

$$15.25 (b) \frac{3.6 \times 10^{-5} \text{ mol Ag}^+}{2} \times 0.10 \neq 3.6 \times 10^{-10} \text{ mol Ag}^+$$

$$(c) 6.6 \times 10^{-10} \frac{\text{mol Ag}^+}{L} \times 0.10 \neq 6.6 \times 10^{-7} \text{ mol Ag}^+$$

15.27	$\Delta H^\circ (\text{kJ/mol})$	$\Delta S^\circ (\text{J/mol}\cdot\text{K})$	$\Delta G^\circ (\text{kJ/mol})$	K	spontaneous?
Ku	16.9	76.4	-6.0	>1	yes
AgCl	65.5	33.0	55.7	<<1	no
$\text{Cd}(\text{OH})_2$	-17.6	-158.1	29.5	<<1	no

$$15.28 \quad \Delta G^\circ = -RT \ln K \quad \Rightarrow \quad K = \exp\left(-\frac{\Delta G^\circ}{RT}\right), \text{ so}$$

$$\text{AgCl: } K_{sp} = \exp\left[-\frac{55700 \text{ J mol}^{-1}}{(8.3145 \text{ J mol}^{-1}\text{K}^{-1})(298 \text{ K})}\right] = 1.7 \times 10^{-10}$$

$$\text{Cd}(\text{OH})_2: K_{sp} = \exp\left[-\frac{29500 \text{ J mol}^{-1}}{(8.3145 \text{ J mol}^{-1}\text{K}^{-1})(298 \text{ K})}\right] = 6.8 \times 10^{-6}$$

15.33 when the can is sealed, the equilibrium:
 $\text{CO}_2(\text{aq}) \rightleftharpoons \text{CO}_2(\text{g})$ lies to the left.

A sudden reduction of CO_2 pressure drives the reaction to the right.

$$\text{M}^2+0.1 \times 0.2 = \frac{0.1 \times 0.2}{0.1 \times 0.2} = \frac{0.2}{0.2} = [0.2] \quad \text{Starting concentration: M}^2 = 0.2 \text{ M}$$

$$\left[\frac{0.2}{x}\right]^2 = [0.2] \quad \text{or} \quad \frac{0.04}{x^2} = [0.2] = \left[\frac{0.04}{x^2}\right] = [0.2] = (0.2)^2 = 0.04$$

$x = 0.1 \times 3.3 = 0.33 \text{ M}$ M^2 is not significant enough to affect the result.