• A 50.0 mL solution contained 10.00 g of NaOH in water at 25.00 °C. When it was added to a 250.0 mL solution of 0.200 M HCl at 25.00 °C in a “coffee cup” calorimeter, the temperature of the solution rose to 33.95 °C. Assuming the specific heat of the solution is 4.18 J K⁻¹ g⁻¹, that the calorimeter absorbs a negligible amount of heat, and that the density of the solution is 1.00 g mL⁻¹, calculate \( \Delta H_r \) (in kJ mol⁻¹) for the following reaction.  
\[ \text{H}^+ (aq) + \text{OH}^- (aq) \rightarrow \text{H}_2\text{O}(l) \]

The total volume = (50.0 + 250.0) mL = 300.0 mL. If the density = 1.00 g mL⁻¹, this has a mass of 300.0 g. The heat change is given by:

\[
q = c \times m \times \Delta T
= (4.18 \text{ J K}^{-1} \text{ g}^{-1}) \times (300.0 \text{ g}) \times ((33.95 - 25.00) \text{ K}) = 11223 \text{ J or 11.22 kJ}
\]

NaOH has a molar mass of (22.99 (Na) + 16.00 (O) + 1.008 (H)) g mol⁻¹ = 40.00 g mol⁻¹ so 10.00 g corresponds to 0.2500 mol.

250.0 mL of a 0.200 M solution of HCl contains (0.2500 mL) × (0.200 mol L⁻¹) = 0.0500 mol.

As the reaction is a 1:1 reaction between NaOH and HCl, only 0.0500 mol of the NaOH can react. The heat change is therefore associated with 1 mol is therefore:

\[
\text{heat change for 1 mol} = \frac{11223 \text{ J}}{0.0500 \text{ mol}} = 225 \text{ kJ mol}^{-1}
\]

The reaction is exothermic as the temperature rises. The enthalpy change is therefore:

\[ \Delta_r H = -225 \text{ kJ mol}^{-1} \]

When the experiment was repeated using 12.00 g of NaOH in water, the temperature increase was the same. Explain.

As noted above, 10.00 g of NaOH contains more moles than 250.0 mL of 0.200 M HCl. 12.00 g of NaOH therefore contains even more moles and so is still in excess. This excess cannot react and so there is no change in the amount of heat produced.
Water solutions of NaOH (100 mL, 2.0 M) and HCl (100 mL, 2.0 M), both at 24.6 °C, were mixed together in a coffee cup calorimeter. The temperature of the solution rose to 38.0 °C during the reaction process. Write a balanced chemical equation to describe the reaction in the calorimeter.

\[ \text{H}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{H}_2\text{O}(l) \]

Is the process an endothermic or exothermic reaction? Temperature increases so exothermic

Assuming a perfect calorimeter, determine the standard enthalpy change for the neutralisation reaction. Assume the density of water is 1.00 g mL⁻¹. The heat capacity of water is 4.18 J K⁻¹ g⁻¹.

The total volume = 100.0 + 100.0 = 200.0 mL. If the density = 1.00 g mL⁻¹, this has a mass of 200.0 g. The heat change is given by:

\[ q = c \times m \times \Delta T \]
\[ = (4.18 \text{ J K}^{-1} \text{ g}^{-1}) \times (200.0 \text{ g}) \times ((38.0 - 24.6) \text{ K}) = 11202 \text{ J or 11.2 kJ} \]

Both solutions have the same concentration. 200 mL of a 2.0 M solution contains \((0.200 \text{ L}) \times (2.0 \text{ mol L}^{-1}) = 0.40 \text{ mol}\).

The reaction is a 1:1 reaction between NaOH and HCl and equal amounts of each are present. The reaction of 0.20 mol of NaOH with 0.20 mol of HCl generates 11.2 kJ so the heat change for the reaction of 1 mol with 1 mol is:

heat change for 1 mol = \( \frac{11202 \text{ J}}{0.20 \text{ mol}} = 56 \text{ kJ mol}^{-1} \)

The reaction is exothermic as the temperature rises. The enthalpy change is therefore:

\[ \Delta_r H = -56 \text{ kJ mol}^{-1} \]
A 0.50 g sample of ammonium nitrate, \( \text{NH}_4\text{NO}_3(s) \), was dissolved in 35.0 g of water in a coffee cup calorimeter. The temperature of the solution dropped from 22.7 to 21.6 °C. Write a balanced equation to describe the reaction in the calorimeter.

\[
\text{NH}_4\text{NO}_3(s) \rightarrow \text{NH}_4^+(aq) + \text{NO}_3^-(aq)
\]

Describe this process as either endothermic or exothermic.

Assuming a perfect calorimeter what is the heat of solution of ammonium nitrate, expressed in kJ mol\(^{-1}\)? Assume the density of the solution is 1.00 g mL\(^{-1}\) and that the heat capacity of the solution is 4.18 J K\(^{-1}\) g\(^{-1}\).

The molar mass of \( \text{NH}_4\text{NO}_3 \) is \((14.01 \text{ (N) + 4 } \times 1.008 \text{ (H) + 14.01 \text{ (N) + 3 } 	imes 16.00 \text{ (O)) g mol}^{-1} = 80.052 \text{ g mol}^{-1} \). The sample of 0.50 g therefore corresponds to:

\[
\text{number of moles} = \frac{0.50 \text{ g}}{80.052 \text{ g mol}^{-1}} = 0.0062 \text{ mol}
\]

The total mass of \( \text{NH}_4\text{NO}_3 \) and water is \((0.50 + 35.0) \text{ g} = 35.5 \text{ g} \). The heat change is given by:

\[
q = c \times m \times \Delta T
= (4.18 \text{ J K}^{-1} \text{ g}^{-1}) \times (35.5 \text{ g}) \times ((22.7 - 21.6) \text{ K}) = 163 \text{ J or 0.163 kJ}
\]

This is the heat change produced by 0.0062 mol. The heat change produced by 1 mol is therefore \( \frac{163 \text{ J}}{0.0062 \text{ mol}} = 26 \text{ kJ mol}^{-1} \).

The reaction is endothermic as the temperature drops. The enthalpy change is therefore:

\[
\Delta_r H = +26 \text{ kJ mol}^{-1}
\]

Heat radiating fins are used to dissipate heat and prevent damage to electronic components. Is it better to make the fins out of aluminium or iron? Give reasons for your answer.

Data: Specific heat of Al = 0.900 J K\(^{-1}\) g\(^{-1}\) Specific heat of Fe = 0.444 J K\(^{-1}\) g\(^{-1}\)

The fins are required to remove heat from the electronic components. Aluminium has a higher heat capacity so it can absorb more heat per gram.