Chemistry 2720 Fall 2005 Quiz 5 Solution

When calcium hydroxide dissolves in water, it dissociates into its ions:

$$\mathrm{Ca(OH)_{2(s)}} \rightarrow \mathrm{Ca^{2+}_{(aq)}} + 2\mathrm{OH^{-}_{(aq)}}.$$

The equilibrium constant for this process is

$$K_{\mathrm{sp}} = (a_{\mathrm{Ca^{2+}}})(a_{\mathrm{OH^-}})^2.$$

We need to convert the concentrations to mol/L:

$$\begin{aligned} 0^\circ\mathrm{C}: \quad & s = \tfrac{1.85\,\mathrm{g/L}}{74.093\,\mathrm{g/mol}} = 0.0250\,\mathrm{mol/L}. \\ 100^\circ\mathrm{C}: \quad & s = \tfrac{0.77\,\mathrm{g/L}}{74.093\,\mathrm{g/mol}} = 0.010\,\mathrm{mol/L}. \end{aligned}$$

These concentrations are well in excess of those which would be generated by the autoionization of water. Therefore $a_{\mathrm{Ca^{2+}}} = s$ and $a_{\mathrm{OH^-}} = 2s$. The equilibrium constant at the two temperatures is therefore

$$\begin{aligned} K_0 &= (0.0250\,\mathrm{mol/L})(0.0499\,\mathrm{mol/L})^2 = 6.23 \times 10^{-5}. \\ K_{100} &= (0.010\,\mathrm{mol/L})(0.021\,\mathrm{mol/L})^2 = 4.5 \times 10^{-6}. \end{aligned}$$

We can now calculate $\Delta\bar{H}^\circ$ from the formula relating equilibrium constants to temperature:

$$\begin{aligned} \Delta\bar{H}^\circ &= \frac{R\ln(K_2/K_1)}{T_1^{-1} - T_2^{-1}} \\ &= \frac{(8.314\,472\,\mathrm{J\,K^{-1}mol^{-1}})\ln\dfrac{4.5\times10^{-6}}{6.23\times10^{-5}}}{\dfrac{1}{273.15\,\mathrm{K}} - \dfrac{1}{373.15\,\mathrm{K}}} \\ &= -22\,\mathrm{kJ/mol}. \end{aligned}$$