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- Give the equilibrium concentration of $Ni^{2+}(aq)$ ions in a solution formed by dissolving 0.15 mol of $NiCl_2$ in 0.500 L of 2.00 M KCN solution. The K_{stab} of $[Ni(CN)_4]^{2-} = 1.7 \times 10^{30}$.

When the $NiCl_2$ is added to 0.500 L, the concentration of $Ni^{2+}(aq)$:

concentration = $[Ni^{2+}(aq)]$ = number of moles / volume
= 0.15 mol / 0.500 L = 0.30 M

As K_{stab} is so large, essentially *all* of this will be complexed by the excess $CN^-(aq)$ ions:

$$Ni^{2+}(aq) + 4CN^-(aq) \rightleftharpoons [Ni(CN)_4{}^{2-}(aq)]$$

As essentially all of the $Ni^{2+}(aq)$ becomes $[Ni(CN)_4{}^{2-}(aq)]$:

$[Ni(CN)_4{}^{2-}(aq)] = 0.30$ M
$[CN^-(aq)] = (2.00 - 4 \times 0.30)$ M = 0.80 M

K_{stab} is the equilibrium constant for the reaction so:

$$K_{stab} = \frac{[Ni(CN)_4^{2-}(aq)]}{[Ni^{2+}(aq)][CN^-(aq)]^4}$$

If the *tiny* amount of uncomplexed $Ni^{2+}(aq)$ has a concentration of x M then:

$$K_{stab} = \frac{(0.30)}{(x)(0.80)^4} = 1.7 \times 10^{30} \quad \text{so } x = 4.3 \times 10^{-31} \text{ M}$$

Answer: **4.3×10^{-31} M**