds will **not** be more soluble if the pH of a saturated solution is lowered?

    A) Ag_2S B) AgI C) $NiCO_3$ D) BaC_2O_4 E) $Cr(OH)_3$

Selective Precipitation: Precipitation Reactions

Classwork

39. If 0.1 M aqueous solutions of the following pairs of substances are combined, which pair will yield a precipitate?
    A) sodium sulfide & ammonium chloride
    B) sodium sulfide & iron (III) chloride
    C) sodium hydroxide & potassium nitrate
    D) nickel (II) nitrate & magnesium chlorate
    E) potassium chloride & aluminum nitrate

40. What is the formula of the precipitate formed when 0.1M aqueous solutions of silver nitrate and potassium phosphate are mixed?
    A) $AgPO_4$ C) Ag_3PO_4 E) $Ag_3(PO_4)_3$
    B) Ag_2PO_4 D) $Ag(PO_4)_3$

41. Will a silver iodate precipitate form when 100 mL of 0.010 M $AgNO_3$ is mixed with 10.0 mL of 0.015 M $NaIO_3$? (K_{sp} of $AgIO_3$ is 3.1×10^{-8})

    A) No, because Q < K
    B) Yes, because Q < K
    C) No, because Q > K
    D) Yes, because Q > K
    E) Not enough information

Homework

42. If 0.1 M aqueous solutions of the following pairs of substances are combined, which pair will **not** yield a precipitate?
    A) silver nitrate & ammonium chloride
    B) potassium carbonate & iron (III) perchlorate
    C) sodium fluoride & lead (II) nitrate

D) nickel (II) nitrate & magnesium perchlorate
E) potassium chloride & mercury (I) nitrate

43. Will a silver sulfate precipitate form when 50.0 mL of 0.050 M $AgNO_3$ is mixed with 75.0 mL of 0.0050 M Na_2SO_4? (K_{sp} of Ag_2SO_4 is 1.5×10^{-5})
A) No, because Q < K
B) Yes, because Q < K
C) No, because Q > K
D) Yes, because Q > K
E) Not enough information

44. Will a precipitate form when 0.96g $(NH_4)_2CO_3$ is mixed with 0.20 g $CaBr_2$ in 10 L of solution? (Ksp= 4.5×10^{-9} for $CaCO_3$)
A) No, because Q < K
B) Yes, because Q < K
C) No, because Q > K
D) Yes, because Q > K
E) Not enough information

Selective Precipitation: Separation of Ions

Classwork

45. A solution contains 2.0×10^{-4} M Ag^+ and 1.5×10^{-3} M Pb^{2+}. If sodium iodide, NaI is added, what [I-] will cause the first precipitate? (Ksp= 8.3×10^{-17} for AgI; Ksp= 7.9×10^{-9} for PbI_2)
A) 2.0×10^{-4} *M* I^- will cause AgI to precipitate first
B) 4.2×10^{-13} *M* I^- will cause AgI to precipitate first
C) 5.3×10^{-6} *M* I^- will cause PbI_2 to precipitate first
D) 2.3×10^{-3} *M* I^- will cause PbI_2 to precipitate first
E) Not enough information

46. Referring to the questions above, how much of the first ion will be present when the second ion begins to precipitate?

47. A solution contains 1.0×10^{-5} M Pb^{2+} and 2.2×10^{-4} M Mg^{2+}. If lithium carbonate, $LiCO_3$ is added, what $[CO_3^{2-}]$ will cause the first precipitate? (Ksp= 7.4×10^{-14} for $PbCO_3$; Ksp= 5.0×10^{-10} for $MgCO_3$)
    A) 7.4×10^{-9} *M* CO_3^{2-} will cause $PbCO_3$ to precipitate first
    B) 4.2×10^{-13} *M* CO_3^{2-} will cause $PbCO_3$ to precipitate first
    C) 5.3×10^{-6} *M* CO_3^{2-} will cause $MgCO_3$ to precipitate first
    D) 2.3×10^{-6} *M* CO_3^{2-} will cause $MgCO_3$ to precipitate first
    E) Not enough information

48. Referring to the questions above, how much of the first ion will be present when the second ion begins to precipitate?

Homework

49. A solution contains 0.010 M Ba^{2+} and 0.010 M Sr^{2+}. If sodium sulfate, Na_2SO_4 is slowly added, what $[SO_4^{2-}]$ will cause the first precipitate? (Ksp= 1.1×10^{-10} for $BaSO_4$; Ksp= 3.2×10^{-7} for $SrSO_4$)
    A) 1.1×10^{-10} *M* SO_4^{2-} will cause $BaSO_4$ to precipitate first
    B) 1.1×10^{-8} *M* SO_4^{2-} will cause $BaSO_4$ to precipitate first
    C) 3.2×10^{-5} *M* SO_4^{2-} will cause $SrSO_4$ to precipitate first
    D) 3.2×10^{-7} *M* SO_4^{2-} will cause $SrSO_4$ to precipitate first
    E) Not enough information

50. Referring to the questions above, how much of the first ion will be present when the second ion begins to precipitate?

51. A solution contains 0.0250 M Ca^{2+} and 1.5×10^{-3} M Fe^{2+}. If potassium hydroxide, KOH is added, what $[OH^-]$ will cause the first precipitate? (Ksp= 6.5×10^{-6} for $Fe(OH)_2$; Ksp= 7.9×10^{-16} for $Fe(OH)_2$)
    A) 1.6×10^{-2} *M* OH^- will cause $Ca(OH)_2$ to precipitate first
    B) 1.1×10^{-8} *M* OH^- will cause $Ca(OH)_2$ to precipitate first
    C) 1.5×10^{-5} *M* OH^- will cause $Fe(OH)_2$ to precipitate first
    D) 7.3×10^{-7} *M* OH^- will cause $Fe(OH)_2$ to precipitate first
    E) Not enough information

52. Referring to the questions above, how much of the first ion will be present when the second ion begins to precipitate?

II- Conceptual Questions Ksp and Solubility

1) The best explanation for the solubility of MnS in dilute HCl is that

A) The solubility product of $MnCl_2$ is less than that of MnS

B) Concentration of Mn^{2+} is lowered by the formation of complexs ions with chloride ions

C) Concentration of sulfide ions is lowered by oxidation of free sulfur

D) Concentration of sulfide ions is lowered by the formation of the weak H_2S

E) Because of the formation of more $MnCl_2$

2) The solubility of AgI in NaI solution is less than that in pure water because

A) AgI forms complexs with NaI

B) of common ion effect

C) solubility product of AgI is less than that of NaI

D) the temperature of the solution decreases

E) none of the above

3) The solubility product of CuS, Ag_2S and HgS are 10-31, 10-44 and 10-54 respectively. The solubility of these sulfides are in the order

A) Ag_2S> CuS >HgS

B) Ag_2S > HgS > CuS
C) HgS> Ag_2S > CuS
D) CuS > Ag_2S > HgS
E) HgS > CuS > Ag_2S

4) What is the correct expression for the solubility product of SnS?
A) $[Sn^{2+}]\ [S^{2-}]^2$
B) $[Sn^{2+}]\ [S^{2-}]$
C) $[Sn^{2+}]\ [2S^{2-}]^2$
D) $[Sn^{2+}]\ [2S^{2-}]$
E) $[Sn^{2+}]^2\ [S^{2-}]^2$

5) The solubility of $CaCO_3$ in water is 3.05×10^{-4} moles/L. Its solubility product will be
A) 6.1×10^{-4}
B) 9.3
C) 3.05×10^{-4}
D) 9.3×10^{-8}
E) 15.4

6) The solubility of A_2X_3 is y mol $dm^{-3.}$ Its solubility product is
A) $6y^4$
B) $64y^4$
C) $36y^5$
D) $108y^5$
E) $6y^5$

7) Consider the following solubility data for various chromates at 25°C.

	K_{sp}
Ag_2CrO_4	9.0×10^{-12}
$BaCrO_4$	2.0×10^{-10}
$PbCrO_4$	1.8×10^{-14}

The chromate that is the most soluble in water at 25°C on a molar basis is:
A) Ag_2CrO_4
B) $BaCrO_4$
C) $PbCrO_4$
D) impossible to determine
E) none of these

8) What is the molar solubility, of $Ba_3(PO_4)_2$ in terms of K_{sp}?
A) $K_{sp}^{1/2}$
B)) $K_{sp}^{1/5}$
C) $[K_{sp}/27]^{1/5}$
D) $[K_{sp}/108]^{1/5}$
E) $[K_{sp}/4]^5$

9) Ag_3PO_4 would be least soluble at 25°C in
A) 0.1 M $AgNO_3$
B) 0.1 M HNO_3
C) pure water
D) 0.1 M Na_3PO_4
E) solubility in (a), (b), (c), or (d) is not different

10) When we mix together, from separate sources, the ions of a slightly soluble ionic salt, the salt will precipitate if Q _____ K_{sp}, and will continue to precipitate until Q _____ K_{sp}.
A) is greater than; equals
B) is less than; is greater than
C) is less than; equals
D) equals; is less than
E) equals; is greater than

11) Which of the following pairs of compounds gives a precipitate when aqueous solutions of them are mixed? Assume that the concentrations of all compounds are 1.0 M immediately after mixing.
A) $CuBr_2$ and K_2CO_3
B) HNO_3 and NH_4I
C) $BaCl_2$ and $KClO_4$
D) Na_2CO_3 and H_2SO_4
E) KCl and KNO_3

12) Which of the following occurs when excess of concentrated $NH_{3\ (aq)}$ is mixed with 0.1M $Cu(NO_3)_{2\ (aq)}$?
A) A dark red precipitate forms and settles out
B) Separate layers of immiscible liquids form with a blue layer on top.
C) The color of the solution turns from light blue to dark blue.
D) Bubbles of ammonia gas form.
E) The pH of the solution decreases.

13) A yellow precipitate forms when 0.2 M NaI is added to a 0.2M solution of which of the following ions?
A) Zn^{2+}
B) Pb^{2+}
C) $CrO_4^{\ 2-}$
D) SO_4^{2-}
E) OH^-

14) Which solution below could selectively precipitate lead alone from the mixture if it has $Fe^{2+,}$ Cu^{2+} and Pb^{2+} ions?
A) sodium sulfide
B) sodium hydroxide
C) dilute Hydrochloric acid
D) dilute Nitric acid
E) dilute Ammonia

15) When 50 ml each of 0.1 M $Li_3(PO_4)$ and $Ag(NO_3)$ are mixed together, yellow precipitate of silver phosphate is produced. Which of the following ions is the major component in the solution?
A) $[PO_4^{3-}]$
B) $[NO_3^-]$
C) $[Ag^+]$
D) $[Li^{3+}]$
E) They are all equal

16) If 100ml each of 0.1M of $Ba(OH)_2$ and Na_2 (SO_4) are mixed together, what would be the concentration of the hydroxide ions in solution?

A) 0.2 M
B) 0.1M
C) 1.0M
D) .02M
E) 0.01M

17) Referring to the above question, what would be the concentration of $[SO_4^{2-}]$ ions in the resulting solution?

A) 0.1M
B) 0.5M
C) 1.0M
D) Negligible amount
E) 0.05M

18) When aqueous NH_3 is first added to a solution containing $Ni^{2+,}$ a precipitate forms, but when an excess of aqueous NH3 is added, the precipitate dissolves. Which of the below explains why the precipitate dissolves?

A) Ni^{2+} forms hydrogen bonds with NH_3
B) Ni^{2+} forms hydrogen bonds with NH_3
C) Ni^{2+} forms complex ion with NH_3
D) Ni^{2+} is oxidized to Ni^{3+}
E) Ni^{2+} is reduced to Ni^{1+}

19) Which of the following ions are generally insoluble in cold water?

A) Acetate
B) Potassium
C) Phosphate
D) Nitrate
E) Ammonium

20) A student mixes equal volumes of 1.0M solutions of Tin chloride and 1.0 M copper sulfate and observes that no precipitate forms. Then the student mixes equal volumes of 1.0 M solutions of sodium sulfide and copper sulfide and observes the formation of a precipitate. The formula of the precipitate must be

A) CuS
B) $CuCl_2$
C) CuCl
D) $Na_2(SO_4)_2$
E) $NaSO_4$

Solubility product, Ksp - Free Response Questions

1) The solubility of iron (II) hydroxide, $Fe(OH)_2$, is 1.43×10^{-3} gram per liter at 25˚C.

   a) Write a balanced equation for the solubility equilibrium.
   b) Write the expression for the solubility product constant, K_{sp}, and calculate its value.
   c) Calculate the pH of a saturated solution of $Fe(OH)_2$ at 25˚C.
   d) A 50.0 milliliter sample of 3.00×10^{-3} molar $FeSO_4$ solution is added to 50.0 milliliters of 4.00×10^{-6} molar NaOH solution. Does a precipitate of $Fe(OH)_2$ form? Explain and show calculations to support your answer.

2) Solve the following problem related to the solubility equilibria of some metal hydroxides in aqueous solution.

(a)The solubility of $Cu(OH)_2(s)$ is 1.72×10^{-6} gram per 100 milliliters of solution at 25˚C.

i. Write the balanced chemical equation for the dissociation of $Cu(OH)_2(s)$ in aqueous solution.
ii. Calculate the solubility (in moles per liter) of $Cu(OH)_2$ at 25˚C.
iii. Calculate the value of the solubility-product constant, K_{sp}, for $Cu(OH)_2$ at 25˚C.

(b)The value of the solubility-product constant, K_{sp}, for $Zn(OH)_2$ is $7.7\text{x}10^{-17}$ at 25˚C.
i. Calculate the solubility (in moles per liter) of $Zn(OH)_2$ at 25˚C in a solution with a pH of 9.35.
ii. At 25˚C, 50.0 milliliters of 0.100-molar $Zn(NO_3)_2$ is mixed with 50.0 milliliters of 0.300-molar NaOH. Calculate the molar concentration of $Zn^{2+}(aq)$ in the resulting solution once equilibrium has been established. Assume that volumes are additive.

3) $MgF_2(s) \leftrightarrow Mg^{2+}{}_{(aq)} + 2\ F^-{}_{(aq)}$

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.

a) Write the expression for the solubility-product constant, K_{sp} and calculate its value at 18° C.
b) 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.
c) 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°C. Calculations to support your prediction must be shown.
d) At 27°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

4)

a) At 283K, $8.9\ \text{x}10^{-5}$ g of AgCl (s) will dissolve in 100 ml of water.

i. Write the equation for the dissociation of AgCl in water.
ii. Calculate the solubility in mol/L of AgCl in water at 283K
iii. Calculate the value of the solubility –product constant, K_{sp} for AgCl.

b) If 60.0 ml of 0.0400M NaCl is added to 60ml of 0.0300M $Pb(NO_3)_2$, will a precipitate form? Volumes are additive. Show the calculations.
c) Calculate the equilibrium value of [Pb2+] in 1L of saturated PbCl2 solution to which 0.250mol of NaCl (s) has been added.(no volume change occurs) Ksp $PbCl_2$ = $1.6\ \text{x}10^{-5}$

5) Several reactions are carried out using AgBr, a cream-colored silver salt for which the value of the solubility product constant, K_{sp}, is 5.0×10^{-13} at 298 K.

a) Write the expression for the solubility-product constant, K_{sp}, of AgBr.

b) Calculate the value of $[Ag^+]$ in 50.0 mL of a saturated solution of AgBr at 298 K.

c) A 50.0 mL sample of distilled water is added to the solution described in part (b), which is in a beaker with some solid AgBr at the bottom. The solution is stirred and equilibrium is reestablished. Some solid AgBr remains in the beaker. Is the value of $[Ag^+]$ greater than, less than, or equal to the value you calculated in part (b)? Justify your answer.

d) Calculate the minimum volume of distilled water, in liters, necessary to completely dissolve a 5.0 g sample of AgBr(*s*) at 298 K. (The molar mass of AgBr is 188 g mol^{-1})

e) A student mixes 10.0 mL of 1.5×10^{-4} *M* $AgNO_3$ with 2.0 mL of 5.0×10^{-4} *M* NaBr and stirs the resulting mixture. What will the student observe? Justify your answer with calculations.

f) The color of another salt of silver, AgI(*s*), is yellow. A student adds a solution of NaI to a test tube containing a small amount of solid, cream-colored AgBr. After stirring the contents of the test tube, the student observes that the solid in the test tube changes color from cream to yellow.

   i. Write the chemical equation for the reaction that occurred in the test tube.

   ii. Which salt has the greater value of K_{sp}: AgBr or AgI? Justify your answer.

6)

(a) For a 1:1 compound, $K_{sp} = s^2$

i. What is the relationship for Al_2S_3? K_{sp} =

ii. What is the relationship for Ag_3PO_4? K_{sp} =

iii. What is the relationship for $Al(OH)_3$? K_{sp} =

iv. What is the relationship for Ag_2SO_3 K_{sp} =

a) The K_{sp} of CuCN is 3.2×10^{-20}. What is the molar solubility of CuCN?

b) The $[F^-]$ in a saturated solution of BaF_2 is 1.5×10^{-2} <u>M</u>. What is the K_{sp} of BaF_2?

c) The K_{sp} of AuI is 1.6×10^{-23}. What is the molar solubility of AuI?

d) The K_{sp} of ZnS is 2.0×10^{-25}. What is the molar solubility of ZnS in 0.10 M K_2S?

7) Ag_2CrO_4 □□ $2Ag^+$ (aq) + CrO_4^{2-}(aq) $K_{sp} = 2.6 \times 10^{-12}$

a) Write the equilibrium – constant expression for the dissolving of Ag_2CrO_4 (s)

b) Calculate the concentration in mol/L of Ag^+ (aq) in a saturated solution of Ag_2CrO_4 at 25°C

c) Calculate the maximum mass in grams of Ag_2CrO_4 that can dissolve in 100 ml of water at 25°C.

d) A 0.100 mol sample of solid Ag NO_3 is added to a 1.00 L saturated solution of Ag_2CrO_4. Assuming no volume change, does $[CrO4^{2-}]$ increase, decrease or remain the same? Justify your answer.

In a saturated solution of Ag_3PO_4 at 250C, the concentration of Ag(aq) is 5.3 x10 -5M. The equilibrium constant expression for the dissolving of Ag_3PO_4 (s) in water is below.

$$Ksp = [Ag^+]^3 [PO4^{3-}]$$

a) Write the balanced equation for the dissociation of Ag_3PO_4 in water.
b) Calculate the value of K_{sp} for Ag_3PO_4 at 25˚C
c) A 1L sample of saturated Ag_3PO_4 solution is allowed to evaporate at 25˚C to a final volume of 500 mL. What is $[Ag^+]$ in the solution? Justify your answer.

8) 1.2M NaI is added to a 1L of solution containing 0.002M $Ag(NO_3)$ and 0.5M $Pb(NO_3)_2$. Assume the volume change is negligible. Ksp AgI = $8.3\ 10^{-17}$, Ksp PbI_2 = $7.9\ 10^{-9}$

a) Write down the individual reactions between NaI and $Ag(NO_3)$ and PbI_2.
b) What is/are the solid product/s if there is any would form?
c) Which one of the solid product will precipitate first from the solution?

Answers

Introduction to Solubility Equilibria

1) $AgI_{(s)} \leftrightarrow Ag^{+}_{(aq)} + I^{-}_{(aq)}$
$K_{sp} = [Ag^{+}][I^{-}]$

2) $CdCO_{3\,(s)} \leftrightarrow Cd^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$
$K_{sp} = [Cd^{2+}][CO_3^{2-}]$

3) $Ag_2CrO_{4\,(s)} \leftrightarrow 2Ag^{+}_{(aq)} + CrO_4^{2-}{}_{(aq)}$
$K_{sp} = [Ag^{+}]^2[CrO_4^{2-}]$

4) $Fe_3(PO_4)_{2(s)} \leftrightarrow 3Fe^{2+}_{(aq)} + 2PO_4^{3-}{}_{(aq)}$
$K_{sp} = [Fe^{2+}]^3[PO_4^{2-}]^2$

5) $Fe(OH)_{3(s)} \leftrightarrow Fe^{3+}_{(aq)} + 3OH^{-}_{(aq)}$
$K_{sp} = [Fe^{3+}]\,[OH^{-}]^3$

6) C. The salt with the smallest K_{sp} value is the least soluble if all the salts have a 1:1 ratio of cation to anion.

Calculating K_{sp} from Solubility

7) E $K_{sp} = [Ag^{+}][Cl^{-}] = (1.34 \times 10^{-5})^2 = 1.8 \times 10^{-10}$

8) D $K_{sp} = [Pb^{2+}][Cl^{-}]^2 = (1.6 \times 10^{-2})(3.2 \times 10^{-2})^2 = 1.6 \times 10^{-5}$

9) B $K_{sp} = [Ag^{+}]^2[SO_4^{2-}] = (4.4 \times 10^{-5})^2(2.2 \times 10^{-5}) = 4.3 \times 10^{-14}$

10) C $K_{sp} = [Au^{3+}][Cl^{-}]^3 = (3.3 \times 10^{-7})(9.9 \times 10^{-7})^3 = 3.2 \times 10^{-25}$

11) 0.67g x 1 mol/136.2g = 4.9×10^{-3} mol/liter
$K_{sp} = [Ca^{2+}][SO_4^{2-}] = 4.9 \times 10^{-3})^2 = 2.4 \times 10^{-5}$

12) 0.015g x 1 mol/331.8g = 4.5×10^{-5} mol/liter
$K_{sp} = [Ag^{+}]^2[CrO_4^{2-}] = (9.0 \times 10^{-5})^2(4.5 \times 10^{-5}) = 3.7 \times 10^{-13}$

13) $K_{sp} = [Li^{+}]^2[CO_3^{2-}] = (0.30)^2(0.15) = 0.0135$

14) B $K_{sp} = 1.8 \times 10^{-5}$

Calculating Solubility from K_{sp}

15) C $K_{sp} = [Ag^{+}][Br^{-}] = x^2$ $x = \sqrt{5.0 \times 10^{-13}} = 7.1 \times 10^{-7}$ mol/L

16) A $K_{sp} = [Pb^{2+}]^2[I^{-}] = 4x^3$ $x = (K_{sp}/4)^{1/3} = 1.5 \times 10^{-3}$ mol/L

17) $Ag_2S_{(s)} \leftrightarrow 2Ag^{+}_{(aq)} + S^{2-}_{(aq)}$
$K_{sp} = (2x)^2(x) = 4x^3$
$6 \times 10^{-51} = 4x^3$
$x = 1.1 \times 10^{-17}$ mol/L

18) $AlCl_{3(s)} \leftrightarrow Al^{3+}_{(aq)} + 3Cl^{-}_{(aq)}$
$K_{sp} = (x)(3x)^3 = 27x^4$
$27 \times 10^{-64} = 27x^4$
$x = 1 \times 10^{-16}$ mol/L

19. C $K_{sp} = [Ba^{2+}][CO_3^{2-}] = x^2$
$x = \sqrt{5.0 \times 10^{-9}} = 7.1 \times 10^{-5}$ mol/L

20. A $K_{sp} = [Ag^{+}]^2[CO_3^{2-}] = (2x)^2(x) = 4x^3$

$x = (K_{sp}/4)^{1/3} = 1.3 \times 10^{-4}$ mol/L

21. $K_{sp} = [Zn^{2+}][OH^-]^2 = (x)(2x)^2 = 4x^3$

    $x = (K_{sp}/4)^{1/3} = 4.2 \times 10^{-6}$ mol/L

22. $K_{sp} = [La^{3+}][F^-]^3 = (x)(3x)^3 = 27x^4$

    $x = (K_{sp}/27)^{1/4} = 9.3 \times 10^{-6}$ mol/L

Factors Affecting Solubility: Common Ion Effect

23. B
24. A
25. B
26. D $K_{sp} = [Pb^{+2}]\ [Cl^-]^2$

    $[Cl^-] = \sqrt{}K_{sp}/[Pb^{+2}]$

    $[Cl^-] = \sqrt{}\ 1.0 \times 10^{-4}/ 0.100 = 3.2 \times 10^{-2}$
27. E
28. B
29. D
30. A

    $K_{sp} = [Ag^+]^2[CO_3^{2-}]$

    $[Ag^+] = \sqrt{}K_{sp}/[\ CO_3^{2-}]$

    $[Ag^+] = \sqrt{}\ 8.1 \times 10^{-12}/ 0.025 = 1.8 \times 10^{-5}$

Factors Affecting Solubility: pH

31. C

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

    $x = (K_{sp}/4)^{1/3} = 5.3 \times 10^{-6}$ mol/L
32. B

    pH = 10.0, pOH = 4.0 $[OH^-] = 1.0 \times 10^{-4}$

    $K_{sp} = [Ni^{2+}][OH^-]^2$

    $6.0 \times 10^{-16} = [Ni^{2+}]\ (1.0 \times 10^{-4})^2$

    $[Ni^{2+}] = 6.0 \times 10^{-16}/(1.0 \times 10^{-4})^2$

    $[Ni^{2+}] = 6.0 \times 10^{-8}$
33. E pH = 11.5, pOH = 2.5 $[OH^-] = 3.2 \times 10^{-3}$

    $K_{sp} = [Ni^{2+}][OH^-]^2$

    $6.0 \times 10^{-16} = [Ni^{2+}]\ (3.2 \times 10^{-3})^2$

    $[Ni^{2+}] = 6.0 \times 10^{-16}/(3.2 \times 10^{-3})^2$

    $[Ni^{2+}] = 5.9 \times 10^{-11}$
34. E
35. B $K_{sp} = [Cu^{2+}][OH^-]^2 = (x)(2x)^2 = 4x^3$

    $x = (K_{sp}/4)^{1/3} = 2.3 \times 10^{-7}$ mol/L

36. B

    pH = 9.5, pOH = 4.5 $[OH^-] = 3.2 \times 10^{-5}$

    $K_{sp} = [Cu^{2+}][OH^-]^2$

    $4.8 \times 10^{-20} = [Cu^{2+}]\ (3.2 \times 10^{-5})^2$

    $[Cu^{2+}] = 4.8 \times 10^{-20}/(3.2 \times 10^{-5})^2$

    $[Cu^{2+}] = 4.7 \times 10^{-11}$

37.

pH = 12.2, pOH = 1.8 $[OH^-] = 1.6 \times 10^{-2}$
$K_{sp} = [Cu^{2+}][OH^-]^2$
$4.8 \times 10^{-20} = [Cu^{2+}] (1.6 \times 10^{-2})^2$
$[Cu^{2+}] = 4.8 \times 10^{-20}/(1.6 \times 10^{-2})^2$
$[Cu^{2+}] = 1.9 \times 10^{-16}$

38. B

Selective Precipitation: Precipitation Reactions

39. B
40. C
41. $Q = [Ag^+] [IO_3^-]$
$[Ag^+] = (0.010)(0.10L)/(.110 L) = 9.09 \times 10^{-3}$
$[IO_3^-] = (0.015)(0.01L)/(.110L) = 1.36 \times 10^{-3}$

$Q = (9.09 \times 10^{-3}) (1.36 \times 10^{-3}) = 1.24 \times 10^{-5}$
Since Q > K, then there will be a precipitate of $Ag(IO_3)$
42. D
43. $Q = [Ag^+]^2 [SO_4^{2-}]$
$[Ag^+] = (0.050)(0.50L)/(.125 L) = 0.20M$
$[IO_3^-] = (0.005)(0.75L)/(.125L) = 0.03M$

$Q = (0.20)^2(0.03) = 1.2 \times 10^{-3}$
Since Q > K, then there will be a precipitate of Ag_2SO_4

44. 0.96 g $(NH_4)_2CO_3$ x 1 mol/96 g = 1 mol/10 L = 0.10 M CO_3^{2-}
0.20 g $CaBr_2$ x 1 mol/199.9 g = 1.0×10^{-3} M Ca^{2+}
$Q = [Ca^{2+}][CO_3^{2-}] = 1.0 \times 10^{-3})(0.10) = 1 \times 10^{-4}$
Since Q > K, then there will be a precipitate of $CaCO_3$

Selective Precipitation: Separation of Ions

45. B

$[I^-]$ needed to precipitate AgI:
$[I^-] = Ksp / [Ag^+] = 8.3 \times 10^{-17} / 2.0 \times 10^{-4} = 4.2 \times 10^{-13}$ mol/L

$[I^-]$ needed to precipitate PbI_2:

$[I^-] = \sqrt{}Ksp / [Pb^{2+}] = \sqrt{}7.9 \times 10^{-9} / 1.5 \times 10^{-3} = 2.3 \times 10^{-3}$ M
Since a lower iodide ion concentration is needed for AgI, it will precipitate first.

46. Ag^+ precipitates first. The question asks what will be the concentration of Ag^+ when Pb^{2+} begins to precipitate.
$K_{sp} = [Ag^+] [I^-] = 8.3 \times 10^{-17}$
$[Ag^+] = 8.3 \times 10^{-17} / [I^-]$
$[I^-]$ needed to precipitate Pb^{2+} is 2.3×10^{-3}.

$[Ag^+] = 8.3 \times 10^{-17} / 2.3 \times 10^{-3}$

$[Ag^+] = 3.6 \times 10^{-14}$ when the PbI_2 begins to precipitate.

47. A

$[CO_3^{2-}]$ needed to precipitate $PbCO_3$:
$[CO_3^{2-}] = Ksp / [Pb^{2+}] = 7.4 \times 10^{-14} / 1.0 \times 10^{-5} = 7.4 \times 10^{-9}$ mol/L

$[CO_3^{2-}]$ needed to precipitate $MgCO_3$:
$[CO_3^{2-}] = Ksp / [Mg^{2+}] = 5.0 \times 10^{-10} / 2.2 \times 10^{-4} = 2.3 \times 10^{-6}$ mol/L

Since a lower carbonate ion concentration is needed for $PbCO_3$, it will precipitate first.

48. Pb^{2+} precipitates first. The question asks what will be the concentration of Pb^{2+} when Mg^{2+} begins to precipitate.
$K_{sp} = [Pb^{2+}] [CO_3^{2-}] = 7.4 \times 10^{-14}$
$[Pb^{2+}] = 7.4 \times 10^{-14} / [CO_3^{2-}]$
$[CO_3^{2-}]$ needed to precipitate Mg^{2+} is 2.3×10^{-6} mol/L

$[Pb^{2+}] = 7.4 \times 10^{-14} / 2.3 \times 10^{-6}$ mol/L

$[Pb^{2+}] = 3.2 \times 10^{-8}$ mol/L when the $MgCO_3$ begins to precipitate.

49. B

$[SO_4^{2-}]$ needed to precipitate $BaSO_4$:
$[SO_4^{2-}] = Ksp / [Ba^{2+}] = 1.1 \times 10^{-10} / 0.010 = 1.1 \times 10^{-8}$ mol/L

$[SO_4^{2-}]$ needed to precipitate $SrSO_4$:

$[SO_4^{2-}] = Ksp / [Sr^{2+}] = 3.2 \times 10^{-7} / 0.010 = 3.2 \times 10^{-5}$ mol/L

Since a lower sulfate ion concentration is needed for $BaSO_4$, it will precipitate first.

50. Ba^{2+} precipitates first. The question asks what will be the concentration of Ba^{2+} when Sr^{2+} begins to precipitate.
$K_{sp} = [Ba^{2+}] [SO_4^{2-}] = 1.1 \times 10^{-10}$
$[Ba^{2+}] = 1.1 \times 10^{-10} / [SO_4^{2-}]$
$[SO_4^{2-}]$ needed to precipitate Sr^{2+} is 3.2×10^{-5} mol/L

$[Ba^{2+}] = 1.1 \times 10^{-10} / 3.2 \times 10^{-5} = 3.4 \times 10^{-6}$ mol/L

51. D

$[OH^-]$ needed to precipitate $Ca(OH)_2$:
$[OH^-] = \sqrt{}Ksp / [Ca^{2+}]$

$[OH^-] = \sqrt{6.5 \times 10^{-6}/0.0250}$
$[OH^-] = 0.016$ mol/L

$[OH^-]$ needed to precipitate $Fe(OH)_2$:
$[OH^-] = \sqrt{Ksp / [Fe^{2+}]}$
$[OH^-] = \sqrt{7.9 \times 10^{-16}/1.5 \times 10^{-3}}$
$[OH^-] = 7.3 \times 10^{-7}$ mol/L

Since a lower hydroxide ion concentration is needed for $Fe(OH)_2$, it will precipitate first.

52. Fe^{2+} precipitates first. The question asks what will be the concentration of Fe^{2+} when Ca^{2+} begins to precipitate.
$K_{sp} = [Fe^{2+}] [OH^-]^2 = 7.9 \times 10^{-16}$
$[Fe^{2+}] = K_{sp} / [OH^-]^2$
$[OH^-]$ needed to precipitate Ca^{2+} is 0.016 mol/L

$[Fe^{2+}] = 7.9 \times 10^{-16} / (0.016)^2 = 3.1 \times 10^{-12}$ mol/L

Conceptual questions:

1) D	8) D	15) D
2) B	9) A	16) C
3) A	10) A	17) D
4) B	11) A	18) C
5) D	12) C	19) C
6) A	13) B	20) A
7) A	14) C	

Answers for Free Response Questions

1)
(a) $Fe(OH)_{2(s)} \leftrightarrow Fe^{2+}_{(aq)} + 2\ OH^{-+}_{(aq)}$
(b)

$$\frac{1.43\times10^{-3}\text{ g}}{\text{L}} \times \frac{1\text{ mol}}{89.9\text{ g}} = 1.59\times10^{-5}\ \text{mol}/\text{L}\ Fe(OH)_2$$

$= 1.59\text{x}10^{-5}$ M $= [Fe^{2+}]$
$= 3.18\text{x}10^{-5}$ M $= [OH^-]$ $[OH^-] = 2\,[Fe^{2+}]$

$K_{sp} = [Fe^{2+}][OH^-]^2 = (1.59\text{x}10^{-5})(3.18\text{x}10^{-5})^2$

$= 1.61\text{x}10^{-14}$

$[H^+] = 1.0 \text{ x } 10^{-14}/3.18 \text{ x } 10^{-5} = 3.14 \text{ x } 10^{-10}$

(c) $pH = -\log[H^+] = 9.50$

OR

$pOH = -\log[OH^-] = -\log(3.18\text{x}10^{-8}) = 4.50$

$pH = 14 - pOH = 9.50$

(d) 50.0 mL of $3.00\text{x}10^{-3}$ M Fe^{2+} diluted of 100.0 mL = $1.50\text{x}10^{-3}$ M Fe^{2+}

50.0 mL of $4.00\text{x}10^{-6}$ M OH^- diluted of 100.0 mL = $2.00\text{x}10^{-6}$ M OH^-

$Q = [Fe^{2+}][OH^-]^2 = (1.50\text{x}10^{-3})(2.00\text{x}10^{-6})^2$

$= 6.00\text{x}10^{-15}$

Precipitate will <u>NOT</u> form since $Q < K_{sp}$

2)

(a)

i. $Cu(OH)_{2(s)} \rightarrow Cu^{2+}{}_{(aq)} + 2\ OH^-{}_{(aq)}$

ii. $\frac{1.72 \text{ x } 10^{-6}}{0.100\text{ L}} \text{ x } \frac{1\text{ mol}}{97.5\text{ g}} = 1.76 \text{ x } 10^{-7}\ {}^{mol}/_{L}$

iii. $K_{sp} = [Cu^{2+}][OH^-]^2$

$= [1.76\text{x}10^{-7}][3.53\text{x}10^{-7}]^2 = 2.20\text{x}10^{-20}$

(b)

i. $Zn(OH)_{2(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2\ OH^-{}_{(aq)}$

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

pH 9.35 = pOH 4.65; $[OH^-] = 10^{-pOH}$

$[OH^-] = 10^{-4.65} = 2.24\text{x}10^{-5}\ M$

$[Zn^{2+}]$ = solubility of $Zn(OH)_2$ in ${}^{mol}/_{L}$

$[Zn^{2+}] = K_{sp} / [OH^-]^2 = (7.7 \text{ x } 10^{-17})/(2.24 \text{ x } 10^{-5})^2 = 1.5 \text{ x } 10^{-7}$ M

ii. $[Zn^{2+}]_{init} = 0.100M \text{ x } \frac{50\text{ mL}}{100\text{ mL}} = 0.0500\ M$

$[OH^-]_{init} = 0.300M \text{ x } \frac{50\text{ mL}}{100\text{ mL}} = 0.150\ M$

X = conc. loss to get to equilibrium

$K_{sp} = 7.7\text{x}10^{-17} = (0.0500 - X)(0.150 - 2X)^2$ X = 0.05

All the Zn^{2+} ions get precipitated.

OR

Zn^{2+} + $2(OH)^-$ $\rightarrow Zn(OH)_2$

0.05M 0.15M initially

All the Zn2+ ions will be precipitated as the Zn^{2+} ions is the limiting here. In that situation we have an equilibrium as $Zn(OH)_2 \rightarrow Zn^{2+} + 2(OH)^-$ and the Ksp expression should be followed.

$Ksp = 7.7\ 10^{-17} = 4x^3$, $x = [Zn^{2+}] = 2.7\ 10^{-6}\ M$

3)

(a) $Ksp = [Mg^{2+}][F^-]^2$

$= (1.21 \times 10^{-3})\ (2 \times 1.21 \times 10^{-3})^2$

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

(b) $Ksp = [Mg^{2+}]\ (0.1)^2$

$7.09 \times 10\text{-}9 = [Mg^{2+}]\ (0.1)^2$

$[Mg^{2+}] = (7.09 \times 10^{-9}) / (0.01)$

$= 7.09 \times 10^{-7}$ M

(c) $[Mg^{2+}] = 3.00 \times 10^{-3} \times 0.1L\ /0.3l = 0.001M$

$[F^-] = 2.00 \times 10^{-3} \times 0.2L / 0.3L = 0.00133M$

Q = Ion Product = $[Mg^{2+}]\ [F^-]^2$

$= (1.00 \times 10^{-3})\ (1.33 \times 10^{-3})^2$

$= 1.77 \times 10^{-9}$

Since $Q < K_{sp}$, no precipitate will form.

Note: conclusion must be consistent with Q value.

d) Solubility of MgF_2 decreases with the increasing temperature, thus dissolution process is exothermic.

$MgF_2 \rightarrow Mg_2 + + 2F^- + Q$ (or H)

Reason:

i) Increased temperature puts a stress on the system (LeChâtelier). The system will reduce the stress by shifting the equilibrium in the endothermic (left) direction.

4) (a) i) $AgCl(s) \rightarrow Ag^+(aq) + Cl^-(aq)$

ii) $\dfrac{8.9 \times 10^{-5}g}{143.32\ g/mol} = 6.2 \times 10^{-7}$ (in 100 mL)

$(6.2 \times 10^{-7}\ mol/100\ mL)\ (1,000\ mL/\ 1\ L) = 6.2 \times 10^{-6}$

iii) $K_{sp} = [Ag^+][Cl^-] = (6.2 \times 10^{-6})^2 = 3.8 \times 10^{-11}$

(b) Find Q and see if it is bigger than the K_{sp}.

$[Cl^-]=0.0200M$, $[Pb^{2+}]=0.0150M$ (Note since the volume is doubled the concentrations are cut in half—$M_1V_1=M_2V_2$)

$PbCl_2 \leftrightarrow Pb^{2+} + 2Cl^-$, $Q = [Pb^{2+}][Cl^-]^2 = [0.0150][0.0200]^2 = 6.00 \times 10^{-6}$

Since $6.00 \times 10^{-6} < 1.6 \times 10^{-5}$, no precipitate will form.

(c) Use common ion effect. You will know the [Cl-] and then solve for the $[Pb^{2+}]$ using the K_{sp} equation.

$K_{sp}=1.6 \text{ x } 10^{-5}=[Pb^{2+}][0.25M]^2$

Therefore, $[Pb^{2+}]=2.56 \text{ x } 10^{-4}$ M

5)
(a) $K_{sp} = 5.0 \text{ x } 10^{-13} = [Ag^+][Br^-]$

(b) $[Ag^+]= [Br^-] = (Ksp)^{1/2} = 7.1 \text{ x } 10^{-7}$ M (saturated)

(c) Adding distilled water to a solution with solid (it is already saturated)will not change the concentration. It will remain the same. The value of [Ag+] after addition of distilled water is equal to the value in part (b). The concentration of ions in solution in equilibrium with a solid does <u>not</u> depend on the volume of the solution.

(d) moles AgBr = 5.0/188= 0.027 moles
$5.0 \text{ x } 10^{-13} = (0.027/V)(0.027/V) = 7.3 \text{ x } 10^{-4}/V^2$
$V = 3.8 \text{ x } 10^4$ L
(e) moles $Ag^+ = 0.010 \text{ x } 1.5 \text{ x } 10^{-4} = 1.5 \text{ x } 10^{-6}$
moles $Br^- = 2.0 \text{ x } 10^{-3}$ L x $5.0 \text{ x } 10^{-4}$ M= $1.0 \text{ x } 10^{-6}$
total volume = 0.012 L
$[Br^-] = 1.0 \text{ x } 10^{-6} / 0.012$ L=$8.3 \text{ x } 10^{-5}$ M
$[Ag^+] = 1.5 \text{ x } 10^{-6} / 0.012$ L=$1.3 \text{ x } 10^{-4}$ M
$Q = 1.3 \text{ x } 10^{-4} \text{ x } 8.3 \text{ x } 10^{-5} = 1.0 \text{ x } 10^{-8}$
Q > Ksp so precipitation should occur
(f) ~~Na +~~ I- (aq) + AgBr (s) → AgI (s) + Br- (aq) + ~~Na+~~ solid AgI - yellow colored, is precipitated. The solubility of AgBr is greater than the solubility of AgI. In presence of Br- and I- ions, the less soluble AgI got precipitated.

6)
a) i) $108x^5$
ii) $6912x^4$
iii) $27x^4$
iv) $4x^3$

b) $K_{sp} = 3.2 \text{ x } 10^{-20} = x^2$; $x = 1.8 \text{ x } 10^{-10}$ mol/L

c) Recall that the K_{sp} expression can be written as $K_{sp} = [Ba^{2+}][F^-]^2$ for BaF_2. One mole of BaF_2 creates two moles of F^- and one mole of Ba^{2+} per mol of BaF_2 dissolved.
$x = [F^-]/2 = .0075$
$K_{sp} = 4x^3 = 4(0.0075^3) = 1.7 \text{ x } 10^{-6}$
OR
$Ba^{2+}] = ½ [F-] = ½ \text{ x } 0.015 = 0.0075M$ substitute in ksp
$0.0075 \text{ x } 0.015^2 \quad 1.7 \text{ x } 10^{-6}$

d) Molar solubility = $(K_{sp})^{1/2}$ = 4 x 10^{-12}

e) K_{sp} = [Zn^{2+}][S^{2-}] = (x)(.1) = 2 x 10^{-25} ; x = 2 x 10^{-24}

7)

a) Ag2CrO_4 ↔ 2Ag^+ (aq) + CrO_4^{2-}(aq) ; Ksp = 2x^2 . x = 4x^3

b) 2.6 x10^{-12} = 4x^3 ; x^3 = (2.6 x10^{-12})/4 ; x = 3(2.6 x10^{-12} $)^{1/2}$ = 8.7x 10^{-5}

[Ag^+] = 2x = 2 x 8.7x 10^{-5} = 1.7 x10^{-4}

c) X = [CrO_4^{2-}] ; 1mol Ag2CrO_4 = 1 mol [CrO_4^{2-}]

(8.7 x 10^{-5} x 0.1L)/1L x 331.7g/mol = 0.0029g/L

d) 0.1 mol AgNO3 + 1L of Ag2CrO4

Ag2CrO_4 ↔ 2Ag^+ (aq) + CrO_4^{2-}(aq)

When more Ag^+ ions are added, the equilibrium will shift to the left producing solid Ag_2CrO_4 . Therefore [CrO4^{2-}] will decrease.

e) Ag3PO4 ↔ 3 Ag^+ + $(PO4)^{3-}$

f) K_{sp} = [Ag^+ $]^3$ [PO4^{3-}] = [5.3 x10^{-5} $]^3$ x (5.3 x10^{-5})/3

= 2.6 x 10^{-18}

g)The [Ag^+] will remain the same as long as the solution is saturated. It is independent of the volume of the saturated solution.

8) a) NaI + Ag(NO_3) ↔ AgI(s) + $NaNO_3$(aq)

2NaI + Pb$(NO_3)_2$ ↔ PbI_2(s) + 2$NaNO_3$ (aq)

b) AgI; PbI_2

c) AgI: 8.3 x 10^{-17} = [I^-][0.002 M] ; [I^-] = 4.15 x 10^{-14}

PbI_2:7.9 x 10^{-9} = [$I^-$$]^2$[0.5 M] ; [$I^-$] = 1.26 x 10^{-4}

AgI will precipitate first