JEE Mains 2019 Chapter wise Question Bank

Chemical Kinetics - Questions

Q1

The following results were obtained during kinetic studies of the reaction;

$2A + B \rightarrow$ Products

Experiment	[A] (in mol L^{-1})	[B] (in mol L^{-1})	Initial Rate of reaction (in mol L^{-1} min^{-1})
I	0.10	0.20	6.93×10^{-3}
II	0.10	0.25	6.93×10^{-3}
III	0.20	0.30	1.386×10^{-2}

The time (in minutes) required to consume half of A is:

(1) 5 (2) 10

(3) 1 (4) 100

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Q2

For the reaction, $2A + B \rightarrow$ products, when the concentrations of A and B both were doubled, the rate of the reaction increased from 0.3 mol L^{-1} s^{-1} to 2.4 mol $L^{-1}s^{-1}$. When the concentration of A alone is doubled, the rate increased from 0.3 and $L^{-1}s^{-1}$ to 0.6 mol $L^{-1}s^{-1}$. Which one of the following statements is correct?

(1) Total order of the reaction is 4

(2) Order of the reaction with respect to B is 2

(3) Order of the reaction with respect to B is 1

(4) Order of the reaction with respect to A is 2

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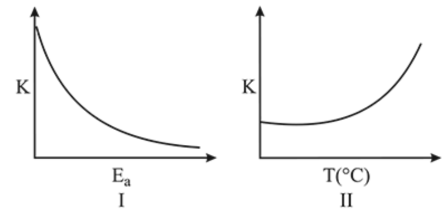
Q3

Consider the given plots for a reaction obeying Arrhenius equation (0 °C < T < 300 °C) : (K and E_a are rate constant and activation energy, respectively)



Choose the correct option:

(1) I is right but II is wrong

(2) Both I and II are correct

(3) I is wrong but II is right

(4) Both I and II are wrong

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Q4

For an elementary chemical reaction, $A_2 \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} 2A$, the expression for $\frac{d[A]}{dt}$ is:

(1) $k_1[A_2] - k_{-1}[A]^2$ (2) $2k_1[A_2] - k_{-1}[A]^2$

(3) $k_1[A_2] + k_{-1}[A]^2$ (4) $2k_1[A_2] - 2k_{-1}[A]^2$

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Q5

If a reaction follows the Arrhenius equation, the plot lnk vs $\frac{1}{(RT)}$ gives straight line with a gradient (– y) unit. The energy required to activate the reactant is:

(1) y/R unit (2) y unit

(3) yR unit (4) – y unit

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Q6

The reaction $2X \rightarrow B$ is a zeroth order reaction. If the initial concentration of X is 0.2 M, the half-life is 6 h. When the initial concentration of X is 0.5 M, the time required to reach its final concentration of 0.2 M will be :

(1) 9.0 h (2) 12.0 h
(3) 18.0 h (4) 7.2 h

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Q7

Decomposition of X exhibits a rate constant of 0.05 μg/year. How many years are required for the decomposition of 5 μg of X into 2.5 μg?

(1) 50 (2) 25
(3) 20 (4) 40

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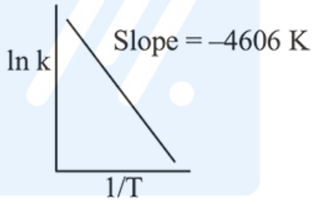
Q8

For a reaction, consider the plot of ln k versus 1/T given in the figure. If the rate constant of this reaction at 400 K is $10^{-5}\ s^{-1}$, then the rate constant at 500 K is :



(1) $10^{-6}\ s^{-1}$ (2) $2 \times 10^{-4}\ s^{-1}$
(3) $10^{-4}\ s^{-1}$ (4) $4 \times 10^{-4}\ s^{-2}$

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Q9

For the reaction $2A + B \rightarrow C$, the values of initial rate at different reactant concentrations are given in the table below.
The rate law for the reaction is :

[A] (mol L^{-1})	[B] (mol L^{-1})	Initial Rate (mol $L^{-1}\ s^{-1}$)
0.05	0.05	0.045
0.10	0.05	0.090
0.20	0.10	0.72

(1) Rate = $k[A][B]^2$ (2) Rate = $k[A]^2[B]^2$
(3) Rate = $k[A][B]$ (4) Rate = $k[A]^2[B]$

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Q10

For a reaction scheme $A \xrightarrow{k_1} B \xrightarrow{k_2} C$, if the rate of formation of B is set to be zero then the concentration of B is given by :

(1) $(k_1 - k_2)[A]$ (2) $k_1 k_2 [A]$
(3) $(k_1 + k_2)[A]$ (4) $\left(\frac{k_1}{k_2}\right)[A]$

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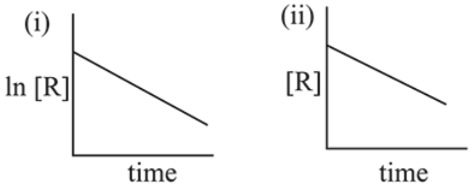
Q11

The given plots represents the variation of the concentration of a reactant R with time for two different reactions (i) and (ii). The respective orders of the reactions are:



(1) 1, 0 (2) 1, 1 (3) 0, 1 (4) 0, 2

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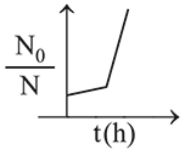
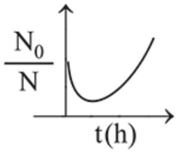
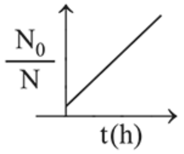
Q12

A bacterial infection in an internal wound grows as $N'(t) = N_0 \exp(t)$, where the time t is in hours. A dose of antibiotic, taken orally, needs 1 hour to reach the wound. Once it reaches there, the bacterial population goes down as $\frac{dN}{dt} = -5N^2$. What will be the plot of $\frac{N_0}{N}$ vs. t after 1 hour?

(1)


(2)


(3)


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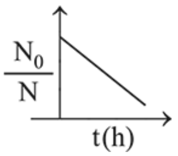

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Q13

For the reaction of H_2 with I_2, the rate constant is $2.5 \times 10^{-4}\ dm^3\ mol^{-1}\ s^{-1}$ at 327 °C and $1.0\ dm^3\ mol^{-1}\ s^{-1}$ at 527 °C. The activation energy for the reaction, in kJ mol^{-1} is :
($R = 8.314 J\ K^{-1}\ mol^{-1}$)

(1) 166 (2) 150 (3) 72 (4) 59

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Q14

In the following reaction : $xA \rightarrow yB$

$$\log_{10}\left[-\frac{d[A]}{dt}\right] = \log_{10}\left[\frac{d[B]}{dt}\right] + 0.3010$$

'A' and 'B' respectively can be:

(1) *n*-Butane and Iso-butane

(2) C_2H_2 and C_6H_6

(3) C_2H_4 and C_4H_8

(4) N_2O_4 and NO_2

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Q15

NO_2 required for a reaction is produced by the decomposition of N_2O_5 in CCl_4 as per the equation, $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$.

The initial concentration of N_2O_5 is 3.00 mol L^{-1} and it is 2.75 mol L^{-1} after 30 minutes. The rate of formation of NO_2 is :

(1) 4.167×10^{-3} mol L^{-1} min^{-1}

(2) 1.667×10^{-2} mol L^{-1} min^{-1}

(3) 8.333×10^{-3} mol L^{-1} min^{-1}

(4) 2.083×10^{-3} mol L^{-1} min^{-1}

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Chemical Kinetics - Answers

Q1

(1) From experiment 1 and II, it is observed that order of reaction w.r.t. (3) is zero.

From experiment II and III, α can be calculated as:

$$\frac{1.386\times10^{-2}}{6.93\times10^{-3}}=\left(\frac{0.2}{0.1}\right)^{\alpha}$$

$\therefore\quad \alpha=1$

Now, $\quad \text{Rate}=K[A]^1$

or, $\quad 6.93\times10^{-3}=K(0.1)$

$K=6.93\times10^{-2}$

For the reaction, $2A+B\rightarrow$ Products

$$2Kt=\ln\frac{[A]_0}{[A]}$$

$$\therefore\quad t_{1/2}=\frac{0.693}{2K}=\frac{0.693}{0.693\times10^{-2}\times2}$$

$t_{1/2}=5$

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Q2

(2) $r=K[A]^x[B]^y$

$$\frac{r_2}{r_1}=2^x.2^y=8\Rightarrow x+y=3$$

$$\frac{r_3}{r_1}=2x=2\Rightarrow x=1$$

$\therefore\quad y=2$

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Q3

(2) From Arrhenius equation,

$$K=Ae^{-E_a/RT}$$

So, as E_a increases, $e^{-E_a/RT}$ decreases, K decreases

and as T increases, $\frac{E_a}{RT}$ decreases, $e^{-E_a/RT}$ increases,

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Q4

(4) Given: $A_2 \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} 2A$

Now, $-\frac{1}{2}\frac{d[A]}{dt}=k_{-1}[A]^2-k_1[A_2]$

$$\frac{d[A]}{dt}=-2k_{-1}[A]^2+2k_1[A_2]$$

$$\Rightarrow\quad \frac{d[A]}{dt}=2k_1[A_2]-2k_{-1}[A]^2$$

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Q5

(2) From Arrhenius equation,

$$k=Ae^{-E_a/RT}$$

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

slope $=-y$ (given)

$-y=-E_a$

$\Rightarrow\quad E_a=y$

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Q6

(3) For the reaction $2X \rightarrow B$, follow zeroth order

Rate equation is

$Kt = [A]_0 - [A]$

For the half-life; $t = t_{1/2}$ and $[A] = 0.1$

$K\, t_{1/2} = 0.2 - 0.1$

$$\frac{0.2-0.1}{6} = \frac{0.1}{6}\,\mathrm{M\,hr^{-1}}$$

$\therefore$ Time required to reach from 0.5 M to 0.2 M

$Kt = [A]_0 - [A]$

$$\frac{0.1}{6} \times t = (0.5 - 0.2);\ t = 18 \text{ hour}$$

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Q7

(1) Rate constant of decomposition of $X = 0.05$ mg/year. Unit of rate constant confirms that the decomposition of X is a zero order reaction.

For zero order kinetics,

$$[X] = [X]_0 - kt$$

$$kt = [X]_0 - [X]$$

$$t = \frac{[X]_0 - [X]}{k}$$

$$t = \frac{5-2.5}{0.05}$$

$$= \frac{2.5}{0.05} = 50 \text{ years}$$

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Q8

(3) From Arrhenius equation,

$$\ln K = \ln A - \frac{E_a}{RT}$$

$$\text{Slope} = \frac{-E_a}{R} = -4606 \text{ K}$$

$$\log\left(\frac{K_2}{K_1}\right) = \frac{E_a}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\log\left(\frac{K_2}{10^{-5}}\right) = \frac{1}{2.303} \times 4606\left(\frac{1}{400} - \frac{1}{500}\right)$$

$$= \frac{4606 \times (100)}{2.303 \times 400 \times 500} = 1$$

$$\Rightarrow \quad \log\left(\frac{K_2}{10^{-5}}\right) = 1;\ \frac{K_2}{10^{-5}} = \text{Antilog}(1)$$

$$\frac{K_2}{10^{-5}} = 10$$

$$\Rightarrow \quad K_2 = 10^{-5} \times 10 = 10^{-4}\ \mathrm{s^{-1}}$$

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Q9

(1) $2A + B \longrightarrow C$

Rate $= k[A]^x\,[B]^y$

Exp-1, $0.045 = k[0.05]^x\,[0.05]^y$ ...(i)

Exp-2, $0.090 = k[0.1]^x\,[0.05]^y$ ...(ii)

Exp-3, $0.72 = k[0.2]^x\,[0.1]^y$ ...(iii)

Divide equation (i) by equation (ii)

$$\frac{0.045}{0.090} = \left(\frac{1}{2}\right)^x \Rightarrow x = 1$$

Divide equation (i) by equation (iii)

$$\frac{0.045}{0.72} = \left(\frac{0.05}{0.1}\right)^y \left(\frac{0.05}{0.2}\right)^1$$

$$\frac{0.045}{0.72} \times \frac{0.2}{0.05} = \left(\frac{0.05}{0.1}\right)^y$$

$$\left(\frac{1}{2}\right)^2 = \left(\frac{1}{2}\right)^y \Rightarrow y = 2$$

Rate law $= k[A]^1\,[B]^2$.

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Q10

(4) $A \xrightarrow{k_1} B \xrightarrow{k_2} C$

$$\frac{d[B]}{dt} = k_1[A] - k_2[B] = 0$$

$$[B] = \frac{k_1[A]}{k_2}$$

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Q11

(1) In graph (i), ln [Reactant vs time is linear with positive intercept and negative slope. Hence it is 1st order In graph (ii), [Reactant] vs time is linear with positive intercept and negative slope. Hence, it is zero order.

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Q12

(3) When drug is administered bacterial growth is given by $\frac{dN}{dt} = -5N^2$

On integrating the above equation,

$\Rightarrow \frac{N_0}{N_t} = 1 + 5tN_0$.

The above equation is similar to straight line equation with positive slope.

Thus $\frac{N_0}{N_t}$ increases linearly with t.

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Q13

(1) $K = e^{-\frac{E_a}{RT}}$ or $\log K = \frac{-E_a}{2.303RT}$

$$\text{So, } \log\frac{K_2}{K_1} = \frac{E_a}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\log\frac{1}{2.5\times10^{-4}} = \frac{E_a}{8.314\times2.303}\left(\frac{1}{600} - \frac{1}{800}\right)$$

$$3.6 = \frac{E_a}{8.314\times2.303}\times\frac{200}{600\times800}$$

$E_a = 165.4$ kJ/mol ≈ 166 kJ/mol

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Q14

(3) $xA \rightarrow yB$

$$\therefore \quad \frac{-dA}{xdt} = \frac{1}{y}\frac{dB}{dt}$$

$$\frac{-dA}{dt} = \frac{dB}{dt}\times\frac{x}{y}$$

$$\log\left[\frac{-dA}{dt}\right] = \log\left[\frac{dB}{dt}\right] + \log\left(\frac{x}{y}\right)$$

Comparing this equation with the equation given in question. We get,

$\log\frac{x}{y} = 0.3010$ or $\log\frac{x}{y} = \log 2$

$$\therefore \quad \frac{x}{y} = 2$$

$\therefore$ The reaction is of type $2A \rightarrow B$.
Hence, option (3) is correct.

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Q15

(2) $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$

$$\text{rate of reaction} = \frac{-1}{2}\frac{d[N_2O_5]}{dt} = \frac{1}{4}\frac{d[NO_2]}{dt} = \frac{d[O_2]}{dt}$$

According to the question

$$\frac{-d[N_2O_5]}{dt} = -\frac{(2.75-3)}{30} = \frac{0.25}{30}\text{M min}^{-1}$$

$$\therefore \frac{d[NO_2]}{dt} = 2\times\frac{-d[N_2O_5]}{dt} = 2\times\frac{0.25}{30}$$

$= 1.67 \times 10^{-2}$ M min^{-1}

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