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AP Chemistry

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Period _____

Ksp Problems Worksheet #4

Review over all types of Ksp Problems

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

$Ksp = [Ag^+][Br^-]$

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

$AgBr \rightleftharpoons Ag^+ + Br^-$

Ag^+: 0, +x, x; Br^-: 0, +x, x

$Ksp = 5.0 \times 10^{-13} = [Ag^+][Br^-]$

$5.0 \times 10^{-13} = [x][x]$

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

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

$AgBr \rightleftharpoons Ag^+ + Br^-$

Ag^+: 0, +x, x; Br^-: .015M, +x, .015+x

$5.0 \times 10^{-13} = [x][.015+x]$

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

4. Would a precipitate be observed if the following were mixed?
   100.0 mL of 0.00015 M $AgNO_3$ and 20.0 mL of 0.00050 M NaBr

$[Ag^+] = \frac{(.00015)(.1000)}{.120} = 1.25 \times 10^{-4}$ M

$[Br^-] = \frac{(.00050)(.0200)}{.120} = 8.33 \times 10^{-5}$ M

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

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

Q > Ksp so a ppt will form

5. A person adds solid $AgNO_3$ 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

$AgCl \rightleftharpoons Ag^+ + Cl^-$

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

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

AgBr

$AgBr \rightleftharpoons Ag^+ + Br^-$

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

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

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

6. What is the molar solubility of $CaSO_4$ in pure water at 25°C?

$CaSO_4 \rightleftharpoons Ca^{2+} + SO_4^{2-}$

	Ca^{2+}	SO_4^{2-}
	0	0
	+x	+x
	x	x

$K_{sp} = 2.4 \times 10^{-5} = [Ca^{2+}][SO_4^{2-}]$

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

$x = .0049$ M = max molarity of $CaSO_4$

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

$CaF_2 \rightleftharpoons Ca^{2+} + 2F^-$

	Ca^{2+}	$2F^-$
	0	0
	+x	+2x
	x	2x

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

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

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

$x = 2.1 \times 10^{-4}$ M = max molarity of CaF_2

8. At 10°C, 8.9×10^{-5} g of 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 AgCl(*s*) in water.

$AgCl \rightleftharpoons Ag^+ + Cl^-$

(ii) Calculate the maximum molarity of AgCl(*s*) in water at 10°C.

moles AgCl that dissolves in 100. mL of H_2O

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

fwt of AgCl

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

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

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

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 $PbI_{2(s)}$ to distilled water to form 1.0 L of solution at 25°C. The concentration of $Pb^{2+}(aq)$ in the saturated solution is 0.0013 M. The chemical equation for the dissolution of $PbI_{2(s)}$ in water is shown below.

$$PbI_{2(s)} \leftrightarrows Pb^{2+}(aq) + 2\, I^-(aq)$$

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

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

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

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

$[I^-] = .0026$ M

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

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

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

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

(b) A saturated solution is prepared by adding $PbI_{2(s)}$ to distilled water to form 2.0 L of solution at 25°C. What are the molar concentrations of $Pb^{2+}(aq)$ and $I^-(aq)$ 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 $[Pb^{2+}] = .0013$ M

$[I^-] = .0026$ 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 $Pb^{2+}(aq)$ in the solution increase, decrease, or remain the same? Justify your answer.

$PbI_2 \rightleftarrows Pb^{2+} + 2I^-$

Solid NaI is a delightful source of I^-. The addn 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 8.2×10^{-6} M $Ba(NO_3)_2$ is added to 500. mL of 8.2×10^{-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.

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

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

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

Ksp for BaCrO4 (listed above) is 1.2×10^{-10}

Because Q < Ksp will not see a ppt