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HL Answers to Activation energy questions

1. i. Since the units of the rate constant are $mol^{-1} dm^3 s^{-1}$ the overall order of the reaction must be two (i.e. second order).

ii. $\ln k = -E_a/RT + \ln A$

At 700 K: ln 1.30 = – (133.8 x 1000)/(8.314 x 700) + ln *A*

Hence ln *A* = ln 1.30 + (133.8 x 1000)/(8.314 x 700)

At T, ln 20.0 = – (133.8 x 1000)/(8.314 x T) + ln *A*

= – (133.8 x 1000)/(8.314 x T) + ln1.30 + (133.8 x 1000)/(8.314 x 700)

Hence ln 20.0 – ln1.30 = (133.8 x 1000)/(8.314) x (1/700 – 1/T)

2.7334 = 22.990 – 16093/T

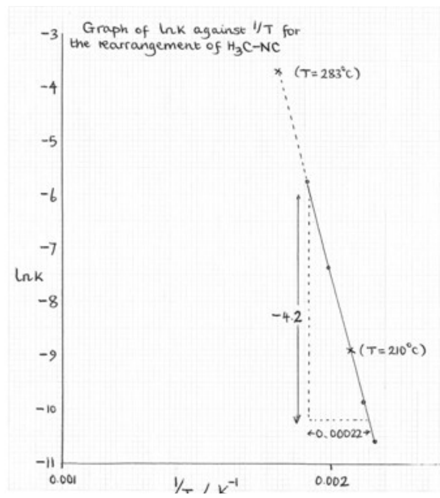
T = 16093/20.257 = **794 K (521 °C)**

2. i. Since the units of the rate constant are s^{-1} the overall order of the reaction must be one (i.e. first order).

ii. Rate = $k[H_3CNC]$

iii. The temperatures need to be converted into Kelvin then a graph of ln *k* against 1/*T* must be plotted. The gradient is equal to $-E_a/R$

1/Temperature / K^{-1}	**ln *k***
2.16×10^{-3}	– 10.59
2.12×10^{-3}	– 9.85
1.99×10^{-3}	– 7.37
1.91×10^{-3}	– 5.76



The gradient = – 4.2/0.00022 = – 19091 = $-E_a/R$

E_a = 19091 x 8.314 = 159000 J = **159 kJ mol^{-1}**

(Note that if the value for *A* was required the ln *k* axis would need to be extended so the line could be extrapolated to give the value of ln*k* when 1/*T* = zero. Alternatively the value for E_a can be put in the equation $\ln k = -E_a/RT + \ln A$ for one of the values of *T* and *A* can be calculated directly.)

iv. (i) By interpolating the graph for when *T*= 210 °C ln *k* = – 8.8 so *k* = **1.5×10^{-4} s^{-1}**.

(ii) By extrapolating the graph for when *T*= 283 °C ln *k* = – 3.7 so *k* = **2.5×10^{-2} s^{-1}**.