Marks 4

- Barium sulfate is used as a contrast agent for X-ray images of intestines. What is the solubility product constant, K_{sp}, for $BaSO_4$, given that a maximum of 1.167×10^{-8} g will dissolve in 500 mL of water?

The formula mass of $BaSO_4$ is 137.34 (Ba) + 32.07 (S) + 4 × 16.00 = 233.41 g mol^{-1}. 1.167×10^{-8} g therefore corresponds to:

$$\textbf{number of moles} = \frac{\textbf{mass}}{\textbf{formula mass}} = \frac{1.167 \times 10^{-8}}{233.41} = 5.000 \times 10^{-11}$$

Barium sulfate dissolves according to the equilibrium:

$$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq) \qquad K_{sp} = [Ba^{2+}(aq)][SO_4^{2-}(aq)]$$

As 5.000×10^{-11} mol dissolves in 500 mL,

$$[Ba^{2+}(aq)] = [SO_4^{2-}(aq)] = \frac{\textbf{number of moles}}{\textbf{volume}} = \frac{5.000 \times 10^{-11}}{0.500} = 1.00 \times 10^{-10}\ \text{M}$$

Hence,

$$K_{sp} = (1.00 \times 10^{-10}) \times (1.00 \times 10^{-10}) = 1.00 \times 10^{-20}$$

Answer: **1.00×10^{-20}**

What advantage would there be in administering $BaSO_4$ as a slurry that also contains 0.5 M Na_2SO_4?

As indicated by the very small solubility above, the equilibrium

$$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$$

lies far to the left. Adding more $SO_4^{2-}(aq)$ ions pushes it further to the left.

This acts to remove $Ba^{2+}(aq)$ ions from solution. This is advantageous as barium ions are highly toxic.

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