

STARTER FOR 10!!!

1.5. Arrhenius and rate

1. A reaction between A and B was found to be first order with respect to both.
- (a) Write the rate equation for the reaction.....(1 mark)
- (b) The rate constant k varies with temperature. Use the data together with the rate equation you have written in part (a) to complete the table below; (4 marks)

Temperature / K	[A] / mol dm ⁻³	[B] / mol dm ⁻³	Rate / mol dm ⁻³ s ⁻¹	k / mol ⁻¹ dm ³ s ⁻¹
288	0.015	0.030	2.70×10^{-8}	
313	0.015	0.015	2.16×10^{-7}	
338	0.030	0.030	9.18×10^{-6}	
363	0.030	0.015	3.52×10^{-5}	

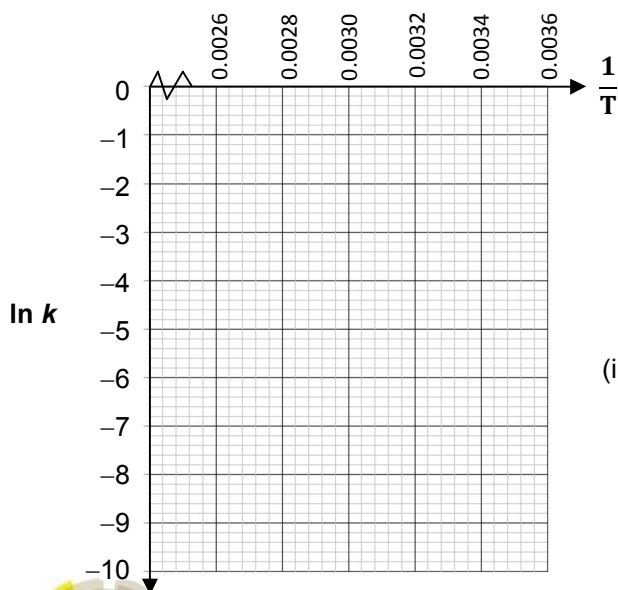
- (c) A scientist wishes to use this data to determine the activation energy for the reaction. He can do this using the Arrhenius equation;

$$k = A e^{-E_A/RT} \quad \text{where}$$

k is the rate constant
 E_A is the activation energy
 T is the temperature in Kelvin
 R is the gas constant ($8.31 \text{ J K}^{-1} \text{ mol}^{-1}$)
 A is a constant

A more useful form of the same equation is, $\ln k = \ln A - \frac{E_A}{R} \left(\frac{1}{T}\right)$. A graph of $\ln k$ vs $\frac{1}{T}$ is therefore a straight line with a gradient equal to $-\frac{E_A}{R}$.

- (i) Use the axes below to plot a graph of $\ln k$ vs $\frac{1}{T}$ for the data calculated in part (b). (4 marks)



- (ii) Therefore calculate the value of the activation energy, E_A , for this reaction.

.....
 (1 mark)



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1. Kinetics answers

1.3. Measuring reaction rate in the lab

- Change in concentration of sulfur = $(1 \times 10^{-4} \text{ mol} - 0 \text{ mol}) / 0.05 \text{ dm}^3 = 2 \times 10^{-3} \text{ mol dm}^{-3}$ (1 mark)
Initial rate of production of sulfur = $2 \times 10^{-3} \text{ mol dm}^{-3} / 56 \text{ s} = \underline{3.6 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}}$ (1 mark)
- Moles of iodine in reaction mixture at start = $0.002 \text{ dm}^3 \times 0.005 \text{ mol dm}^{-3} = 1 \times 10^{-5} \text{ mol}$ (1 mark)
Change in concentration of iodine = $(1 \times 10^{-5} \text{ mol} - 0 \text{ mol}) / 0.025 \text{ dm}^3 = 4 \times 10^{-4} \text{ mol dm}^{-3}$ (1 mark)
Initial rate of loss of iodine = $4 \times 10^{-4} \text{ mol dm}^{-3} / 279 \text{ s} = \underline{1.4 \times 10^{-6} \text{ mol dm}^{-3} \text{ s}^{-1}}$ (1 mark)
- Moles in 14 cm^3 of hydrogen at RTP = $0.014 \text{ dm}^3 / 24 \text{ dm}^3 = 5.8 \times 10^{-4} \text{ mol}$ (1 mark)
Moles of acid used up to produce this many moles of hydrogen = $5.8 \times 10^{-4} \text{ mol} \times 2 = 1.17 \times 10^{-3} \text{ mol}$ (1 mark)
Change in concentration of acid = $1.17 \times 10^{-3} \text{ mol} / 0.025 \text{ dm}^3 = 0.047 \text{ mol dm}^{-3}$ (1 mark)
Initial rate of loss of hydrochloric acid = $0.047 \text{ mol dm}^{-3} / 10 \text{ s} = \underline{0.0047 \text{ mol dm}^{-3} \text{ s}^{-1}}$ (1 mark)

1.4 Determining the rate equation

- Rate = $k[A]^2$; $k = 280 \text{ mol}^{-1} \text{ dm}^3 \text{ s}^{-1}$
(2 marks for the identification of the correct order wrt A and B, 1 mark for k with correct units)
- Rate = $k[B]$; $k = 2.5 \times 10^{-3} \text{ s}^{-1}$
(2 marks for the identification of the correct order wrt A and B, 1 mark for k with correct units)
- Rate = $k[X]^2[Y]$; $k = 2.08 \text{ mol}^{-2} \text{ dm}^6 \text{ s}^{-1}$
(3 marks for the identification of the correct order wrt X, Y and Z, 1 mark for k with correct units)

1.5 Arrhenius and rate

- (a) Rate = $k[A][B]$ (1 mark)

(b)

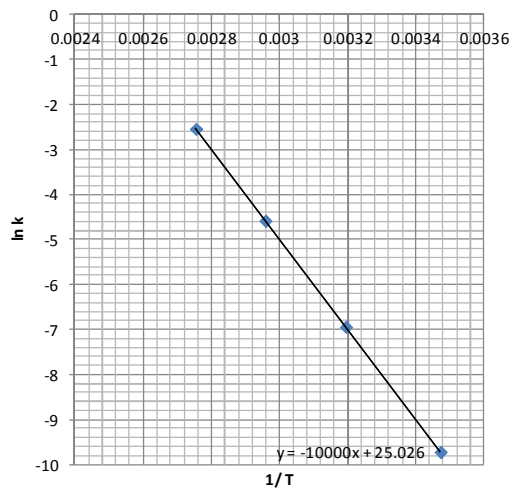
Temperature / K	[A] / mol dm ⁻³	[B] / mol dm ⁻³	Rate / mol dm ⁻³ s ⁻¹	k / mol ⁻¹ dm ³ s ⁻¹
288	0.015	0.030	2.70×10^{-8}	$\underline{6.0 \times 10^{-5}}$
313	0.015	0.015	2.16×10^{-7}	$\underline{9.6 \times 10^{-4}}$
338	0.030	0.030	9.18×10^{-6}	$\underline{0.010}$
363	0.030	0.015	3.52×10^{-5}	$\underline{0.078}$



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1. Kinetics answers

(c) (i)



1/T	ln K
0.00347	-9.72
0.00319	-6.95
0.00296	-4.59
0.00275	-2.55

(4 marks)

(ii) Gradient = - 10000, $\therefore \frac{E_A}{R} = 10000$. Therefore $E_A = \underline{83100 \text{ J or } 83.1 \text{ kJ}}$

(1 mark)