Marks 6

- The thermite reaction is written below. Show that the heat released in this reaction is sufficient for the iron to be produced as molten metal.

$$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(l)$$

Assume that the values in the table are independent of temperature.

Substance	Enthalpy of formation, $\Delta_f H^o$ kJ mol^{-1}	Molar heat capacity, C_p J K^{-1} mol^{-1}	Melting point ^{o}C	Enthalpy of fusion kJ mol^{-1}
Al	0	24	660	11
Al_2O_3	–1676	79	2054	109
Fe	0	25	1535	14
Fe_2O_3	–824	104	1565	138

Assume 1 mol of reactants at initial temperature of 25 °C. Need to show that ΔH for the reaction is *greater* than the amount of energy required to melt 2 mol of Fe(s) and heat all the products (2 mol of Fe(s) + 1 mol of Al_2O_3(s)) to the melting point of Fe.

$$\Delta H = \sum\Delta_f H(\text{products}) - \sum\Delta_f H(\text{reactants})$$

$$= \Delta_f H(Al_2O_3(s)) + 2\Delta_f H(Fe(s)) - (2\Delta_f H(Al(s)) + \Delta_f H(Fe_2O_3(s)))$$

$$= [(-1676 + 2 \times 0) - (-824 + 2 \times 0)] \text{ kJ mol}^{-1}$$

$$= -852 \text{ kJ mol}^{-1}$$

ΔH to heat 2 mol of Fe(s) to its melting point

$$\Delta H = n_{Fe(s)} \times C_p(Fe(s)) \times \Delta T$$

$$= (2 \text{ mol}) \times (25 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1535\text{–}25) \text{ K} = 75.5 \text{ kJ}$$

ΔH to heat 1 mol of Al_2O_3(s) to melting point of Fe(s)

$$\Delta H = n_{Al_2O_3(s)} \times C_p(Al_2O_3(s)) \times \Delta T$$

$$= (1 \text{ mol}) \times (79 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1535\text{–}25) \text{ K} = 119 \text{ kJ}$$

ΔH to melt 2 mol of Fe(s)

$$\Delta H = 2 \times \Delta_{fus}H^\circ = (2 \text{ mol}) \times (14 \text{ kJ mol}^{-1}) = 28 \text{ kJ}$$

Total energy required to melt the iron = (75.5 + 119 + 28) kJ = +222.5 kJ.

The energy generated by the reaction is more than enough to melt the iron.