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- Nitrous oxide decomposes at 25 °C according to the following equation.

$$2N_2O(g) \rightleftharpoons 2N_2(g) + O_2(g) \qquad K_p = 1.8 \times 10^{36}$$

What is the value for K_p at 40 °C?

Using $\Delta_r G° = -RT\ln K_p$, at 25 °C (= 298 K):

$\Delta_r G° = -(8.314\ \text{J K}^{-1}\ \text{mol}^{-1}) \times (298\ \text{K}) \times \ln(1.8 \times 10^{36}) = -207\ \text{kJ mol}^{-1}$

Assuming that $\Delta_r G°$ does not change over the temperature range, at 40 °C = 313 K:

$$\ln K_p = \frac{\Delta_r G°}{RT} = \frac{(210 \times 10^3\ \text{J mol}^{-1})}{(8.314\ \text{J K}^{-1}\ \text{mol}^{-1})(298\ \text{K})} = 79$$

$K_p = 3.3 \times 10^{34}$

Answer: **3.3×10^{34}**

Is the reaction endothermic or exothermic? Give a reason for your answer.

The equilibrium constant decreases when the temperature increases. This is consistent with an exothermic reaction since these become less favourable as the temperature increases (Le Chatelier's principle).