

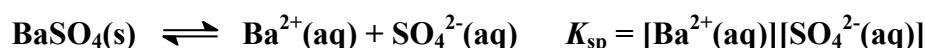
- Barium sulfate is used as a contrast agent for X-ray images of intestines. What is the solubility product constant, K_{sp} , for BaSO_4 , given that a maximum of 1.167×10^{-8} g will dissolve in 500 mL of water?

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

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

Barium sulfate dissolves according to the equilibrium:



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

$$[\text{Ba}^{2+}(\text{aq})] = [\text{SO}_4^{2-}(\text{aq})] = \frac{\text{number of moles}}{\text{volume}} = \frac{5.000 \times 10^{-11}}{0.500} = 1.00 \times 10^{-10} \text{ M}$$

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



lies far to the left. Adding more $\text{SO}_4^{2-}(\text{aq})$ ions pushes it further to the left.

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

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