Marks 7

- The K_{sp} of AgBr is 5.0×10^{-13}. The K_{stab} of $[Ag(S_2O_3)_2]^{3-}$ is 4.7×10^{13}. Calculate the value of the equilibrium constant for the dissolution of AgBr in $Na_2S_2O_3$ solution.

The reactions corresponds to K_{sp} and K_{stab} can be added together to give the reaction for the dissolution of AgBr in $Na_2S_2O_3$ solution:

$$AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq) \qquad K_{sp} = 5.0 \times 10^{-13}$$

$$Ag^+(aq) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Ag(S_2O_3)_2]^{3-}(aq) \qquad K_{stab} = 4.7 \times 10^{13}$$

$$AgBr(s) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Ag(S_2O_3)_2]^{3-}(aq) + Br^-(aq) \qquad K = K_{sp} \times K_{stab}$$

The equilibrium constant for the overall reaction is the product of the equilibrium constants for the individual reactions:

$$K = K_{sp} \times K_{stab} = (5.0 \times 10^{-13}) \times (4.7 \times 10^{13}) = 24$$

Answer: **24**

Calculate the solubility of AgBr in 2.0 M $Na_2S_2O_3$.

The solubility can be calculated using a reaction table, assuming x mol dissolves:

	AgBr(s)	$2S_2O_3^{2-}(aq)$	$\rightleftharpoons$	$[Ag(S_2O_3)_2]^{3-}(aq)$	$Br^-(aq)$
initial	excess	2.0		0	0
change	$-x$	$-2x$		$+x$	$+x$
final	excess	$2.0 - 2x$		x	x

$$K = \frac{[Ag(S_2O_3)_2]^{3-}(aq)][Br^-(aq)]}{[S_2O_3{}^{2-}(aq)]^2}$$

$$= \frac{(x)(x)}{(2.0-2x)^2} = \frac{x^2}{(2.0-2x)^2} = 24$$

Taking square roots of both sides gives:

$$\frac{x}{(2.0-2x)} = (24)^{1/2} \qquad x = 0.91 \text{ mol L}^{-1}$$

Answer: **0.91 mol L^{-1}**

The K_{stab} for $[Ag(S_2O_3)_2]^{3-}$ is much greater than the K_{stab} for $[Ag(NH_3)_2]^+$. Explain why this is so.

$S_2O_3^{2-}$ is a stronger ligand than NH_3, presumably because of its negative charge.