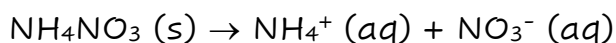


Calorimetry

1. Octane, C_8H_{18} , has a specific heat of $2.22 \text{ J/(g}\cdot\text{K)}$. What quantity of heat, in kJ, is required to raise the temperature of 125.00 g of octane from $15.0 \text{ }^\circ\text{C}$ to $28.0 \text{ }^\circ\text{C}$?

$$q = c \times m \times \Delta T; \quad \Delta T = 28.0 \text{ }^\circ\text{C} - 15.0 \text{ }^\circ\text{C} = 13.0 \text{ }^\circ\text{C}$$
$$q = 2.22 \text{ J/(g}\cdot\text{K)} \times 125.00 \text{ g} \times 13.0 \text{ }^\circ\text{C} = \mathbf{3.61 \times 10^3 \text{ J}}$$

2. Consider the equation below:



A 4.25 g sample of solid ammonium nitrate is dissolved in 60.0 g of water in a coffee cup calorimeter. The temperature decreases from $22.5 \text{ }^\circ\text{C}$ to 17.4°C . Calculate ΔH , in kJ/mol of NH_4NO_3 , for this dissolution process. Assume the specific heat of the solution is that of water, $4.184 \text{ J/(g}\cdot\text{K)}$. Is this an endothermic or an exothermic process?

$$q_{\text{solution}} = c \times m \times \Delta T; \quad \text{mass of solution} = 60.0 \text{ g} + 4.25 \text{ g} = 64.25 \text{ g}$$
$$\Delta T = 17.4 \text{ }^\circ\text{C} - 22.5 \text{ }^\circ\text{C} = -5.1 \text{ }^\circ\text{C}$$

$$q_{\text{solution}} = 4.184 \text{ J/(g}\cdot\text{K)} \times 64.25 \text{ g} \times -5.1 \text{ }^\circ\text{C} = -1370.99 \text{ J}$$

The solution (the surroundings) loses heat to the system. The system is the NH_4NO_3 . $q_{\text{system}} = +1370.99 \text{ J}$; the dissolution is **endothermic**.

A 4.25 g sample requires 1370.99 J of heat to dissolve. We need the kJ/mol. $M_m (\text{NH}_4\text{NO}_3) = 80.043 \text{ g/mol}$

$$\frac{1.37 \text{ kJ}}{4.25 \text{ g}} \times \frac{80.043 \text{ g}}{1 \text{ mol}} = \mathbf{25.8 \text{ kJ/mol}}$$

3. A 1.00 g sample of pine nuts was burned in a bomb calorimeter containing 250.0 grams of water at an initial temperature of $22.5 \text{ }^\circ\text{C}$. Once the reaction was completed, the temperature of the water was $49.2 \text{ }^\circ\text{C}$. The heat capacity of the calorimeter is $8.74 \text{ J/}^\circ\text{C}$. Calculate the heat of combustion for the pine nuts in kJ/g. How many Cal (food calories) is $100. \text{ g}$ of pine nuts?

$$q_{\text{rxn}} = -q_{\text{water}} - q_{\text{bomb}} \quad \Delta T = 49.2 \text{ }^\circ\text{C} - 22.5 \text{ }^\circ\text{C} = 26.7 \text{ }^\circ\text{C}$$

$$q_{\text{rxn}} = -(4.184 \text{ J/(g}\cdot\text{K)} \times 250.0 \text{ g} \times 26.7^\circ\text{C}) - 8.74 \text{ J/}^\circ\text{C} \times 26.7 \text{ }^\circ\text{C} =$$
$$-28161.6 \text{ J per } 1.00 \text{ g of pine nuts} = \mathbf{28.2 \text{ kJ/g.}}$$

$$-28161.6 \frac{\text{J}}{\text{g}} \times \frac{\text{cal}}{4.184 \text{ J}} \times \frac{1 \text{ Cal}}{1000 \text{ cal}} = 6.73 \frac{\text{Cal}}{\text{g}} \quad \mathbf{100 \text{ g would be } 673 \text{ Cal.}}$$