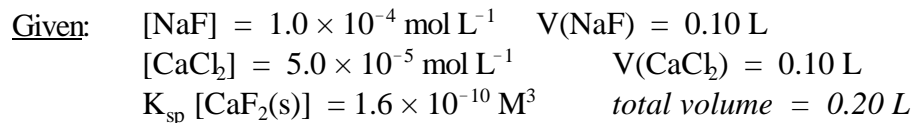
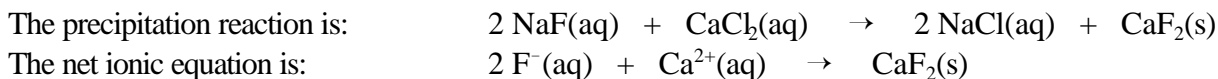


**PART C  
QUESTION 7**



To show: whether precipitation occurs



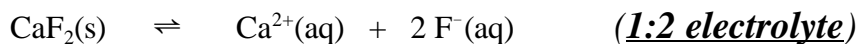
$$\begin{aligned} n(\text{NaF}) &= c(\text{NaF}) \times V(\text{NaF}) = (1.0 \times 10^{-4} \text{ mol L}^{-1}) \times (0.10 \text{ L}) = \underline{1.0} \times 10^{-5} \text{ mol} \\ n(\text{CaCl}_2) &= c(\text{CaCl}_2) \times V(\text{CaCl}_2) = (5.0 \times 10^{-5} \text{ mol L}^{-1}) \times (0.10 \text{ L}) = \underline{5.0} \times 10^{-6} \text{ mol} \end{aligned}$$

$$\text{total volume} = 0.10 \text{ L NaF}(\text{aq}) + 0.10 \text{ L CaCl}_2(\text{aq}) = \underline{0.20 \text{ L}}$$

In 0.20 L total volume,

$$[\text{Ca}^{2+}(\text{aq})] = (5.0 \times 10^{-6} \text{ mol}) \div (0.20 \text{ L}) = \underline{2.5} \times 10^{-5} \text{ mol L}^{-1}$$

$$[\text{F}^-(\text{aq})] = (1.0 \times 10^{-5} \text{ mol}) \div (0.20 \text{ L}) = \underline{5.0} \times 10^{-5} \text{ mol L}^{-1}$$



$$\begin{aligned} Q_{\text{sp}} [\text{CaF}_2(\text{s})] &= [\text{Ca}^{2+}(\text{aq})] [\text{F}^-(\text{aq})]^2 \\ &= (2.5 \times 10^{-5} \text{ mol L}^{-1}) (5.0 \times 10^{-5} \text{ mol L}^{-1})^2 \\ &= \underline{6.25} \times 10^{-14} \text{ M}^3 \end{aligned}$$

If  $Q_{\text{sp}} < K_{\text{sp}}$ , then the reaction proceeds toward products.  
*(i.e. to the **right**,  $\therefore$  **no precipitation**)*  
 If  $Q_{\text{sp}} > K_{\text{sp}}$ , then the reaction proceeds to form reactants.  
*(i.e. to the **left**,  $\therefore$  **precipitation forms**)*  
 If  $Q_{\text{sp}} = K_{\text{sp}}$ , then the system reaches equilibrium.

$$\therefore Q_{\text{sp}} (6.25 \times 10^{-14} \text{ M}^3) < K_{\text{sp}} (1.6 \times 10^{-10} \text{ M}^3)$$

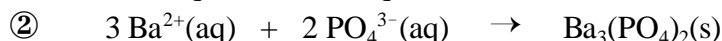
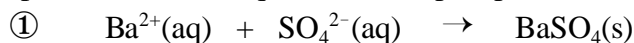
$\therefore \text{CaF}_2$  precipitate will not form.  
*(i.e. the solution is **too dilute** for a precipitate to form.)*

**PART C**  
**QUESTION 8**

Given: BaCl<sub>2</sub>(aq) solution  
 $[K_2SO_4] = 0.050 \text{ M} = [SO_4^{2-}(\text{aq})]$   
 $[Na_3PO_4] = 0.010 \text{ M} = [PO_4^{3-}(\text{aq})]$   
 $K_{sp} [BaSO_4(\text{s})] = 8.7 \times 10^{-11}$   
 $K_{sp} [Ba_3(PO_4)_2(\text{s})] = 6.0 \times 10^{-39}$

To find: Which precipitate begins to form first?

The two possible net ionic equations of the precipitation reactions are:



① The barium compound requiring the *lowest* concentration of Ba<sup>2+</sup> for precipitation will precipitate first.

∴ first calculate the concentrations of Ba<sup>2+</sup> required to precipitate barium sulfate (BaSO<sub>4</sub>) and barium phosphate (Ba<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>).

② Q<sub>sp</sub> must *equal to or greater than* K<sub>sp</sub> for precipitation to occur.



$$K_{sp} [BaSO_4(\text{s})] = [Ba^{2+}(\text{aq})] [SO_4^{2-}(\text{aq})] = 8.7 \times 10^{-11} \text{ M}^2$$

$$\begin{aligned} \therefore [Ba^{2+}(\text{aq})] \text{ for } BaSO_4(\text{s}) \text{ to precipitate} &= K_{sp} [BaSO_4(\text{s})] \div [SO_4^{2-}(\text{aq})] \\ &= (8.7 \times 10^{-11} \text{ M}^2) \div (0.050 \text{ M}) \\ &= \underline{1.74 \times 10^{-9} \text{ M}} \end{aligned}$$



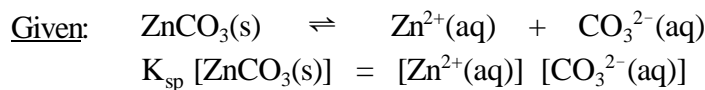
$$K_{sp} [Ba_3(PO_4)_2(\text{s})] = [Ba^{2+}(\text{aq})]^3 [PO_4^{3-}(\text{aq})]^2 = 6.0 \times 10^{-39} \text{ M}^3$$

$$\begin{aligned} \therefore [Ba^{2+}(\text{aq})] \text{ for } Ba_3(PO_4)_2(\text{s}) \text{ to precipitate} &= (K_{sp} [Ba_3(PO_4)_2(\text{s})] \div [PO_4^{3-}(\text{aq})]^2)^{\frac{1}{3}} \\ &= (6.0 \times 10^{-39} \text{ M}^3 \div (0.010 \text{ M})^2)^{\frac{1}{3}} \\ &= \underline{3.91 \times 10^{-12} \text{ M}} \end{aligned}$$

∴ Ba<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>(s) requires the *lower* [Ba<sup>2+</sup>] for precipitation,

∴ **Ba<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>(s)** will therefore precipitate first.

**PART C  
QUESTION 9**



To find: condition that shifts the  $\text{ZnCO}_3$  solubility equilibrium to the **RIGHT**

(a) increase  $[\text{Zn}^{2+}]$

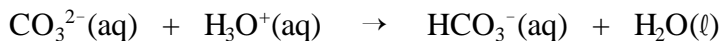
$$Q_{\text{sp}} [\text{ZnCO}_3(\text{s})] = [\text{Zn}^{2+}] [\text{CO}_3^{2-}]$$

If  $Q_{\text{sp}} < K_{\text{sp}}$ , then the reaction proceeds toward products.  
(i.e. to the **right**,  $\therefore$  **no precipitation**)  
If  $Q_{\text{sp}} > K_{\text{sp}}$ , then the reaction proceeds to form reactants.  
(i.e. to the **left**,  $\therefore$  **precipitation forms**)  
If  $Q_{\text{sp}} = K_{\text{sp}}$ , then the system reaches equilibrium.

Increase in  $[\text{Zn}^{2+}]$  will make  $Q_{\text{sp}} > K_{\text{sp}}$ , therefore, the equilibrium will shift to the **LEFT** to produce more  $\text{ZnCO}_3(\text{s})$ .

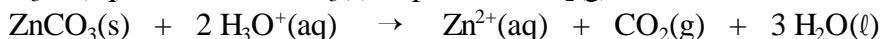
(b) **addition of HCl**

(1)  $\text{CO}_3^{2-}(\text{aq})$  is a **basic** anion, which reacts with strong acid ( $\text{H}_3\text{O}^+$ ) to form  $\text{HCO}_3^-(\text{aq})$  and water.



(i.e. With the increase in  $[\text{H}_3\text{O}^+]$  in acid solution,  $[\text{CO}_3^{2-}(\text{aq})]$  *decreases*, and the solubility equilibrium shifts to the **right**, causing  $\text{ZnCO}_3(\text{s})$  to dissolve.)

(2) Experimentally,  $\text{H}_3\text{O}^+(\text{aq})$  dissolves  $\text{ZnCO}_3(\text{s})$  to produce  $\text{CO}_2(\text{g})$  and water.



Based on the above two factors, the equilibrium will shift to the **RIGHT**.

(c) addition of NaOH

Increase in **bases** such as  $\text{OH}^-(\text{aq})$  or  $\text{CO}_3^{2-}(\text{aq})$  will make  $Q_{\text{sp}} > K_{\text{sp}}$ , and the equilibrium will shift to the **LEFT** to produce more  $\text{ZnCO}_3(\text{s})$ .

(d) addition of solid  $\text{ZnCO}_3$

The position of an equilibrium does not depend on the amounts of pure solids present. Because  $\text{ZnCO}_3(\text{s})$  does not appear in the  $K_{\text{sp}}$  expression, there is **no effect** on the position of equilibrium.