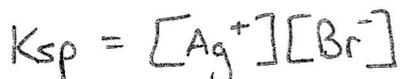
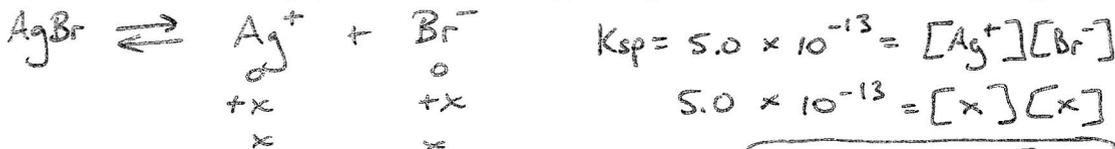


Ksp Problems Worksheet #4
Review over all types of Ksp Problems

1. Write the equilibrium expression for the solubility product constant (Ksp) for AgBr.



2. What is the molar solubility (maximum molarity) of Ag⁺ in a saturated solution of AgBr?



$$x = 7.1 \times 10^{-7} M$$

3. What is the molar solubility of AgBr in a 0.015 M KBr solution?



$$x = 3.3 \times 10^{-11} M$$

4. Would a precipitate be observed if the following were mixed?

100.0 mL of 0.00015 M AgNO₃ and 20.0 mL of 0.00050 M NaBr

$$[Ag^+] = \frac{(0.00015)(.1000)}{.120} = 1.25 \times 10^{-4} M$$

$$Q = [1.25 \times 10^{-4}][8.33 \times 10^{-5}]$$

$$Q = 1.0 \times 10^{-8}$$

$$[Br^-] = \frac{(0.00050)(.0200)}{.120} = 8.33 \times 10^{-5} M$$

Q > Ksp so a ppt will form

5. A person adds solid AgNO₃ crystals into a large beaker containing a mixture of 0.0022 M KCl and 0.00011M KBr until a precipitate begins to form. Was the observed precipitate AgCl or AgBr?

AgCl



$$K_{sp} = 1.8 \times 10^{-10} = [Ag^+][.0022]$$

$$[Ag^+] = 8.2 \times 10^{-8} M$$

AgBr



$$K_{sp} = 5.0 \times 10^{-13} = [Ag^+][.00011]$$

$$[Ag^+] = 4.5 \times 10^{-9} M$$

AgBr will ppt first because the molarity of Ag⁺ will increase from zero to 4.5 x 10⁻⁹ M before it reaches the molarity of 8.2 x 10⁻⁸ M to precipitate the rest of AgCl

6. What is the molar solubility of CaSO_4 in pure water at 25°C ? $\text{CaSO}_4 \rightleftharpoons \text{Ca}^{2+} + \text{SO}_4^{2-}$

$$K_{sp} = 2.4 \times 10^{-5} = [\text{Ca}^{2+}][\text{SO}_4^{2-}]$$

$$2.4 \times 10^{-5} = [x][x]$$

$$x = .0049 \text{ M} = \text{max molarity of } \text{CaSO}_4$$

7. What is the molar solubility of CaF_2 in pure water at 25°C ?

$$K_{sp} = 3.9 \times 10^{-11} = [\text{Ca}^{2+}][\text{F}^-]^2$$

$$3.9 \times 10^{-11} = [x][2x]^2$$

$$3.9 \times 10^{-11} = 4x^3$$

$$x = 2.1 \times 10^{-4} \text{ M} = \text{max molarity of } \text{CaF}_2$$

8. At 10°C , $8.9 \times 10^{-5} \text{ g}$ of $\text{AgCl}_{(s)}$ will dissolve in 100. mL of water. *Cannot use our pink sheet since its values are based on a temperature of 25°C .*

(i) Write the equation for the dissociation of $\text{AgCl}_{(s)}$ in water.



(ii) Calculate the maximum molarity of $\text{AgCl}_{(s)}$ in water at 10°C .

moles AgCl that dissolves in 100. mL of H_2O

$$\frac{8.9 \times 10^{-5}}{143.5} = 6.2 \times 10^{-7} \text{ moles AgCl}$$

→
wt of AgCl

$$[\text{AgCl}] = \frac{6.2 \times 10^{-7}}{.100 \text{ Liters}} = 6.2 \times 10^{-6} \text{ M}$$

(iii) Calculate the value of the solubility-product constant, K_{sp} for $\text{AgCl}_{(s)}$ at 10°C .

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

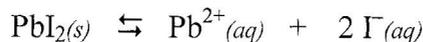
since our problem(ii) calculated the maximum molarity for AgCl , we know the molarities for the ions.

$$K_{sp} = [6.2 \times 10^{-6}][6.2 \times 10^{-6}]$$

$$K_{sp} = 3.8 \times 10^{-11}$$

9. Answer the following questions that relate to the solubility of salts of lead and barium.

- (a) A saturated solution is prepared by adding excess $\text{PbI}_2(s)$ to distilled water to form 1.0 L of solution at 25°C . The concentration of $\text{Pb}^{2+}(aq)$ in the saturated solution is 0.0013 M . The chemical equation for the dissolution of $\text{PbI}_2(s)$ in water is shown below.



- (i) Write the equilibrium-constant expression for the equation.

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$$

- (ii) Calculate the molar concentration of $\text{I}^{-}(aq)$ in the solution.

Problem states that $[\text{Pb}^{2+}] = .0013\text{ M}$, therefore the molarity of I^{-} would be twice that #.

$$[\text{I}^{-}] = .0026\text{ M}$$

- (iii) Calculate the value of the equilibrium constant, K_{sp} .

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$$

$$K_{sp} = [.0013][.0026]^2$$

$$K_{sp} = 8.8 \times 10^{-9}$$

- (b) A saturated solution is prepared by adding $\text{PbI}_2(s)$ to distilled water to form 2.0 L of solution at 25°C . What are the molar concentrations of $\text{Pb}^{2+}(aq)$ and $\text{I}^{-}(aq)$ in the solution? Justify your answer.

Saturated solution of PbI_2 has a molarity of 0.0013 M (as stated @ top of page). Regardless of the volume of the solution, the molarities will be $[\text{Pb}^{2+}] = .0013\text{ M}$
 $[\text{I}^{-}] = .0026\text{ M}$

- (c) Solid NaI is added to a saturated solution of PbI_2 at 25°C . Assuming that the volume of the solution does not change, does the molar concentration of $\text{Pb}^{2+}(aq)$ in the solution increase, decrease, or remain the same? Justify your answer.



Solid NaI is a delightful source of I^{-} . The addition of I^{-} to a saturated solution of PbI_2 would cause the equilibrium to shift to the left, reducing the concentration of Pb^{2+} . (Le Chat Principle)

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(d) When a 500. mL sample of $8.2 \times 10^{-6} M$ $Ba(NO_3)_2$ is added to 500. mL of $8.2 \times 10^{-6} M$ Na_2CrO_4 , no precipitate is observed. The value of K_{sp} for the salt $BaCrO_4$ is 1.2×10^{-10} .

(i) Calculate the molar concentrations of $Ba^{2+}_{(aq)}$ and $CrO_4^{2-}_{(aq)}$ in the combined 1.00 L of solution.

$$[Ba^{2+}] = \frac{(8.2 \times 10^{-6})(.500)}{1.00} = 4.1 \times 10^{-6} M$$

$$[CrO_4^{2-}] = \frac{(8.2 \times 10^{-6})(.500)}{1.00} = 4.1 \times 10^{-6} M$$

(ii) Use the molar concentrations of $Ba^{2+}_{(aq)}$ ions and $CrO_4^{2-}_{(aq)}$ ions as determined above to show why a precipitate does not form. You must include a calculation as part of your answer.

$$Q = [Ba^{2+}][CrO_4^{2-}]$$

$$Q = [4.1 \times 10^{-6}][4.1 \times 10^{-6}]$$

$$Q = 1.7 \times 10^{-11}$$

K_{sp} for $BaCrO_4$ (listed above) is 1.2×10^{-10}

Because $Q < K_{sp}$ will not see a ppt
