

- 1) Suggest experimental means by which the following reactions can be monitored for rate.



monitor gas p.

- 2) Iron (III) chloride is reduced by tin (II) chloride in the following reaction.



The concentration of Fe^{3+} ion at the beginning of an experiment was 0.03586M and after 4.00 min it was 0.02715M. What is the average rate of reaction of FeCl_3 in this time? What is the rate of formation of tin (IV) chloride? (ans: (E) 0.00218M/min; 0.00109M/min)

$$\frac{(0.0356 - 0.02715) \text{ M}}{(4 - 0) \text{ min}} = \boxed{0.00218 \text{ M/min}}$$

$$\frac{1 \text{ SnCl}_4}{2 \text{ FeCl}_3} \times 0.00218 \text{ M/min} = \boxed{0.00109 \text{ M/min}}$$

- 3) The rate of the following reaction in aqueous solution is monitored by measuring the rate of formation of I_3^- . Data obtained are listed in the table below.



Exp	$[\text{S}_2\text{O}_8^{2-}]$, M	$[\text{I}^-]$, M	Initial rate M/s
1	0.038	0.060	1.4×10^{-5}
2	0.076	0.060	2.8×10^{-5}
3	0.076	0.120	5.6×10^{-5}

- a) Determine the order of the reaction with respect to $\text{S}_2\text{O}_8^{2-}$, and with respect to I^- and overall.

(ans: 1, 1, 2)

- b) What is the value of rate constant, k? (ans: $6.1 \times 10^{-3} \text{ M}^{-1} \text{ s}^{-1}$)

- c) What would be the initial rate of reaction if $[\text{S}_2\text{O}_8^{2-}] = 0.083 \text{ M}$ and $[\text{I}^-] = 0.115 \text{ M}$? (ans: $5.8 \times 10^{-5} \text{ M/s}$)

$$\text{(a)} \quad \frac{\text{exp 2}}{\text{exp 1}} \quad \frac{0.076}{0.038} = \frac{2.8 \times 10^{-5}}{1.4 \times 10^{-5}} \quad \left| \quad \frac{\text{exp 3}}{\text{exp 2}} \quad \frac{0.12}{0.06} = \frac{5.6 \times 10^{-5}}{2.8 \times 10^{-5}} \right.$$

$$[\text{S}_2\text{O}_8^{2-}] \quad 2 = 2 \quad \left. \quad [\text{I}^-] \quad 2 = 2 \right.$$

$$\text{Rate} = k [\text{S}_2\text{O}_8^{2-}] [\text{I}^-] \quad ; \quad \boxed{\text{overall order} = 2}$$

$$\text{(b)} \quad k = \frac{\text{Rate}}{[\text{S}_2\text{O}_8^{2-}] [\text{I}^-]} = \frac{1.4 \times 10^{-5} \text{ M/s}}{0.038 \text{ M} \times 0.06 \text{ M}} = \boxed{6.14 \times 10^{-3} / \text{M s}}$$

$$\text{(c)} \quad \text{Rate} = 6.14 \times 10^{-3} / \text{M s} \times 0.083 \text{ M} \times 0.115 \text{ M} = \boxed{5.86 \times 10^{-5} \text{ M/s}}$$

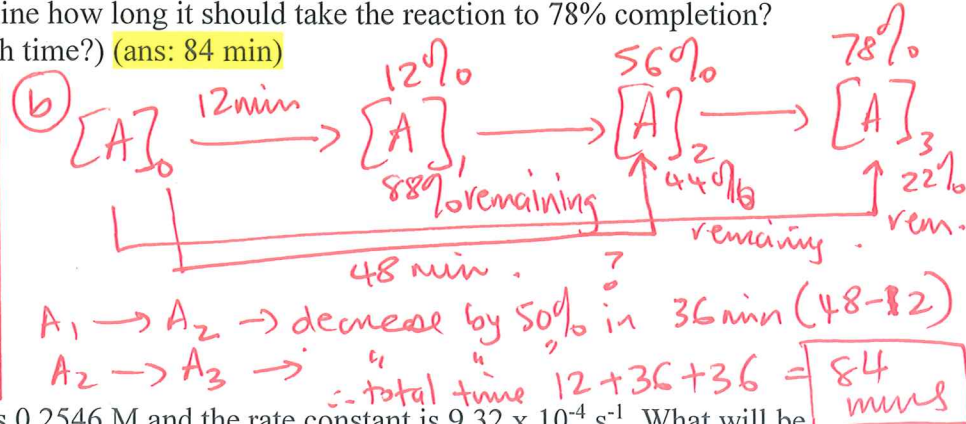
4) A first order A \longrightarrow products, has a rate constant k, of 0.0462 min^{-1} .

a) What is the [A] at the time when the reaction is proceeding at a rate of $0.0150 \text{ M min}^{-1}$? (ans: 0.325 M)

b) It takes 12 mins for the reaction to 12% completion and 48 min to 56% completion. Without doing detailed calculations, determine how long it should take the reaction to 78% completion?

(Hint: what is the % remaining each time?) (ans: 84 min)

$$\begin{aligned} \text{(a) Rate} &= k[A] \\ [A] &= \frac{\text{Rate}}{k} \\ &= \frac{0.0150 \text{ M/min}}{0.0462} \\ &= \boxed{0.325 \text{ M}} \end{aligned}$$



5) The initial concentration of H_2O_2 is 0.2546 M and the rate constant is $9.32 \times 10^{-4} \text{ s}^{-1}$. What will be the $[\text{H}_2\text{O}_2]$ at $t = 35 \text{ s}$? (ans: 0.246 M) (1st order)

$$\begin{aligned} \ln \frac{[\text{H}_2\text{O}_2]_t}{[\text{H}_2\text{O}_2]_0} &= -kt \\ \ln \frac{0.2546}{0.2546} &= -9.32 \times 10^{-4} / \text{s} \times 35 \text{ s} \\ \text{inv. log both sides} &\rightarrow \rightarrow \boxed{0.03262} \end{aligned}$$

$$\begin{aligned} \frac{[\text{H}_2\text{O}_2]_t}{0.2546} &= 0.9679 \\ [\text{H}_2\text{O}_2]_t &= 0.9679 \times 0.2546 \\ &= \boxed{0.246 \text{ M}} \end{aligned}$$

6) The thermal decomposition of phosphine (PH_3) into phosphorous and molecular hydrogen is a first order reaction.



The half life of the reaction is 35.0 s at 680°C . Calculate

a) The first order rate constant for the reaction. ((B)ans: $0.0198/\text{s}$)

b) The time required for 95% of the phosphine to decompose. ((B)ans: 151 s)

$$\begin{aligned} \text{(a)} \quad k &= \frac{0.693}{t_{1/2}} \\ &= \frac{0.693}{35 \text{ s}} \\ &= \boxed{0.0198 / \text{s}} \end{aligned}$$

$$\begin{aligned} \text{(b) phosphine left} &= 5\% \\ \frac{[A]_t}{[A]_0} &= \frac{5\%}{100\%} = \frac{0.05}{1} \\ \ln \frac{[A]_t}{[A]_0} &= -kt \\ \ln \frac{0.05}{1} &= -0.0198 \times t \\ -2.996 &= -0.0198 t \\ \frac{-2.996}{-0.0198} &= t = \boxed{151 \text{ s}} \end{aligned}$$

- 7) Given the same concentrations, the reaction at 250°C, is 1.50×10^3 times faster than the same reaction at 150°C. Calculate the activation energy for this reaction. ((B)ans: 135KJ/mol)



$$T_2 = 250 + 273 = 523\text{K} ; T_1 = 150 + 273 = 423\text{K}$$

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\ln 1.5 \times 10^3 = \frac{E_a}{8.314\text{J/molK}} \times \left(\frac{1}{423} - \frac{1}{523} \right)$$

$$9.31 = \frac{E_a}{8.314\text{J/molK}} \times 4.52 \times 10^{-4}/\text{K}$$

$$E_a = \frac{7.31 \times 8.314\text{J/molK}}{4.52 \times 10^{-4}/\text{K}} = 1.35 \times 10^5\text{J/mol}$$

OR $\boxed{135\text{KJ/mol}}$