

Ionic Equilibria in Water



#12

Calculate the solubility of AgBr in $1.00 \frac{\text{mol}}{\text{L}} \text{NH}_3$.

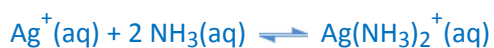
$$K_{sp\text{AgBr}} = 5.0 \times 10^{-13}$$

$$\beta_2_{\text{Ag}(\text{NH}_3)_2^+} = 1.7 \times 10^7$$

Solution



$$K_{sp\text{AgBr}} = [\text{Ag}^+][\text{Br}^-] = 5.0 \times 10^{-13}$$



$$\beta_2_{\text{Ag}(\text{NH}_3)_2^+} = 1.7 \times 10^7$$

Global reaction:



$$K = K_{sp} \times \beta_2 = \frac{[\text{Ag}(\text{NH}_3)_2^+][\text{Br}^-]}{[\text{NH}_3]^2} = 8.5 \times 10^{-6}$$

$\frac{\text{mol}}{\text{L}}$	AgBr	NH ₃	Ag(NH ₃) ₂ ⁺	Br ⁻
Initial	?	1.00	-	-
Δ	-S = -x	-2x	+x	+x
Equilibrium	?	1.00 - 2x	x	x

$$K = \frac{[\text{Ag}(\text{NH}_3)_2^+][\text{Br}^-]}{[\text{NH}_3]^2} = 8.5 \times 10^{-6} = \frac{x^2}{(1.00 - 2x)^2}$$

$$\Rightarrow x = 2.9 \times 10^{-3} \frac{\text{mol}}{\text{L}}$$

$$\Rightarrow S = 2.9 \times 10^{-3} \frac{\text{mol}}{\text{L}}$$