

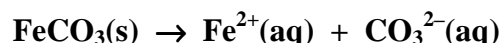
- Explain briefly why the  $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$  cation has a  $K_a$  of  $6 \times 10^{-3}$ , whilst the  $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$  cation has a  $K_a$  of  $4 \times 10^{-9}$ .

**The  $\text{Fe}^{3+}$  ion is much smaller and has a much higher charge density than the  $\text{Fe}^{2+}$  ion. This results in a stronger Fe–O bond and a weakening of the O–H bonds. The high charge density on the  $\text{Fe}^{3+}$  ion pulls electron density from the attached  $\text{OH}_2$  ligands. This leads to the release of  $\text{H}^+(\text{aq})$ .**

**As the O–H bonds are weaker, it is more acidic: it has a much *greater* value for  $K_a$ .**

**Marks**  
**4**

- Write a balanced chemical equation representing the dissolution of  $\text{FeCO}_3$  in water at pH 7.



Ignoring any hydrolysis of the ions, calculate the solubility (in  $\text{g L}^{-1}$ ) of  $\text{FeCO}_3$  in water at pH 7. The solubility product constant,  $K_{\text{sp}}$ , for  $\text{FeCO}_3$  is  $2.1 \times 10^{-11}$ .

From the equation above,  $K_{\text{sp}} = [\text{Fe}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})]$ . If  $s$  mol of  $\text{FeCO}_3$  dissolves in 1.0 L,  $[\text{Fe}^{2+}(\text{aq})] = [\text{CO}_3^{2-}(\text{aq})] = s$  M. Hence:

$$K_{\text{sp}} = [\text{Fe}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})] = (s)(s) = s^2 = 2.1 \times 10^{-11} \text{ or } s = 4.6 \times 10^{-6} \text{ M}$$

The formula mass of  $\text{FeCO}_3$  is  $(55.85 (\text{Fe}) + 12.01 (\text{C}) + 3 \times 16.00 (\text{O})) \text{ g mol}^{-1} = 115.86 \text{ g mol}^{-1}$ . From above,  $4.6 \times 10^{-6}$  mol of  $\text{FeCO}_3$  dissolves in 1.0 L. This corresponds to:

$$\begin{aligned} \text{mass} &= \text{number of moles} \times \text{formula mass} \\ &= (4.6 \times 10^{-6} \text{ mol}) \times (115.86 \text{ g mol}^{-1}) = 5.3 \times 10^{-4} \text{ g} \end{aligned}$$

This is the mass that dissolves in 1.0 L. The solubility is  $5.3 \times 10^{-4} \text{ g L}^{-1}$ .

Answer:  $5.3 \times 10^{-4} \text{ g L}^{-1}$

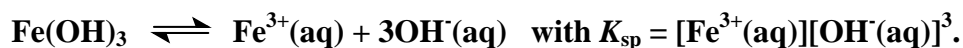
**4**

- The concentration of iron in the ocean is one of the primary factors limiting the growth rates of some basic life forms. The pH of the oceans before the Industrial Revolution was around 8.22. What was the maximum concentration of  $\text{Fe}^{3+}(\text{aq})$  in the ocean at this pH? The  $K_{\text{sp}}$  of  $\text{Fe}(\text{OH})_3$  is  $1 \times 10^{-39}$ .

As  $\text{pH} + \text{pOH} = 14.00$ ,  $\text{pOH} = 14.00 - 8.22 = 5.78$ .

By definition,  $\text{pOH} = -\log_{10}[\text{OH}^-(\text{aq})]$  and so  $[\text{OH}^-(\text{aq})] = 10^{-5.78}$ .

$\text{Fe}(\text{OH})_3$  dissolves according to the equilibrium:



As  $K_{\text{sp}} = 1 \times 10^{-39}$  and  $[\text{OH}^-(\text{aq})] = 10^{-5.78}$ :

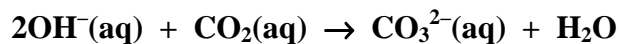
$$[\text{Fe}^{3+}(\text{aq})] = K_{\text{sp}} / [\text{OH}^-(\text{aq})]^3 = (1 \times 10^{-39}) / (10^{-5.78})^3 = 2 \times 10^{-22} \text{ M}$$

Answer:  $2 \times 10^{-22} \text{ M}$

ANSWER CONTINUES ON THE NEXT PAGE

Industrialisation has led to an increase in atmospheric CO<sub>2</sub>. What effect has this had on the amount of Fe<sup>3+</sup>(aq) in sea water?

**CO<sub>2</sub> dissolves in water to give acidic solution that reacts with OH<sup>-</sup> ions.**



**From Le Chatelier's principle, the decrease in [OH<sup>-</sup>(aq)] will result in an increase in [Fe<sup>3+</sup>(aq)].**

**Equivalently, if [OH<sup>-</sup>(aq)] is decreased, [Fe<sup>3+</sup>(aq)] must increase as  $K_{\text{sp}}$  is a constant.**