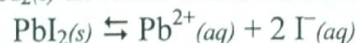


16 • Solubility Equilibria**PRACTICE FRQ**

Answer the following questions that relate to solubility of salts of lead and barium.

- (a) A saturated solution is prepared by adding excess $\text{PbI}_2(s)$ to distilled water to form 1.0 L of solution at 25°C . The concentration of $\text{Pb}^{2+}(aq)$ in the saturated solution is found to be $1.3 \times 10^{-3} \text{ M}$. The chemical equation for the dissolution of $\text{PbI}_2(s)$ in water is shown below.



- (i) Write the equilibrium-constant expression for the equation.

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$$

- (ii) Calculate the molar concentration of $\text{I}^{-}(aq)$ in the solution.

$$[\text{Pb}^{2+}] = 1.3 \times 10^{-3} \text{ M}$$

$$[\text{I}^{-}] = 2.6 \times 10^{-3} \text{ M}$$

- (iii) Calculate the value of the equilibrium constant, K_{sp} .

$$\begin{aligned} K_{sp} &= (1.3 \times 10^{-3})(2.6 \times 10^{-3})^2 \\ &= 8.788 \times 10^{-9} \\ &= \boxed{8.8 \times 10^{-9}} \end{aligned}$$

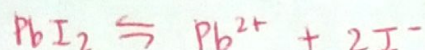
- (b) A saturated solution is prepared by adding $\text{PbI}_2(s)$ to distilled water to form 2.0 L of solution at 25°C . What are the molar concentrations of $\text{Pb}^{2+}(aq)$ and $\text{I}^{-}(aq)$ in the solution? Justify your answer.

$$[\text{Pb}^{2+}] = 1.3 \times 10^{-3} \text{ M}$$

$$[\text{I}^{-}] = 2.6 \times 10^{-3} \text{ M}$$

The molar concentrations do not change because the temp. is still at 25°C . Solubility is temperature dependent.

- (c) Solid NaI is added to a saturated solution of PbI_2 at 25°C . Assuming that the volume of the solution does not change, does the molar concentration of $\text{Pb}^{2+}(aq)$ in the solution increase, decrease, or remain the same? Justify your answer.



the addition of $\text{NaI}(s)$ increases the $[\text{I}^{-}]$ in solution and shifts the equilibrium towards the reactants.

Therefore, the $[\text{Pb}^{2+}]$ decreases.

(d) The value of K_{sp} for the salt BaCrO_4 is 1.2×10^{-10} . When a 500. mL sample of $8.2 \times 10^{-6} \text{ M}$ $\text{Ba}(\text{NO}_3)_2$ is added to 500. mL of $8.2 \times 10^{-6} \text{ M}$ Na_2CrO_4 , no precipitate is observed.

(i) Assuming that volumes are additive, calculate the molar concentrations of $\text{Ba}^{2+}_{(aq)}$ and $\text{CrO}_4^{2-}_{(aq)}$ in the 1.00 L of solution.

$$[\text{Ba}^{2+}] = \frac{(0.500 \text{ L})(8.2 \times 10^{-6} \text{ M})}{1.00 \text{ L}} = 4.1 \times 10^{-6} \text{ M}$$

$$[\text{CrO}_4^{2-}] = \frac{(0.500 \text{ L})(8.2 \times 10^{-6} \text{ M})}{1.00 \text{ L}} = 4.1 \times 10^{-6} \text{ M}$$

(ii) Use the molar concentrations of $\text{Ba}^{2+}_{(aq)}$ ions and $\text{CrO}_4^{2-}_{(aq)}$ ions as determined above to show why a precipitate does not form. You must include a calculation as part of your answer.

$$\begin{aligned} Q_{sp} &= [\text{Ba}^{2+}][\text{CrO}_4^{2-}] \\ &= (4.1 \times 10^{-6})(4.1 \times 10^{-6}) \\ &= 1.681 \times 10^{-11} \end{aligned}$$

$$K_{sp} = 1.2 \times 10^{-10}$$

$$Q_{sp} < K_{sp}$$

Therefore, more BaCrO_4 can dissolve in the solution and a precipitate does not form.