Use the thermochemical data provided to calculate the heat of reaction of the following reaction:

$$\text{PCl}_3(l) + \text{Cl}_2(g) \rightarrow \text{PCl}_5(s)$$

Data: 

1. $\text{P}_4(s) + 6\text{Cl}_2(g) \rightarrow 4\text{PCl}_3(l) \quad \Delta H^\circ = -1280 \text{ kJ mol}^{-1}$
2. $\text{P}_4(s) + 10\text{Cl}_2(g) \rightarrow 4\text{PCl}_5(s) \quad \Delta H^\circ = -1774 \text{ kJ mol}^{-1}$

The reactions given in the data correspond to the enthalpy of formation of PCl$_3$(l) and PCl$_5$(s). As $\Delta H^\circ_{\text{rxn}} = \sum n\Delta H^\circ_{\text{f(products)}} - \sum n\Delta H^\circ_{\text{f(reactants)}}$ and the enthalpy of formation of Cl$_2$(g) is zero:

$$\Delta H^\circ_{\text{rxn}} = (-1774) - (-1280) = -494 \text{ kJ mol}^{-1}$$

ANSWER: $-494 \text{ kJ mol}^{-1}$

Identify one property of a molecule necessary for it to be considered a “greenhouse gas”.

To be a greenhouse gas, a molecule must be able to absorb infrared light: it must undergo a vibration which causes a change in the dipole moment.
The normal boiling point of chloroform is 61.7 °C and its enthalpy of vaporisation is 31.4 kJ mol⁻¹. Calculate the entropy of vaporisation for chloroform at 1 atm.

The entropy change for vaporisation is:

$$\Delta_{vap} S = \frac{\Delta_{vap} H}{T_{vap}} = \frac{31.4 \times 10^3}{(61.7 + 273)} = 93.8 \text{ J K}^{-1} \text{ mol}^{-1}$$

Answer: 93.8 J K⁻¹ mol⁻¹
When 156 g of aluminium metal at 50.0 °C is added to 100 g of water at 20.0 °C, the final temperature becomes 30.0 °C. The heat capacity of water is 4.18 J K⁻¹ g⁻¹. Calculate the specific heat capacity of aluminium.

Assuming equilibrium has been established and no heat is lost to the surroundings, the heat lost by the aluminium is gained by the water.

As $q = c \times m \times \Delta T$:

$$c_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}} = c_{\text{aluminium}} \times m_{\text{aluminium}} \times \Delta T_{\text{aluminium}}$$

or

$$4.18 \times 100 \times (30.0 - 20.0) = c_{\text{aluminium}} \times 156 \times (50.0 - 30.0)$$

so,

$$c_{\text{aluminium}} = 1.34 \text{ J g}^{-1} \text{ mol}^{-1}$$

Answer: 1.34 J g⁻¹ mol⁻¹

What would the final temperature have been if the 156 g of aluminium metal at 50 °C had been added to iced water containing 10 g ice and 90 g water at 0 °C? ∆$H$ for the freezing of water is –6.02 kJ mol⁻¹.

In this case, some of the heat will be used to melt the ice. As the molar mass of H₂O is 2 × 1.008 (H) + 16.00 = 18.016, the number of moles in 10 g of ice is:

$$\text{moles of ice} = \frac{10}{18.016} = 0.555 \text{ mol}$$

Hence, the heat required to melt 10 g of ice is $(6.02 \times 10³) \times 0.555 = 3340 \text{ J}$.

If $T$ is the final temperature of the water, the overall heat change for the water is the heat required to melt 10 g of ice plus the heat required to warm 100 g of water from 0 °C to $T$:

$$q_{\text{water}} = q_{\text{melt}} + (c_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}}) = 3340 + (4.18 \times 100 \times (T - 0))$$

This heat is provided by cooling the aluminium from 50 °C to $T$:

$$q_{\text{aluminium}} = c_{\text{aluminium}} \times m_{\text{aluminium}} \times \Delta T_{\text{aluminium}} = 1.34 \times 156 \times (50 - T)$$

As $q_{\text{water}} = q_{\text{aluminium}}$, $3340 + (4.18 \times 100 \times T) = 1.34 \times 156 \times (50 - T)$

This rearranges to give $T = 11 \degree \text{C}$

Answer: 11 °C
In an experiment, 50.0 mL of 1.00 M HNO₃ was combined with 50.0 mL of 0.540 M NaOH in a calorimeter. Give an equation for the reaction that took place.

\[ \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \rightarrow 2\text{H}_2\text{O}(l) \]

The temperature of the solution was found to increase by 2.98 °C. If the heat capacity of the calorimeter was 80.0 J K⁻¹ and the heat capacity of the final solution was 426 J K⁻¹, determine the molar heat of reaction.

As \( q = C \times \Delta T \) for both the calorimeter and the solution,

\[ q = (80.0 \times 2.98) + (426 \times 2.98) = 1510 \text{ J} \]

As the reaction is a 1:1 reaction and fewer moles of OH⁻ are present, it is rate limiting. The number of moles of OH⁻ is:

\[ \text{moles of OH}^- = \frac{50.0}{1000.0} \times 0.540 = 0.027 \text{ mol} \]

This number of moles generates 1510 J of heat. The reaction is exothermic as the temperature increases. The enthalpy change for one mole is therefore:

\[ \Delta_{\text{rxn}}H = \frac{-1510}{0.027} = -55800 \text{ J mol}^{-1} = -55.8 \text{ kJ mol}^{-1} \]

Answer: \(-55.8 \text{ kJ mol}^{-1}\)

The average bond enthalpy of the O-H bond is 463 kJ mol⁻¹. Explain briefly why the heat of neutralisation calculated in the first part of this question differs significantly from this value.

The bond enthalpy refers to the reaction \( \text{OH}(g) \rightarrow \text{O}(g) + \text{H}(g) \) rather than the transfer of a proton from \( \text{H}_3\text{O}^+ \) to \( \text{OH}^- \) in aqueous solution.