Ionic Equilibriums in Water

#12

Calculate the solubility of AgBr in 1.00 $\frac{\text{mol}}{\text{L}}$ NH_3.

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

$$\beta_{2_{Ag(NH_3)_2^+}} = 1.7 \times 10^{7}$$

Solution

$AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)$ $\qquad K_{sp_{AgBr}} = \left[Ag^+\right]\left[Br^-\right] = 5.0 \times 10^{-13}$

$Ag^+(aq) + 2\ NH_3(aq) \rightleftharpoons Ag(NH_3)_2^+(aq)$ $\qquad \beta_{2_{Ag(NH_3)_2^+}} = 1.7 \times 10^{7}$

Global reaction:

$AgBr(s) + 2\ NH_3(aq) \rightleftharpoons Ag(NH_3)_2^+(aq) + Br^-(aq)$

$$K = K_{sp} \times \beta_2 = \frac{\left[Ag(NH_3)_2^{\ +}\right]\left[Br^-\right]}{[NH_3]^2} = 8.5 \times 10^{-6}$$

$\frac{\text{mol}}{\text{L}}$	AgBr	NH_3	$Ag(NH_3)_2^+$	Br^-
Initial	?	1.00	–	–
Δ	$-S = -x$	$-2x$	$+x$	$+x$
Equilibrium	?	$1.00 - 2x$	x	x

$$K = \frac{\left[Ag(NH_3)_2^{\ +}\right]\left[Br^-\right]}{[NH_3]^2} = 8.5 \times 10^{-6} = \frac{x^2}{(1.00 - 2x)^2}$$

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

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