CHEMICAL KINETICS

1. 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

2. 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 unit (2) –y unit

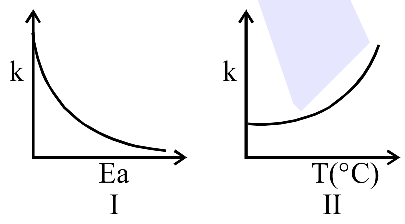
(3) yR unit (4) y/R unit

3. 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) 18.0 h (2) 7.2 h

(3) 9.0 h (4) 12.0 h

4. 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) Both I and II are wrong

(2) I is wrong but II is right

(3) Both I and II are correct

(4) I is right but II is wrong

5. For an elementary chemical reaction,

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

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

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

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

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

6. For the reaction, $2A + B \rightarrow$ products, when the concentrations of A and B both wrere 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 mol $L^{-1}s^{-1}$ to0.6 mol $L^{-1}s^{-1}$ Which one of the following statements is correct ?

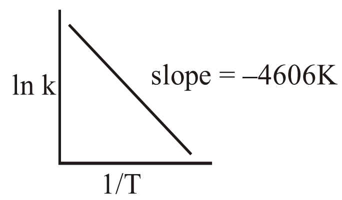
(1) Order of the reaction with respect to Bis2

(2) Order of the reaction with respect to Ais2

(3) Total order of the reaction is 4

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

7. 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) 2×10^{-4} s^{-1} (2) 10^{-4} s^{-1}

(3) 10^{-6} s^{-1} (4) 4×10^{-4} s^{-1}

8. The following results were obtained during kinetic studies of the reaction :

$2A + B \rightarrow$ Products

Experment	[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) 10 (2) 5 (3) 100 (4) 1

9. 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 :

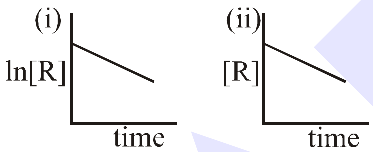
[A] (mol L^{-1})	[B] (mol L^{-1})	Initial Rate (mol $L^{-1}s^{-1}$)
0.05	0.05	0.045
0.10	0.05	0.090
0.20	0.10	0.72

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

10. 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) $\left(\frac{k_1}{k_2}\right)[A]$ (2) $(k_1 + k_2)$ [A]
(3) $k_1k_2[A]$ (4) $(k_1 - k_2)$ [A]

11. The given plots represent 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

12. For the reaction of H_2 with I_2, the rate constant is 2.5×10^{-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=8.314J K^{-1} mol^{-1})

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

13. 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_4 and C_4H_8
(3) N_2O_4 and NO_2
(4) C_2H_2 and C_6H_6

14. 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) 2.083×10^{-3} mol L^{-1} min^{-1}
(2) 4.167×10^{-3} mol L^{-1} min^{-1}
(3) 8.333×10^{-3} mol L^{-1} min^{-1}
(4) 1.667×10^{-2} mol L^{-1} min^{-1}

SOLUTION

1. Ans.(1)

Rate constant (K) = 0.05 μg/year

means zero order reaction

$$t_{1/2} = \frac{a_0}{2K} = \frac{5\mu g}{2\times 0.05\,\mu g/\text{year}} = 50 \text{ year}$$

2. Ans. (1)

3. Ans. (1)

For zero order

$[A_0]-[A_t] = kt$

$0.2 - 0.1 = k\times 6$

$$k = \frac{1}{60} \text{ M/hr}$$

and $0.5-0.2 = \frac{1}{60}\times t$

t = 18 hrs.

4. Ans. (4)

On increasing E_a, k decreases.

In plot II initially k is shown to be almost constant with temperature while as in moderate temperature range increase of k is very sharp, therefore plot II is incorrect.

5. Ans. (3)

Ans.(3)

$$A_2 \underset{K_{-1}}{\overset{K_1}{\rightleftharpoons}} 2A$$

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

6. Ans. (1)

$r = K[A]^x[B]^y$

$\Rightarrow 8 = 2^3 = 2^{x+y}$

$\Rightarrow x + y = 3$ ...(1)

$\Rightarrow 2 = 2^x$

$\Rightarrow x = 1, y = 2$

Order w.r.t. A = 1

Order w.r.t. B = 2

7. Ans.(2)

$$\ell n\frac{K_2}{K_1} = \frac{E_a}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$$

$$2.303\log\frac{K_2}{10^{-5}} = 4606\left[\frac{1}{400} - \frac{1}{500}\right]$$

$\Rightarrow K_2 = 10^{-4}\ s^{-1}$

8. Ans. (2)

$6.93 \times 10^{-3} = K \times (0.1)^x\,(0.2)^y$

$6.93 \times 10^{-3} = K \times (0.1)^x\,(0.25)^y$

So y = 0

and $1.386 \times 10^{-2} = K \times (0.2)^x\,(0.30)^y$

$$\frac{1}{2} = \left(\frac{1}{2}\right)^x \quad \boxed{x = 1}$$

So $r = K \times (0.1) \times (0.2)^0$

$6.93 \times 10^{-3} = K \times 0.1 \times (0.2)^0$

$$\boxed{K = 6.93\times 10^{-2}}$$

$$t_{1/2} = \frac{0.693}{2K} = \frac{0.693}{0.693\times 10^{-1}\times 2} = \frac{10}{2} = 5$$

9. Ans.(3)

Sol. $r = K\,[A]^x\,[B]^y$

$0.045 = K\,(0.05)^x\,(0.05)^y$ ....(1)

$0.090 = K\,(0.10)^x\,(0.05)^y$ ....(2)

$0.72 = K\,(0.20)^x\,(0.10)^y$ ....(3)

From (1) ÷ (2), $\frac{0.045}{0.090} = \left(\frac{0.05}{0.10}\right)^x \Rightarrow x = 1$

From (2) ÷ (3), $\frac{0.090}{0.720} = \left(\frac{0.10}{0.20}\right)^x\cdot\left(\frac{0.05}{0.10}\right)^y \Rightarrow y = 2$

Hence, $r = K\,[A]\,[B]^2$

Correct option : (3)

10. Ans.(1)

Sol. $A \xrightarrow{K_1} B \xrightarrow{K_2} C$

$$\frac{d[B]}{dt} = 0 = K_1[A] - K_2[B]$$

$$\Rightarrow [B] = \frac{K_1}{K_2}[A]$$

11. Ans.(1)

Sol. (i) $\ln[R] = \ln[R]_0 - Kt$ (Ist order)

$[R] = [R]_0 - Kt$ (zero order)

∴ Ans.(1)

12. Ans.(2)

Sol. $H_2(g) + I_2(g) \rightarrow 2HI(g)$

Apply Arrhenius equation

$$\log\frac{K_2}{K_1} = \frac{E_a}{2.303R}\left(\frac{1}{600} - \frac{1}{800}\right)$$

$$\log\frac{1}{2.5\times10^{-4}} = \frac{E_a}{2.303\times8.31}\left(\frac{200}{600\times800}\right)$$

$\therefore\ E_a \approx 166\text{kJ/mol}$

13. Ans.(2)

Sol. $$\log\frac{-d[A]}{dt} = \log\frac{d[B]}{dt} + 0.3010$$

$$\frac{-d[A]}{dt} = 2\times\frac{d[B]}{dt}$$

$$\frac{1}{2}\times\frac{-d[A]}{dt} = \frac{d[B]}{dt}$$

$2A \longrightarrow B$

$2C_2H_4 \longrightarrow C_4H_8$

14. Ans.(4)

Sol. $2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$

t=0 3.0M

t=30 2.75 M

$$\frac{-\Delta[N_2O_5]}{\Delta t} = \frac{0.25}{30}$$

$$\frac{1}{2}\times\frac{-\Delta[N_2O_5]}{\Delta t} = \frac{1}{4}\times\frac{\Delta[NO_2]}{\Delta t}$$

$$\frac{\Delta[NO_2]}{\Delta t} = \frac{0.25}{30}\times2 = 1.66\times10^{-2}\ \text{M/min}$$