Name: ………………………………………. ( ) **Class:** …………….
Date: …………….

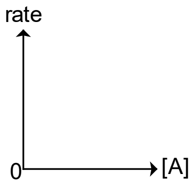
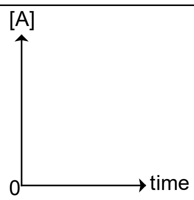
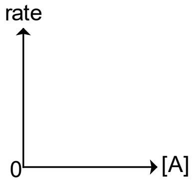
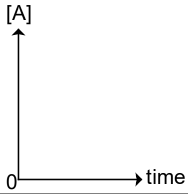
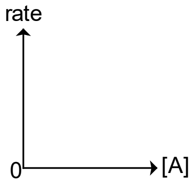
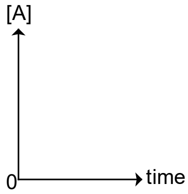
Raffles Institution
Year 6 H2 Chemistry 2025

Set A T2W3 – Kinetics

Pre-ChemFocus Activity

- Revise key concepts in Kinetics: half-life, “clock” and continuous experiments.
- To view infopack at IVY > C2025 – H2 CHEMISTRY > Pages > [ChemFocus] Y6 T2W3 Kinetics (and Y5 T3W9, if needed) > **FIRST VIDEO ONLY**

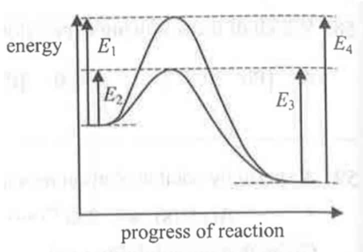
Notes:

overall order of reaction	rate equation	units of *k*	rate-[reactant] graph	[reactant]-time graph	half-life (circle)
zero	rate = *k*				constant / not constant
one	rate = *k*[A]				constant / not constant
two	rate = $k[A]^2$				constant / not constant

Self-Check Questions

Check your answers at IVY > C2025 – H2 CHEMISTRY > Pages > [ChemFocus] Y6 T2W3 Kinetics > Remaining videos

1 The diagram shows the energy profile of a reaction that occurs with and without a catalyst.



Which of the following can be deduced from the diagram?

A E_4 is the .

B The forward catalysed reaction is endothermic.

C The enthalpy change of reaction is (E_2 – E_3).

D The enthalpy change of reaction is decreased by using a catalyst.

2 Which changes would cause **both** of the rate constants k_1 and k_{-1} to be increased?

$$\textbf{reactants} \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} \textbf{products}$$

1 introducing a catalyst

2 heating the equilibrium mixture

3 increasing the concentrations of the reactants

A **1** , **2** and **3**

B **1** and **2** only

C **2** and **3** only

D **1** only

3 A chemical plant illegally dumped some radioactive waste in a landfill. This waste contains two radioactive isotopes **M** and **N**. The half-life of **M** is 3 days whereas that of **N** is 2 days. The authorities found out about this illegal dumping only when the waste had been in the landfill for 12 days. They did an immediate analysis on a sample of the waste and found equal amounts of **M** and **N**.

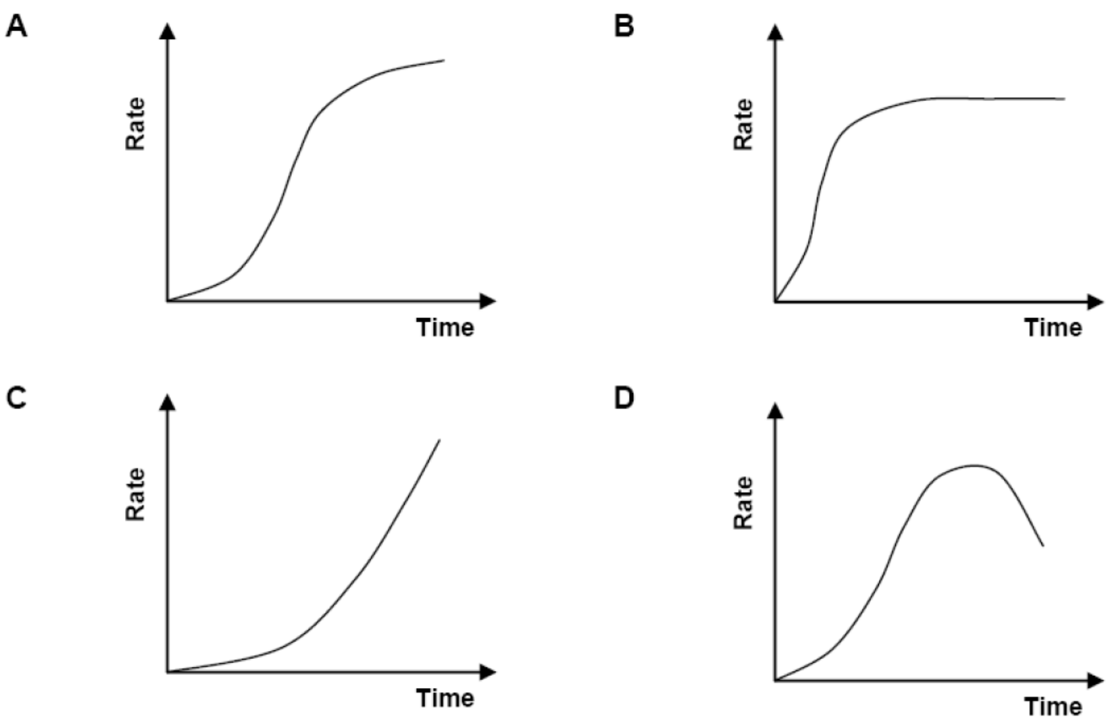
Considering that the decay of radioactive isotopes follows first-order kinetics, what is the molar ratio of **M** to **N** if the waste had been in the landfill for 6 days?

	M	:	N
A	1	:	4
B	1	:	2
C	2	:	1
D	4	:	1

4 The reaction between potassium manganate(VII) and ethanedioic acid is an example of auto-catalytic reactions, in which one of the products catalyses the reaction.

$$2MnO_4^-(aq) + 5H_2C_2O_4(aq) + 6H_3O^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 14H_2O(l)$$

Which graph correctly represents the kinetics of this reaction?



Tutorial Discussion Questions (to complete before Chemfocus session)

1a **[modified NYJC 2018/I/9]**

The decomposition of hydrogen peroxide follows first order kinetics.

$$2H_2O_2 \longrightarrow 2H_2O + O_2$$

A certain solution of hydrogen peroxide undergoes complete decomposition to liberate 96 cm^3 of oxygen gas. It is found that at 20 °C, 48 cm^3 of oxygen was collected in 35 min.

Use the following equation, how long will it take for 80 cm^3 of the gas to be produced?

$$\frac{(H_2O_2)_t}{(H_2O_2)_0} = \left(\frac{1}{2}\right)^n$$

where: $(H_2O_2)_0$ is the amount of H_2O_2 when time = 0,
$(H_2O_2)_t$ is the amount of H_2O_2 left at time, t,
n is the number of half-lives elapsed.

A 87.5 min

B 90.5 min

C 97.5 min

D 105 min

2a [N2013/3/5(c),(d)]

The Harcourt and Esson reaction is that between hydrogen peroxide and acidified potassium iodide. The rate of reaction can be followed by measuring the amount of iodine produced after various times, from which the concentration of H_2O_2 remaining can be calculated.

The following reaction mixture was prepared for experiment 1.

initial $[H^+]$ = 0.200 mol dm^{-3}
initial $[I^-]$ = 0.200mol dm^{-3}
initial $[H_2O_2]$ = 0.0200 mol dm^{-3}

The following table shows $[H_2O_2]$ at various times.

time /s	$[H_2O_2]$ /mol dm^{-3}
0	0.0200
80	0.0167
183	0.0135
315	0.0103
490	0.0071
760	0.0039

(i) Plot these data on suitable axes below.

[2]

(ii) Use your graph to determine

(1) the order of reaction with respect to $[H_2O_2]$,
(2) the initial rate, in mol dm^{-3} s^{-1},

showing all your working and drawing clearly any construction lines on your graph.

[4]

Further experiments were carried out changing $[H^+]$ and $[I^-]$, but keeping the initial $[H_2O_2]$ the same as before. The following results were obtained.

experiment	initial $[H^+]$ /mol dm^{-3}	initial $[I^-]$ /mol dm^{-3}	initial rate /mol dm^{-3} s^{-1}
2	0.400	0.200	8.4×10^{-5}
3	0.300	0.200	6.3×10^{-5}
4	0.200	0.100	2.1×10^{-5}

(iii) Determine the orders with respect to $[H^+]$ and $[I^-]$. Explain your reasoning.

[2]

(iv) Hence, write the rate equation for the reaction, and calculate a value for the rate constant. Include units in your answer.

[2]

<u>Tutorial Practice Questions</u> (to be attempted **<u>IN CLASS</u>**)

1b **[N2020/1/13]**

Hydrogen peroxide solution decomposes. The equation for this reaction is shown.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

A 200 cm^3 sample of hydrogen peroxide solution is warmed. After 120 minutes, 6.00 dm^3 of oxygen gas, measured at r.t.p., is collected. Under these conditions, the reaction has a constant half-life of 40 minutes.

What is the initial concentration of the hydrogen peroxide solution?

A 0.57 mol dm^{-3}

B 1.4 mol dm^{-3}

C 2.5 mol dm^{-3}

D 2.9 mol dm^{-3}

2b [RI 2022 Prelim P2 Q2(b)]

Nitrogen dioxide undergoes the following gas phase reaction with hydrogen chloride, as shown in equation 1.1.

equation 1.1 $2NO_2(g) + HCl(g) \longrightarrow NOCl(g) + HNO_3(g)$

To study the kinetics for this reaction, three separate experiments were carried out in a vessel of fixed volume at a constant temperature of 500 K. The initial concentrations of NO_2 and HCl are shown in Table 1.1.

Table 1.1

experiment	initial $[NO_2]$ / mol dm^{-3}	initial $[HCl]$ / mol dm^{-3}
1	1.00	0.05
2	0.50	0.05
3	0.50	0.01

To monitor the progress of the reaction for each experiment, the total pressure in the vessel was measured at regular time intervals and the corresponding concentrations of $NOCl$ were calculated. Fig. 1.1 shows how the concentration of $NOCl$ varies with time for experiments 1 and 2.

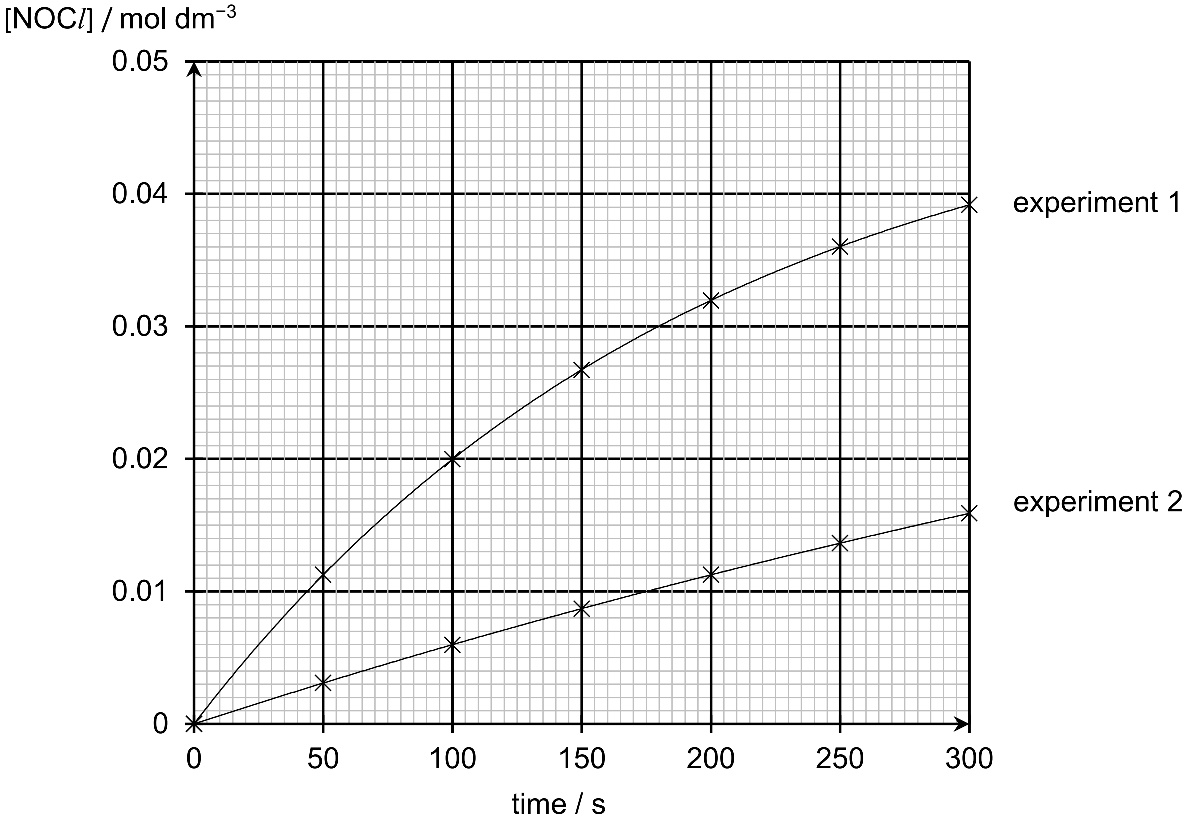


Fig. 1.1

(i) Using Fig. 1.1, determine the initial rate of reaction for experiments 1 and 2, showing your working clearly on the graph.

Hence, deduce the order of reaction with respect to NO_2.

[4]

(ii) State the final concentration of $NOCl$ in the vessel for experiment 1 if the reaction were to proceed to completion.

……………………………………………………………………………………….[1]

(iii) Hence, determine the order of reaction with respect to HCl.

Show your working, including construction lines on Fig. 1.1.

[2]

(iv) Using your answer in **(iii)**, predict the half-lives of experiments 2 and 3, giving your reasoning.

[2]

Self-Check Answers

1 **Answer: C**.

Option **A** is wrong because E4 is the activation energy for the reverse uncatalysed reaction since the activation energy of the forward reaction is higher.

Option **B** is wrong because in the forward reaction, the products have lower energy than the reactants. Hence, the forward reaction is exothermic.

Option **D** is wrong because a catalyst has no effect on the enthalpy change of reaction.

2 **Answer: B**.

Option 1 is correct because a catalyst lowers the activation energy of both the forward and reverse reaction by the same amount, causes both k_1 and k_{-1} to increase. ($k = Ae^{\frac{-E_a}{RT}}$)

Option 2 is correct because heating the equilibrium mixture would increase the rate of both forward and reverse reaction, causing both k_1 and k_{-1} to increase.

Option 3 is wrong because concentration has no effect on the rate constant of the reaction.

3 **Answer: B**

After 12 days, there were equal amounts of **M** and **N**. We can assume that there is 1 mol of **M** and 1 mol of **N** after **12** days.

M has a half-life of 3 days. 12 days would have allowed for 4 half-lives to pass.
The initial amount of **M** would be $2^4(1)$ = 16 mol
4 half-lives : 16 mol $\longrightarrow$ 8 mol $\longrightarrow$ 4 mol $\longrightarrow$ 2 mol $\longrightarrow$ 1 mol

N has a half-life of 2 days. 12 days would have allowed for 6 half-lives to pass.
The initial amount of **N** would be $2^6(1)$ = **64 mol**
6 half-lives : **64 mol** $\rightarrow$ 32 mol $\rightarrow$ 16 mol $\rightarrow$ 8 mol $\rightarrow$ 4 mol $\rightarrow$ 2 mol $\rightarrow$ 1 mol

What is the mole ratio of **M** to **N** after 6 days?
For **M**, in 6 days, 2 half-lives would have passed i.e. 16 mol $\longrightarrow$ 8 mol $\longrightarrow$ **4 mol**
For **N**, in 6 days, 3 half-lives would have passed i.e.
64 mol $\rightarrow$ 32 mol $\rightarrow$ 16 mol $\rightarrow$ **8 mol**
Final mole ratio **M** to **N** = 4 : 8 = **1** : **2**

4 **Answer: D**

This is an auto-catalytic reaction where Mn^{2+} produced in the reaction acts catalyses further reaction between MnO_4^- and $C_2O_4^{2-}$. All four options involve rate-time graphs, which allow you to read the rate as the y-values.

Initially, at time = 0, the rate is slow and near zero as it has a high activation energy. (All options show near zero rate at time = 0)

The reaction rate then increases as the Mn^{2+} ions produced act as the autocatalyst, providing an alternative pathway with lower activation energy. (All options show an increase in rate as time goes by)

Towards the end of reaction, the concentration of reactants falls to a low level, the rate of reaction will decrease even though there is an adequate supply of catalyst. (**Option D** shows the rate decreasing towards the end of reaction).

<u>Set A Tutorial Discussion Questions</u> (to complete before Chemfocus session)

1a Answer: B

Since O_2 is the product and volume of O_2 obtained upon complete decomposition = 96 cm^3.

		$t_{1/2}$	
Volume of O_2 formed / cm^3	0	$\longrightarrow$	½(96) = **48**

Since 48 cm^3 of O_2 was collected in 35 min, the half life = 35 min.

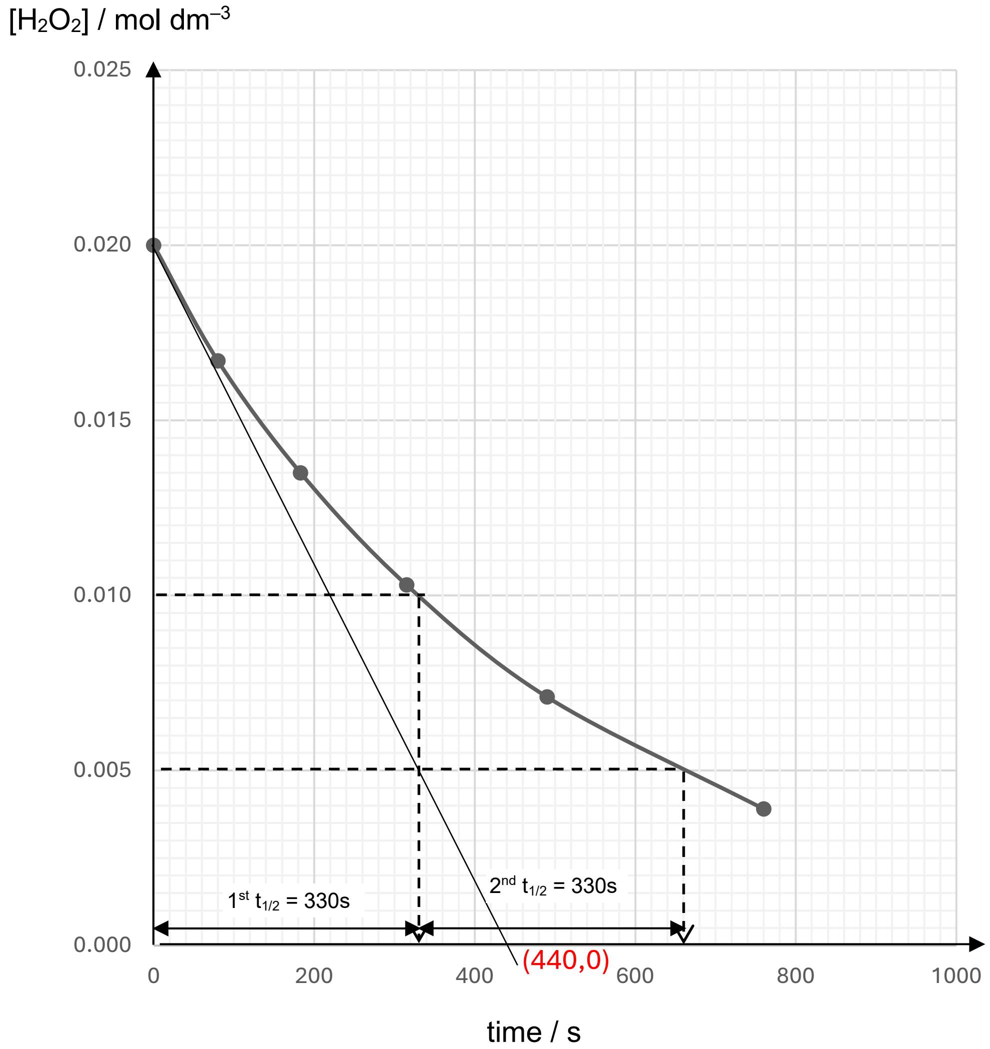
Total amount of O_2 evolved after complete decomposition = 96/24000 = 0.00400 mol
Since $2H_2O_2 \longrightarrow 2H_2O + O_2$, initial amount of H_2O_2 present = 2(0.00400) = 0.00800 mol

Amount of O_2 in 80 cm^3 of O_2 = 80/24000 = 0.00333 mol
Amount of H_2O_2 reacted to produce 80 cm^3 of O_2 = 2(0.00333) = 0.00666 mol
Amount of H_2O_2 remaining after producing 80 cm^3 of O_2 = 0.00800 – 0.00666 = 0.00134 mol

$0.00134 = (\frac{1}{2})^n(0.00800)$
$(\frac{1}{2})^n = 0.1675$
$n\lg(0.5) = \lg(0.1675)$
$n = \lg(0.1675)/\lg(0.5) = 2.578$

time for 80 cm^3 of O_2 to be produced = 2.578(35) = 90.5 min.

2a **(i)**



(ii) (1)

From the graph, 1st $t_{½}$ = 2nd $t_{½}$ = 330 s

Hence, the reaction is first order wrt H_2O_2.

(2)

Initial rate = –gradient of tangent to graph at time = 0

= 0.02 / 440 = 4.55×10^{-5} mol dm^{-3} s^{-1}

(iii) Comparing experiments 2 and 3.

When initial $[H^+]$ x 1.33 , initial rate x 1.33.

$\Rightarrow$ rate $\propto [H^+]$

$\Rightarrow$ Order of reaction with respect to H^+ = 1

Let rate = $k[H_2O_2][I^-]^a[H^+]$, where *a* is the constant to be determined

From experiments 2 and 4,

$$\frac{\text{initial rate of experiment 2}}{\text{initial rate of experiment 4}} = \frac{8.4 \times 10^{-5}}{2.1 \times 10^{-5}} = \frac{k(0.400)(0.20)^a[H_2O_2]}{k(0.200)(0.10)^a[H_2O_2]}$$

$\Rightarrow 2 = 2^a$

$\Rightarrow$ order of reaction with respect to I^-, a = 1

(iv) rate = $k[H_2O_2][I^-][H^+]$

Using experimental data from 1st experiment,

$8.4 \times 10^{-5} = k\ (0.0200)(0.400)(0.200)$

k = $\underline{0.0525\ \text{mol}^{-2}\ \text{dm}^6\ \text{s}^{-1}}$

Set A Tutorial Practice Questions (to be attempted **IN CLASS**)

1b **Answer: D**

Since $t_{½}$ = 40 min, 120 min = 3 $t_{½}$

At r.t.p.,

$n(O_2)$ formed at 120 min = 6.00 / 24 = 0.250 mol

mole ratio of $H_2O_2 : O_2$ = 2 : 1

Let the initial amount of H_2O_2 be x mol.

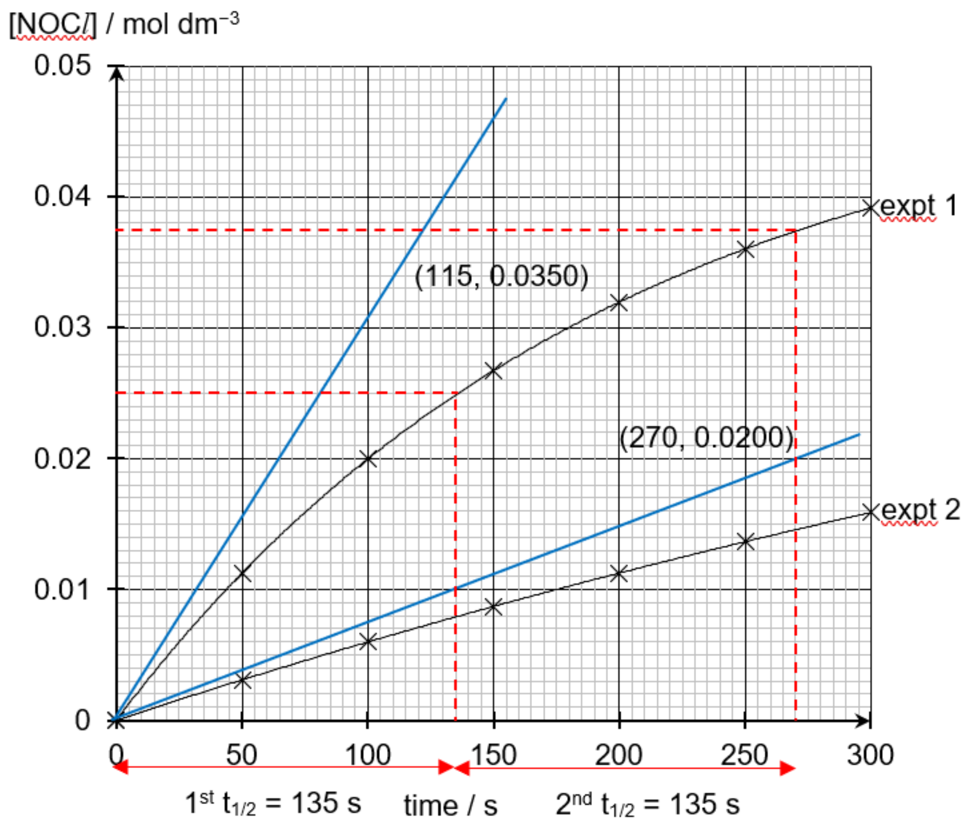
	amt of H_2O_2 / mol	amt of O_2 / mol
t = 0 min	x	0
t = 40 min	$\frac{1}{2}x$	$\frac{1}{2}(x - \frac{1}{2}x)$
t = 80 min	$\frac{1}{4}x$	$\frac{1}{2}(x - \frac{1}{4}x)$
t = 120 min	$\frac{1}{8}x$	$\frac{1}{2}(x - \frac{1}{8}x)$ = 0.250

x = 0.5714 mol

Initial concentration of H_2O_2 = 0.5714 / 200 × 1000

= 2.9 mol dm^{-3}

2b (i)



initial rate of reaction for experiment 1 = $\dfrac{0.0350 - 0}{115 - 0}$ = 3.04×10^{-4} mol dm^{-3} s^{-1}

initial rate of reaction for experiment 2 = $\dfrac{0.0200 - 0}{270 - 0}$ = 7.41×10^{-5} mol dm^{-3} s^{-1}

$\dfrac{\text{initial rate of expt 1}}{\text{initial rate of expt 2}} = \dfrac{3.04 \times 10^{-4}}{7.41 \times 10^{-5}} = 4.10 \approx \underline{4}$ (to nearest integer)

Comparing experiments 1 and 2, when [NO_2] $\times$ 2, rate $\times$ 4.
$\Rightarrow$ rate $\propto$ [NO_2]^2,
$\Rightarrow$ reaction is <u>second order with respect to NO_2.</u>

(ii) Final [NOC*l*] = <u>0.0500 mol dm^{-3}</u>

(iii) For experiment 1, final [NOC*l*] = 0.05 mol dm^{-3}.

1st $t_{½}$ $\Rightarrow$ time taken for [HC*l*] to decrease from 0.05 mol dm^{-3} to 0.025 mol dm^{-3}
$\Rightarrow$ time taken for [NOC*l*] to increase from 0 mol dm^{-3} to 0.025 mol dm^{-3}
$\Rightarrow$ 135 s

2nd $t_{½}$ $\Rightarrow$ time taken for [HC*l*] to decrease from 0.025 mol dm^{-3} to 0.0125 mol dm^{-3}

$\Rightarrow$ time taken for [NOC*l*] to increase from 0.025 mol dm^{-3} to 0.0375 mol dm^{-3}

$\Rightarrow$ 135 s

Since <u>$t_{½}$ is constant, reaction is first order with respect to HC*l*</u>.

(iv) $t_{\frac{1}{2}} = \dfrac{\ln 2}{k'}$

where $k' = k[NO_2]^2$

$135 = \dfrac{\ln 2}{k(1)^2}$

$135 = \frac{\ln 2}{k}$

Since $[NO_2]$ is the same for experiments 2 and 3,

$t_{\frac{1}{2}} = 135\left(\dfrac{1}{0.5^2}\right)$

$= 540$ s