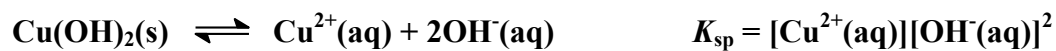


- What is the solubility of  $\text{Cu}(\text{OH})_2$  in  $\text{mol L}^{-1}$ ?  $K_{\text{sp}}(\text{Cu}(\text{OH})_2)$  is  $1.6 \times 10^{-19}$  at  $25^\circ\text{C}$ .

**Marks**  
**2**

**The dissolution reaction and associated solubility product are:**



**If  $x$  mol dissolve in one litre,  $[\text{Cu}^{2+}(\text{aq})] = x \text{ M}$  and  $[\text{OH}^{-}(\text{aq})] = 2x$ . Hence:**

$$K_{\text{sp}} = (x)(2x)^2 = 4x^3 = 1.6 \times 10^{-19}$$

$$x = 3.4 \times 10^{-7} \text{ M}$$

**Answer:  $3.4 \times 10^{-7} \text{ M}$**

**Marks**  
**6**

- All forms of life depend on iron and the concentration of iron in the oceans and elsewhere is one of the primary factors limiting the growth rates of the most basic life forms. One reason for the low availability of iron(III) is the insolubility of the hydroxide,  $\text{Fe}(\text{OH})_3$ , which has a  $K_{\text{sp}}$  of only  $2 \times 10^{-39}$ .

Calculate the maximum possible concentration of  $\text{Fe}^{3+}(\text{aq})$  in the pre-industrial era ocean which had a pH of about 8.2.

**When pH = 8.2, pOH = 14.0 – 8.2 = 5.8. As pOH =  $-\log_{10}[\text{OH}^-(\text{aq})]$ :**

$$[\text{OH}^-(\text{aq})] = 10^{-5.8} \text{ M}$$

**$\text{Fe}(\text{OH})_3(\text{s})$  dissolves according to the chemical equation:**



**The solubility product is therefore given by:**

$$K_{\text{sp}} = [\text{Fe}^{3+}(\text{aq})][\text{OH}^-(\text{aq})]^3$$

**As  $[\text{OH}^-(\text{aq})] = 10^{-5.8} \text{ M}$ :**

$$[\text{Fe}^{3+}(\text{aq})] = K_{\text{sp}} / [\text{OH}^-(\text{aq})]^3 = 2 \times 10^{-39} / (10^{-5.8})^3 \text{ M} = 5 \times 10^{-22} \text{ M}$$

$$[\text{Fe}^{3+}(\text{aq})] = 5 \times 10^{-22} \text{ M}$$

How many  $\text{Fe}^{3+}(\text{aq})$  ions are present in a litre of seawater at this pH?

**From above,  $[\text{Fe}^{3+}(\text{aq})] = 5 \times 10^{-22} \text{ M} = 5 \times 10^{-22} \text{ mol L}^{-1}$ . Hence, a litre of seawater contains  $5 \times 10^{-22} \text{ mol}$ .**

**The number of ions of  $\text{Fe}^{3+}$  is therefore:**

$$\text{number of ions} = (5 \times 10^{-22} \text{ mol}) \times (6.022 \times 10^{23} \text{ mol}^{-1}) = 300$$

Answer: **300**

The pH of the ocean is predicted to drop to 7.8 by the end of this century as the concentration of  $\text{CO}_2$  in the atmosphere increases. What percentage change in the concentration of  $\text{Fe}^{3+}(\text{aq})$  will result from this fall in pH?

**When pH = 7.8, pOH = 14.0 – 7.8 = 6.2 and  $[\text{OH}^-(\text{aq})] = 10^{-6.2} \text{ M}$ . Hence:**

$$[\text{Fe}^{3+}(\text{aq})] = K_{\text{sp}} / [\text{OH}^-(\text{aq})]^3 = 2 \times 10^{-39} / (10^{-6.2})^3 \text{ M} = 8 \times 10^{-21} \text{ M}$$

**The percentage increase is therefore:**

$$\text{percentage change} = \frac{(8 \times 10^{-21} - 5 \times 10^{-22})}{5 \times 10^{-22}} \times 100 \% = 1500 \%$$

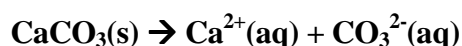
Answer: **1500 %**

**Marks**  
**4**

- The ocean contains a variety of forms of  $\text{CO}_3^{2-}$  and  $\text{CO}_2$  with a variety of acid-base and solubility equilibria determining their concentrations. There is concern that increasing levels of  $\text{CO}_2$  will lead to increased dissolution of  $\text{CaCO}_3$  and critically affect the survival of life forms that rely on a carbonaceous skeleton.

Calculate the concentrations of  $\text{Ca}^{2+}$  and  $\text{CO}_3^{2-}$  in a saturated solution of  $\text{CaCO}_3$ . (The  $K_{\text{sp}}$  of  $\text{CaCO}_3$  is  $3.3 \times 10^{-9}$ .)

The dissolution of  $\text{CaCO}_3$  follows the reaction,



If the molar solubility of  $\text{CaCO}_3$  is  $S$  then  $[\text{Ca}^{2+}(\text{aq})] = [\text{CO}_3^{2-}(\text{aq})] = S$ .

The solubility product is given by:

$$K_{\text{sp}} = [\text{Ca}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})] = (S)(S) = S^2$$

As  $K_{\text{sp}} = 3.3 \times 10^{-9}$ ,

$$S^2 = 3.3 \times 10^{-9} \text{ or } S = 5.7 \times 10^{-5} \text{ M}$$

$$[\text{Ca}^{2+}] = 5.7 \times 10^{-5} \text{ M}$$

$$[\text{CO}_3^{2-}] = 5.7 \times 10^{-5} \text{ M}$$

Calculate the pH of such a solution. (The  $\text{p}K_{\text{a}}$  of  $\text{HCO}_3^-$  is 10.33).

$\text{CO}_3^{2-}$  is a weak base and will react with water to produce  $\text{HCO}_3^-$ :

	$\text{CO}_3^{2-}$	$\text{H}_2\text{O}$	$\rightleftharpoons$	$\text{OH}^-$	$\text{HCO}_3^-$
<b>initial</b>	$5.7 \times 10^{-5}$	large		0	0
<b>change</b>	$-x$	negligible		$+x$	$+x$
<b>final</b>	$5.7 \times 10^{-5} - x$	large		$x$	$x$

The equilibrium constant  $K_{\text{b}}$  is given by:

$$K_{\text{b}} = \frac{[\text{OH}^-(\text{aq})][\text{HCO}_3^-(\text{aq})]}{[\text{CO}_3^{2-}(\text{aq})]} = \frac{x^2}{5.7 \times 10^{-5} - x}$$

As  $\text{p}K_{\text{a}} + \text{p}K_{\text{b}} = 14.00$ ,  $\text{p}K_{\text{b}} = 14.00 - 10.33 = 3.67$ . As  $\text{p}K_{\text{b}} = -\log K_{\text{b}}$ , so

$$K_{\text{b}} = 10^{-3.67}.$$

**ANSWER CONTINUES ON THE NEXT PAGE**

As the concentration of the base is so small, the 'small  $x$ ' approximation cannot be used and it is necessary to solve the quadratic equation. From above,

$$x^2 + 10^{-3.67}x - (5.7 \times 10^{-5} \times 10^{-3.67}) = 0$$

Solving this using the quadratic formula gives  $x = 4.67 \times 10^{-5} = [\text{OH}^-(\text{aq})]$ . Hence,

$$\text{pOH} = -\log_{10}[\text{OH}^-(\text{aq})] = 4.33$$

Finally, since  $\text{pH} + \text{pOH} = 14.00$ ,

$$\text{pH} = 14.00 - 4.33 = 9.67$$

$$\text{pH} = 9.67$$

**THIS QUESTION CONTINUES ON THE NEXT PAGE**

The pH of surface ocean water is currently 8.10 (having fallen from a pre-industrial era level of 8.16), the concentration of  $\text{HCO}_3^-$  is  $2.5 \times 10^{-3}$  M, and it is saturated with  $\text{CaCO}_3$ . Calculate the concentration of  $\text{Ca}^{2+}$  in these conditions.

**Marks**  
**4**

**From 2008-N-2,**

$$K_b = \frac{[\text{OH}^-(\text{aq})][\text{HCO}_3^-(\text{aq})]}{[\text{CO}_3^{2-}(\text{aq})]} = 10^{-3.67}$$

If pH = 8.10 then pOH = 14.00 – 8.10 = 5.90 and so  $[\text{OH}^-(\text{aq})] = 10^{-5.90}$  M.

If  $[\text{HCO}_3^-] = 2.5 \times 10^{-3}$  M, then

$$[\text{CO}_3^{2-}] = \frac{[\text{OH}^-(\text{aq})][\text{HCO}_3^-(\text{aq})]}{K_b} = \frac{(10^{-5.90})(2.5 \times 10^{-3})}{(10^{-3.67})} = 1.47 \times 10^{-5}$$

**From 2008-N-2,  $K_{sp} = [\text{Ca}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})] = 3.3 \times 10^{-9}$ . Hence,**

$$[\text{Ca}^{2+}(\text{aq})] = K_{sp} / [\text{CO}_3^{2-}(\text{aq})] = 3.3 \times 10^{-9} / (1.47 \times 10^{-5}) = 2.2 \times 10^{-4} \text{ M}$$

$$[\text{Ca}^{2+}] = 2.2 \times 10^{-4} \text{ M}$$

The pH is expected to drop to about 7.8 by the end of the century as  $\text{CO}_2$  levels increase further. What effect will this have on the solubility of  $\text{CaCO}_3$  in sea water? Use chemical equations to assist with explaining your answer.

**The solubility of  $\text{CaCO}_3$  will increase.**

**At a lower pH,  $[\text{OH}^-(\text{aq})]$  will be lower. The equilibrium below will be shifted to the right:**



**With lower  $[\text{CO}_3^{2-}(\text{aq})]$ ,  $[\text{Ca}^{2+}(\text{aq})]$  will increase as  $[\text{Ca}^{2+}(\text{aq})] = K_{sp} / [\text{CO}_3^{2-}(\text{aq})]$ . The solubility will be increased.**

**Marks**  
**2**

- In order to reduce the incidence of dental cavities, water is fluoridated to a level of  $1 \text{ mg L}^{-1}$ . In regions where the water is “hard” the calcium concentration is typically  $100 \text{ mg L}^{-1}$ . Given that the  $K_{\text{sp}}$  of calcium fluoride is  $3.9 \times 10^{-11} \text{ M}^3$ , would it precipitate in these conditions? Show all working.

**For the dissolution of  $\text{CaF}_2(\text{s}) \rightleftharpoons \text{Ca}^{2+}(\text{aq}) + 2\text{F}^{-}(\text{aq})$ ,  $Q_{\text{sp}} = [\text{Ca}^{2+}(\text{aq})][\text{F}^{-}(\text{aq})]^2$ .  
If  $Q_{\text{sp}} > K_{\text{sp}}$ , then  $\text{CaF}_2(\text{s})$  will precipitate. If  $Q_{\text{sp}} < K_{\text{sp}}$ , then  $\text{CaF}_2(\text{s})$  will completely dissolve.**

**The atomic mass of calcium is  $40.08 \text{ g mol}^{-1}$ . A  $100 \text{ mg L}^{-1}$  solution contains  $\frac{100 \times 10^{-3}}{40.08} \text{ mol}$  and has a concentration of  $\frac{100 \times 10^{-3}}{40.08} \text{ M}$ .  $[\text{Ca}^{2+}(\text{aq})] = 2.495 \times 10^{-3} \text{ M}$ .**

**The atomic mass of fluorine is  $19.00 \text{ g mol}^{-1}$ . A  $1 \text{ mg L}^{-1}$  solution contains  $\frac{1 \times 10^{-3}}{19.00} \text{ mol}$  and has a concentration of  $\frac{1 \times 10^{-3}}{19.00} \text{ M}$ .  $[\text{F}^{-}(\text{aq})] = 5.263 \times 10^{-5} \text{ M}$ .**

**Hence,  $Q_{\text{sp}} = (2.495 \times 10^{-3} \text{ M}) \times (5.263 \times 10^{-5} \text{ M})^2 = 6.911 \times 10^{-12} \text{ M}^3$ . As  $Q_{\text{sp}} < K_{\text{sp}}$ , calcium fluoride will not precipitate.**

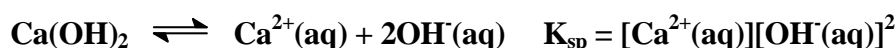
**Answer: No, it does not precipitate**

- 2.00 g of solid calcium hydroxide is added to 1.00 L of water. What proportion of the calcium hydroxide remains undissolved when the system has reached equilibrium?  
 $K_{sp}(\text{Ca}(\text{OH})_2) = 6.5 \times 10^{-6} \text{ M}^3$

The formula mass of  $\text{Ca}(\text{OH})_2$  is  $(40.08 \text{ Ca}) + 2 \times (16.00 \text{ O}) + 1.008 \text{ (H)}) = 74.096$ .  
 2.00 g of  $\text{Ca}(\text{OH})_2$  therefore corresponds to:

$$\text{amount of } \text{Ca}(\text{OH})_2 = \frac{\text{mass}}{\text{formula mass}} = \frac{2.00}{74.096} = 0.0270 \text{ mol}$$

The solubility equilibrium and constant are given by:



If  $S$  mol dissolves in 1.00 L then  $[\text{Ca}^{2+}(\text{aq})] = S$  and  $[\text{OH}^{-}(\text{aq})] = 2S$ . Thus,

$$K_{sp} = (S)(2S)^2 = 4S^3 = 6.5 \times 10^{-6} \quad \text{so } S = 0.0118 \text{ mol}$$

The amount that remains undissolved is  $(0.0270) - (0.0118) = 0.0152 \text{ mol}$ . The

proportion that is undissolved is  $\frac{0.0152}{0.0270} \times 100\% = 56\%$ .

Answer: 56%

What volume (in mL) of 10.0 M nitric acid must be added to this mixture in order to just dissolve all of the calcium hydroxide? Assume the volume of the nitric acid is small and can be ignored in the calculation of the total volume.

If all of the  $\text{Ca}(\text{OH})_2$  dissolves then  $[\text{Ca}^{2+}(\text{aq})] = 0.0270 \text{ M}$ . The  $[\text{OH}^{-}(\text{aq})]$  required to achieve this is given by:

$$K_{sp} = [\text{Ca}^{2+}(\text{aq})][\text{OH}^{-}(\text{aq})]^2 = (0.0270) \times [\text{OH}^{-}(\text{aq})]^2 = 6.5 \times 10^{-6}$$

$$[\text{OH}^{-}(\text{aq})] = 0.0155 \text{ M}$$

As dissolution of 0.0270 mol of  $\text{Ca}(\text{OH})_2$  produces  $(2 \times 0.0270) = 0.0540 \text{ mol}$  of  $\text{OH}^{-}$ , the remainder has been neutralized by the added nitric acid:

$$\text{number of moles of nitric acid added} = 0.0540 - 0.0155 = 0.0384 \text{ mol}$$

The volume of 10.0 M nitric acid which contains this amount is given by:

$$\text{volume of nitric acid} = \frac{\text{number of moles}}{\text{concentration}} = \frac{0.0384}{10.0} = 3.84 \times 10^{-3} \text{ L} = 3.84 \text{ mL}$$

Answer: 3.84 mL

**Marks**  
**2**

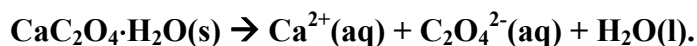
- Calcium oxalate is a major constituent of kidney stones. Calculate the solubility product constant for calcium oxalate given that a saturated solution of the salt can be made by dissolving 0.0061 g of  $\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}(\text{s})$  in 1.0 L of water.

**The molar mass of  $\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}$  is:**

$$(40.08(\text{Ca}) + 2 \times 12.01(\text{C}) + 5 \times 16.00(\text{O}) + 2 \times 1.008(\text{H})) \text{ g mol}^{-1} = 146.116 \text{ g mol}^{-1}$$

Hence, 0.0061 g corresponds to  $\frac{0.0061 \text{ g}}{146.116 \text{ g mol}^{-1}} = 4.2 \times 10^{-5} \text{ mol}$ . As this amount dissolves in 1.0 L, the molar solubility =  $S = 4.2 \times 10^{-5} \text{ M}$ .

**The dissolution equilibrium is:**



**As one mol of cation and one mol of anion is produced, the solubility product is:**

$$K_{\text{sp}} = [\text{Ca}^{2+}(\text{aq})][\text{C}_2\text{O}_4^{2-}(\text{aq})] = (S)(S) = S^2 = (4.2 \times 10^{-5})^2 = 1.7 \times 10^{-9}$$

Answer:  $1.7 \times 10^{-9}$

**3**

- A sample of 2.0 mg of  $\text{Cu}(\text{OH})_2$  is added to 1.0 L of a solution buffered at a pH of 8.00. Will all of the  $\text{Cu}(\text{OH})_2$  dissolve? Show all working. (The  $K_{\text{sp}}$  of  $\text{Cu}(\text{OH})_2$  is  $4.8 \times 10^{-20} \text{ M}^3$ .)

As  $\text{pH} + \text{pOH} = 14.00$ ,  $\text{pOH} = 14.00 - 8.00 = 6.00$ . Hence,  $[\text{OH}^-(\text{aq})] = 10^{-6} \text{ M}$ .

**The dissolution equilibrium is:**  $\text{Cu}(\text{OH})_2(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$

Hence, if  $S$  is the molar solubility,  $K_{\text{sp}} = [\text{Cu}^{2+}(\text{aq})][\text{OH}^-(\text{aq})]^2 = S \times [\text{OH}^-(\text{aq})]^2$ .

$$\text{As } K_{\text{sp}} = 4.8 \times 10^{-20}, S = \frac{4.8 \times 10^{-20}}{(10^{-6})^2} = 4.8 \times 10^{-8} \text{ M}$$

**The molar mass of  $\text{Cu}(\text{OH})_2$  is  $(63.55 (\text{Cu}) + 2 \times (16.00 (\text{O}) + 1.008 (\text{O})) \text{ g mol}^{-1} = 97.566 \text{ g mol}^{-1}$ .**

**The solubility in  $\text{g L}^{-1}$  is therefore  $(4.8 \times 10^{-8}) \times 97.566 = 4.7 \times 10^{-6}$ .**

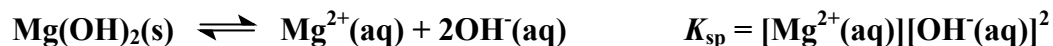
**Hence, only  $4.7 \times 10^{-3} \text{ mg}$  will dissolve.**

Answer: **NO**



In the presence of excess hydroxide ion,  $\text{Mg}^{2+}$  can be precipitated as  $\text{Mg}(\text{OH})_2(\text{s})$ . What amount (in mol) of solid sodium hydroxide must be added to a 0.10 M solution of  $\text{Mg}(\text{NO}_3)_2$  to just cause precipitation of  $\text{Mg}(\text{OH})_2(\text{s})$ . The solubility product constant of  $\text{Mg}(\text{OH})_2$  is  $7.1 \times 10^{-12} \text{ M}^3$ .

The solubility equilibrium and product for  $\text{Mg}(\text{OH})_2(\text{s})$  are:



With  $[\text{Mg}^{2+}(\text{aq})] = 0.10 \text{ M}$ , precipitation will occur when:

$$[\text{OH}^{-}(\text{aq})]^2 \geq \frac{K_{\text{sp}}}{[\text{Mg}^{2+}(\text{aq})]} = \frac{(7.1 \times 10^{-12})}{(0.10)} \text{ so } [\text{OH}^{-}(\text{aq})] \geq 8.4 \times 10^{-6} \text{ M}$$

(As the volume of the solution is not specified, the number of moles of  $\text{NaOH}(\text{s})$  cannot be given. A 1 L solution would require  $8.4 \times 10^{-6} \text{ mol}$ .)

ANSWER:  $8.4 \times 10^{-6} \text{ M}$

In a separate experiment, the  $\text{Mg}(\text{OH})_2$  is precipitated by adding 0.10 mol of  $\text{Mg}(\text{NO}_3)_2$  to 1.0 L of a 0.10 M  $\text{NH}_3$  solution. What amount (in mol) of  $\text{NH}_4\text{Cl}$  must be added to this solution to just dissolve the precipitate? The  $\text{p}K_{\text{a}}$  of  $\text{NH}_4\text{Cl}$  is 9.24.

With  $[\text{Mg}^{2+}(\text{aq})] = 0.10 \text{ M}$ , dissolution will start to occur when:

$$[\text{OH}^{-}(\text{aq})]^2 \leq \frac{K_{\text{sp}}}{[\text{Mg}^{2+}(\text{aq})]} = \frac{(7.1 \times 10^{-12})}{(0.10)} \text{ so } [\text{OH}^{-}(\text{aq})] \leq 8.4 \times 10^{-6} \text{ M}$$

This  $[\text{OH}^{-}(\text{aq})]$  corresponds to  $\text{pOH} = -\log_{10}([\text{OH}^{-}(\text{aq})]) = -\log_{10}(8.4 \times 10^{-6}) = 5.1$ .  
Using  $\text{pH} = 14.0 - \text{pOH}$ ,  $\text{pH} = (14.0 - 5.1) = 8.9$ .

When  $\text{NH}_4\text{Cl}$  is added, the solution contains an acid ( $\text{NH}_4\text{Cl}$ ) and its conjugate base ( $\text{NH}_3$ ). The solution contains an acid ( $\text{NH}_4\text{Cl}$ ) and its conjugate base ( $\text{NH}_3$ ). The Henderson-Hasselbalch equation can be used to work out the required  $[\text{NH}_4\text{Cl}]$  with  $[\text{NH}_3] = 0.10 \text{ M}$  and  $\text{pH} = 8.9$ :

$$\text{pH} = \text{p}K_{\text{a}} + \log_{10} \left( \frac{[\text{base}]}{[\text{acid}]} \right) = 9.24 + \log_{10} \left( \frac{0.10}{[\text{NH}_4\text{Cl}]} \right) = 8.9$$

$$\left( \frac{0.10}{[\text{NH}_4\text{Cl}]} \right) = 10^{(8.9 - 9.24)} \text{ so } [\text{NH}_4\text{Cl}] = 0.2 \text{ M}$$

This molarity is for a 1.0 L solution so that 0.2 mol of  $\text{NH}_4\text{Cl}$  are required.

ANSWER: 0.2 mol

- The solubility product constant of  $\text{Fe}(\text{OH})_3$  is  $1 \times 10^{-39} \text{ M}^4$ . What is the concentration of  $\text{Fe}^{3+}(\text{aq})$  in equilibrium with  $\text{Fe}(\text{OH})_3$  at pH 7.0?

**Marks**  
**4**

As  $\text{pH} + \text{pOH} = 14.0$  and  $\text{pOH} = -\log_{10}([\text{OH}^-(\text{aq})])$ ,  $[\text{OH}^-(\text{aq})] = 10^{-7.0}$ .

The dissolution reaction and solubility product are:



With  $K_{\text{sp}} = 1 \times 10^{-39}$  and  $[\text{OH}^-(\text{aq})] = 10^{-7.0}$ ,

$$[\text{Fe}^{3+}(\text{aq})] = \frac{K_{\text{sp}}}{[\text{OH}^-(\text{aq})]^3} = \frac{(1 \times 10^{-39})}{(10^{-7.0})^3} = 1 \times 10^{-18} \text{ M}$$

ANSWER:  $1 \times 10^{-18} \text{ M}$

To what value does the pH need to be increased to decrease the concentration of  $\text{Fe}^{3+}(\text{aq})$  to a single  $\text{Fe}^{3+}(\text{aq})$  ion per litre of solution?

A single  $\text{Fe}^{3+}(\text{aq})$  ion corresponds to:

$$\text{number of moles} = \frac{\text{number of ions}}{\text{Avogadro's number}} = \frac{1}{6.022 \times 10^{23}} = 1.66 \times 10^{-24} \text{ mol}$$

If this is in a litre of solution,  $[\text{Fe}^{3+}(\text{aq})] = 1.66 \times 10^{-24} \text{ M}$ .

As  $K_{\text{sp}} = [\text{Fe}^{3+}(\text{aq})][\text{OH}^-(\text{aq})]^3$ ,

$$[\text{OH}^-(\text{aq})]^3 = \frac{K_{\text{sp}}}{[\text{Fe}^{3+}(\text{aq})]} = \frac{(1 \times 10^{-39})}{(1.66 \times 10^{-24})} = 6 \times 10^{-16} \text{ M}^3$$

$$[\text{OH}^-(\text{aq})] = 8 \times 10^{-6} \text{ M}$$

Thus,  $\text{pOH} = -\log_{10}([\text{OH}^-(\text{aq})]) = -\log_{10}(8 \times 10^{-6}) = 5$  and as  $\text{pH} = 14 - \text{pOH}$ ,

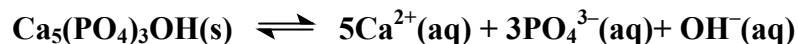
$$\text{pH} = 14 - 5 = 9$$

ANSWER: 9

**Marks**  
**2**

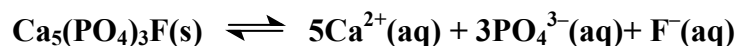
- Teeth are made from hydroxyapatite,  $\text{Ca}_5(\text{PO}_4)_3\text{OH}$ . Why does an acidic medium promote tooth decay and how can the decay be stopped using fluoridation of drinking water? Use chemical equations where appropriate.

**Hydroxyapatite dissolves according to the equation:**



**This equilibrium lies far to the left so that hydroxyapatite does not dissolve to any significant degree. However, the addition of  $\text{H}_3\text{O}^{+}$  (i.e. an acidic medium) will decrease  $[\text{OH}^{-}(\text{aq})]$  and push the reaction to the right (Le Chatelier's principle).**

**Fluoridation of water promotes the replacement of  $\text{OH}^{-}$  with  $\text{F}^{-}$  to form  $\text{Ca}_5(\text{PO}_4)_3\text{F}(\text{s})$ . This compound is much less water soluble than  $\text{Ca}_5(\text{PO}_4)_3\text{OH}$ :**



**$\text{F}^{-}$  is a weaker base than  $\text{OH}^{-}$  so this equilibrium is less affected by the addition of  $\text{H}_3\text{O}^{+}$ . As the enamel is less soluble, this promotes less tooth decay.**