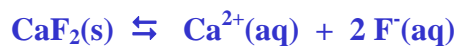


Solubility Key

1. Write the reaction and K_{sp} expression for barium sulfate.



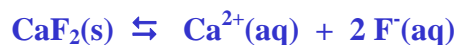
2. The K_{sp} for calcium fluoride is 3.9×10^{-11} . Calculate its solubility. Write the reaction for full credit.



$$K_{sp} = (x)(2x)^2 = 4x^3 = 3.9 \times 10^{-11} \quad \text{solve } x = 2.1363 \times 10^{-4} \text{ mol/L}$$

$$\text{Now turn molar solubility into solubility } 2.1363 \times 10^{-4} \text{ mol/L } (78.078 \text{ g/mol}) = 0.017 \text{ g/mol}$$

3. Calculate the molar solubility of calcium fluoride in a solution containing 0.010M calcium nitrate. Write the reaction for full credit.



$$\text{I -} \quad \quad \quad 0.010 \quad \quad 0$$

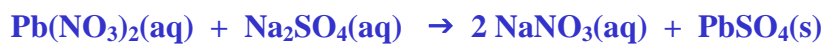
$$\text{C -} \quad \quad \quad x \quad \quad 2x$$

$$\text{E -} \quad \quad \quad (.010+x) \quad \quad (2x)$$

$$K_{sp} = 3.9 \times 10^{-11} = (.010)(2x)^2 = 0.04x^2$$

$$x = 3.1 \times 10^{-5} \text{ mol/L}$$

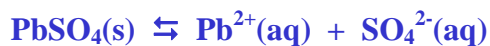
4. Will a precipitate form if 0.10 L of $8.0 \times 10^{-3}\text{M}$ lead(II) nitrate is added to 0.40 L of $5.0 \times 10^{-3}\text{M}$ sodium sulfate? Write all balanced reactions for full credit.



The precipitate ion molarities are needed to see if they exceed K_{sp} or not

$$[\text{Pb}^{2+}] = .10\text{L } (8.0 \times 10^{-3}\text{mol/L}) = .00080 \text{ moles} / .50\text{L} = .0016\text{M}$$

$$[\text{SO}_4^{2-}] = .40\text{L } (5.0 \times 10^{-3}\text{mol/L}) = .0020 \text{ moles} / .50\text{L} = .0040\text{M}$$



$$Q = (.0016)(.0040) = 6.4 \times 10^{-6} \quad \text{but} \quad K_{sp} = 1.8 \times 10^{-8} \quad \text{so } Q > K_{sp} \text{ so solid ppt forms}$$