

**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<sup>+</sup> 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 AgNO<sub>3</sub> 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 AgNO<sub>3</sub> 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



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

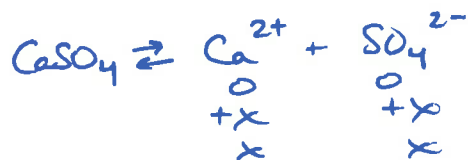
$$[Ag^+] = 4.5 \times 10^{-9} M$$

AgBr will ppt first because the molarity of Ag<sup>+</sup> will increase from zero to 4.5 × 10<sup>-9</sup> M before it reaches the required 8.2 × 10<sup>-8</sup> M to observe the ppt of AgCl

6. What is the molar solubility of  $\text{CaSO}_4$  in pure water at  $25^\circ\text{C}$ ?

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

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



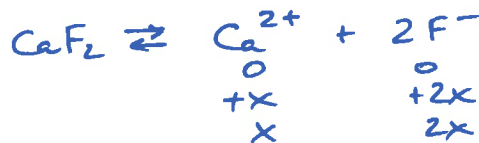
$$x = .0049 \text{ M} = \text{max molarity of } \text{CaSO}_4$$

7. What is the molar solubility of  $\text{CaF}_2$  in pure water at  $25^\circ\text{C}$ ?

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

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

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



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

8. At  $10^\circ\text{C}$ ,  $8.9 \times 10^{-5}$  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  $25^\circ\text{C}$ .*

(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  $10^\circ\text{C}$ .

moles  $\text{AgCl}$  that dissolves in 100. mL of  $\text{H}_2\text{O}$

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

→  
wt of  $\text{AgCl}$

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

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

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

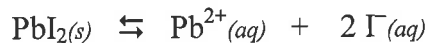
since our problem(ii) calculated the maximum molarity for  $\text{AgCl}$ , we know the molarities for the ions.

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

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

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  $25^\circ\text{C}$ . 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  $\text{I}^-$  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  $25^\circ\text{C}$ . What are the molar concentrations of  $\text{Pb}^{2+}(aq)$  and  $\text{I}^-(aq)$  in the solution? Justify your answer.

Saturated solution of  $\text{PbI}_2$  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  $\text{NaI}$  is added to a saturated solution of  $\text{PbI}_2$  at  $25^\circ\text{C}$ . 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  $\text{NaI}$  is a delightful source of  $\text{I}^-$ . The addition of  $\text{I}^-$  to a saturated solution of  $\text{PbI}_2$  would cause the equilibrium to shift to the left, reducing the concentration of  $\text{Pb}^{2+}$ . (Le Chat Principle)

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(d) When a 500. mL sample of  $8.2 \times 10^{-6} M \text{Ba}(\text{NO}_3)_2$  is added to 500. mL of  $8.2 \times 10^{-6} M \text{Na}_2\text{CrO}_4$ , no precipitate is observed. The value of  $K_{sp}$  for the salt  $\text{BaCrO}_4$  is  $1.2 \times 10^{-10}$ .

(i) Calculate the molar concentrations of  $\text{Ba}^{2+}_{(aq)}$  and  $\text{CrO}_4^{2-}_{(aq)}$  in the combined 1.00 L of solution.

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

$$[\text{CrO}_4^{2-}] = \frac{(8.2 \times 10^{-6})(.500)}{1.00} = 4.1 \times 10^{-6} M$$

(ii) Use the molar concentrations of  $\text{Ba}^{2+}_{(aq)}$  ions and  $\text{CrO}_4^{2-}_{(aq)}$  ions as determined above to show why a precipitate does not form. You must include a calculation as part of your answer.

$$Q = [\text{Ba}^{2+}][\text{CrO}_4^{2-}]$$

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

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

$K_{sp}$  for  $\text{BaCrO}_4$  (listed above) is  $1.2 \times 10^{-10}$

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