

REVIEW QUESTIONS

Chapter 17

1. Calculate the molar solubility of AgBr ($K_{sp} = 5.0 \times 10^{-13}$) in 0.50 M NaBr solution.

	AgBr (s)	→	Ag ⁺ (aq) +	Br ⁻ (aq)
Initial	----		0	0.50
Δ	-x		+x	+x
Final	-----		x	0.50 + x

$$K_{sp} = [\text{Ag}^+][\text{Br}^-] = (x)(0.50 + x) = 5.0 \times 10^{-13}$$

$$\text{solubility} = x = \frac{5.0 \times 10^{-13}}{0.50} = 1.0 \times 10^{-12} \text{ M}$$

2. AgNO₃ is added to a solution that is 0.10 M in NaCl and 0.010 M K₂CrO₄. Assume no dilution caused by the addition of AgNO₃. Given the K_{sp} values below:

$$K_{sp} \text{ for AgCl} = 1.6 \times 10^{-10}$$

$$K_{sp} \text{ for Ag}_2\text{CrO}_4 = 9.0 \times 10^{-12}$$

- a) Which precipitates first, AgCl or Ag₂CrO₄? Calculate the [Ag⁺] when precipitation first begins.

	AgCl (s)	→	Ag ⁺ (aq) +	Cl ⁻ (aq)
Initial	----		0	0.10
Δ	-x		+x	+x
Final	-----		x	0.10 + x

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (x)(0.10 + x) = 1.6 \times 10^{-10} \quad [\text{Ag}^+] = \frac{1.6 \times 10^{-10}}{0.10} = 1.6 \times 10^{-9} \text{ M}$$

	Ag ₂ CrO ₄ (s)	→	2 Ag ⁺ (aq) +	CrO ₄ ²⁻ (aq)
Initial	----		0	0.010
Δ	-x		+2x	+x
Final	-----		2x	0.010 + x

$$K_{sp} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}] = (2x)^2(0.010 + x) = 9.0 \times 10^{-12} \quad [x] = \sqrt{\frac{9.0 \times 10^{-12}}{4(0.010)}} = 1.5 \times 10^{-5} \text{ M}$$

$$[\text{Ag}] = 2x = 3.0 \times 10^{-5} \text{ M}$$

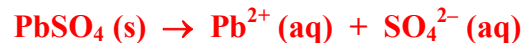
The precipitation occurs for the salt which requires the smallest [Ag⁺] at equilibrium. Therefore **AgCl precipitates first** at a [Ag⁺] of $1.6 \times 10^{-9} \text{ M}$

b) What is the $[Cl^-]$ when Ag_2CrO_4 first begins to precipitate?

Calculations above show that in order of Ag_2CrO_4 to precipitate, the $[Ag^+]$ must equal 3.0×10^{-5} M. Therefore,

$$K_{sp} = [Ag^+][Cl^-] = (3.0 \times 10^{-5}) [Cl^-] = 1.6 \times 10^{-10} \quad [Cl^-] = \frac{1.6 \times 10^{-10}}{3.0 \times 10^{-5}} = 5.3 \times 10^{-6} \text{ M}$$

3. A solution is made by mixing 10.0 mL of 0.10 M $Pb(NO_3)_2$ and 10.0 mL of 0.0010 M Na_2SO_4 . Will a precipitate form? (K_{sp} for $PbSO_4 = 1.06 \times 10^{-8}$)



$$[Pb^{2+}] = \frac{(0.10 \text{ M})(10.0 \text{ mL})}{(20.0 \text{ mL})} = 0.0500 \text{ M}$$

$$[SO_4^{2-}] = \frac{(0.0010 \text{ M})(10.0 \text{ mL})}{(20.0 \text{ mL})} = 5.00 \times 10^{-4} \text{ M}$$

$$Q_{sp} = [Pb^{2+}][SO_4^{2-}] = (0.0500)(5.00 \times 10^{-4}) = 2.50 \times 10^{-5}$$

Since $Q_{sp} > K_{sp}$, precipitation will occur

4. The solubility of iron (II) hydroxide, $Fe(OH)_2$, is 1.43×10^{-3} g/L.

a) Calculate the K_{sp} for iron (II) hydroxide.



$$[Fe^{2+}] = [Fe(OH)_2] = \frac{1.43 \times 10^{-3} \text{ g}}{1 \text{ L}} \times \frac{1 \text{ mol}}{89.85 \text{ g}} = 1.59 \times 10^{-5} \text{ M}$$

$$[OH^-] = 2 (1.59 \times 10^{-5}) = 3.18 \times 10^{-5} \text{ M}$$

$$K_{sp} = [Fe^{2+}][OH^-]^2 = (1.59 \times 10^{-5})(3.18 \times 10^{-5})^2 = 1.61 \times 10^{-14}$$

b) Calculate pH of a saturated solution of iron (II) hydroxide.

$$\text{From part (a)} \quad [OH^-] = 3.18 \times 10^{-5} \text{ M}$$

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{3.18 \times 10^{-5}} = 3.14 \times 10^{-10} \text{ M}$$

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

- c) A 50.0 mL sample of 3.00×10^{-3} M FeSO_4 solution is added to 50.0 mL of 4.00×10^{-6} M NaOH solution. Does a precipitate form?

$$[\text{Fe}^{2+}] = \frac{(3.00 \times 10^{-3} \text{ M})(50.0 \text{ mL})}{(100.0 \text{ mL})} = 1.50 \times 10^{-3} \text{ M}$$

$$[\text{OH}^-] = \frac{(4.00 \times 10^{-6} \text{ M})(50.0 \text{ mL})}{(100.0 \text{ mL})} = 2.00 \times 10^{-6} \text{ M}$$

$$Q_{\text{sp}} = [\text{Fe}^{2+}][\text{OH}^-]^2 = (1.50 \times 10^{-3})(2.00 \times 10^{-6})^2 = 6.00 \times 10^{-15}$$

Since $Q_{\text{sp}} < K_{\text{sp}}$, precipitation will not occur

5. Lead iodate, $\text{Pb}(\text{IO}_3)_2$, is a slightly soluble salt with a K_{sp} of 2.6×10^{-13} at 25°C . To 35.0 mL of 0.150 M $\text{Pb}(\text{NO}_3)_2$ solution is added 15.0 mL of 0.800 M KIO_3 . A precipitate of $\text{Pb}(\text{IO}_3)_2$ results. What are the $[\text{Pb}^{2+}]$ and $[\text{IO}_3^-]$ in the final solution?

$$[\text{Pb}^{2+}] = \frac{(0.150 \text{ M})(35.0 \text{ mL})}{(50.0 \text{ mL})} = 0.105 \text{ M}$$

$$[\text{IO}_3^-] = \frac{(0.800 \text{ M})(15.0 \text{ mL})}{(50.0 \text{ mL})} = 0.240 \text{ M}$$

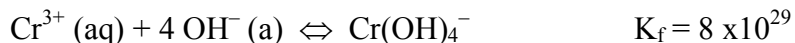
	$\text{Pb}(\text{IO}_3)_2$ (s)	\rightarrow	Pb^{2+} (aq)	+	2IO_3^- (aq)
Initial	----		0.105		0.240
Δ_1	+0.105		-0.105		-0.210
Precipitate	-----		0		0.030
Δ_2	-x		+x		+2x
Finish	-----		x		0.030+2x

$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{IO}_3^-]^2 = (x)(0.030 + 2x)^2 = 2.6 \times 10^{-13}$$

$$[\text{Pb}^{2+}] = x = \frac{2.6 \times 10^{-13}}{(0.030)^2} = 2.9 \times 10^{-10} \text{ M}$$

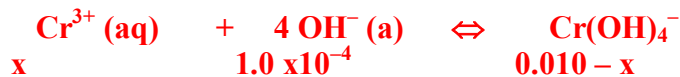
$$[\text{IO}_3^-] = 0.030 + 2x = 0.030 \text{ M}$$

6. What is the Cr^{3+} concentration when 0.010 mol of $\text{Cr}(\text{NO}_3)_3$ is dissolved in a liter of solution buffered at pH of 10.0. Cr^{3+} forms a complex ion with hydroxide shown below:



$$\text{pOH} = 14.0 - 10.0 = 4.0 \quad [\text{OH}^-] = \text{antilog pOH} = 1.0 \times 10^{-4}$$

Due to the large K_f all of Cr^{3+} is converted to the complex ion, and some subsequently dissociates back to Cr^{3+} . Then at equilibrium,



$$K_f = \frac{[\text{Cr}(\text{OH})_4^-]}{[\text{Cr}^{3+}][\text{OH}^-]^4} = \frac{(0.010 - x)}{x(1.0 \times 10^{-4})^4} = 8 \times 10^{29}$$

$$[\text{Cr}^{3+}] = x = 1.2 \times 10^{-16} \text{ M}$$

Note: Since the solution is buffered, $[\text{OH}^-]$ will remain constant during the reaction.

7. A 0.10-mol sample of AgNO_3 is dissolved in 1.00 L of 1.00 M NH_3 . If 0.010 mol of NaCl is added to this solution, will AgCl ($K_{sp} = 1.8 \times 10^{-10}$) precipitate? (Ag^+ and NH_3 form the complex ion $[\text{Ag}(\text{NH}_3)_2]^+$ with $K_f = 1.6 \times 10^7$)

First we need to determine the concentration of Ag^+ in the solution, and then calculate Q_{sp} to determine if a precipitate forms.

Due to the large K_f value, assume all of the Ag^+ reacts to form the complex and then some dissociates back.

	$\text{Ag}^+ (\text{aq})$	$+ 2 \text{NH}_3 (\text{aq})$	\rightarrow	$\text{Ag}(\text{NH}_3)_2^+ (\text{aq})$
Initial	0.10	1.00		0
Δ_1	-0.10	-0.20		+0.10
Complex	0	0.80		0.10
Δ_2	+x	+2x		-x
Finish	x	0.80+2x		0.10-x

$$K_f = \frac{[\text{Ag}(\text{NH}_3)_2^+]}{[\text{Ag}^+][\text{NH}_3]^2} = \frac{(0.10 - x)}{x(0.80 + 2x)^2} = 1.6 \times 10^7$$

$$[\text{Ag}^+] = x = \frac{0.20}{(0.80)^2(1.6 \times 10^7)} = 9.8 \times 10^{-9} \text{ M}$$

$$Q_{sp} = [\text{Ag}^+][\text{Cl}^-] = (9.8 \times 10^{-9})(0.010) = 9.8 \times 10^{-11} < 1.8 \times 10^{-10}$$

AgCl will NOT precipitate