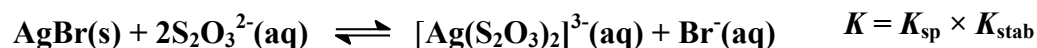
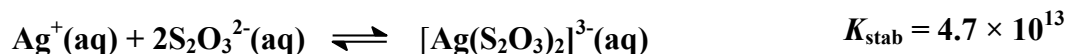


- The  $K_{sp}$  of AgBr is  $5.0 \times 10^{-13}$ . The  $K_{stab}$  of  $[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$  is  $4.7 \times 10^{13}$ . Calculate the value of the equilibrium constant for the dissolution of AgBr in  $\text{Na}_2\text{S}_2\text{O}_3$  solution.

Marks  
7

The reactions corresponds to  $K_{sp}$  and  $K_{stab}$  can be added together to give the reaction for the dissolution of AgBr in  $\text{Na}_2\text{S}_2\text{O}_3$  solution:



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  $\text{Na}_2\text{S}_2\text{O}_3$ .

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

	AgBr(s)	$2\text{S}_2\text{O}_3^{2-}(\text{aq})$	$\rightleftharpoons$	$[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}(\text{aq})$	$\text{Br}^-(\text{aq})$
initial	excess	2.0		0	0
change	-x	-2x		+x	+x
final	excess	$2.0 - 2x$		x	x

$$K = \frac{[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}(\text{aq})[\text{Br}^-(\text{aq})]}{[\text{S}_2\text{O}_3^{2-}(\text{aq})]^2}$$

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

Taking square roots of both sides gives:

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

Answer: 0.91 mol L<sup>-1</sup>

The  $K_{stab}$  for  $[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$  is much greater than the  $K_{stab}$  for  $[\text{Ag}(\text{NH}_3)_2]^+$ . Explain why this is so.

$\text{S}_2\text{O}_3^{2-}$  is a stronger ligand than  $\text{NH}_3$ , presumably because of its negative charge.