THERMODYNAMICS (E)

Magnesium nitrate decomposes when heated. $2Mg(NO_3)_2(s) \rightarrow 2MgO(s) + 4NO_2(g) + O_2(g)$

substance	$Mg(NO_3)_2(s)$	$MgO(s)$	$NO_2(g)$	$O_2(g)$
$\Delta_f H^\circ$ (kJ mol^{-1})	–790	–602	+33.9	
S° (J mol^{-1} K^{-1})	+65.7	+27.0	+240	+205

a Calculate the enthalpy change for this reaction.

$$\Delta H = [\text{Sum } \Delta_f H \text{ products}] - [\text{Sum } \Delta_f H \text{ reactants}]$$
$$= [2(-602) + 4(33.9)] - [2(-790)]$$
$$= +512 \text{ kJ mol}^{-1}$$

b Calculate the entropy change for this reaction.

$$\Delta S = [\text{Sum S products}] - [\text{Sum S reactants}]$$
$$= [2(27.0) + 4(240) + 205] - [2(65.7)]$$
$$= +1088 \text{ J mol}^{-1} \text{ K}^{-1}$$

c Is this reaction feasible at 298K? Explain your answer.

$$\Delta G = \Delta H - T\Delta S$$
$$= 512 - 298\left(\frac{1088}{1000}\right)$$
$$= +188 \text{ kJ mol}^{-1}$$

reaction is not feasible at 298 K as ΔG is positive

d Calculate the temperature at which the reaction becomes feasible.

when $\Delta G = 0$ $\quad \Delta H - T\Delta S = 0$

$$T = \frac{\Delta H}{\Delta S} = \frac{512}{\frac{1088}{1000}} = 471 \text{ K}$$

e Explain why the feasibility changes with temperature.

reaction is feasible when $\Delta G \leq 0$ where $\Delta G = \Delta H - T\Delta S$

as ΔH is positive and ΔS is positive, at low temperature ΔG is positive as ΔH is greater than $T\Delta S$

at higher temperatures, ΔG is negative as $T\Delta S$ is greater than ΔH