Marks 4

- Nitroglycerine, $C_3H_5(NO_3)_3$, decomposes to form N_2, O_2, CO_2 and H_2O according to the following equation.

$$4C_3H_5(NO_3)_3(l) \rightarrow 6N_2(g) + O_2(g) + 12CO_2(g) + 10H_2O(g)$$

If 15.6 kJ of energy is evolved by the decomposition of 2.50 g of nitroglycerine at 1 atm and 25 °C, calculate the enthalpy change, $\Delta H°$, for the decomposition of 1.00 mol of this compound under standard conditions.

The molar mass of $C_3H_5(NO_3)_3$ is:

$$(3 \times 12.01\ (C) + 5 \times 1.008\ (H) + 3 \times 14.01\ (N) + 9 \times 16.00\ (O))\ \text{g mol}^{-1}$$
$$= 227.1\ \text{g mol}^{-1}$$

2.50 g therefore corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{2.50\ \text{g}}{227.1\ \text{g mol}^{-1}} = 0.0110\ \text{mol}$$

As this amount leads to 15.6 kJ being evolved, the enthalpy change for the decomposition of 1.00 mol is:

$$\Delta H° = 15.6\ \text{kJ} / 0.0110\ \text{mol} = -1420\ \text{kJ mol}^{-1}$$

Answer: **–1420 kJ mol^{-1}**

Hence calculate the enthalpy of formation of nitroglycerine under standard conditions.

Data:

	$\Delta_f H°$ (kJ mol^{-1})
$H_2O(g)$	–242
$CO_2(g)$	–394

The balanced reaction above is for the decomposition of 4 mol of nitroglycerine. Hence, $\Delta_{rxn}H° = 4 \times -1420\ \text{kJ mol}^{-1} = -5680\ \text{kJ mol}^{-1}$.

Using $\Delta_{rxn}H° = \Sigma m\Delta_f H°(\text{products}) - \Sigma n\Delta_f H°(\text{reactants})$, the enthalpy change for the above reaction is:

$$\Delta_{rxn}H° = [12\Delta_f H°(CO_2(g)) + 10\Delta_f H°(H_2O(g))] - [4\Delta_f H°(C_3H_5(NO_3)_3(l))]$$

Hence:

$$-5680\ \text{kJ mol}^{-1} = [(12 \times -394 + 10 \times -242)\ \text{kJ mol}^{-1}] - [4\Delta_f H°(C_3H_5(NO_3)_3(l))]$$

$$\Delta_f H°(C_3H_5(NO_3)_3(l)) = -367\ \text{kJ mol}^{-1}$$

Answer: **–367 kJ mol^{-1}**