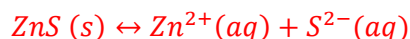


General Chemistry II

RR # 12
Summer 2022

1. What mass of ZnS ($K_{sp} = 2.5 \times 10^{-22}$) will dissolve in 300.0 mL of 0.050 M $Zn(NO_3)_2$? Ignore the basic properties of S^{2-} . "X" is negligible.



	$[Zn^{2+}]$	$[S^{2-}]$
I	0.050	0
C	+x	+x
E	0.050+x	x

$$K_{sp} = 2.5 \times 10^{-22} = [Zn^{2+}][S^{2-}] = x(0.050)$$

$$x = 5.0 \times 10^{-21} M = [S^{2-}]$$

$$\frac{5.0 \times 10^{-21} \text{ mol } S^{2-}}{L} * 0.300 L = 15 \times 10^{-21} \text{ mol } S^{2-} * \frac{1 \text{ mol } ZnS}{1 \text{ mol } S^{2-}} * \frac{97.45g}{1 \text{ mol } ZnS} = 1.46 \times 10^{-19}$$

2. Predict whether the equivalence point of each of the following titration is below, above, or at pH 7.
- $NaHCO_3$ titrated with NaOH basic
 - NH_3 titrated with HCl acidic
 - KOH titrated with HBr neutral

3. Consider a beaker containing a saturated solution of CaF_2 in equilibrium with undissolved CaF_2 . The molar solubility of CaF_2 at 35 degrees Celsius is 1.24×10^{-3} M.

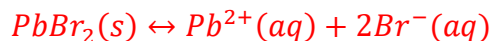
- a. What are the expression and the value for K_{sp} ?

$$K_{sp} = [Ca^{2+}][F^{2-}]^2 = [1.24 \times 10^{-3}][2.48 \times 10^{-3}]^2 = 7.63 \times 10^{-9}$$

- b. Solid $CaCl_2$ is then added to the solution. Will the amount of $CaF_2(s)$ increase, decrease, or remain the same? Explain

It increases. Because Ca^{2+} , one of the products of the dissociation of CaF_2 , is already present in solution, owing to $CaCl_2$, less CaF_2 will dissolve, leading to more solid CaF_2 in the mixture.

- PbBr₂**
4. The molar solubility of **PbBr₂** is **2.17 x 10⁻³ M** at a certain temperature. Calculate **K_{sp}** for **PbBr₂**.



Recall that the molar solubility refers to the solubility of the solid compound. Set up an ICE chart.

	Pb ²⁺	2Br ⁻
I	0	0
C	+x	+2x
E	x	2x

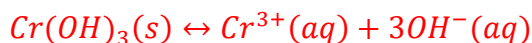
Because the stoichiometry of the equation indicates that PbBr₂ (s) and Pb²⁺ (aq) are in 1:1 ratio, the given molar solubility will be equal to x. Thus,

$$[Pb^{2+}]_{eq} = 2.17 * 10^{-3} M$$

$$[Br^{-}]_{eq} = 4.34 * 10^{-3} M$$

$$K_{sp} = [Pb^{2+}][Br^{-}]^2 = (2.17 * 10^{-3} M)(4.34 * 10^{-3} M)^2 = 4.09 * 10^{-8}$$

5. What is the pH at which **Cr(OH)₃**, **K_{sp} = 6.3 x 10⁻³¹**, just starts to precipitate from a **1.0 x 10⁻¹² M Cr³⁺** solution?



Precipitation will occur when the reaction quotient, Q, just begins to exceed K_{sp}. You are given the concentration of Cr³⁺, so you can solve for the equilibrium OH⁻ concentration.

$$K_{sp} = [Cr^{3+}][OH^{-}]^3$$

$$6.3 * 10^{-31} = (1.0 * 10^{-12} M)[OH^{-}]^3$$

$$[OH^{-}]_{eq} = 8.57 * 10^{-7} M$$

$$pOH = -\log([OH^{-}]_{eq}) = -\log(8.57 * 10^{-7} M) = 6.07$$

$$pH = 14 - pOH = 14 - 6.07 = \mathbf{7.93}$$

6. Solution A is 1.0 L of pure water. Solution B is 1.0 L of $3.4 \times 10^{-2} \text{ M NaCl}$. How many more moles of AgCl ($K_{sp} = 1.77 \times 10^{-10}$) dissolve in solution A than solution B? (Hint: for solution B, you can consider the change to be negligible compared to the initial concentration of Cl^-)

Solution A:



	Ag^+	Cl^-
I	0	0
C	+x	+x
E	x	x

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.77 \times 10^{-10} = x^2$$

$$x = 1.33 \times 10^{-5} \text{ M} = \text{molar solubility of AgCl}$$

$$\frac{1.33 \times 10^{-5} \text{ mol AgCl}}{1 \text{ L}} * 1.0 \text{ L} = 1.33 \times 10^{-5} \text{ mol AgCl dissolved}$$

Now, consider solution B. The Na^+ plays no role chemically in this case (spectator).



	Ag^+	Cl^-
I	0	0.034
C	+x	+x
E	x	0.034+x

← we get this number from that fact that sol'n B is 1 L of $3.4 \times 10^{-2} \text{ M NaCl}$ which means we have $3.4 \times 10^{-2} \text{ mol}$ of Cl^-

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.77 \times 10^{-10} = (x)(0.034 + x)$$

You can assume that "x" is negligible compared to the 0.034 so:

$$1.77 \times 10^{-10} = 0.034x$$

$$x = 5.21 \times 10^{-9} \text{ M} = \text{molar solubility of AgCl}$$

$$\frac{5.21 \times 10^{-9} \text{ M mol AgCl}}{1 \text{ L}} * 1.0 \text{ L} = 5.21 \times 10^{-9} \text{ mol AgCl dissolved}$$

LAST PART OF SOLUTION ON NEXT PAGE!!!!

Lastly, find the difference in moles.

$$\begin{aligned} &1.33 * 10^{-5} \text{ mol AgCl dissolved in A} - 5.21 * 10^{-9} \text{ mols AgCl dissolved in B} \\ &= \mathbf{1.33 * 10^{-5} \text{ more mol AgCl dissolved in A}} \end{aligned}$$