Determination of Enthalpy of Neutralization by Calorimetry

Problem 7-24

The heat of neutralization of HCl(aq) by NaOH(aq) is -55.84 kJ/mol H2O produced. If 50.00 mL of 1.05 M NaOH is added to 25.00 mL of 1.86 M HCl, with both solutions originally at 24.72°C, what will be the final temperature? Assume that no heat is lost to the surrounding air and that the solution produced in the neutralization reaction has a density of 1.02 g/mL and a specific heat of 3.89 J g^{-1}°C^{1}.

Solution:

This is a combination of a limiting reactant problem and calorimetry. Let's start with the stoichiometry.

$$\mathrm{HCl(aq)} + \mathrm{NaOH(aq)} \rightarrow \mathrm{NaCl(aq)} + \mathrm{H_2O(l)}$$

$$n_{\mathrm{HCl}} = 0.02500\ \mathrm{L} \times 1.86\ \mathrm{M} = 0.04650\ \mathrm{mol\ HCl}$$

$$n_{\mathrm{NaOH}} = 0.05000\ \mathrm{L} \times 1.05\ \mathrm{M} = 0.05250\ \mathrm{mol\ NaOH}$$

Since the stoichiometry is 1:1, by inspection the HCl is the limiting reactant.

$$n_{\mathrm{H_2O}} = 0.04650\ \mathrm{mol\ H_2O\ produced}$$

Now for the calorimetry. Write the 1st Law equation for the system:

$$q_{\mathrm{rxn}} + q_{\mathrm{soln}} = 0$$

Expand the heat terms

$$\Delta H_{\mathrm{neutralization}} \times n_{\mathrm{H_2O}} + m_{\mathrm{soln}} c_{\mathrm{soln}} \Delta T_{\mathrm{soln}} = 0$$

The mass of the solution is calculated from density and volume:

$$m_{\mathrm{soln}} = 75.00\ \mathrm{mL} \times 1.02\,\tfrac{\mathrm{g}}{\mathrm{mL}} = 76.50\ \mathrm{g\ soln}$$

Finally, do the algebra:

$$\left(-55.84\times10^{3}\,\tfrac{\mathrm{J}}{\mathrm{mol}}\right)\left(0.04650\ \mathrm{mol\ H_2O}\right) + \left(76.50\ \mathrm{g}\right)\left(3.89\,\tfrac{\mathrm{J}}{\mathrm{g°C}}\right)\left(T_{\mathrm{f}} - 24.72\mathrm{°C}\right) = 0$$

$$\left(297.59\,\tfrac{\mathrm{J}}{\mathrm{°C}}\right)\left(T_{\mathrm{f}} - 24.72\mathrm{°C}\right) = 2596.56\ \mathrm{J}$$

$$\left(T_{\mathrm{f}} - 24.72\mathrm{°C}\right) = 8.725\mathrm{°C}$$

$$T_{\mathrm{f}} = 33.45\mathrm{°C}$$