The conversion of hydroquinone ($C_6H_6O_2$(aq)) to quinone ($C_6H_4O_2$(aq)) is involved in many important biochemical reactions. The bombardier beetle, for example, uses the explosive reaction between hydroquinone and hydrogen peroxide (as described by the equation below) as a defence mechanism.

$$C_6H_6O_2(aq) + H_2O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H_2O(l)$$

From the following reaction data, calculate $\Delta H_{rxn}$ for the reaction between 1.00 mol of hydroquinone and 1.00 mol of hydrogen peroxide.

- $C_6H_6O_2(aq) \rightarrow C_6H_4O_2(aq) + H_2(g)$  \hspace{1cm} $\Delta H_{rxn} = +177.4$ kJ mol$^{-1}$
- $O_2(g) + 2H_2O(l) \rightarrow 2H_2O_2(aq)$  \hspace{1cm} $\Delta H_{rxn} = +189.1$ kJ mol$^{-1}$
- $H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g)$  \hspace{1cm} $\Delta H_{rxn} = +285.8$ kJ mol$^{-1}$

Use the answer you obtained above to calculate the heat liberated (in joules) in the oxidation of $3.86 \times 10^{-4}$ mol of hydroquinone to quinone.

Answer:

Calculate the temperature rise of 0.250 g of water for this quantity of heat. (The heat capacity of water, $C_p = 4.184$ J K$^{-1}$ g$^{-1}$)

Answer:
The combustion of hydrazine, \( \text{N}_2\text{H}_4 \), with oxygen is described by the following equation:

\[
\text{N}_2\text{H}_4(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \quad \Delta H^\circ = -623 \text{ kJ mol}^{-1}
\]

Given that \( \Delta H^\circ_f \) of \( \text{H}_2\text{O}(\text{l}) \) is \(-286 \text{ kJ mol}^{-1} \), find the standard enthalpy of formation of \( \text{N}_2\text{H}_4(\text{l}) \).

The combustion of 1.00 mol of \text{N}_2\text{H}_4(\text{l}) can also be accomplished using \( \text{N}_2\text{O}_4(\text{l}) \) as the oxidant, whereupon 629 kJ of energy is released at standard temperature and pressure. What is the standard enthalpy of formation of \( \text{N}_2\text{O}_4(\text{l}) \)?