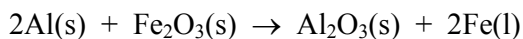


- The thermite reaction is written below. Show that the heat released in this reaction is sufficient for the iron to be produced as molten metal.



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

Substance	Enthalpy of formation, $\Delta_f H^\circ$ kJ mol ⁻¹	Molar heat capacity, C_p J K ⁻¹ mol ⁻¹	Melting point °C	Enthalpy of fusion kJ mol ⁻¹
Al	0	24	660	11
Al ₂ O ₃	-1676	79	2054	109
Fe	0	25	1535	14
Fe ₂ O ₃	-824	104	1565	138

Marks

6

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 Al₂O₃(s)) to the melting point of Fe.

$$\begin{aligned} \Delta H &= \sum \Delta_f H(\text{products}) - \sum \Delta_f H(\text{reactants}) \\ &= \Delta_f H(\text{Al}_2\text{O}_3(\text{s})) + 2\Delta_f H(\text{Fe}(\text{s})) - (2\Delta_f H(\text{Al}(\text{s})) + \Delta_f H(\text{Fe}_2\text{O}_3(\text{s}))) \\ &= [(-1676 + 2 \times 0) - (-824 + 2 \times 0)] \text{ kJ mol}^{-1} \\ &= -852 \text{ kJ mol}^{-1} \end{aligned}$$

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

$$\begin{aligned} \Delta H &= n_{\text{Fe}(\text{s})} \times C_p(\text{Fe}(\text{s})) \times \Delta T \\ &= (2 \text{ mol}) \times (25 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1535 - 25) \text{ K} = 75.5 \text{ kJ} \end{aligned}$$

ΔH to heat 1 mol of Al₂O₃(s) to melting point of Fe(s)

$$\begin{aligned} \Delta H &= n_{\text{Al}_2\text{O}_3(\text{s})} \times C_p(\text{Al}_2\text{O}_3(\text{s})) \times \Delta T \\ &= (1 \text{ mol}) \times (79 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1535 - 25) \text{ K} = 119 \text{ kJ} \end{aligned}$$

ΔH to melt 2 mol of Fe(s)

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

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.