



EXERCISE #O-I

Rate of Reaction

1. The rate of a reaction is expressed in different ways as follows:

$$+\frac{1}{2} \frac{d[C]}{dt} = -\frac{1}{3} \frac{d[D]}{dt} = +\frac{1}{4} \frac{d[A]}{dt} = -\frac{d[B]}{dt}$$

The reaction is:

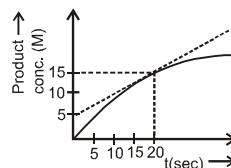
- | | |
|--|--|
| <p>(A) $4A + B \longrightarrow 2C + 3D$</p> <p>(C) $A + B \longrightarrow C + D$</p> | <p>(B) $B + 3D \longrightarrow 4A + 2C$</p> <p>(D) $B + D \longrightarrow A + C$</p> |
|--|--|
2. For the reaction $2A + 3B \rightarrow 4C$
the rate of reaction may be represented as :
- | | |
|--|---|
| <p>(A) $r = -2 \frac{d(A)}{dt} = -\frac{d(B)}{dt} = 4 \frac{d(C)}{dt}$</p> <p>(C) $r = -\frac{1}{2} \frac{d(A)}{dt} = \frac{1}{3} \frac{d(B)}{dt} = \frac{1}{4} \frac{d(C)}{dt}$</p> | <p>(B) $r = -2 \frac{d(A)}{dt} = -4 \frac{d(B)}{dt} = 3 \frac{d(C)}{dt}$</p> <p>(D) $r = -\frac{1}{2} \frac{d(A)}{dt} = -\frac{1}{3} \frac{d(B)}{dt} = \frac{1}{4} \frac{d(C)}{dt}$</p> |
|--|---|
3. In a reaction $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ the rate of appearance of NH_3 is $2.5 \times 10^{-4} \text{ mol L}^{-1} \text{ sec}^{-1}$.
The Rate of reaction & rate of disappearance of H_2 will be (In $\text{mol L}^{-1} \text{ sec}^{-1}$)
- | | |
|---|---|
| <p>(A) $3.75 \times 10^{-4}, 1.25 \times 10^{-4}$</p> <p>(C) $1.25 \times 10^{-4}, 3.75 \times 10^{-4}$</p> | <p>(B) $1.25 \times 10^{-4}, 2.5 \times 10^{-4}$</p> <p>(D) $5.0 \times 10^{-4}, 3.75 \times 10^{-4}$</p> |
|---|---|
4. For the reaction $4A + B \rightarrow 2C + 2D$
The incorrect statement is :-
- | |
|--|
| <p>(A) The rate of disappearance of B is one fourth the rate of disappearance of A</p> <p>(B) The rate of appearance of C is half the rate of disappearance of B</p> <p>(C) The rate of formation of D is half the rate of consumption of A</p> <p>(D) The rates of formation of C and D are equal</p> |
|--|
5. In the following reaction : $xA \longrightarrow yB$

$$\log \left[-\frac{d[A]}{dt} \right] = \log \left[\frac{d[B]}{dt} \right] + 0.3$$

where -ve sign indicates rate of disappearance of the reactant. Thus, x : y is :

- (A) 1 : 2 (B) 2 : 1 (C) 3 : 1 (D) 3 : 10

6. Rate of formation of product at $t = 20$ seconds is :



- (A) 0.5 MS^{-1} (B) 1 M S^{-1} (C) 1.5 M S^{-1} (D) 2 MS^{-1}

Rate law and order of reaction

7. If concentration of reactants is increased by 'x' then the k becomes –

- (A) $\ln \frac{k}{x}$ (B) $\frac{k}{x}$ (C) $k + x$ (D) k

8. For the reaction, $2\text{NO(g)} + 2\text{H}_2\text{(g)} \longrightarrow \text{N}_2\text{(g)} + 2\text{H}_2\text{O(g)}$ the rate expression can be written in the following ways:

$$\begin{aligned} d[\text{N}_2] / dt &= k_1 [\text{NO}][\text{H}_2]; \\ d[\text{H}_2\text{O}] / dt &= k[\text{NO}][\text{H}_2]; \\ -d[\text{NO}] / dt &= k'_1 [\text{NO}][\text{H}_2]; \\ -d[\text{H}_2] / dt &= k''_1 [\text{NO}][\text{H}_2] \end{aligned}$$

The relationship between k , k_1 , k'_1 and k''_1 . is :

- (A) $k = k_1 = k'_1 = k''_1$ (B) $k = 2k_1 = k'_1 = k''_1$
 (C) $k = 2k'_1 = k_1 = k''_1$ (D) $k = k_1 = k'_1 = 2k''_1$

9. For a reaction $\text{A} + \text{B} \rightarrow \text{products}$, the rate of the reaction was doubled when the concentration of A was doubled, the rate was again doubled when the conc. of A & B were doubled the order of the reaction with respect to A & B are:-

- (A) 1, 1 (B) 2, 0 (C) 1, 0 (D) 0, 1

10. For the irreversible process, $\text{A} + \text{B} \longrightarrow \text{products}$, the rate is first-order w.r.t. A and second-order w.r.t. B. If 1.0 mol each of A and B introduced into a 1.0 L vessel, and the initial rate was $1.0 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1}$, rate when half reactants have been turned into products is :

- (A) $1.25 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1}$ (B) $1.0 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1}$
 (C) $2.50 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1}$ (D) $2.0 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1}$

11. For the reaction $2\text{NO}_2 \longrightarrow \text{N}_2\text{O}_2 + \text{O}_2$, rate expression is as follows :



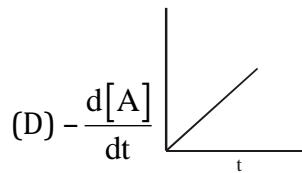
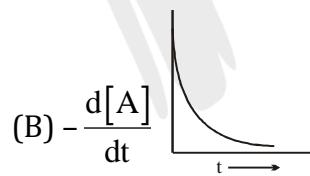
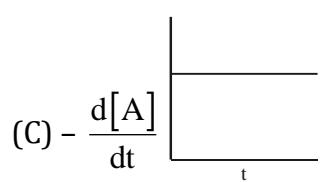
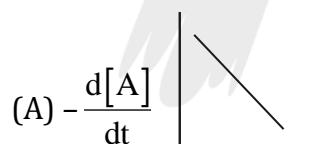
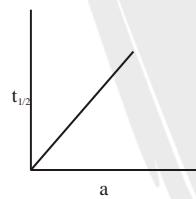
$-\frac{d[\text{NO}_2]}{dt} = k [\text{NO}_2]^n$, where $k = 3 \times 10^{-3} \text{ mol}^{-1} \text{ L sec}^{-1}$. If the rate of formation of oxygen is $1.5 \times 10^{-4} \text{ mol L}^{-1} \text{ sec}^{-1}$, then the molar concentration of NO_2 in mole L^{-1} is

- (A) 1.5×10^{-4} (B) 0.0151 (C) 0.214 (D) 0.316

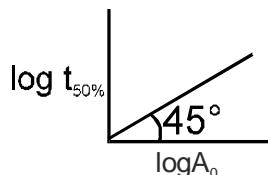
Zero Order Reactions

12. If the rate of the reaction is equal to the rate constant, the order of the reaction is:-
 (A) 0 (B) 1 (C) 2 (D) 3
13. The rate constant of a zero order reaction is $0.2 \text{ mol dm}^{-3} \text{ h}^{-1}$. If the concentration of the reactant after 30 minutes is 0.05 mol dm^{-3} . Then its initial concentration would be :
 (A) 0.15 mol dm^{-3} (B) 1.05 mol dm^{-3} (C) 0.25 mol dm^{-3} (D) 4.00 mol dm^{-3}
14. The rate constant of the reaction $\text{A} \rightarrow 2\text{B}$ is $1.0 \times 10^{-3} \text{ mol L}^{-1} \text{ min}^{-1}$, if the initial concentration of A is 1.0 mole L^{-1} what would be the concentration of B after 100 minutes.
 (A) 0.1 mol lit^{-1} (B) 0.2 mol lit^{-1} (C) 0.9 mol lit^{-1} (D) 1.8 mol lit^{-1}
15. Consider the reaction $\text{A} \rightarrow \text{B}$, graph between half life ($t_{1/2}$) is and initial concentration (a) of the reactant

Hence graph between $-\frac{d[\text{A}]}{dt}$ and time will be



16. What will be the order of reaction and rate constant for a chemical change having $\log t_{50\%}$ vs $\log A_0$ (A_0 = Initial concentration of reactant) curves as :

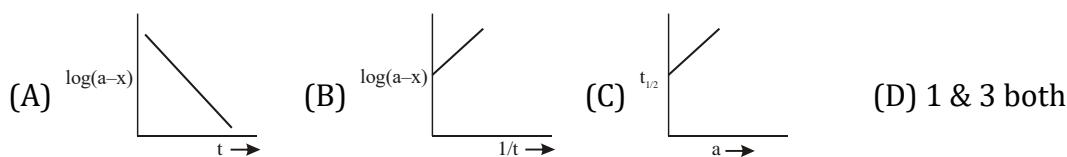


- (A) 0, 1/2 (B) 1, 1 (C) 2, 2 (D) 3, 1

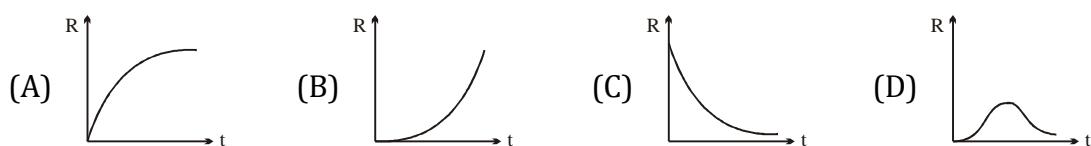
First order reaction

17. The data of $t_{1/2}$ at different initial concentrations for a given reaction, was found to be constant. The order of reaction is :-
- (A) 0 (B) 1 (C) 2 (D) 3
18. If the first order reaction involves gaseous reactants and gaseous products the units of its rate are -
- (A) sec^{-1} (B) atm. sec (C) atm.sec^{-1} (D) $\text{atm}^2 \text{ sec}^2$
19. A reaction is found to have the rate constant $X \text{ sec}^{-1}$ by what factor the rate is increased if initial conc. of A is tripled
- (A) 3 (B) 9 (C) X (D) Remains same
20. The rate constant of a first order reaction is $4 \times 10^{-3} \text{ sec}^{-1}$. At a reactant concentration of 0.02 M, the rate of reaction would be -
- (A) $8 \times 10^{-5} \text{ M sec}^{-1}$ (B) $4 \times 10^{-3} \text{ M sec}^{-1}$ (C) $2 \times 10^{-1} \text{ M sec}^{-1}$ (D) $4 \times 10^{-1} \text{ M sec}^{-1}$
21. In a first order reaction the concentration of the reactant is decreased from 1.0 M to 0.25 M in 20 min. The rate constant of the reaction would be -
- (A) 10 min^{-1} (B) 6.931 min^{-1} (C) 0.6931 min^{-1} (D) 0.06931 min^{-1}
22. A first order reaction has a half life period of 69.3 sec. At $0.10 \text{ mol lit}^{-1}$ reactant concentration rate will be -
- (A) $10^{-4} \text{ M sec}^{-1}$ (B) $10^{-3} \text{ M sec}^{-1}$ (C) $10^{-1} \text{ M sec}^{-1}$ (D) $6.93 \times 10^{-1} \text{ M sec}^{-1}$
23. What fraction of a reactant left showing first order reaction after 40 minute if $t_{1/2}$ is 20 minute
- (A) 1/4 (B) 1/2 (C) 1/8 (D) 1/6

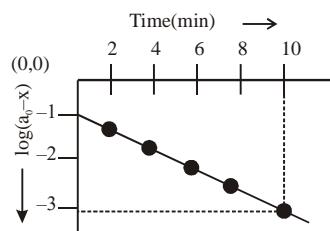




30. If decomposition reaction $A(g) \rightarrow B(g)$ follows first order kinetics then the graph of rate of formation (R) of B against time t will be



31. For the first order decomposition of $\text{SO}_2\text{Cl}_2(\text{g})$, $\text{SO}_2\text{Cl}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) + \text{Cl}_2(\text{g})$ a graph of $\log(a_0 - x)$ vs t is shown in figure. What is the rate constant (sec^{-1})?



- (A) 0.2 (B) 4.6×10^{-1} (C) 7.7×10^{-3} (D) 1.15×10^{-2}

Second Order Reaction

32. Which is not true for a second order reaction ?

- (A) It can have rate constant $1 \times 10^{-2} \text{ L mol}^{-1} \text{ s}^{-1}$
 (B) Its half-life is inversely proportional to its initial concentration
 (C) Time to complete 75% reaction is twice of half-life

$$(D) t_{50\%} = \frac{1}{k \times A_0}$$

33. A reaction, which is second order, has a rate constant of $0.002 \text{ L mol}^{-1} \text{ s}^{-1}$. If the initial conc. of the reactant is 0.2 M. how long will it take for the concentration to become 0.0400 M ?

- (A) 1000 sec (B) 400 sec (C) 200 sec (D) 10,000 sec

34. The rate constant for a second order reaction is $8 \times 10^{-5} \text{ M}^{-1} \text{ min}^{-1}$: How long will it take a 1M solution to be reduced to 0.5 M.

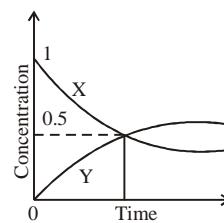
- (A) $8.665 \times 10^3 \text{ min}$ (B) $8 \times 10^{-3} \text{ min}$ (C) $1.25 \times 10^4 \text{ min}$ (D) $4 \times 10^{-5} \text{ min}$

nth order and pseudo order reaction

35. Unit of rate constant of n^{th} order reaction is.

- (A) litre $^{1-n}$ mol $^{1-n}$ sec $^{-1}$ (B) mol $^{n-1}$ litre $^{1-n}$ sec $^{-1}$
 (C) mol $^{1-n}$ litre $^{n-1}$ sec $^{-1}$ (D) None

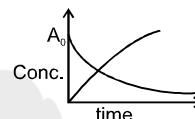
36. The accompanying figure depicts the change in concentration of species X and Y for the reaction $X \rightarrow Y$ as a function of time the point of intersection of the two curves represents.



- (A) $t_{1/2}$
(C) $t_{2/3}$

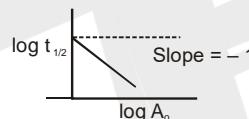
- (B) $t_{3/4}$
(D) Data insufficient to predict

37. At the point of intersection of the two curves shown, the conc. of B is given by....for, $A \rightarrow nB$:



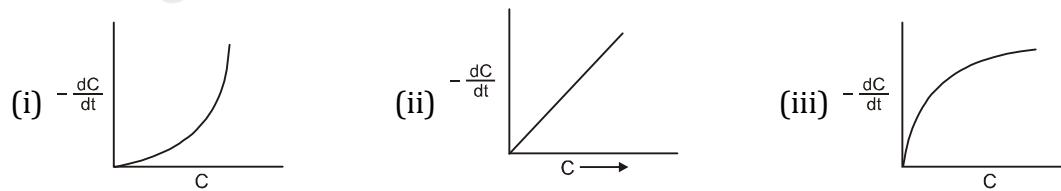
- (A) $\frac{nA_0}{2}$
(B) $\frac{A_0}{n-1}$
(C) $\frac{nA_0}{n+1}$
(D) $\left(\frac{n-1}{n+1}\right)A_0$

38. A graph between $\log t_{1/2}$ and $\log a$ (abscissa) a being the initial concentration of A_0 in the reaction For reaction $A \rightarrow$ Product, the rate law is :



- (A) $\frac{-d[A]}{dt} = k$
(B) $\frac{-d[A]}{dt} = k [A]$
(C) $\frac{-d[A]}{dt} = k [A]^2$
(D) $\frac{-d[A]}{dt} = k [A]^3$

39. In three different reactions, involving a single reactant in each case, a plot of rate of the reaction on the y-axis, versus concentration of the reactant on the x-axis, yields three different curves shown below.



What are the possible orders of the reactions (i), (ii), (iii).

- (A) 1, 2, 3
(B) 2, 1, 1/2
(C) 0, 1, 2
(D) 0, 1, 1/2
40. The rate law of the reaction $A + 2B \rightarrow$ product(P) is given by $\frac{d[P]}{dt} = K[A]^2 [B]$. If A is taken in large excess, the order of the reaction will be -
(A) Zero
(B) 1
(C) 2
(D) 3



Experimental determination of order

- 42.** The data for the reaction $A + B \rightarrow C$ is:

Exp.	[A] ₀	[B] ₀	initial rate
1	0.012	0.035	0.10
2	0.024	0.035	0.80
3	0.012	0.070	0.10
4	0.024	0.070	0.80

The rate law is

- (A) $r = k[B]^3$ (B) $r = k[A]^3$ (C) $r = k[A][B]^4$ (D) $r = k[A]^2[B]^2$

- 43.** The kinetic data for the given reaction $A(g) + 2B(g) \rightarrow C(g)$ is provided in the following table for three experiments at 300 K.

Ex. No.	[A/M]	[B/M]	[Initial rate ($M\sec^{-1}$)]
1	0.01	0.01	6.930×10^{-6}
2	0.02	0.01	1.386×10^{-5}
3	0.02	0.02	1.386×10^{-5}

In another experiment starting with initial concentration of 0.5 and 1 M respectively for A and B at 300 K. Find the rate of reaction after 50 minutes from start of experiment (in m/sec)?

- (A) 6.93×10^{-4} (B) 0.25×10^{-7} (C) 4.33×10^{-5} (D) 3.46×10^{-9}

- 44.** At certain temperature, the half life period for the thermal decomposition of a gaseous substance depends on the initial partial pressure of the substance as follows

Pressure (mmHg)	500	250
$t_{1/2}$ (in min.)	235	950

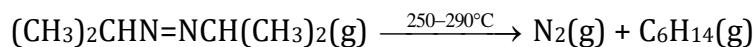
Find the order of reaction [Given $\log(23.5) = 1.37$; $\log(95) = 1.97$; $\log 2 = 0.30$]

Experimental determination of First order rate constant



45. At 373 K, a gaseous reaction $A \rightarrow 2B + C$ is found to be of first order. Starting with pure A, the total pressure at the end of 10 min was 176 mm of Hg and after a long time when A was completely dissociated, it was 270 mm of Hg. The pressure of A at the end of 10 minutes was :

- 46.** Azo isopropane decomposes according to the equation :-



It is found to be a first order reaction. If initial pressure is P_0 and pressure of the mixture at time t is (P_t) then rate constant k would be :-

$$(A) k = \frac{2.303}{t} \log \frac{P_0}{2P_0 - P_t}$$

$$(B) k = \frac{2.303}{t} \log \frac{P_0 - P_t}{P_0}$$

$$(C) k = \frac{2.303}{t} \log \frac{P_0}{P_0 - P_t}$$

$$(D) k = \frac{2.303}{t} \log \frac{2P_0}{2P_0 - P_t}$$

- 47.** Half life of reaction : $\text{H}_2\text{O}_2(\text{aq}) \longrightarrow \text{H}_2\text{O}(\text{l}) + \frac{1}{2}\text{O}_2(\text{g})$ is independent of initial concentration of H_2O_2 . Volume of O_2 gas after 20 minute is 5L at 1 atm and 27°C and after completion of reaction 50L. The rate constant is

$$(A) \frac{1}{20} \log 10 \text{ min}^{-1}$$

$$(B) \frac{2.303}{20} \log 10 \text{min}^{-1}$$

$$(C) \frac{2.303}{20} \log \frac{50}{45} \text{min}^{-1}$$

$$(D) \frac{2.303}{20} \log \frac{45}{50} \text{ min}^{-1}$$

- 48.** In the reaction $\text{NH}_4\text{NO}_2 \text{ (aq.)} \rightarrow \text{N}_2 \text{ (g)} + 2 \text{ H}_2\text{O \text{ (l)}}$ the volume of N_2 after 20 min and after a long time is 40 ml and 70 ml respectively. The value of rate constant is :

$$(A) (1/20) \log (7/4) \text{ min}^{-1}$$

$$(B) (2.303 / 1200) \log (7/3) \text{ sec}^{-1}$$

$$(C) (1/20) \log (7/3) \text{ min}^{-1}$$

$$(D) (2.303 / 20) \log (11/7) \text{ min}^{-1}$$

- 49.** For a reaction $A \longrightarrow B + C$, it was found that at the end of 10 minutes from the start the total optical rotation of the system was 50° and when the reaction is complete, it was 100° . Assuming that only B and C are optically active and dextrorotatory. Calculate the rate constant of this first order reaction.

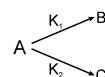
(A) 0.693 min⁻¹ (B) 0.0693 sec⁻¹ (C) 0.0693 min⁻¹ (D) 0.00693 sec⁻¹

- 50.** The following data were obtained in an experiment on inversion of cane sugar (a first order kinetics)

Time (min)	0	10	After a long time
Total angle of rotation (degree)	+40	+15	- 10
The rate constant (in second ⁻¹) is [$\ln 2 = 0.693$]			
(A) 0.0693	(B) 1.15×10^{-3}	(C) 0.693	(D) 1.15×10^{-2}

Parallel first order reaction

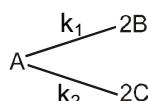
- 51.** The substance undergoes first order decomposition. The decomposition follows two parallel first order reactions as :



$$k_1 = 1.26 \times 10^{-4} \text{ sec}^{-1} \text{ and } k_2 = 3.8 \times 10^{-5} \text{ sec}^{-1}$$

The percentage distribution of B and C

- 52.** For a hypothetical elementary reaction,



where $\frac{k_1}{k_2} = \frac{1}{2}$

Initially only 2 moles of A are present. The total number of moles of A, B and C at the end of 75% of reaction are :

Collision Theory

53. According to collision theory of reaction rates –

 - (A) Every collision between reactant leads to chemical reaction
 - (B) Rate of reaction depends upon shape of molecule
 - (C) All reactions which occur in gaseous phase are zero order reaction
 - (D) Rate of reaction is directly proportional to collision frequency.

- 54.** Activation energy of a reaction is –

 - (A) The energy released during the reaction
 - (B) The energy evolved when activated complex is formed
 - (C) Minimum kinetic energy needed to overcome the potential barrier of reaction
 - (D) The energy needed to form one mole of the product



55. The minimum energy for molecules to enter into chemical reaction is called.
 (A) Kinetic energy (B) Potential energy (C) Threshold energy (D) Activation energy
56. For producing the effective collisions, the colliding molecules must possess:-
 (A) A certain minimum amount of energy
 (B) Energy equal to or greater than threshold energy
 (C) Proper orientation
 (D) Threshold energy as well as proper orientation of collision
57. A large increase in the rate of a reaction for a rise in temperature is due to –
 (A) Increase in the number of collisions (B) Increase in the number of activated molecules
 (C) Lowering of activation energy (D) Shortening of the mean free path

Effect of temperature on rate

58. Which plot the slope of which can give us the value of activation energy :
 (A) k versus T (B) $\frac{1}{k}$ versus T (C) $\log k$ versus $1/T$ (D) C versus T
59. Given that k is the rate constant for some order of any reaction at temp T then the value of $\lim_{T \rightarrow \infty} k$ _____.
 (A) $\frac{2}{2.303}$ (B) A (C) $2.303 A$ (D) $\log A$
60. The rate constant, the activation energy and the Arrhenius parameter (A) of a chemical reaction at 25°C are $3.0 \times 10^{-4} \text{ s}^{-1}$, $104.4 \text{ kJ mol}^{-1}$ and $6.0 \times 10^{14} \text{ s}^{-1}$ respectively. The value of the rate constant at $T \rightarrow \infty$ is
 (A) $2.0 \times 10^{18} \text{ s}^{-1}$ (B) $6.0 \times 10^{14} \text{ s}^{-1}$ (C) infinity (D) $3.6 \times 10^{30} \text{ s}^{-1}$
61. For a certain gaseous reaction, a 10°C rise of temp. from 25°C to 35°C doubles the rate of reaction. What is the value of activation energy :-
 (A) $\frac{10}{2.303R \times 298 \times 308}$ (B) $\frac{2.303 \times 10}{298 \times 308R}$

(C) $\frac{0.693R \times 10}{290 \times 308}$

(D) $\frac{0.693R \times 298 \times 308}{10}$

62. From the following data; the activation energy for the reaction (cal/mol) $H_2 + I_2 \rightarrow 2HI$

T, (in K)

1/T, (in K⁻¹)log₁₀k

769

 1.3×10^{-3}

2.9

667

 1.5×10^{-3}

1.1

(A) 4×10^4 (B) 2×10^4 (C) 8×10^4 (D) 3×10^4

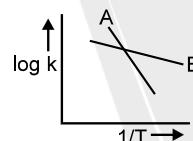
63. A first order reaction is 50% completed in 20 minutes at 27°C and in 5 min at 47°C. The energy of activation of the reaction is

(A) 43.85 kJ/mol (B) 55.14 kJ/mol (C) 11.97 kJ/mol (D) 6.65 kJ/mol

64. For the first order reaction A → B + C, carried out at 27 °C if $3.8 \times 10^{-16} \%$ of the reactant molecules can overcome energy barrier, the E_a (activation energy) of the reaction is [log 3.8 = 0.58]

(A) 12 kJ/mole (B) 831.4 kJ/mole (C) 100 kJ/mole (D) 88.57 kJ/mole

65. The Arrhenius relationship of two different reactions is shown below. Which reaction is faster at a lower temperature and which is more sensitive to changes of temperature ?



(A) B faster, A more sensitive

(B) B in both cases

(C) A in both cases

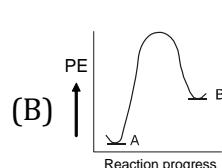
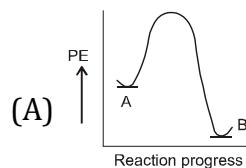
(D) A faster, B more sensitive

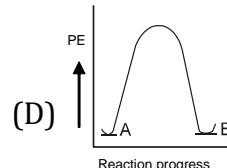
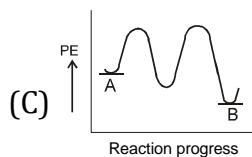
66. When a graph between ln K and 1/T is drawn a straight line is obtained. The point at which line cuts y -axis and x -axis respectively correspond to the temperature :

(A) 0, E_a / 2.303 R log A (B) ∞, E_a / (R ln A)

(C) 0, log A (D) None of these.

67. For a reaction A → B, E_a = 10 kJ mol⁻¹, ΔH = 5 kJ mol⁻¹. Thus, potential energy profile for this reaction is :





Effect of catalyst on rate

Molecularity of reaction

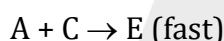
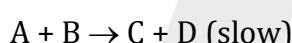
69. For the reaction $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$ the experiment data suggested that $r = k[\text{H}_2][\text{Br}_2]^{1/2}$

The molecularity and order of the reaction are respectively :

- (A) 2, 3/2 (B) 3/2 , 3/2 (C) Not defined, 3/2 (D) 1,1/2

Mechanism of reaction

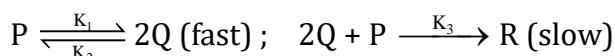
- 79.** Following mechanism has been proposed for a reaction : $2A + B \rightarrow D + E$ Mechanism



The rate law expression for the reaction is :

- (A) $r = k[A]^2[B]$ (B) $r = k[A][B]$ (C) $r = k[A]^2$ (D) $r = k[A][C]$

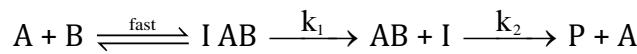
- 71.** The reaction mechanism for the reaction $P \rightarrow R$ is as follows :



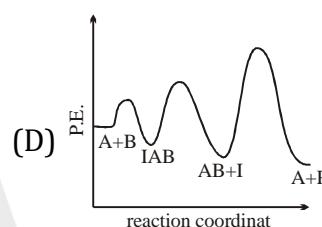
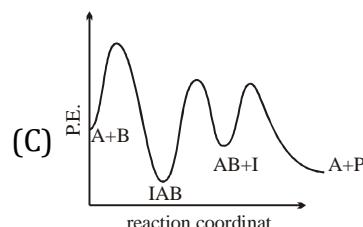
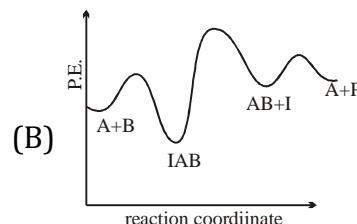
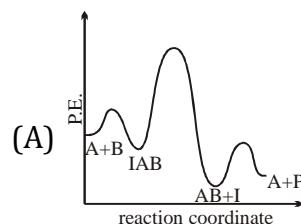
the rate law for the main reaction ($P \rightarrow R$) is:

- (A) $k_1 [P] [Q]$ (B) $k_1 k_2 [P]$ (C) $\frac{k_1 k_3 [P]^2}{k_2}$ (D) $k_1 k_2 [a]$

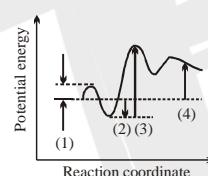
72. The following mechanism has been proposed for the exothermic catalyzed complex reaction.



If k_1 is much smaller than k_2 . The most suitable qualitative plot of potential energy (P.E.) versus reaction coordinate for the above reaction



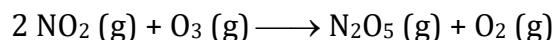
73. Choose the correct set of identifications, for the reaction : $S \xrightarrow{[E]} P$



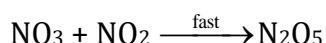
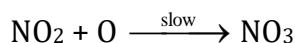
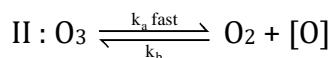
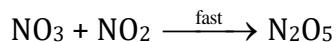
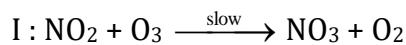
- | (A) | (B) | (C) | (D) |
|---|---|---|---|
| (A) ΔE for $E + S \rightarrow ES$ | E_a for $ES \rightarrow EP$ | $\Delta E_{\text{overall}}$ for $S \rightarrow P$ | E_a for $EP \rightarrow E + P$ |
| (B) E_a for $E + S \rightarrow ES$ | ΔE for $E + S \rightarrow ES$ | E_a for $ES \rightarrow EP$ | $\Delta E_{\text{overall}}$ for $S \rightarrow P$ |
| (C) E_a for $ES \rightarrow EP$ | E_a for $EP \rightarrow E + P$ | $\Delta E_{\text{overall}}$ for $S \rightarrow P$ | ΔE for $EP \rightarrow E + P$ |
| (D) E_a for $E + S \rightarrow ES$ | E_a for $ES \rightarrow EP$ | E_a for $EP \rightarrow E + P$ | $\Delta E_{\text{overall}}$ for $S \rightarrow P$ |
| (E) ΔE for $E + S \rightarrow ES$ | $\Delta E_{\text{overall}}$ for $S \rightarrow P$ | ΔE for $EP \rightarrow E + P$ | E_a for $EP \rightarrow E + P$ |



74. The reaction of NO_2 (g) and O_3 (g) is first-order in NO_2 (g) and O_3 (g)



The reaction can take place by mechanism :

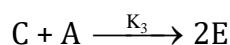
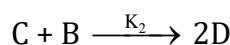
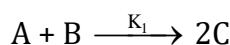


Select correct mechanism.

- (A) I only
 (C) both I and II

- (B) II only
 (D) None of I and II

75. Reaction $\text{A} + \text{B} \longrightarrow \text{D} + \text{E}$ take place as



The rate of disappearance of C is given by

$$(A) -\frac{d[\text{C}]}{dt} = K_2 [\text{B}][\text{C}] + K_3[\text{A}][\text{C}] - 2K_1[\text{A}][\text{B}]$$

$$(B) -\frac{d[\text{C}]}{dt} = K_2[\text{B}][\text{C}] + K_3 [\text{E}] - K_1[\text{C}]$$

$$(C) -\frac{d[\text{C}]}{dt} = K_2[\text{D}]0 + K_3[\text{E}] - K_1[\text{C}]$$

$$(D) -\frac{d}{dt} [\text{C}] = 2K_1[\text{A}][\text{B}] - K_2 [\text{B}][\text{C}] - K_3[\text{A}][\text{C}]$$



ANSWER KEY

EXERCISE #O-I

RATE OF REACTION AND STOICHIOMETRIC COEFFICIENT

- | | | | | | | |
|---------|---------|---------|---------|---------|---------|---------|
| 1. (B) | 2. (D) | 3. (C) | 4. (B) | 5. (B) | 6. (A) | 7. (D) |
| 8. (B) | 9. (C) | 10. (A) | 11. (D) | 12. (A) | 13. (A) | 14. (B) |
| 15. (C) | 16. (A) | 17. (B) | 18. (C) | 19. (A) | 20. (D) | 21. (B) |
| 22. (A) | 23. (A) | 24. (A) | 25. (D) | 26. (C) | 27. (A) | 28. (C) |
| 29. (A) | 30. (C) | 31. (C) | 32. (C) | 33. (D) | 34. (C) | 35. (C) |
| 36. (A) | 37. (C) | 38. (C) | 39. (B) | 40. (B) | 41. (C) | 42. (B) |
| 43. (C) | 44. (D) | 45. (B) | 46. (A) | 47. (C) | 48. (B) | 49. (C) |
| 50. (B) | 51. (B) | 52. (D) | 53. (D) | 54. (C) | 55. (C) | 56. (D) |
| 57. (B) | 58. (C) | 59. (B) | 60. (B) | 61. (D) | 62. (A) | 63. (B) |
| 64. (C) | 65. (A) | 66. (B) | 67. (B) | 68. (D) | 69. (C) | 70. (B) |
| 71. (C) | 72. (A) | 73. (B) | 74. (C) | 75. (A) | | |