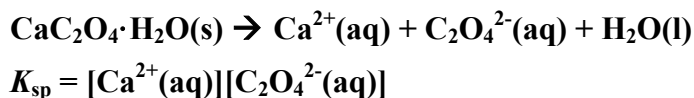


**Marks**  
**9**

- The salt calcium oxalate,  $\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}$ , is sparingly soluble. Write down the chemical equation for its dissolution in water and the expression for  $K_{\text{sp}}$ .



What is the molar solubility of calcium oxalate?  $K_{\text{sp}} = 2.3 \times 10^{-9}$

**If  $x$  mol of the salt dissolves in one litre, then the molar solubility is  $x$  M. If  $x$  mol dissolves in one litre then  $[\text{Ca}^{2+}(\text{aq})] = x$  M and  $[\text{C}_2\text{O}_4^{2-}(\text{aq})] = x$  M.**

$$K_{\text{sp}} = [\text{Ca}^{2+}(\text{aq})][\text{C}_2\text{O}_4^{2-}(\text{aq})] = (x)(x) = x^2 = 2.3 \times 10^{-9}$$

$$x = 4.8 \times 10^{-5} \text{ mol L}^{-1}$$

Answer:  $4.8 \times 10^{-5} \text{ mol L}^{-1}$

If additional calcium oxalate is added to a saturated solution, what is the effect on  $[\text{Ca}^{2+}(\text{aq})]$ ?

**A saturated solid has the maximum possible dissolution. Adding additional solid has no effect on the equilibrium and so no effect on  $[\text{Ca}^{2+}(\text{aq})]$ .**

Following blood donation, a solution of sodium oxalate is added to remove  $\text{Ca}^{2+}(\text{aq})$  ions which cause the blood to clot. The concentration of  $\text{Ca}^{2+}(\text{aq})$  ions in blood is  $9.7 \times 10^{-5} \text{ g mL}^{-1}$ . If 100.0 mL of 0.1550 M  $\text{Na}_2\text{C}_2\text{O}_4$  is added to 100.0 mL of blood, what will be the concentration (in  $\text{mol L}^{-1}$ ) of  $\text{Ca}^{2+}$  ions remaining in the blood?

**The amount of  $\text{Ca}^{2+}$  present in 100.0 mL is  $9.7 \times 10^{-3} \text{ g}$ . As its molar mass is  $40.08 \text{ g mol}^{-1}$ , this corresponds to:**

$$\begin{aligned} \text{number of moles} &= \text{mass} / \text{molar mass} = \\ &= (9.7 \times 10^{-3} \text{ g}) / (40.08 \text{ g mol}^{-1}) = 2.4 \times 10^{-4} \text{ mol} \end{aligned}$$

**The number of moles of  $\text{C}_2\text{O}_4^{2-}(\text{aq})$  added is:**

$$\begin{aligned} \text{number of moles} &= \text{concentration} \times \text{volume} \\ &= (0.1550 \text{ mol L}^{-1}) \times (0.1000 \text{ L}) = 0.01550 \text{ mol} \end{aligned}$$

**The amount of  $\text{C}_2\text{O}_4^{2-}$  is *much* larger than the amount of  $\text{Ca}^{2+}$  present so precipitation of  $\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}(\text{s})$  does not reduce the  $\text{C}_2\text{O}_4^{2-}$  significantly.**

**When the oxalate is added to the blood, the total volume increases to (100.0 + 100.0) mL = 200.0 mL. The concentration of  $\text{C}_2\text{O}_4^{2-}(\text{aq})$  is now:**

**ANSWER CONTINUES ON THE NEXT PAGE**

$$\begin{aligned}\text{concentration} &= \text{number of moles} / \text{volume} \\ &= (0.01550 \text{ mol}) / (0.2000 \text{ L}) = 0.0775 \text{ mol L}^{-1}\end{aligned}$$

Using  $K_{\text{sp}} = [\text{Ca}^{2+}(\text{aq})][\text{C}_2\text{O}_4^{2-}(\text{aq})]$ :

$$[\text{Ca}^{2+}(\text{aq})] = K_{\text{sp}} / [\text{C}_2\text{O}_4^{2-}(\text{aq})] = (2.3 \times 10^{-9} / 0.0775) \text{ M} = 3.0 \times 10^{-8} \text{ M}$$

Answer:  $3.0 \times 10^{-8} \text{ M}$