1. Which statement(s) explain(s) why reaction rates increase as temperature increases?

   ~~1~~ The activation energy is less. activation energy only changes with the addition of a catalyst

   2 Collisions between molecules are more frequent.

   3 A greater proportion of molecules have energy greater than the activation energy.

**A** 1, 2 and 3

**B** Only 1 and 2

**C** Only 2 and 3

**D** Only 1

Your answer C **[1]**

2. This question is about reaction rates.

Aqueous iron(III) ions, $Fe^{3+}$(aq), react with aqueous iodide ions, $I^-$(aq), as shown below.

$2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

A student carries out three experiments to investigate how different concentrations of $Fe^{3+}$(aq) and $I^-$(aq) affect the initial rate of this reaction. The results are shown below.

first order / second order

| Experiment | $[Fe^{3+}(aq)]$ / mol dm$^{-3}$ | $[I^-(aq)]$ / mol dm$^{-3}$ | Initial rate / mol dm$^{-3}$ s$^{-1}$ |
|---|---|---|---|
| 1 | $4.00 \times 10^{-2}$ | $3.00 \times 10^{-2}$ | $8.10 \times 10^{-4}$ |
| 2 ×1 | $8.00 \times 10^{-2}$ ×2 | $3.00 \times 10^{-2}$ ×1 | ×2 $1.62 \times 10^{-3}$ ×2 |
| 3 | $4.00 \times 10^{-2}$ | $6.00 \times 10^{-2}$ | $3.24 \times 10^{-3}$ ×4 |

**(a)*** Determine the rate constant and a possible two-step mechanism for this reaction that are consistent with these results. **[6]**

$\text{rate} = k[Fe^{3+}][I^-]^2$

$k = \frac{8.10 \times 10^{-4}}{(4 \times 10^{-2})(3 \times 10^{-2})^2} = 22.5 \text{ mol}^{-2}\text{dm}^6\text{s}^{-1}$

$\frac{\text{mol dm}^{-3}\text{ s}^{-1}}{(\text{mol dm}^{-3})^3} = \frac{\cancel{\text{mol dm}^{-3}}\text{s}^{-1}}{\text{mol}^{3}_{2}\text{ dm}^{-9\,6}}$

cancel out

Slow: $Fe^{3+}(aq) + 2I^-(aq) \rightarrow [FeI_2]^+$

fast: $[FeI_2]^+ + Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

overall equation:

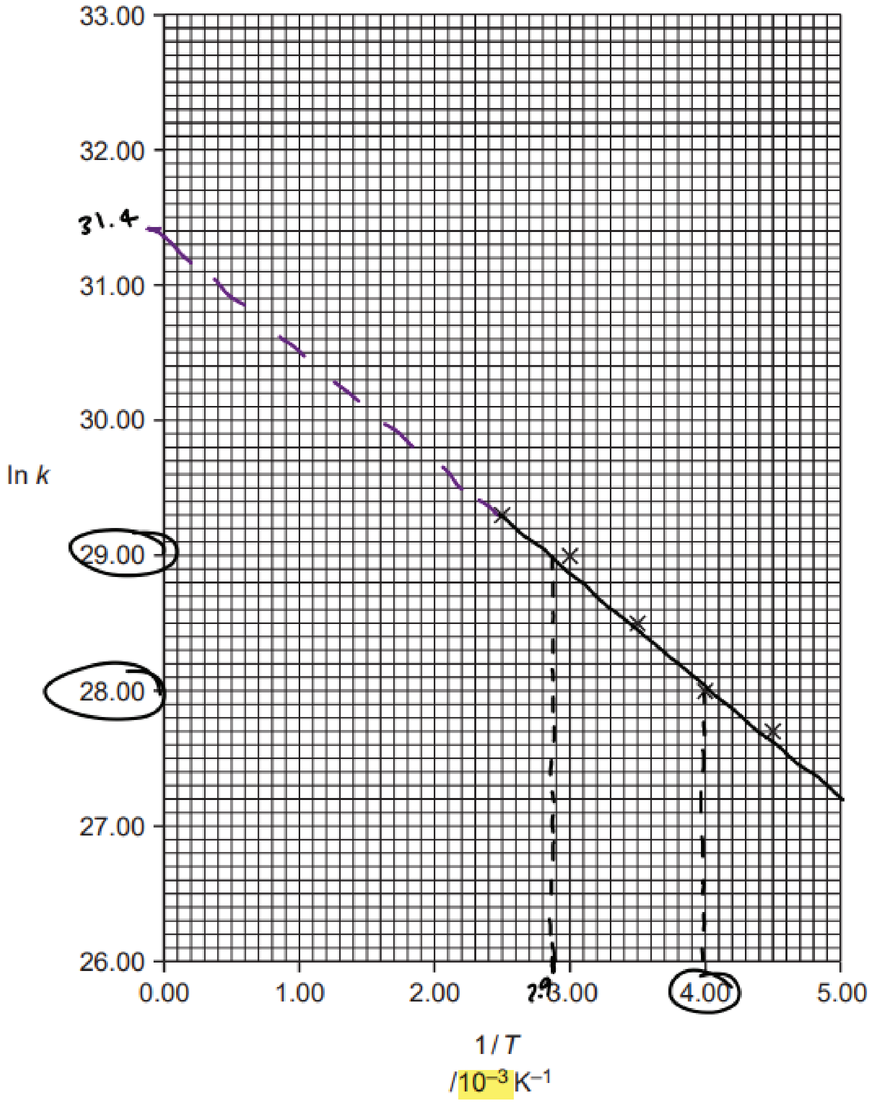
$2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

Additional answer space if required

**(b)** A student carries out an investigation to find the activation energy, $E_a$, and the pre-exponential factor, $A$, of a reaction.

$\ln k = -\frac{E_a}{RT} + \ln A$

The student determines the rate constant, $k$, at different temperatures, $T$.
The student then plots a graph of $\ln k$ against $1/T$ as shown below.



(i) Draw a best-fit straight line and calculate the activation energy, in $J mol^{-1}$.
Give your answer to **three** significant figures.

Show your working.

gradient: $\frac{29 - 28}{2.9\times10^{-3} - 4\times10^{-3}} = -909 = \frac{-Ea}{R}$

$Ea = +909 \times 8.314 = 7557.4$
$= 7560$ (3sf.)

activation energy, $E_a$ = + ......7560...... $J mol^{-1}$ **[3]**

(ii) Use the graph to calculate the value of the pre-exponential factor, *A*.

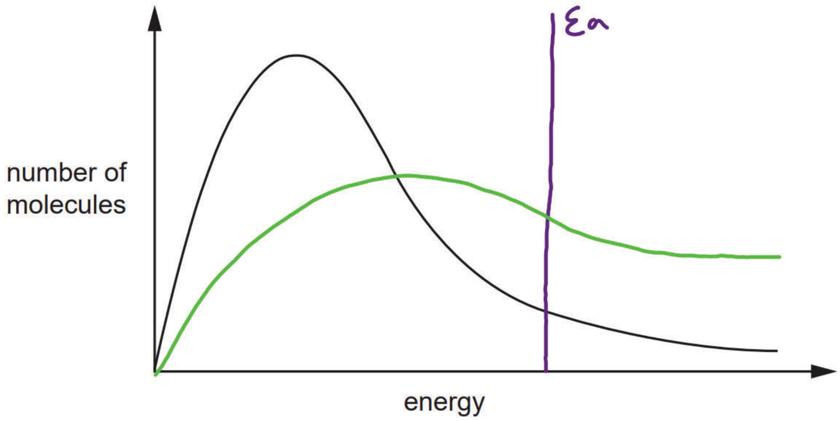
y intercept = ln A

Show your working.

$\ln A = 31.4$
$A = e^{31.4} = 4.33\times10^{13}$

pre-exponential factor, *A* = ......$4.33\times10^{13}$...... **[2]**

3. The diagram represents a Boltzmann distribution curve of molecules at a given temperature.




Which statement for this Boltzmann distribution curve is correct at a higher temperature?

**A** The peak increases in height and moves to the left.

**B** The peak increases in height and moves to the right.

**C** The peak decreases in height and moves to the left.

**D** The peak decreases in height and moves to the right.

Your answer D **[1]**