Marks 3

- The combustion of hydrazine, N_2H_4, with oxygen is described by the following equation:

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(l) \qquad \Delta H° = -623 \text{ kJ mol}^{-1}$$

Given that $\Delta H°_f$ of $H_2O(l)$ is –286 kJ mol^{-1}, find the standard enthalpy of formation of $N_2H_4(l)$.

Using $\Delta_{rxn}H^o = \sum m\Delta_f H^o(\text{products}) - \sum n\Delta_f H^o(\text{reactants})$,

$\Delta_{rxn}H^o = [\Delta_f H^o(N_2(g) + 2\Delta_f H^o(H_2O(l)] - [\Delta_f H^o(N_2H_4(l)) + \Delta_f H^o(O_2(g)]$

As $\Delta_f H^o = 0$ for an element in its standard state, this becomes:

$\Delta_{rxn}H^o = [0 + (2 \times -286)] - [\Delta_f H^o(N_2H_4(l)) + 0] = -623$

Hence,

$\Delta_f H^o(N_2H_4(l)) = +51 \text{ kJ mol}^{-1}$

$\Delta H°_f$ = **+51 kJ mol^{-1}**

The combustion of 1.00 mol of $N_2H_4(l)$ can also be accomplished using $N_2O_4(l)$ as the oxidant, whereupon 629 kJ of energy is released at standard temperature and pressure. What is the standard enthalpy of formation of $N_2O_4(l)$?

The chemical equation for the combustion of $N_2H_4(l)$ with $N_2O_4(l)$ is:

$2N_2H_4(l) + N_2O_4(l) \rightarrow 3N_2(g) + 4H_2O(l)$

As written, this reaction corresponds to burning two moles of $N_2H_4(l)$ and hence $\Delta_{rxn}H^o = 2\Delta_{comb}H^o = 2 \times (-629 \text{ kJ mol}^{-1}) = -1258 \text{ kJ mol}^{-1}$. The negative sign indicates that energy is released during the reaction.

Using $\Delta_{rxn}H^o = \sum m\Delta_f H^o(\text{products}) - \sum n\Delta_f H^o(\text{reactants})$,

$\Delta_{rxn}H^o = [3\Delta_f H^o(N_2(g) + 4\Delta_f H^o(H_2O(l)] - [2\Delta_f H^o(N_2H_4(l)) + \Delta_f H^o(N_2O_4$

$= [(3 \times 0) + (4 \times -286)] - [(2 \times +51) + \Delta_f H^o(N_2O_4(l))] = -1258$

Hence,

$\Delta_f H^o(N_2O_4(l)) = +12 \text{ kJ mol}^{-1}$

$\Delta H°_f$ = **+12 kJ mol^{-1}**