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- Calculate the equilibrium constant for the following reaction.

$$AgI(s) + 2CN^-(aq) \rightleftharpoons [Ag(CN)_2]^-(aq) + I^-(aq)$$

Data: K_{stab} of $[Ag(CN)_2]^-$ = 3×10^{20}; K_{sp} of AgI = 8.3×10^{-17}

The equations for the dissolution of AgI and the stability constant of the complex $[Ag(CN)_2]^-$ are, respectively:

$$\mathbf{AgI(s) \rightleftharpoons Ag^+(aq) + I^-(aq)} \qquad \mathbf{K_{sp} = [Ag^+(aq)][I^-(aq)]}$$

$$\mathbf{Ag^+(aq) + 2CN^-(aq) \rightleftharpoons [Ag(CN)_2]^-(aq)} \qquad \mathbf{K_{stab} = \frac{[[Ag(CN)_2]^-(aq)]}{[Ag^+(aq)][CN^-(aq)]^2}}$$

Addition of these reactions gives the required reaction and so the equilibrium constant for the reaction is the product of the individual equilibrium constants:

$$\mathbf{AgI(s) + 2CN^-(aq) \rightleftharpoons [Ag(CN)_2]^-(aq) + I^-(aq)}$$

$$\mathbf{K = K_{sp} \times K_{stab} = [Ag^+(aq)][I^-(aq)] \times \frac{[[Ag(CN)_2]^-(aq)]}{[Ag^+(aq)][CN^-(aq)]^2}}$$

$$\mathbf{= \frac{[[Ag(CN)_2]^-(aq)][I^-(aq)]}{[CN^-(aq)]^2}}$$

Hence,

$$\mathbf{K = K_{sp} \times K_{stab} = (8.3 \times 10^{-17}) \times (3 \times 10^{20}) = 2 \times 10^4}$$

Answer: $\mathbf{2 \times 10^4}$