Chemical Kinetics

Question1

Consider the following data for the given reaction

$2HI_{(g)} \rightarrow H_{2(g)} + I_{2(g)}$

	1	2	3
$HI(mol\,L^{-1})$	0.005	0.01	0.02
Rate $(mol\,L^{-1}\ s-1)$	7.5×10^{-4}	3.0×10^{-3}	1.2×10^{-2}

[27-Jan-2024 Shift 1]

Answer: 2

Solution:

Let, $R = k[HI]^n$ using any two of given data,

$$\frac{3 \times 10^{-3}}{7.5 \times 10^{-4}} = \left(\frac{0.01}{0.005}\right)^n$$

$n = 2$

Question2

For a reaction taking place in three steps at same temperature, overall rate constant K = $\frac{K_1K_2}{K_3}$. If Ea_1, Ea_2 and Ea_3 are 40, 50 and 60kJ/mol respectively, the overall Ea is kJ/mol.

[29-Jan-2024 Shift 1]

Answer: 30

Solution:

$$K = \frac{K_1 K_2}{K_3} = \frac{A_1 . A_2}{A_3} . e^{-\frac{(E_{a_1} + E_{a_2} - E_{a_3})}{RT}}$$

$$A . e^{-E_a/RT} = \frac{A_1 A_2}{A_3} \cdot e^{-\frac{(E_{a_1} + E_{a_2} - E_{a_3})}{RT}}$$

$$E_a = E_{a_1} + E_{a_2} - E_{a_3} = 40 + 50 - 60 = 30 \text{ kJ/mole.}$$

Question3

The half-life of radioisotopic bromine - 82 is 36 hours. The fraction which remains after one day is _____× 10^{-2}.

(Given antilog 0.2006 = 1.587)

[29-Jan-2024 Shift 2]

Answer: 63

Solution:

Half life of bromine −82 = 36 hours

$$t_{1/2} = \frac{0.693}{K}$$

$$K = \frac{0.693}{36} = 0.01925 \text{hr}^{-1}$$

1^{st} order rxn kinetic equation

$$t = \frac{2.303}{K} \log \frac{a}{a-x}$$

$$\log \frac{a}{a-x} = \frac{t \times K}{2.303} \quad (t = 1 \text{ day } = 24 \text{ hr})$$

$$\log \frac{a}{a-x} = \frac{24 \text{ hr} \times 0.01925 \text{hr}^{-1}}{2.303}$$

$$\log \frac{a}{a-x} = 0.2006$$

$$\frac{a}{a-x} = \text{anti}\log(0.2006)$$

$$\frac{a}{a-x} = 1.587$$

If $a = 1$

$$\frac{1}{1-x} = 1.587 \Rightarrow 1 - x = 0.6301 = \text{ Fraction remain}$$

after one day

Question4

The rate of first order reaction is $0.04 mol L^{-1} s^{-1}$ at 10 minutes and $0.03 mol L^{-1} s^{-1}$ at 20 minutes after initiation. Half life of the reaction is______ minutes. (Given log2 = 0.3010, log3 = 0.4771)

[30-Jan-2024 Shift 1]

Answer: 24

Solution:

$0.04 = k[A]_0 e^{-k \times 10 \times 60}$ ....(1)

$0.03 = k[A]_0 e^{-k \times 20 \times 60}$ ....(2)

(1)/(2)

$\frac{4}{3} = e^{600k(2-1)}$

$\frac{4}{3} = e^{600k}$

$\ln \frac{4}{3} = 600k$

$\ln \frac{4}{3} = 600 \times \frac{\ln 2}{t_{1/2}}$

$t_{1/2} = 600 \frac{\ln 2}{\ln \frac{4}{3}} sec$

$t_{1/2} = 600 \times \frac{\log 2}{\log 4 - \log 3}$ sec. $= 10 \times \frac{0.3010}{0.6020 - 0.477}$ min

$t_{1/2} = 24.08$ min

Ans. 24

Question5

NO_2 required for a reaction is produced by decomposition of N_2O_5 in CCl_4 as by equation $2N_2O_{5(g)} \rightarrow 4NO_{2(g)} + O_{2(g)}$

The initial concentration of N_2O_5 is $3 mol L^{-1}$ and it is $2.75 mol L^{-1}$ after 30 minutes.

The rate of formation of NO_2 is $x \times 10^{-3} mol L^{-1} min^{-1}$, value of x is_____

[30-Jan-2024 Shift 2]

Answer: 17

Solution:

Rate of reaction (ROR)

$$= -\frac{1}{2}\frac{\Delta[N_2O_5]}{\Delta t} = \frac{1}{4}\frac{[NO_2]}{\Delta t} = \frac{\Delta[O_2]}{\Delta t}$$

$$ROR = -\frac{1}{2}\frac{\Delta[N_2O_5]}{\Delta t} = -\frac{1}{2}\frac{(2.75-3)}{30}\,mol\,L^{-1}\,min^{-1}$$

$$ROR = -\frac{1}{2}\frac{(-0.25)}{30}\,mol\,L^{-1}\,min^{-1}$$

$$ROR = \frac{1}{240}\,mol\,L^{-1}\,min^{-1}$$

Rate of formation of $NO_2 = \frac{\Delta[NO_2]}{\Delta t} = 4 \times ROR$

$$= \frac{4}{240} = 16.66 \times 10^{-3} molL^{-1} min^{-1} \simeq 17 \times 10^{-3}.$$

Question6

Integrated rate law equation for a first order gas phase reaction is given by (where P_i is initial pressure and P_t is total pressure at time t)

[31-Jan-2024 Shift 1]

Options:

A.

$$k = \frac{2.303}{t} \times \log \frac{P_i}{(2P_i - P_t)}$$

B.

$$k = \frac{2.303}{t} \times \log \frac{2P_i}{(2P_i - P_t)}$$

C.

$$k = \frac{2.303}{t} \times \log \frac{(2P_i - P_t)}{P_i}$$

D.

$$k = \frac{2.303}{t} \times \frac{P_i}{(2P_i - P_t)}$$

Answer: A

Solution:

$$\begin{array}{cccc} A & \rightarrow & B & + & C \\ P_i & & 0 & & 0 \\ P_i - x & & x & & x \end{array}$$

$P_t = P_i + x$

$P_i - x = P_i - P_t + P_i$

$= 2P_i - P_t$

$$K = \frac{2.303}{t} \log \frac{P_i}{2P_i - P_t}$$

Question7

r = k[A] for a reaction, 50% of A is decomposed in 120 minutes. The time taken for 90% decomposition of A is ____ minutes.

[31-Jan-2024 Shift 2]

Answer: 399

Solution:

$r = k[A]$

So, order of reaction $= 1$

$t_{1/2} = 120 \text{ min}$

For 90% completion of reaction

$$\Rightarrow k = \frac{2.303}{t} \log\left(\frac{a}{a - x}\right)$$

$$\Rightarrow \frac{0.693}{t_{1/2}} = \frac{2.303}{t} \log \frac{100}{10}$$

$\therefore t = 399 \text{ min.}$

Question8

The ratio of $^{14}C/^{12}C$ in a piece of wood is 1/8 part that of atmosphere. If half life of 14C is 5730 years, the age of wood sample is ..... years.

[1-Feb-2024 Shift 1]

Answer: 17190

Solution:

$$\lambda t = \ln \frac{(^{14}C/^{12}C)_{\text{atmosphere}}}{(^{14}C/^{12}C)_{\text{wood sample}}}$$

As per the question,

$$\frac{(^{14}C/^{12}C)_{\text{wood}}}{(^{14}C/^{12}C)_{\text{atmosphere}}} = \frac{1}{8}$$

So, $\lambda t = \ln 8$

$$\frac{\ln 2}{t_{1/2}} t = \ln 8$$

$$t = 3 \times t_{1/2} = 17190 \text{ years}$$

Question9

The following data were obtained during the first order thermal decomposition of a gas A at constant volume:
$A(g) \longrightarrow 2B(g) + C(g)$

S.No	Time/s	Total pressure/(atm)
1.	0	0.1
2.	115	0.28

The rate constant of the reaction is ____ $\times 10^{-2}\ s^{-1}$ (nearest integer)

[1-Feb-2024 Shift 2]

Answer: 2

Solution:

$$\begin{array}{lcccc} & A(g) & \rightarrow 2B(g) & + & C(g) \\ t = 0 & 0.1 & & & \\ t = 115\text{ sec.} & 0.1 - x & 2x & & x \end{array}$$

$$0.1 + 2x = 0.28$$

$$2x = 0.18$$

$$x = 0.09$$

$$K = \frac{1}{115} \ln \frac{0.1}{0.1 - 0.09}$$

$$= 0.0200\ sec^{-1}$$

$$= 2 \times 10^{-2}\ sec^{-1}$$

Question10

The number of correct statement/s from the following is___

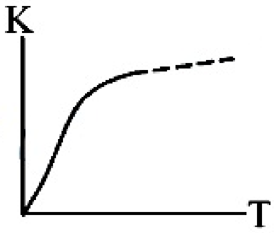
A. Larger the activation energy, smaller is the value of the rate constant.
B. The higher is the activation energy, higher is the value of the temperature coefficient.
C. At lower temperatures, increase in temperature causes more change in the value of k than at higher temperature.
D. A plot of $\ln k$ vs $\frac{1}{T}$ is a straight line with slope equal to $-\frac{Ea}{R}$
[24-Jan-2023 Shift 1]

Answer: 3

Solution:

A : $k = Ae^{-\frac{Ea}{RT}}$
As Ea increases k decreases
B : Temperature coefficient $= \frac{k_{T+10}}{k_T}$

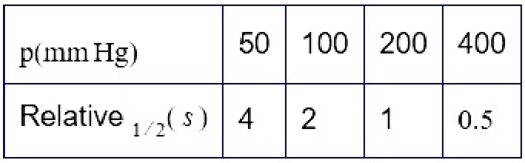
C :



Option (C) is wrong. Δk may be greater or lesser depending on temperature.
D : $\ln k = \ln A - \frac{Ea}{RT}$

Question11

A student has studied the decomposition of a gas AB_3 at 25°C. He obtained the following data.



$p(\text{mm Hg})$	50	100	200	400
Relative $_{1/2}(s)$	4	2	1	0.5

The order of the reaction is
[24-Jan-2023 Shift 2]

Options:

A. 0.5

B. 2

C. 1

D. 0 (zero)

Answer: B

Solution:

$$t_{1/2} \propto (P_o)^{1-n}$$

$$\frac{(t_{1/2})_1}{(t_{1/2})_2} = \frac{(P_0)_1^{\ 1-n}}{(P_{0_2})_2^{\ 1-n}}$$

$$\Rightarrow \left(\frac{4}{2}\right) = \left(\frac{50}{100}\right)^{1-n}$$

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

$$\Rightarrow 2 = (2)^{n-1}$$

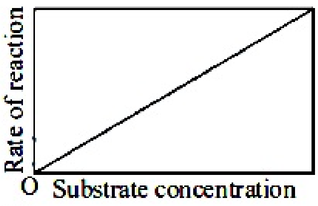
$$\Rightarrow n - 1 = 1$$

$$\Rightarrow n = 2$$

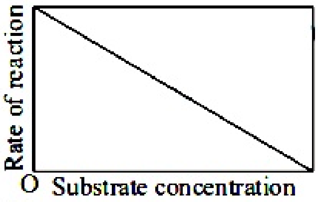
Question12

The variation of the rate of an enzyme catalyzed reaction with substrate concentration is correctly represented by graph

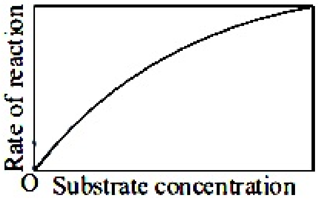
(A)



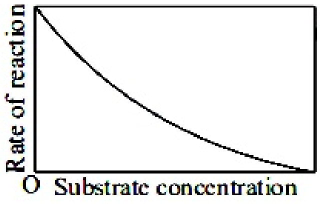
(B)



(C)



(D)



[25-Jan-2023 Shift 1]

Options:

A. b

B. c

C. d

D. a

Answer: B

Solution:

Fact base.

Question13

For the first order reaction A → B Br the half life is 30 mm. The time taken for 75% completion of the reaction is _______ mm. (Nearest mteger)
Given : $\log 2 = 0.3010$
$\log 3 = 0.4771$
$\log 5 = 0.6989$
[25-Jan-2023 Shift 1]

Answer: 60

Solution:

$t_{1/2} = T_{50} = 30\,\text{min}$
$T_{75} = 2t_{1/2} = 30 \times 2 = 60\,\text{min}$

Question14

A first order reaction has the rate constant, $k = 4.6 \times 10^{-3}\ s^{-1}$. The number of correct statement / s from the following is/are ________.
Given : $\log 3 = 0.48$
A. Reaction completes in 1000 s.
B. The reaction has a half-life of 500 s.
C. The time required for 10% completion is 25 times the time required for 90% completion.
D. The degree of dissociation is equal to $(1 - e^{-kt})$.
E. The rate and the rate constant have the same unit.
[25-Jan-2023 Shift 2]

Answer: 1

Solution:

$t_{10\%} = \frac{1}{K}\ln\left(\frac{a}{a-x}\right) = \frac{1}{K}\ln\left(\frac{100}{90}\right)$

$t_{10\%} = \frac{2.303}{K}(\log 10 - \log 9)$

$t_{10\%} = \frac{2.093}{K} \times (0.04)$

Similarly

$t_{90\%} = \frac{1}{K}\ln\left(\frac{100}{10}\right)$

$t_{90\%} = \frac{2.303}{K}$

$\frac{t_{90\%}}{t_{10\%}} = \frac{1}{0.04} = 25$

$e^{kt} = \frac{a}{a-x}$

$\frac{a-x}{a} = e^{-kt}$

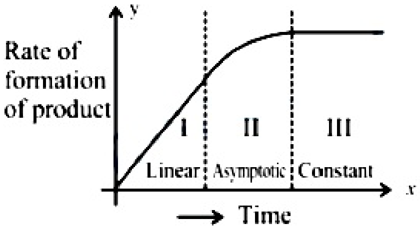
$1 - \frac{x}{a} = e^{-kt}$

$x = a(1 - e^{-kt})$

$\alpha = \frac{x}{a} = (1 - e^{-kt})$

Question15

For certain chemical reaction $X \rightarrow Y$, the rate of formation of product is plotted against the time as shown in the figure. The number of Correct statement/s from the following is ________



(A) Over all order of this reaction is one
(B) Order of this reaction can't be determined
(C) In region-I and III, the reaction is of first and zero order respectively
(D) In region-II, the reaction is of first order
(E) In region-II, the order of reaction is in the range of 0.1 to 0.9.
[29-Jan-2023 Shift 1]

Answer: 2

Solution:

Only option (B) is correct as order cannot be determined

Question16

For conversion of compound A → B, the rate constant of the reaction was found to be 4.6×10^{-5} $\mathrm{Lmol^{-1}\ s^{-1}}$. The order of the reaction is ______. [29-Jan-2023 Shift 2]

Answer: 2

Solution:

As unit of rate constant is $(\text{conc.})^{1-n}$ time^{-1}

$\Rightarrow (\mathrm{L\,mol^{-1}}) \Rightarrow 1 - n = -1$

$n = 2$

Question17

If compound A reacts with B following first order kinetics with rate constant 2.011×10^{-3} $\mathrm{s^{-1}}$. The time taken by A (in seconds) to reduce from 7g to 2g will be _______. (Nearest Integer)
[$\log 5 = 0.698$, $\log 7 = 0.845$, $\log 2 = 0.301$]
[30-Jan-2023 Shift 1]

Answer: 623

Solution:

$A + B \rightarrow P$

$t = t = t \ \ 2g$

at constant volume

$$t = \frac{2.303}{K}\log\frac{[A]_0}{[A]_t}$$

$$= \frac{2\cdot 303}{2\cdot 011\times 10^{-3}}\log\frac{7}{2}$$

$$= \frac{2\cdot 303\times 0\cdot 544}{2\cdot 011\times 10^{-3}}$$

$= 622.989$

≈ 623

Question18

An organic compound undergoes first order decomposition. If the time taken for the 60% decomposition is 540 s, then the time required for 90% decomposition will be is ______ S. (Nearest integer).
Given : $\ln 10 = 2.3$; $\log 2 = 0.3$
[30-Jan-2023 Shift 2]

Answer: 1350

Solution:

$$\frac{t_1}{t_2} = \frac{\frac{1}{K}\ln\frac{a_0}{0.4a_0}}{\frac{1}{K}\ln\frac{a_0}{0.1a_0}}$$

$$\frac{540}{t_2} = \frac{\ln\frac{10}{4}}{\ln 10}$$

$$\frac{540}{t_2} = \frac{\log 10 - \log 4}{\log 10}$$

$$\frac{540}{t_2} = \frac{1 - 0.6}{1}$$

$$\Rightarrow \frac{540}{t_2} = 0.4$$

$$\Rightarrow t_2 = \frac{540}{0.4} = 1350 \text{ sec}$$

Question19

$A \rightarrow B$
The rate constants of the above reaction at 200K and 300K are $0.03\,\text{min}^{-1}$ and $0.05\,\text{min}^{-1}$ respectively. The activation energy for the reaction is ________ J (Nearest integer)
(Given : In 10 = 2.3
$R = 8.3\,\text{JK}^{-1}\text{mol}^{-1}$
$\log 5 = 0.70$
$\log 3 = 0.48$
$\log 2 = 0.30$
[31-Jan-2023 Shift 1]

Answer: 2520

Solution:

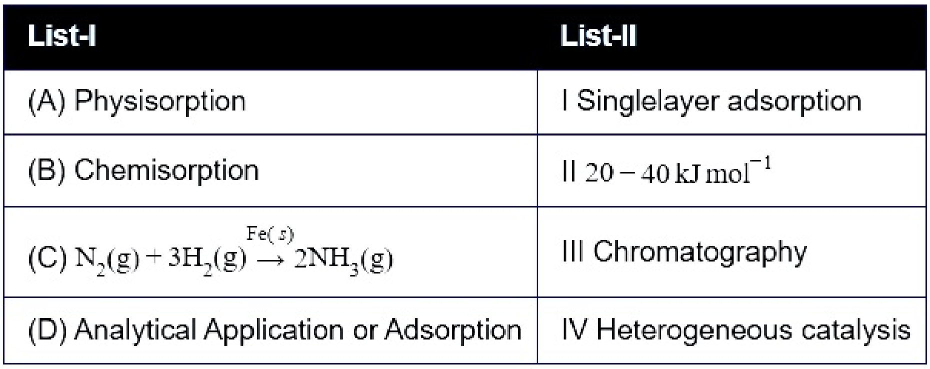
$$\log \frac{K_{300}}{K_{200}} = \frac{E_a}{2.3 \times 8.314}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\log \frac{0.05}{0.03} = \frac{Ea}{2.305 \times 8.314} \times \left[\frac{1}{200} - \frac{1}{300}\right]$$

$E_a = 2519.88J \Rightarrow E_a = 2520J$

Question20

Match List-I with List-II



List-I	List-II
(A) Physisorption	I Singlelayer adsorption
(B) Chemisorption	II $20-40\,\mathrm{kJ\,mol^{-1}}$
(C) $N_2(g)+3H_2(g)\xrightarrow{Fe(s)}2NH_3(g)$	III Chromatography
(D) Analytical Application or Adsorption	IV Heterogeneous catalysis

Choose the correct answer from the options given below :
[31-Jan-2023 Shift 2]

Options:

A. A - II, B - III, C - I, D - IV

B. A - III, B - IV, C - I, D - II

C. A – IV, B – II, C– III, D - I

D. A - II, B - I, C - IV, D - III

Answer: D

Solution:

Solution:
(A) Physisorption $= 20-40\,\mathrm{kJ/mol}$ and Chemisorption $= 80-240\,\mathrm{kJ/mol}$
(B) Physisorption is multi-layered and chemisorption is unimolecular layered.
(C) In heterogeneous catalysis, medium and catalyst are in different phases.
(D) Chromatography uses adsorption to purify/separate mixtures.

Question21

The rate constant for a first order reaction is 20min^{-1}. The time required for the initial concentration of the reactant to reduce to its $\frac{1}{32}$ level is _______ $\times 10^{-2}$ min. (Nearest integer)
(Given :
$\ln 10 = 2.303$
$\log 2 = 0.3010$)
[31-Jan-2023 Shift 2]

:

Answer: 17

Solution:

$$C = \frac{C_o}{2^n} = \frac{C_o}{32}$$

$n = 5$

$t = 5t_{1/2}$

$$= \frac{5 \times 0.693}{20} = \frac{0.693}{4}$$

$= 0.17325 \text{ min} = 17.325 \times 10^{-2} \text{ min}.$

Question22

A and B are two substances undergoing radioactive decay in a container. The half life of A is 15 min and that of B is 5 min. If the initial concentration of B is 4 times that of A and they both start decaying at the same time, how much time will it take for the concentration of both of them to be same? ________ min.
[1-Feb-2023 Shift 1]

Answer: 15

Solution:

$[A]_t = [A]_0 e^{-kt}$

For A : Let $[A]_t$ be y and $[A]_0$ be x; $k = \frac{\ln 2}{t_{1/2}} =$

$\frac{\ln 2}{15 \text{ min}}$

$y = xe\, e^{-kt}$

$= xe^{-\left(\frac{\ln 2}{15}\right)t}$

For B : $[B]_t = [B]_0 e^{-kt}$

Let $[B]_t = y$; $[B]_0 = 4x$; $k = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{5 \text{ min}}$

$y = 4xe^{-\left(\frac{\ln 2}{5}\right)t}$

$\Rightarrow \quad xe^{-\left(\frac{\ln 2}{15}\right)t} = 4xe^{-\left(\frac{\ln 2}{5}\right)t}$

$e^{t\left(\frac{\ln 2}{5} - \frac{\ln 2}{15}\right)} = 4$

$t \times \left[\frac{\ln 2}{5} - \frac{\ln 2}{15}\right] = \ln 4$

$t \times \ln 2\left[\frac{1}{5} - \frac{1}{15}\right] = 2 \ln 2$

$t = 15 \text{ min}$

Question23

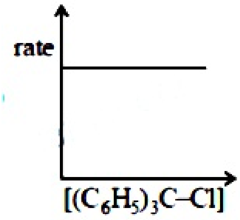
The graph which represents the following reaction is:

$$(C_6H_5)_3C-Cl \xrightarrow[\text{Pyridine}]{OH^-} (C_6H_5)_3C-OH$$

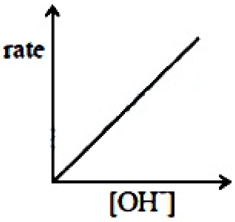
[1-Feb-2023 Shift 2]

Options:

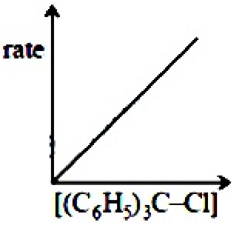
A.



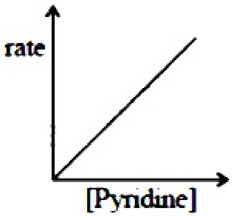
B.



C.



D.



Answer: C

Solution:

Solution:
(It is SN1 reaction so rate of reaction depends on the concentration of alkyl halide only.

Question24

A → B

The above reaction is of zero order. Half life of this reaction is 50 min. The time taken for the concentration of A to reduce to one-fourth of its initial value is ______ min.
(Nearest integer)
[1-Feb-2023 Shift 2]

Answer: 75

Solution:

Assume reaction starts with 1 mole A

$\left(t_{1/2} = \frac{a}{2k}, K = \frac{1}{2 \times 50}. \right.$

For 75% completion

$$a - \frac{a}{4} = kt$$

$$t = \frac{3}{4}\frac{a}{k} = \frac{3}{4} \times \frac{100}{a} = 75$$

Question25

For the adsorption of hydrogen on platinum, the activation energy is 30 kJ mol^{-1} and for the adsorption of hydrogen on nickel, the activation energy is 41.4 kJ mol^{-1}. The logarithm of the ratio of the rates of chemisorption on equal areas of the metals at 300K is
(Nearest integer)
Given: ln 10 = 2.3
R = 8.3JK^{-1}mol^{-1}
[6-Apr-2023 shift 1]

Answer: 2

Solution:

K = Ae

$$K_1 = Ae^{-\frac{E_a}{RT}}$$

$$K_2 = Ae^{-\frac{(E_a)_1}{RT}}$$

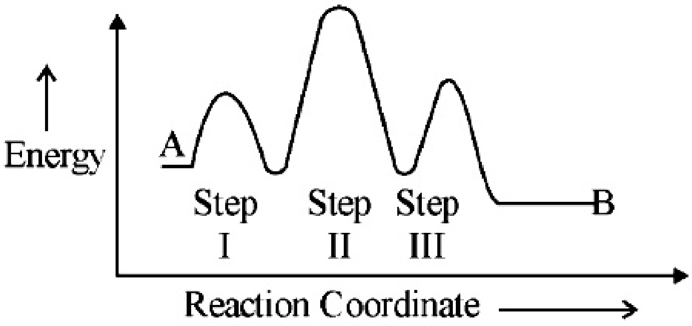
$$\frac{K_2}{K_1} = e^{\frac{(E_a)_2\,)_1 - (E_a)_2}{RT}}$$

$$\log\frac{K_2}{K_1} = \frac{(E_a)_1 - (E_a)_2}{2.3\,RT}$$

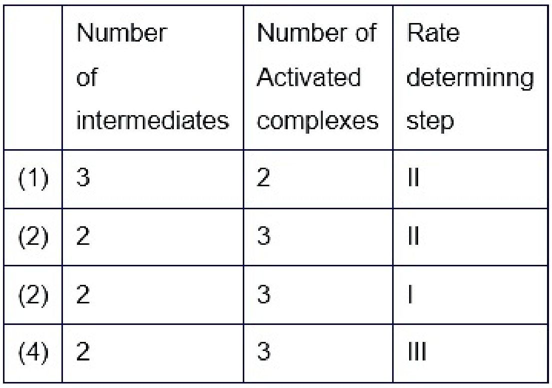
$= \frac{(41.4 - 30) \times 1000}{2.3 \times 8.3 \times 300} = 1.99$

Question26

Consider the following reaction that goes from A to B in three steps as shown below:-



Choose the correct



	Number of intermediates	Number of Activated complexes	Rate determinng step
(1)	3	2	II
(2)	2	3	II
(2)	2	3	I
(4)	2	3	III

[6-Apr-2023 shift 2]

Options:

A. (1)

B. (2)

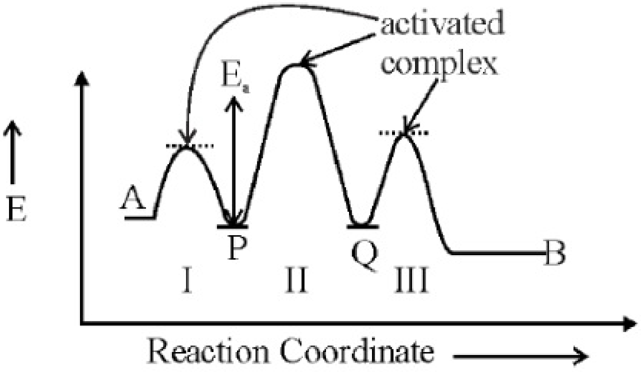
C. (3)

D. (4)

Answer: B

Solution:

Step with highest activation energy is RDS, so step II is RDS
No. of activated complex = 3



P and Q are intermediates
(Number of intermediates = 2)

Question27

The number of given statements/s which is/are correct is ________ [8-Apr-2023 shift 1]

Options:

A. The stronger the temperature dependence of the rate constant, the higher is the activation energy.

B. If a reaction has zero activation energy, its rate is independent of temperature.

C. The stronger the temperature dependence of the rate constant, the smaller is the activation energy

D. If there is no correlation between the temperature and the rate constant then it means that the reaction has negative activation energy.

Answer: B

Solution:

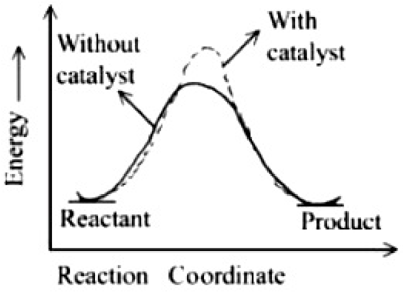
Clearly, if $E_a = 0$, K is temperature independent
if $E_a > 0$, K increase with increase in temperature
if $E_a < 0$, K decrease with increase in temperature

Question28

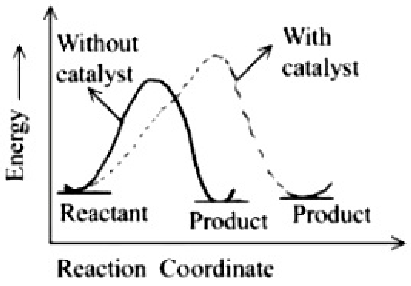
The correct reaction profile diagram for a positive catalyst reaction. [8-Apr-2023 shift 2]

Options:

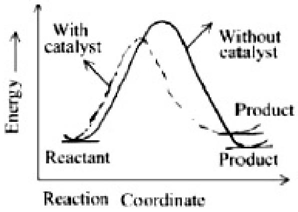
A.



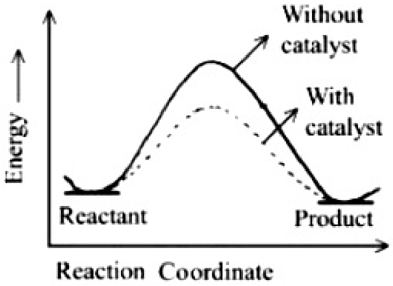
B.



C.



D.



Answer: D

Solution:

Solution:
Catalysts decrease activation energy only.

Question29

A molecule undergoes two independent first order reactions whose respective half lives are 12 min and 3 min. If both the reactions are occurring then the time taken for the 50% consumption of the reactant is _______ min. (Nearest integer) [10-Apr-2023 shift 1]

Answer: 2

Solution:

$k_{\text{eff}} = k_1 + k_2$

$$\frac{\ell n^2}{t_{eff}} = \frac{\ell n^2}{t_1} + \frac{\ell n^2}{t_2}$$

$$\frac{1}{t_{eff}} = \frac{1}{12} + \frac{1}{3} = \frac{1+4}{12} = 5/12$$

$$t_{eff} = \frac{12}{5} = 2.4 = 2$$

Question30

The number of incorrect statement/s from the following is
A. The successive half lives of zero order reactions decreases with time.
B. A substance appearing as reactant in the chemical equation may not affect the rate of reaction
C. Order and molecularity of a chemical reaction can be a fractional number
D. The rate constant units of zero and second order reaction are $mol\,L^{-1}\ s^{-1}$ and $mol^{-1}Ls^{-1}$ respectively.
[10-Apr-2023 shift 2]

Answer: 1

Solution:

Solution:

(A) For zero order $t_{1/2} = \frac{[A]_0}{2K}$ as concentration decreases half life decreases (Correct statement)
(B) If order w.r.t. that reactant is zero then it will not affect rate of reaction. (Correct statement)
(C) Order can be fractional but molecularity can not be (Incorrect statement)
(D) For zero order reaction unit is mol L^{-1-1} and for second order reaction unit is $mol^{-1}Ls^{-1}$ (Correct statement)

Question31

$KClO_3 + 6FeSO_4 + 3H_2SO_4 \rightarrow KCl + 3Fe_2(SO_4)_3 + 3H_2O$
The above reaction was studied at 300K by monitoring the concentration of $FeSO_4$ in which initial concentration was 10M and after half an hour became 8.8M. The rate of production of $Fe_2(SO_4)_3$ is ________ $\times 10^{-6}\,mol\,L^{-1}\ s^{-1}$
[11-Apr-2023 shift 1]

Answer: 333

Solution:

$$\frac{-\Delta \mathrm{FeSO_4}}{\Delta t} = \frac{10-8.8}{30\times 60} = \frac{1.2}{1800}$$

From given equation :

$$-\frac{1}{6}\frac{\Delta \mathrm{FeSO_4}}{\Delta t} = \frac{1}{3}\times(\text{ Rate of production of } \mathrm{Fe_2(SO_4)_3})$$

Rate of production of $\mathrm{Fe_2(SO_4)_3} = \frac{3}{6}\times\frac{1.2}{1800}$

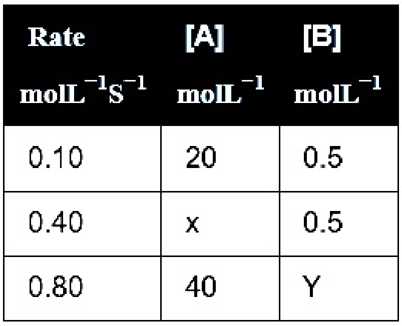
$= \frac{1}{3}\times 10^{-3}$

$= \frac{1000}{3}\times 10^{-6}$

$= 333.33\times 10^{-6}$

Question32

For a chemical reaction A + B→ Product, the order is 1 with respect to A and B.



Rate $\mathrm{molL^{-1}S^{-1}}$	[A] $\mathrm{molL^{-1}}$	[B] $\mathrm{molL^{-1}}$
0.10	20	0.5
0.40	x	0.5
0.80	40	Y

What is the value of x and y ?
[11-Apr-2023 shift 2]

Options:

A. 80 and 2

B. 40 and 4

C. 80 and 4

D. 160 and 4

Answer: A

Solution:

$r = K[A]^1[B]^1$

$0.1 = K(20)^1(0.5)^1$ . . . (i)

$0.40 = K(x)^1(0.5)^1$ . . . (ii)

$0.80 = K(40)^1(y)^1$ . . . (iii)

From (i) and (ii)

$x = 80$

From (i) and (iii)

$y = 2$

Question33

The reaction $2\,NO + Br_2 \rightarrow 2\,NOBr$ takes places through the mechanism given below:

$NO + Br_2 \Leftrightarrow NOBr_2$ (fast)

$NOBr_2 + NO \rightarrow 2\,NOBr$ (slow)

The overall order of the reaction is ______.

[12-Apr-2023 shift 1]

Answer: 3

Solution:

$RDS : NOBr_2 + NO \rightarrow 2\,NOBr$

$r = K[NOBr_2][NO] \ldots\ldots (i)$

$Keq = \frac{[NOBr_2]}{[\ NO_2[B_2].} \ldots\ldots (ii)$

From (i) & (ii)

$R = K\,.\,Keq \cdot [NO][Br_2][NO]$

$R = K'[NO]^2[Br_2]$

Overall order $= 3$

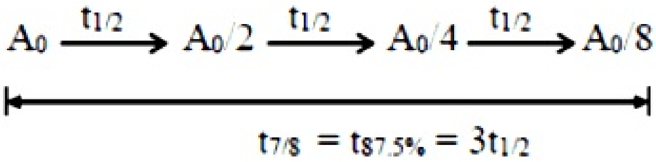
Question34

$t_{87.5}$ is the time required for the reaction to undergo 87.5% completion and t_{50} is the time required for the reaction to undergo 50% completion. The relation between $t_{87.5}$ and t_{50} for a first order reaction is ________ $t_{87.5} = x \times t_{50}$. The value of x is ________ . (Nearest integer)

[13-Apr-2023 shift 1]

Answer: 3

Solution:



Question35

A(g) → 2B(g) + C(g) is first order reaction. The initial pressure of the system was found to be 800 mm Hg which increased to 1600 mm Hg after 10 min. The total pressure of the system after 30 min will be _____ mm Hg. (Nearest integer)
[13-Apr-2023 shift 2]

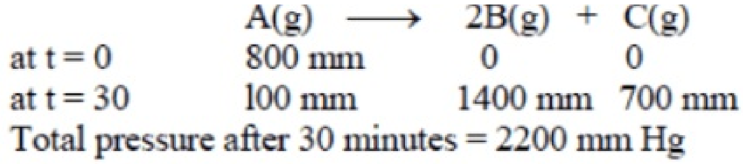
Answer: 2200

Solution:

$t_{\frac{1}{2}} = 10 \text{ minutes}$

$(P_A)_{30\,\text{min}} = (P_A)_0 \left(\frac{1}{2}\right)^{30/10}$

$(P_A)_{30\,\text{min}} = 100 \,\text{mm Hg}$



	$A(g)$	$\longrightarrow$	$2B(g)$	+	$C(g)$
at t = 0	800 mm		0		0
at t = 30	100 mm		1400 mm		700 mm

Total pressure after 30 minutes = 2200 mm Hg

Question36

For a reversible reaction A ⇌ B, the $\Delta H_{\text{forward reaction}} = 20\,\text{kJ mol}^{-1}$. The activation energy of the uncatalysed forward reaction is $300\,\text{kJ mol}^{-1}$. When the reaction is catalysed keeping the reactant concentration same, the rate of the catalyzed forward reaction at 27°C is found to be same as that of the uncatalysed reaction at 327°C. The activation energy of the catalyzed backward reaction is _______ KJ mol^{-1}.
[15-Apr-2023 shift 1]

Answer: 130

Solution:

$E_a = 300\,\mathrm{kJ\,mol^{-1}}$

$\frac{E_a}{T} = \frac{E_a'}{T'}$

(Since rate of catalyzed and uncatalysed reaction is same)

$\frac{300}{600} = \frac{E_{a,f}'}{300}$

$E_{a,f}' = 150$

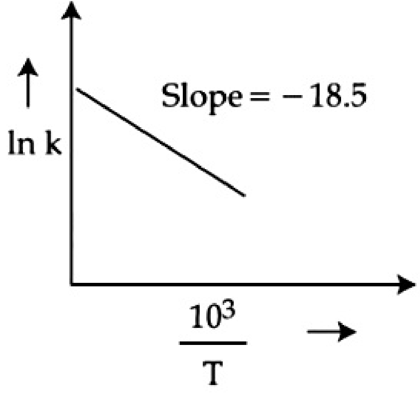
$20 = 150$

$20 = 150 - E_{a,b}'$

$E_{a,b}' = 130$

Question37

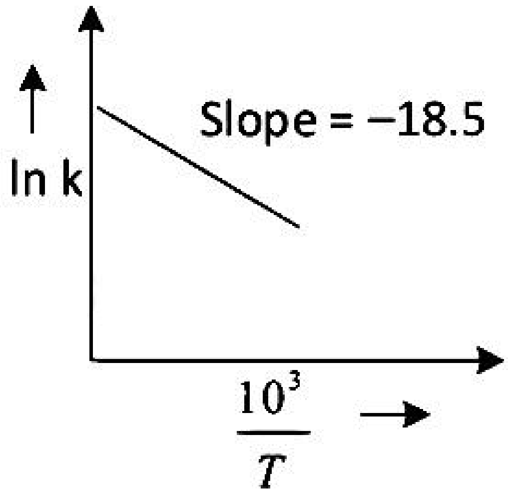
The rate constants for decomposition of acetaldehyde have been measured over the temperature range 700 − 1000K. The data has been analysed by plotting $\ln k$ vs $\frac{10^3}{T}$ graph. The value of activation energy for the reaction is $\mathrm{kJ\,mol^{-1}}$. ________(Nearest integer)
(Given : $R = 8.31\,\mathrm{JK^{-1}mol^{-1}}$)



[24-Jun-2022-Shift-1]

Answer: 154

Solution:



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

∴ Slope of the graph $= -\dfrac{E_a}{R \times 10^3} = -18.5$

$\therefore E_a = 18.5 \times 8.31 \times 1000 \simeq 154\,\text{kJ mol}^{-1}$

Question38

For a given chemical reaction
$\gamma_1 A + \gamma_2 B \rightarrow \gamma_3 C + \gamma_4 D$
Concentration of C changes from $10\,\text{mmol dm}^{-3}$ to $20\,\text{mmol dm}^{-3}$ in 10 seconds. Rate of appearance of D is 1.5 times the rate of disappearance of B which is twice the rate of disappearance A. The rate of appearance of D has been experimentally determined to be $9\,\text{mmol dm}^{-3}\,\text{s}^{-1}$. Therefore, the rate of reaction is ___ $\text{mmol dm}^{-3}\,\text{s}^{-1}$ · (Nearest Integer)
[25-Jun-2022-Shift-1]

Answer: 1

Solution:

$$\text{Rate} = \frac{1}{r_1}\left(\frac{-d[A]}{dt}\right) = \frac{1}{r_2}\left(\frac{-d[B]}{dt}\right) = \frac{1}{r_3}\left(\frac{-d[C]}{dt}\right)$$

$$= \frac{1}{r_4}\left(\frac{d[D]}{dt}\right)$$

$$\frac{d[D]}{dt} = \frac{r_4}{r_2}\left(\frac{-d[B]}{dt}\right)$$

$$\frac{r_4}{r_2} = \frac{3}{2}$$

$$\frac{-d[B]}{dt} = \frac{r_2}{r_1}\left(\frac{-d[A]}{dt}\right) \Rightarrow \frac{r_2}{r_1} = 2$$

$$r_2 = 2r_1$$

$$r_4 = 1.5r_2 = 3r_1$$

$$\frac{d[C]}{dt} = 1\text{m} \cdot \text{mol dm}^{-3}\,\text{sec}^{-1}$$

$$\frac{d[D]}{dt} = \frac{r_4}{r_3} \cdot \frac{d[C]}{dt} \Rightarrow \frac{r_4}{r_3} = 9$$

$$r_4 = 9r_3 = 3r_1$$

$$\Rightarrow r_1 = 3r_3$$

$$3r_3 A + 6r_3 B \rightarrow r_3 C + 9r_3 D$$

∴ rate of reaction $= \dfrac{1}{9} \times 9\text{m} \cdot \text{mol dm}^{-3}\,\text{sec}^{-1}$

$= 1\text{m} \cdot \text{mol dm}^{-3}\,\text{sec}^{-1}$

Question39

At 345K, the half life for the decomposition of a sample of a gaseous compound initially at $55.5\,\text{kPa}$ was 340 s. When the pressure was $27.8\,\text{kPa}$, the half life was found to be 170s. The order of the reaction is

_____. [integer answer]
[25-Jun-2022-Shift-2]

Answer: 0

Solution:

$t_{1/2} \times \frac{1}{[P_0]^{n-1}}$

$\frac{t_1}{t_2} = \frac{(P_2)^{n-1}}{(P_1)^{n-1}}$

$\frac{340}{170} = \left(\frac{27.8}{55.5}\right)^{n-1}$

$\Rightarrow 2 = \frac{1}{(2)^{n-1}}$

$n = 0$

Question40

A flask is filled with equal moles of A and B. The half lives of A and B are 100 s and 50 s respectively and are independent of the initial concentration. The time required for the concentration of A to be four times that of B is____ s.
(Given : $\ln 2 = 0.693$)
[26-Jun-2022-Shift-1]

Answer: 200

Solution:

$k_A = \frac{\ln 2}{100}; k_B = \frac{\ln 2}{50}$

$A_t = A_0 \times e^{-k_A t}$

$A_t = A_0 \times e^{\left(\frac{-\ln 2}{100} \times t\right)}$

$B_t = B_0 \times e^{\left(\frac{-\ln 2}{50} \times t\right)}$

$A_0 = B_0$

$\& A_t = 4B_t$

$e^{-\frac{\ln 2}{100} \times t} = 4 \times e^{-\frac{\ln 2}{50} \times t}$

$e^{\frac{\ln 2}{100} \times t} = 4$

$e^{\frac{\ln 2}{100} \times t} = 4$

$\frac{\ln 2}{100} \times t = \ln 4 = 2 \ln 2$
$t = 200\,\text{sec}$

Question41

Catalyst A reduces the activation energy for a reaction by $10\,\text{kJ mol}^{-1}$ at 300K. The ratio of rate constants, $\frac{{}^{k}T,\ \text{Catalysed}}{{}^{k}T,\ \text{Uncatalysed}}$ is e^{x}. The value of x is_____ [nearest integer]
[Assume that the pre-exponential factor is same in both the cases. Given $R = 8.31 JK^{-1} mol^{-1}$]
[26-Jun-2022-Shift-2]

Answer: 4

Solution:

$$K = Ae\,e^{\frac{-Ea^{2}}{RT}}$$

$$K_{cat} = Ae^{\frac{E_a{}^{1}}{RT}},\quad K_{uncat.}{}^{uthathon} = Ae^{\frac{Ea}{RT}}$$

$$\frac{K_{cat}}{K_{uncat.}} = e^{E_a - \frac{E_a}{1}RT} = e^{\frac{10 \times 1000}{8.3 \times 300}} = e^{4.009} = e^{x}$$

$\therefore x = 4$

Question42

The rate constant for a first order reaction is given by the following equation:
$\ln k = 33.24 - \frac{2.0 \times 10^{4} K}{T}$
The activation energy for the reaction is given by ____kJ mol^{-1}. (In nearest integer) (Given : $R = 8.3 JK^{-1} mol^{-1}$)
[27-Jun-2022-Shift-1]

Answer: 166

Solution:

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

Given: $\ln k = 33.24 - \frac{2.0 \times 10^4}{T}$

$\therefore$ on comparing $\frac{E_A}{R} = 2.0 \times 10^4$

$\therefore E_A = 2.0 \times 10^4 \times R$

$\Rightarrow E_A = 2.0 \times 10^4 \times 8.3\text{J}$

$\Rightarrow E_A = 16.6 \times 10^4\text{J} = 166\,\text{kJ}$

Question43

It has been found that for a chemical reaction with rise in temperature by 9K the rate constant gets doubled. Assuming a reaction to be occurring at 300K, the value of activation energy is found to be _____kJ mol^{-1}. [nearest integer] (Given $\ln 10 = 2.3$, $R = 8.3\text{JK}^{-1}\text{mol}^{-1}$, $\log 2 = 0.30$) [27-Jun-2022-Shift-2]

Answer: 59

Solution:

$T_1 = 300\text{K}$

(Rate constant)

$K_2 = 2K_1$, on increase temperature by 9K

$T_2 = 309\text{K}$

$Ea = ?$

$$\log \frac{K_2}{K_1} = \frac{Ea}{2.3R}\left[\frac{T_2 - T_1}{T_2 \cdot T_1}\right]$$

$$\log 2 = \frac{Ea}{2.3 \times 8.3}\left[\frac{9}{309 \times 300}\right]$$

$$Ea = \frac{0.3 \times 309 \times 300 \times 2.3 \times 8.3}{9}$$

$= 58988.1\text{J / mole}$

$\simeq 59\,\text{kJ / mole}$

Question44

For a first order reaction A → B, the rate constant, $k = 5.5 \times 10^{-14}\text{s}^{-1}$. The time required for 67% completion of reaction is $x \times 10^{-1}$ times the half life of reaction. The value of x is____ (Nearest integer) (Given : $\log 3 = 0.4771$) [28-Jun-2022-Shift-1]

Answer: 16

Solution:

$$t_{67\%} = \frac{1}{k}\ln\left(\frac{1}{1-0.67}\right) = \frac{t_{1/2}}{\ln 2} \times \ln\left(\frac{1}{1-\frac{2}{3}}\right)$$

$$t_{67\%} = \frac{t_{1/2}}{\log 2} \times \log 3 = \frac{t_{1/2} \times 0.4771}{0.301}$$

$\Rightarrow t_{67\%} = 1.585 \times t_{1/2}$

$X \times 10^{-1} = 1.585$

$\Rightarrow X = 15.85$

Ans.16

Question45

A radioactive element has a half life of 200 days. The percentage of original activity remaining after 83 days is_______ (Nearest integer)
(Given : antilog 0.125 = 1.333, antilog 0.693 = 4.93)
[28-Jun-2022-Shift-2]

Answer: 75

Solution:

$$\lambda = \frac{2.303}{t}\log\frac{A_0}{A}$$

$$\frac{0.693}{200} = \frac{2.303}{83}\log\frac{A_0}{A}$$

$$\frac{A}{A_0} = 0.75$$

Hence, percentage of original activity remaining after 83 days is 75%

Question46

The activation energy of one of the reactions in a biochemical process is 532611Jmol^{-1}. When the temperature falls from 310 K to 300K, the change in rate constant observed is $k_{300} = x \times 10^{-3}k_{310}$. The value of x is______
[Given . : $\ln 10 = 2.3$, $R = 8.3JK^{-1}mol^{-1}$]
[29-Jun-2022-Shift-1]

Answer: 1

Solution:

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

$\ln\left(\frac{K_2}{K_1}\right) = \frac{532611}{8.3} \times \left(\frac{10}{310 \times 300}\right)$

where K_2 is at 310K&K_1 is at 300K

$\ln\left(\frac{K_2}{K_1}\right) = 6.9$

$= 3 \times \ln 10$

$\ln\frac{K_2}{K_1} = \ell n 10^3$

$K_2 = K_1 \times 10^3$

$K_1 = K_2 \times 10^3$

So $K = 1$

Question47

The equation
$k = (6.5 \times 10^{12} S^{-1}) e^{-26000K/T}$
is followed for the decomposition of compound A. The activation energy for the reaction is kJ mol^{-1}. [nearest integer]
(Given : $R = 8.314 JK^{-1} mol^{-1}$)
[29-Jun-2022-Shift-2]

Answer: 216

Solution:

$k = Ae^{\frac{-E_a}{RT}}$

$\frac{E_a}{RT} = \frac{26000}{T}$

$E_a = 26000 \times 8.314$

$= 216164 J$

$= 216 kJ$

Question48

The half life for the decomposition of gaseous compound A is 240s when the gaseous pressure was 500 Torr initially. When the pressure was 250 Torr, the half life was found to be 4.0min. The order of the reaction is (Nearest integer)

[25-Jul-2022-Shift-1]

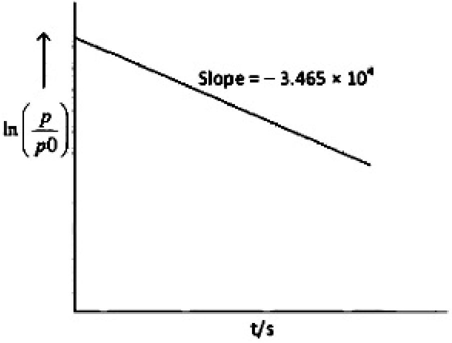
Answer: 1

Solution:

$(t_{1/2})_A = 240\text{s}$ when $P = 500$ torr
$(t_{1/2})_A = 4\,\text{min} = 4 \times 60 = 240\,\text{sec}$ when $P = 250$ torr
If means half-life is independent of concentration of reactant present.
∴ Order of reaction $= 1$

Question49

For the decomposition of azomethane.
$CH_3N_2CH_3(g) \rightarrow CH_3CH_3(g) + N_2(g)$, a first order reaction, the variation in partial pressure with time at 600K is given as



The half life of the reaction is_____ $\times 10^{-5}$s. [Nearest integer]
[25-Jul-2022-Shift-2]

Answer: 2

Solution:

For first order reaction,
$\ln A = \ln A_0 - kt$
Hence Slope $= -k$
$-k = -3.465 \times 10^4$
$k = \dfrac{0.693}{t_{1/2}}$
$3.465 \times 10^4 = \dfrac{0.693}{t_{1/2}}$
$t_{1/2} = 2 \times 10^{-5}\text{s}$

Question50

For a reaction A → 2B + C the half lives are 100s and 50s when the concentration of reactant A is 0.5 and 1.0 mol L^{-1} respectively. The order of the reaction is ______. (Nearest Integer)
[26-Jul-2022-Shift-1]

Answer: 2

Solution:

$t_{1/2} \propto \frac{1}{(a_0)^{n-1}}$
$t_{1/2} = 100\,\text{sec} \;\; a_0 = 0.5$
$t_{1/2} = 50\,\text{sec} \;\; a_0 = 1$
$\frac{100}{50} = \left(\frac{1}{0 \cdot 5}\right)^{n-1}$
$(2) = (2)^{n-1}$
$n - 1 = 1$
$n = 2$

Question51

At 30°C, the half life for the decomposition of AB_2 is 200s and is independent of the initial concentration of AB_2. The time required for 80% of the AB_2 to decompose is Given: $\log 2 = 0.30 \;\; \log 3 = 0.48$
[26-Jul-2022-Shift-2]

Options:

A. 200s

B. 323s

C. 467s

D. 532s

Answer: C

Solution:

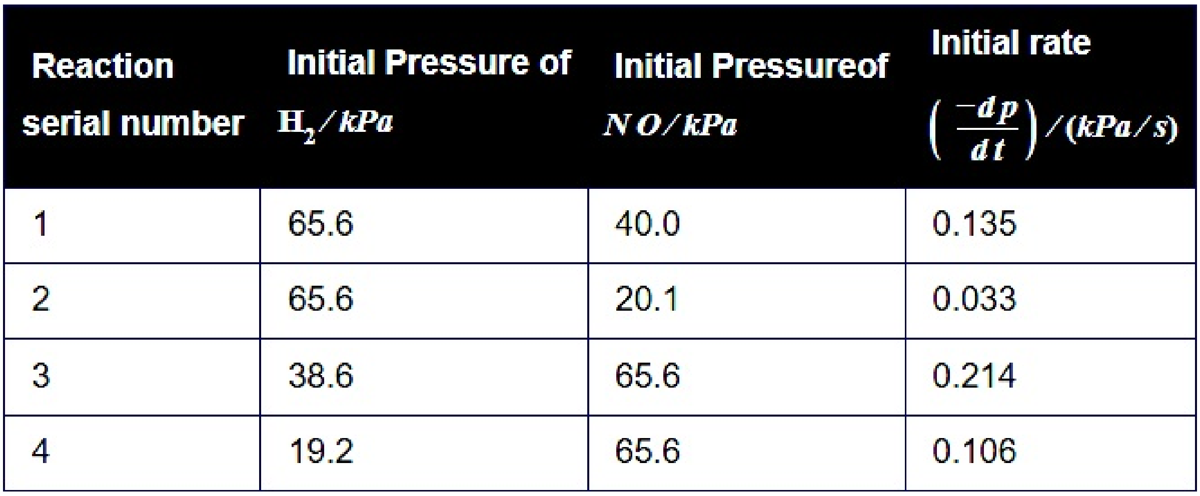
Solution:
Since, half-life is independent of the initial concentration of AB_2. Hence, the reaction is "First Order".
$k = \frac{2.303 \log 2}{t_{1/2}}$
$\frac{2.303 \log 2}{t_{1/2}} = \frac{2.303}{t} \log \frac{100}{(100-80)}$

$$\frac{2.303 \times 0.3}{200} = \frac{2.303}{t} \log 5$$
$t = 467s$

Question52

$2NO + 2H_2 \rightarrow N_2 + 2H_2O$

The above reaction has been studied at 800°C. The related data are given in the table below



Reaction serial number	Initial Pressure of H_2/kPa	Initial Pressureof NO/kPa	Initial rate $\left(\frac{-dp}{dt}\right)/(kPa/s)$
1	65.6	40.0	0.135
2	65.6	20.1	0.033
3	38.6	65.6	0.214
4	19.2	65.6	0.106

The order of the reaction with respect to NO is
[27-Jul-2022-Shift-1]

Answer: 2

Solution:

Let the rate of reaction (r) is as
$r = K[NO]^n[H_2]^m$
From 1^{st} data
$0.135 = K[40]^n \cdot (65.6)^m \quad \ldots (1)$
From 2^{nd} data
$0.033 = K(20.1)^n \cdot (65.6)^m \quad \ldots (2)$
On dividing equation (1) by equation (2)
$$\frac{0.135}{0.033} = \left(\frac{40}{20.1}\right)^n$$
$4 = (2)^n$
$\therefore n = 2$
$\therefore$ Order of reaction w.r.t. NO is 2 .

Question53

$$\underset{\text{Reactant}}{[A]} \rightarrow \underset{\text{Product}}{[B]}$$

If formation of compound [B] follows the first order of kinetics and after 70 minutes the concentration of [A] was found to be half of its initial concentration. Then the rate constant of the reaction is $x \times 10^{-6} s^{-1}$. The value of x is ________.(Nearest Integer)
[27-Jul-2022-Shift-2]

Answer: 165

Solution:

$$K = \frac{0.693}{t_{1/2}} = \frac{0.693}{70 \times 60}$$
$$= \frac{6930}{7 \times 6} \times 10^{-6}$$
$$= 165 \times 10^{-6}\ s^{-1}$$

Question54

For the given first order reaction
$A \rightarrow B$
the half life of the reaction is 0.3010 min. The ratio of the initial concentration of reactant to the concentration of reactant at time 2.0 min will be equal to ________. (Nearest integer)
[28-Jul-2022-Shift-1]

Answer: 100

Solution:

$$t_{1/2} = \frac{0.693}{K} \quad t_{1/2}\ \text{given} \ = 0.3010$$
$$K = \frac{0.693}{0.3010}$$
$$K = 2.30$$
$$K = \frac{2.303}{t} \log \frac{(A_0)}{(A_t)}$$

A_0→ initial concentration of reactant

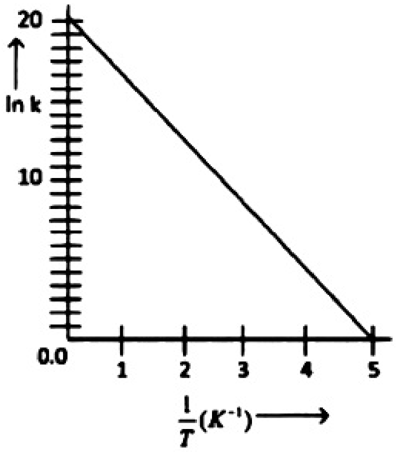
A_t→ concentration of reactant at time t

$$2.303 = \frac{2.303}{2} \log \frac{(A_0)}{(A_t)}$$
$$2 = \log \frac{(A_0)}{(A_t)}$$

$100 = \frac{A_0}{A_t}$

Question55

For a reaction, given below is the graph of $\ln k$ vs $\frac{1}{T}$. The activation energy for the reaction is equal to____ (nearest integer)
(Given : R = $2\,\text{cal}\,K^{-1}\,\text{mol}^{-1}$)



[28-Jul-2022-Shift-2]

Answer: 8

Solution:

Question56

The reaction between X and Y is first order with respect to X and zero order with respect to Y.

Experiment	$\frac{[X]}{\text{mol } L^{-1}}$	$\frac{[Y]}{\text{mol } L^{-1}}$	$\frac{\text{Initial rate}}{\text{mol L}^{-1}\text{min}^{-1}}$
I	0.1	0.1	2×10^{-3}
I	L	0.2	4×10^{-3}
III	0.4	0.4	$M \times 10^{-3}$
IV	0.1	0.2	2×10^{-3}

Examine the data of table and calculate ratio of numerical values of M and L (Nearest Integer)
[29-Jul-2022-Shift-1]

Answer: 40

Solution:

$r = k[X][Y]^0 = k[X]$

Using | \& II

$\frac{4 \times 10^{-3}}{2 \times 10^{-3}} = \left(\frac{L}{0.1}\right) \Rightarrow L = 0.2$

Using I \& III

$\frac{M \times 10^{-3}}{2 \times 10^{-3}} = \frac{0.4}{0.1} \Rightarrow M = 8$

$\frac{M}{L} = \frac{8}{0.2} = 40$

Question57

Assuming 1µg of trace radioactive element X with a half life of 30 years is absorbed by a growing tree. The amount of X remaining in the tree after 100 years is _______ $\times 10^{-1}$µg.
[Given : $\ln 10 = 2.303$; $\log 2 = 0.30$]
[29-Jul-2022-Shift-2]

Answer: 1

Solution:

$t = \frac{1}{\lambda} \ln\left(\frac{a}{a-x}\right)$

$\Rightarrow 100 = \left(\frac{30}{\ln 2}\right)\left[\ln\left(\frac{1}{w}\right)\right]$

$\Rightarrow \left[\frac{100 \times \log 2}{30}\right] = \log\left(\frac{1}{w}\right)$

$\Rightarrow 1 = \log\left(\frac{1}{w}\right)$

$\Rightarrow \frac{1}{w} = 10$

So $w = 0.1$µg

Question58

Sucrose hydrolyses in acid solution into glucose and fructose following first order rate law with a half-life of 3.33h at 25°C. After 9h, the

fraction of sucrose remaining is f . The value of $\log_{10}\left(\frac{1}{f}\right)$ is ……… $\times 10^{-2}$
(Rounded off to the nearest integer).
[Assume, $\ln 10 = 2.303$, $\ln 2 = 0.693$]
[24 Feb 2021 Shift 2]

Answer: 81

Solution:

Given, $\underset{\text{Sucrose}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{1st order}} \underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{Fructose}}{C_6H_{12}O_6}$

$t_{1/2} = \frac{10}{3}h$

$t = 0, \quad a = [A]_0 \quad$ (initial conc.)

At $\quad t = 9h \quad a - x = [A]_t$ [conc. at time t]

For using 1st order equation,

$$K = \frac{2.303}{t}\log\frac{[A]_0}{[A]_t} \Rightarrow \frac{K \times t}{2.303} = \log\frac{[A]_0}{[A]_t}$$

$$\frac{\ln 2 \times 9}{10/3 \times 2303} = \log\left(\frac{1}{F}\right) \Rightarrow \log\left(\frac{1}{F}\right) = 0.8124 \quad \left(\because k = \frac{\ln 2}{t_{1/2}}\right)$$

$$\log\left(\frac{1}{F}\right) = 81.24 \times 10^{-2}$$

$$x = 81.24 \text{ or } x \approx 81$$

Question59

The rate constant of a reaction increases by five times on increase in temperature from 27˚C to 52˚C. The value of activation energy in $kJ\,mol^{-1}$ is …….. (Rounded off to the nearest integer)
[$R = 8.314 J\,K^{-1} mol^{-1}$]
[25 Feb 2021 Shift 2]

Answer: 52

Solution:

$T_1 = (273 + 27) = 300K$, $T_2 = (273 + 52) = 325K$

Given, temperature coefficient of the reaction,

$$\alpha_T = \frac{K_{325}}{K_{300}} = 5$$

$$\log\frac{K_{T_2}}{K_{T_1}} = \frac{E_a}{2.303R} \times \left(\frac{T_2 - T_1}{T_1 T_2}\right)$$

$$\log \frac{K_{325}}{K_{300}} = \frac{E_a}{2.303 \times 8.314}\left(\frac{325-300}{300 \times 325}\right)$$

$$\log 5 = \frac{E_a}{2.303 \times 8319} \times \frac{25}{300 \times 325}$$

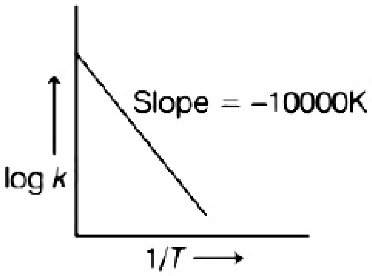
$E_a = 52194.78 \text{J mol}^{-}$

$= 52.194 \text{kJ mol}^{-1}$

$= 52 \text{kJ mol}^{-1}$

Question60

For the reaction, $aA + bB \rightarrow cC + dD$, the plot of $\log k$ vs $\frac{1}{T}$ is given below



The temperature at which the rate constant of the reaction is $10^{-4} s^{-1}$ is .... K
(Rounded off to the nearest integer).
[Given : The rate constant of the reaction is $10^{-5} s^{-1}$ at 500K]
[25 Feb 2021 Shift 1]

Answer: 526

Solution:

According to Arrhenius equation,

$$\log K = \log A - \frac{E_a}{2.303RT}$$

$$\text{Given, Slope} = \frac{E_a}{2.303R} = 10,000$$

$K_1 = 10^{-5}, K_2 = 10^{-4}$

$T_1 = 500K$

$$\log_{10}\left[\frac{K_1}{K_2}\right] = \frac{E a}{2.303R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$$

$$\log\left[\frac{10^{-4}}{10^{-5}}\right] = 10000\left[\frac{1}{500} - \frac{1}{T_2}\right]$$

$$\log 10 = 10000\left[\frac{1}{500} - \frac{1}{T_2}\right]$$

$T_2 = 526.31 \sim eq526K$

Question61

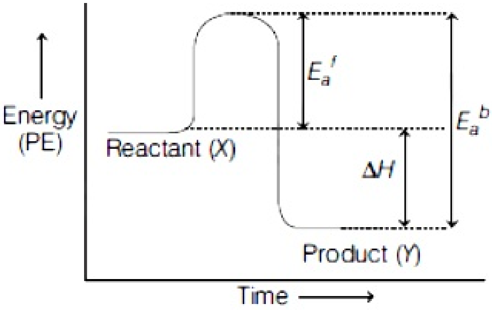
An exothermic reaction $X \rightarrow Y$ has an activation energy 30kJ mol^{-1}. If

energy change ΔE during the reaction is -20kJ mol^{-1} then the activation energy for the reverse reaction in kJ is ………..
[26 Feb 2021 Shift 1]

Answer: 50

Solution:

Solution:
$X \longrightarrow Y$, it is an exothermic reaction whose ΔE or $\Delta H = -20 \text{kJ mol}^{-1}$



Given, activation energy of the forward reaction ($X \rightarrow Y$),
$E_a^f = 30 \text{kJ mol}^{-1}$
Activation energy of the backward or reverse reaction ($Y \rightarrow X$), E_a^b can be calculated as,
$\Delta H = E_a^f - E_a^b$
$\Rightarrow E_a^b = E_a^f - \Delta H = 30 - (-20) = 50 \text{kJ mol}^{-1}$

Question62

A reaction has a half-life of 1min. The time required for 99.9% completion of the reaction is ……… min (Round off to the nearest integer).
[Use : $\ln 2 = 0.69$, $\ln 10 = 2.3$]
[18 Mar 2021 Shift 2]

Answer: 10

Solution:

Using formula,

$$k = \frac{1}{t} \ln\left(\frac{C_0}{C_t}\right)$$

As $k = \dfrac{\ln 2}{t_{1/2}}$

$t_{1/2} = 1 \text{min}$ [Given]

$$\frac{\ln 2}{1} = \frac{1}{t} \ln\left(\frac{100}{0.1}\right)$$

$\Rightarrow\ t = \dfrac{\ln 1000}{\ln 2}$

$t = \dfrac{3 \times 2.3}{0.69} = 10$

Question63

For a certain first order reaction 32% of the reactant is left after 570s. The rate constant of this reaction is ...... $\times 10^{-3} s^{-1}$. (Round off to the nearest integer).
[Given, $\log_{10} 2 = 0.301$, In 10 = 2.303]
[17 Mar 2021 Shift 1]

Answer: 2

Solution:

Solution:
For first order reaction,
$k \times t = \ln \dfrac{A_0}{A_t}$.....(i)
where, rate = rate constant
t = time
Initial conc. of reactant, $A_0 = 100$
Conc. of reactant after ' t ' time, $A_t = 32$
Put all values in equation (i)
t = time
$A_t = 32$
$k \times 570 = \ln \dfrac{100}{32}$
$k \times 570 = 2.303(\log 10^2 - \log 2^5)$
$\Rightarrow\ k = \dfrac{2.303}{570}(2 - 5 \log 2)$
$= \dfrac{2.303}{570}[2 - (5 \times 0.3)]$
$= \dfrac{2.303}{570} \times 0.5$
$k = 2 \times 10^{-3} s^{-1}$

Question64

The decomposition of formic acid on gold surface follows first order kinetics. If the rate constant at 300K is $1.0 \times 10^{-3} s^{-1}$ and the activation energy, $E_a = 11.488 kJ\ mol^{-1}$, the rate constant at 200K is ...... $\times 10^{-5} s^{-1}$ (Round off to the nearest integer).
[Given, $R = 8.314 J\ mol^{-1} K^{-1}$]
[16 Mar 2021 Shift 1]

Answer: 10

Solution:

Solution:
Given, $T_1 = 200K$, $T_2 = 300K$
$E_a =$ Activation energy $= 11.488 kJ / mol$
Rate constant (k_2) at 300K $= 10^{-3} s^{-1}$
To find = rate constant (k_1) at 200K
We know,

$$\log \frac{k_2}{k_1} = \frac{E}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\log \frac{10^{-3}}{k_1} = \frac{11.488 \times 10^3}{2.303 \times 8.314}\left(\frac{1}{200} - \frac{1}{300}\right)$$

$$\log \frac{10^{-3}}{k_1} = 1$$

$$\log \frac{10^{-3}}{k_1} = \log 10$$

$$\Rightarrow k_1 = 10^{-4}$$

$$= 10 \times 10^{-5} s^{-1}$$

Question65

A and B decompose via first order kinetics with half-lives 54.0 min and 18.0min respectively. Starting from an equimolar non-reactive mixture of A and B, the time taken for the concentration of A to become 16 times that of B is .... min. (Round off to the nearest integer).
[16 Mar 2021 Shift 2]

Answer: 108

Solution:

Solution:
Given, $(t_{1/2})_1 = 54 min$
$(t_{1/2})_2 = 18 min$

$\quad$ A $\qquad$ B

$t = 0, 'x' M \quad t = 0, 'x' M$

To calculate, $[A_t] = 16 \times [B_t]$...(i)
time = ?

For Ist order reaction, $[A_t] = \dfrac{A_0}{(2)^n}$

n = number of half-lives
From Eq. (i), $[A_t] = 16 \times [B_t]$

$$\frac{x}{(2)^{n_1}} = \frac{x}{(2)^{n_2}} \times 16$$

$$(2)^{n_2} = (2)^{n_1} \times (2)^4$$

$$\Rightarrow n_2 = n_1 + 4$$

$$\frac{t}{(t_{1/2})_2} = \frac{t}{(t_{1/2})_1} + 4 \Rightarrow t\left(\frac{1}{18} - \frac{1}{54}\right) = 4$$

$$\Rightarrow\ t = \frac{4 \times 18 \times 54}{36} = 108\text{min}$$

Question66

The reaction $2A + B_2 \longrightarrow 2AB$ is an elementary reaction. For a certain quantity of reactants, if the volume of the reaction vessel is reduced by a factor of 3 , the rate of the reaction increases by a factor of (Round off to the nearest integer).
[17 Mar 2021 Shift 2]

Answer: 27

Solution:

Solution:
Given, the reaction $2A + B_2 \longrightarrow 2AB$ is an elementary reaction.
So, rate of reaction
$= k[A]^2[B_2]$
Initial rate $= k(a / V)^2(b/V)^1$
where, a and b are the number of moles of A and B and V is the volume.
On reducing the volume by a factor of 3 , the concentrations of A and B_2 will become 3 times
Final rate $= k(3a / V)^2(3b/V)^1\ 3^2 \times 3k\left(\frac{a}{V}\right)^2\left(\frac{b}{V}\right)^1 = 27\times$ initial rate
Hence, the rate becomes $3^2 \times 3 = 27$ times of initial rate.

Question67

The inactivation rate of a viral preparation is proportional to the amount of virus. In the first minute after preparation, 10% of the virus is inactivated. The rate constant for viral inactivation is __________$\times 10^{-3}\text{min}^{-1}$. (Nearest integer)
[Use : $\ln 10 = 2.303$; $\log_{10}3 = 0.477$;property of logarithm :$\log x^y = y \log x$]
[20 Jul 2021 Shift 1]

Answer: 106

Solution:

As the unit of rate constant is min^{-1} so it must be a first order reaction
$\mathrm{K} \times \mathrm{t} = 2.303 \log \mathrm{A}_0 / \mathrm{A}_t$
in 1 min 10% is in activated so tabing
$\mathrm{A}_0 = 100 \quad \mathrm{A}_t = 90$ in $1 \mathrm{min}$
So $\mathrm{K} \times 1 = 2.303 \times \log \frac{100}{90}$
$= 2.303 \times (\log 10 - 2 \log 3)$
$= 2.303 \times (1 - 2 \times 0.477)$
$= 0.10593$
$= 105.93 \times 10^{-3}$
≈ 106

Question68

For a chemical reaction A → B, it was found that concentration of B is increased by $0.2 \mathrm{mol}\ \mathrm{L}^{-1}$ in 30 min.
The average rate of the reaction is ________ $\times 10^{-1} \mathrm{mol}\ \mathrm{L}^{-1} \mathrm{h}^{-1}$. (in nearest integer)
[25 Jul 2021 Shift 2]

Answer: 4

Solution:

	A	→	B
	$t = 0$		0
	$t = 30 \mathrm{min}$		$0.2 \mathrm{M}$

Av. rate of reaction $= -\frac{\Delta[\mathrm{A}]}{\Delta \mathrm{t}} = \frac{\Delta[\mathrm{B}]}{\Delta \mathrm{t}} = \frac{(0.2 - 0)}{\frac{1}{2}}$

$= 0.4 = 4 \times 10^{-1} \mathrm{mol}\ / \mathrm{L} \times \mathrm{hr}$

Question69

$\mathrm{PCl}_5(\mathrm{g}) \rightarrow \mathrm{PCl}_3(\mathrm{g}) + \mathrm{Cl}_2(\mathrm{g})$
In the above first order reaction the concentration of PCl_5 reduces from initial concentration 50 mol L^{-1} to 10 mol L^{-1} in 120 minutes at $300 \sim \mathrm{K}$. The rate constant for the reaction at $300 \mathrm{K}$ is $\mathrm{x} \times 10^{-2} \mathrm{min}^{-1}$. The value of x is _______.
[Given $\log 5 = 0.6989$]
[20 Jul 2021 Shift 2]

Answer: 1

Solution:

$PCl_5(g) \xrightarrow[300K]{\text{I order}} PCl_3(g) + Cl_2(g)$

$t = 0 \quad 50M$

$t = 120min \quad 10M$

$\Rightarrow K = \frac{2.303}{t}\log\frac{[A_0]}{[A_t]}$

$\Rightarrow K = \frac{2.303}{120}\log\frac{50}{10}$

$\Rightarrow K = \frac{2.303}{120} \times 0.6989 = 0.013413 min^{-1}$

$= 1.3413 \times 10^{-2} min^{-1}$

1.34⇒ Nearest integer = 1

Question70

For the first order reaction A → 2B, 1 mole of reactant A gives 0.2 moles of B after 100 minutes. The half life of the reaction is …… min. (Round off to the nearest integer).

[Use : $\ln 2 = 0.69$, $\ln 10 = 2.3$

Properties of logarithms : $\ln x^y = y \ln x$;

$\ln\left(\frac{x}{y}\right) = \ln x - \ln y$]

(Round off to the nearest integer)

[27 Jul 2021 Shift 2]

Answer: 300

Solution:

	A	→	2B
$t = 0$	1 mole		0
$t = 100\,min$	1–x		2x
	$= 0.9\,mol$		$= 0.2\,mol$

Now, $t = \frac{t_{1/2}}{\ln 2} \times \frac{[A_0]}{[A_t]}$

$100 = \frac{t_{1/2}}{\ln 2} \times \ln\frac{1}{0.9} \Rightarrow t_{1/2} = 690 min.$

(taking $\ln 3 = 1.11$)

Ans. 600 to 700

Question71

For a reaction of order n, the unit of the rate constant is :

[27 Jul 2021 Shift 1]

Options:

A. $\mathrm{mol}^{1-n}\mathrm{L}^{1-n}\mathrm{s}$

B. $\mathrm{mol}^{1-n}\mathrm{L}^{2n}\mathrm{s}^{-1}$

C. $\mathrm{mol}^{1-n}\mathrm{L}^{n-1}\mathrm{s}^{-1}$

D. $\mathrm{mol}^{1-n}\mathrm{L}^{1-n}\mathrm{s}^{-1}$

Answer: C

Solution:

Rate $= k[A]^n$
comparing units
$$\frac{(\mathrm{mol}\ /\ \mathrm{l})}{\mathrm{sec}} = k\left(\frac{\mathrm{mol}}{\mathrm{l}}\right)^n$$
$$\Rightarrow k = \mathrm{mol}^{(1-n)}\mathrm{l}^{(n-1)}\mathrm{s}^{-1}$$

Question72

For the reaction $A + B \rightleftharpoons 2C$ the value of equilibrium constant is 100 at 298K. If the initial concentration of all the three species is 1M each, then the equilibrium concentration of C is $x \times 10^{-1}$M. The value of x is ________ . (Nearest integer)
[25 Jul 2021 Shift 1]

Answer: 25

Solution:

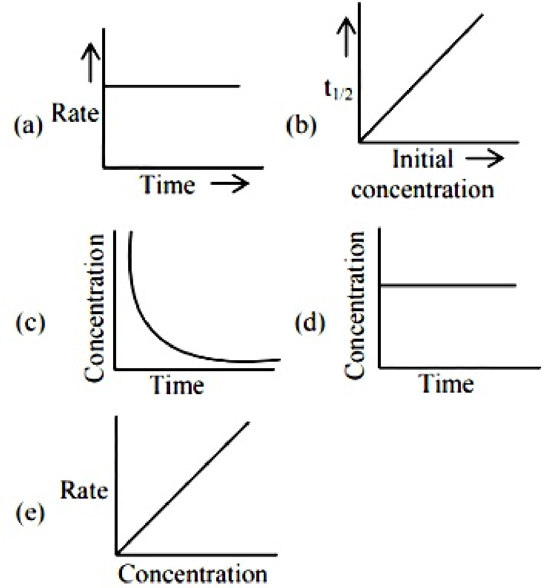
$$\begin{array}{ccccc} A & + & B & \rightleftharpoons & 2C \\ 1 & & 1 & & 1 \\ -x & & -x & & 2x \\ \hline 1-x & & 1-x & & 1+2x \end{array}$$

$$K = \frac{[C]_{eq}^{2}}{[A]_{eq}[B]_{cq}} = \frac{(1+2x)^2}{(1-x)(1-x)}$$
$$100 = \left(\frac{1+2x}{1-x}\right)^2$$
$$\left(\frac{1+2x}{1-x}\right) = 10$$
$$x = \frac{3}{4}$$
$$[\ C]e_{q.} = 1 + 2x$$

$= 1 + 2\left(\frac{3}{4}\right)$
$= 2.5\mathrm{M}$
$25 \times 10^{-1}\mathrm{M}$

Question73

For the following graphs,



Choose from the options given below, the correct one regarding order of reaction is :

[25 Jul 2021 Shift 1]

Options:

A. (b) zero order (c) and (e) First order

B. (a) and (b) Zero order (e) First order

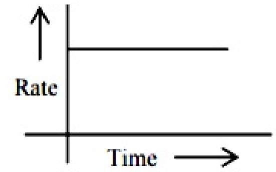
C. (b) and (d) Zero order (e) First order

D. (a) and (b) Zero order (c) and (e) First order

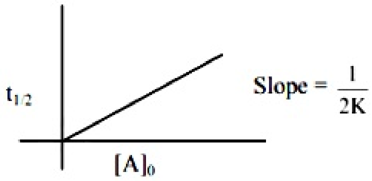
Answer: A

Solution:

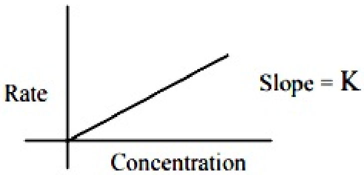
Solution:
For zero order reaction's rate $= \mathrm{K}\,[\text{Reactant }]^{0}$
$\Rightarrow \mathrm{r} = \mathrm{k}$



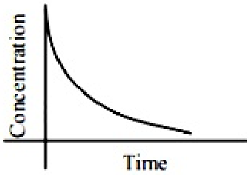
$t_{1/2} = \frac{[A]_0}{2K}$ for zero order



For first order reaction →
r = K [Concentration]



Reactant concentration after time $t \rightarrow C_t = C_0 e^{-kt}$



Reactant concentration after time $t \rightarrow C_t = C_0 e^{-kt}$

Question74

$N_2O_{5(g)} \rightarrow 2NO_{2(g)} + \frac{1}{2}O_{2(g)}$
In the above first order reaction the initial concentration of N_2O_5 is 2.40×10^{-2} mol L^{-1} at 318K. The concentration of N_2O_5 after 1 hour was 1.60×10^{-2} mol L^{-1}. The rate constant of the reaction at318K is s ______ $\times 10^{-3}$ min^{-1}. (Nearest integer)
[Given : $\log 3 = 0.477$, $\log 5 = 0.699$]
[22 Jul 2021 Shift 2]

Answer: 7

Solution:

$$K = \frac{2.303}{t} \log \frac{[N_2O_5]_0}{[N_2O_5]_t}$$
$$= \frac{2.303}{60} \log \frac{2.4}{1.6} = 6.76 \times 10^{-3} \text{min}^{-1} \approx 7 \times 10^{-3} \text{min}^{-1}$$

Question75

The reaction that occurs in a breath analyser, a device used to determine the alcohol level in a person's blood stream is
$2K_2Cr_2O_7 + 8H_2SO_4 + 3C_2H_6O \rightarrow 2Cr_2(SO_4)_3 + 3C_2H_4O_2 + 2K_2SO_4 + 11H_2O$
If the rate of appearance of $Cr_2(SO_4)_3$ is $2.67\,mol\,min^{-1}$ at a particular time, the rate of disappearance of C_2H_6O at the same time is .... $mol\,min^{-1}$. (Nearest integer)
[27 Aug 2021 Shift 1]

Answer: 4

Solution:

Solution:
Rate of appearence of 2 moles of $Cr_2(SO_4)_3$,
$\frac{d}{d t}[Cr_2(SO_4)_3] = 2.67\,mol / min$
Rate of disappearence of 3 moles of C_2H_6O, $\frac{d\,[C_2H_6O]}{d\,t}$
$= \frac{3}{2}\frac{d\,[Cr_2(SO_4)_3]}{d\,t} = \frac{2.67}{2} \times 3$
= 4.005 = 4 mol/min

Question76

The reaction rate for the reaction
$[PtCl_4]^{2-} + H_2O \rightleftharpoons [Pt(H_2O)\,Cl_3]^{-} + Cl^{-}$ was measured as a function of concentrations of different species. It was observed that
$\frac{-d\,[(PtCl_4)^{2-}]}{d\,t} = 4.8 \times 10^{-5}[(PtCl_4)^{2-}] = 2.4 \times 10^{-3}[\,Pt(H_2O)\,Cl_3\,\}^{-}\,][Cl^{-}]$
where, square brackets are used to denote molar concentrations. The equilibrium constant, K_c = ... (Nearest integer)
[26 Aug 2021 Shift 2]

Answer: 50

Solution:

Consider, at equilibrium $\frac{d\,[(PtCl_4)^{2-}]}{d\,t} = 0$
$\therefore 4.8 \times 10^{-5}[(PtCl_4)^{2-}] = 2.4 \times 10^{-3}[\{Pt(H_2O)\,Cl_3\}^{-}][Cl^{-}]$

Equilibrium constant,

$$K_C = \frac{[PtCl_4]^{2-}}{[Pt(H_2O)\,Cl_3]^{-}[Cl^{-}]}$$

$$= \frac{2.4 \times 10^{-3}}{4.8 \times 10^{-5}} = 0.50 \times 100 = 50$$

Question77

For a first order reaction, the ratio of the time for 75% completion of a reaction to the time for 50% completion is.............. (Integer answer) [31 Aug 2021 Shift 1]

Answer: 2

Solution:

Solution:

For a first order reaction, rate constant is given as,

$$k = \frac{2.303}{t}\log\frac{a}{a-x}$$

where, a = initial concentration of reactant and

a - x − = concentration of reactant after time t.

For 50% completion,

$$k = \frac{2.303}{t_{50\%}}\log\frac{100}{100-50} \text{ ...(i)}$$

For 75% completion,

$$k = \frac{2.303}{t_{75\%}}\log\frac{100}{100-75} \text{ ...(ii)}$$

From Eqs. (i) and (ii),

$$\frac{2.303}{t_{50\%}}\log\frac{100}{50} = \frac{2.303}{t_{75\%}}\log\frac{100}{25}$$

$$\frac{0.3010}{t_{50\%}} = \frac{0.6020}{t_{75\%}}$$

$$t_{75\%} = 2t_{50\%}$$

$$\therefore \frac{t_{75\%}}{t_{50\%}} = \frac{2}{1}$$

Question78

The first order rate constant for the decomposition of $CaCO_3$ at 700K is $6.36 \times 10^{-3} s^{-1}$ and activation energy is $209\,kJ\,mol^{-1}$. Its rate constant (in s^{-1}) at 600K is $x \times 10^{-6}$. The value of x is .... (Nearest integer) [Given, $R = 8.31 JK^{-1} mol^{-1}$, $\log 6.36 \times 10^{-3} = -2.19$, $10^{-4.79} = 1.62 \times 10^{-5}$] [27 Aug 2021 Shift 2]

Answer: 16

Solution:

Solution:
According to Arrhenius theory,
$k = Ae^{-E_a / RT}$
(k is the rate constant, A is Arrhenius factor, $E_a =$ Activation energy, T is absolute temperature and R is gas constant.)

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

At temperature T_1,

$$\ln k_1 = \ln A - \frac{E_a}{RT_1} \text{ ...(i)}$$

At temperature T_2,

$$\ln k_2 = \ln A - \frac{E_a}{RT_2} \text{ ...(ii)}$$

Subtracting Eqs. (ii) from (i),

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

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

$$2.303 \log \frac{k_1}{k_2} = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$\log \frac{k_1}{k_2} = \frac{E_a}{2.303R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

At, $T_1 = 700\text{K}$, $k_1 = 6.36 \times 10^{-3} \text{s}^{-1}$, $E_a = 209\text{kJmol}^{-1}$,
$R = 8.314\text{JK}^{-1}\text{mol}^{-1}$
At, $T_2 = 600\text{K}$, $k_2 = x \times 10^{-6}\text{s}^{-1}$

$$\log \frac{6.36 \times 10^{-3}}{x \times 10^{-6}} = \frac{209}{2.303 \times 8.314}\left(\frac{1}{600} - \frac{1}{700}\right)$$

$$= \frac{209}{2.303 \times 8.314}\left(\frac{100}{420000}\right)$$

x = 16.

Question79

The following data was obtained for chemical reaction given below at 975K.
$2\,NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$

	[NO] mol L^{-1}	[H_2] mol L^{-1}	Rate mol L^{-1}
A.	8×10^{-5}	8×10^{-5}	7×10^{-9}
B.	24×10^{-5}	8×10^{-5}	21×10^{-8}
C.	24×10^{-5}	32×10^{-5}	8.4×10^{-8}

The order of the reaction with respect to NO is .... [Integer answer] [26 Aug 2021 Shift 1]

Answer: 1

Solution:

Rate of the reaction $= \mathrm{K[NO]^x[H_2]^y}$

where Kis rate constant and x, y are order of reaction w.r.t. NO and H_2 respectively.

From observations A and B

$7 \times 10^{-9} = \mathrm{K}(8 \times 10^{-5})^x(8 \times 10^{-5})^y$ ...(i)

and $2.1 \times 10^{-8} = \mathrm{K}(24 \times 10^{-5})^x(8 \times 10^{-5})^y$ ...(ii)

Dividing Eqs. (i) by (ii)

$$\frac{7 \times 10^{-9}}{2.1 \times 10^{-8}} = \frac{\mathrm{K}(8 \times 10^{-5})^x(8 \times 10^{-5})^y}{\mathrm{K}(24 \times 10^{-5})^x(8 \times 10^{-5})^y}$$

$$\frac{1}{3} = \left(\frac{1}{3}\right)^x$$

x = 1

∴ Order of reaction w.r.t. NO is 1.

Question80

For the reaction A → B, the rate constant k (in S^{-1}) is given by $\log_{10}k = 20.35 - \frac{(2.47 \times 10^3)}{T}$.The energy of activation in $\mathrm{kJ\,mol^{-1}}$ is....... (Nearest integer)

[Given : $R = 8.314 \mathrm{JK^{-1}mol^{-1}}$]

[31 Aug 2021 Shift 2]

Answer: 47

Solution:

According to Arrhenius theory,

$$\log k = \log A \frac{-E_a}{2.303\,RT} \text{ ...(i)}$$

$$\log_{10}k = 20.35 - \frac{2.47 \times 10^3}{T} \text{ ...(ii)}$$

$$\frac{-2.47 \times 10^3}{T} = -\frac{E_a}{2.303\,RT}$$

Comparing Eqs. (i) and (ii),

$E_a = 2.47 \times 2.303 \times 8.314$

$= 47.29\,\mathrm{kJ / mol} \simeq 47\,\mathrm{kJ / mol}$

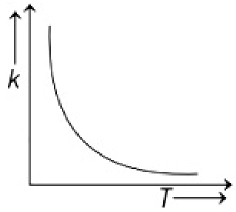
Question81

Which one of the following given graphs represents the variation of rate constant (k) with temperature (T) for an endothermic reaction?

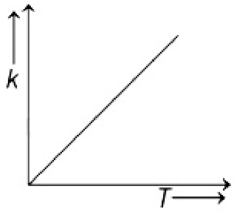
[1 Sep 2021 Shift 2]

Options:

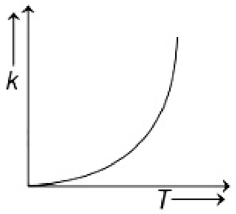
A.



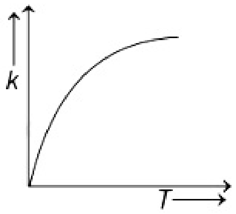
B.



C.



D.



Answer: C

Solution:

Solution:
According to Arrhenius equation,
$k = Ae^{-E_a / RT}$
As the temperature increases, rate constant k also increases whichwill generally speed up the process (increase the rate).
So, graphwill exponentially increases as in option (c).

Question82

Gaseous cyclobutene isomerizes to butadiene in a first order process which has a ' k ' value of $3.3 \times 10^{-4} s^{-1}$ at 153°C. The time in minutes it takes for the isomerization to proceed 40% to completion at this temperature is_____ (Rounded off to the nearest integer)
[2021]

Answer: 26

Solution:

$$k = \frac{1}{t} \cdot \ln \frac{[A]_0}{[A]_t}$$

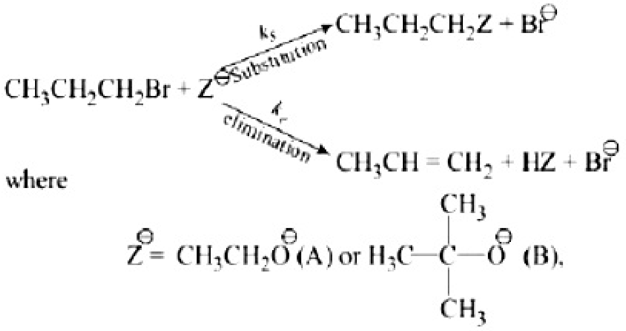
$$\Rightarrow t = \frac{1}{k} \ln \left(\frac{100}{60} \right) \Rightarrow t = \frac{1}{3.3 \times 10^{-4}}$$

$$\Rightarrow t = 1548.49 \text{sec}$$

$$\Rightarrow t = 25.81 \text{min} \approx 26 \text{min}.$$

Question83

For the following reactions



k_s and k_e, are, respectively, the rate constants for substitution and elimination, and $\mu = \frac{k_s}{k_e}$, the correct option is

[Jan. 07, 2020 (II)]

Options:

A. $\mu_B > \mu_A$ and $k_e(A) > k_e(B)$

B. $\mu_A > \mu_B$ and $k_e(B) > k_e(A)$

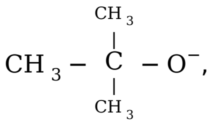
C. $\mu_B > \mu_A$ and $k_e(B) > k_e(A)$

D. $\mu_A > \mu_B$ and $k_e(A) > k_e(B)$

Answer: B

Solution:

Solution:
(b) Among the given bases (A) and (B), t -butoxide being bulky base favours elimination reaction and ethoxide favours substitution reaction.

∴ when $Z^{\circ} = CH_3CH_2O^-$, (substitution reaction favoured) and when Z =  $CH_3-\underset{CH_3}{\overset{CH_3}{C}}-O^-$, (elimination reaction favoured)
Hence, $\mu_A > \mu_B$ and $k_e(B) > k_e(A)$

Question84

For the reaction $2H_2(g) + 2NO(g) \rightarrow N_2(g) + 2H_2O(g)$ the observed rate expression is, rate $= k_e[NO]^2[H_2]$. The rate expression for the reverse reaction is:
[Jan. 07, 2020 (II)]

Options:

A. $k_b[N_2][H_2O]^2$

B. $k_b[N_2][H_2O]^2 / NO]$

C. $k_b[N_2][H_2O]$

D. $k_b[N_2][H_2O]^2 / [H_2]$

Answer: D

Solution:

Solution:
Rate of forward reaction $= k_f[NO]^2[H_2]^2$
Observed rate $= k_f[NO]^2[H_2]$
Observed rate $= \dfrac{\text{Rate of forward reaction}}{[H_2]}$
$\therefore$ Rate of backward reaction $= \dfrac{k_b[N_2][H_2O]^2}{[H_2]}$

Question85

During the nuclear explosion, one of the products is ^{90}Sr with half life of 6.93 years. If 1μg of ^{90}Sr was absorbed in the bones of a newly born baby in place of Ca, how much time, in years, is required to reduce it by 90% if it is not lost metabolically.
[NV, Jan. 07, 2020 (I)]

Answer: 23.03

Solution:

$t_{1/2} = 6.93$ years,
$a = 10^{-6}g$
$t_{1/2} = \dfrac{0.693}{K}$
$\Rightarrow K = \dfrac{0.693}{t_{1/2}} = \dfrac{0.693}{6.93} = 0.1$

For I st order reaction,

$$K = \frac{2.303}{t}\log\frac{a}{a-x}$$

$$t = \frac{2.303}{K}\log\frac{a}{a-x}$$

$$= \frac{2.303}{0.1}\log\frac{10^{-6}}{10^{-7}}$$

$$= \frac{2.303}{0.1} = 23.03 \text{ years}$$

Question86

For following reactions:

$$\mathbf{A}\xrightarrow{700K}\mathbf{Product}$$

$$\mathbf{A}\xrightarrow[catalyst]{500K}\mathbf{Product}$$

it was found that the E_a is decrease by 30kJ / mol in the presence of catalyst. If the rate remains unchanged, the activationenergy for catalysed reaction is (Assume pre exponential factor is same): [Jan. 09, 2020(I)]

Options:

A. 75kJ / mol

B. 105kJ / mol

C. 135kJ / mol

D. 198kJ / mol

Answer: A

Solution:

$$A\xrightarrow{700K\,(k_1)}\text{Product}$$

$$A\xrightarrow[\text{Catalyst}]{500K\,(k_2)}\text{Product}$$

Given: $k_1 = k_2$

$$Ae^{-\frac{E_{a_1}}{RT_1}} = Ae^{-\frac{E_{a_2}}{RT_2}}$$

$$\frac{E_{a_1}}{T_1} = \frac{E_{a_2}}{T_2}$$

$$E_{a_2} = E_{a_1} - 30$$

$$\Rightarrow E_{a1} = E_{a_2} + 30$$

$$\frac{E_{a_2}+30}{T_1} = \frac{E_{a_2}}{T_2}$$

$$\frac{E_{a_2}+30}{700} = \frac{E_{a_2}}{500}$$

$$\Rightarrow 150 = 2E_{a_2}$$

$$E_{a_2} = 75kJ / mol$$

Question87

A sample of milk splits after 60 min. at 300K and after 40 min. at 400K when the population of lactobacillus acidophilus in it doubles. The activation energy (in kJ / mol) for this process is closest to (Given, $\mathbf{R = 8.3 J\,mol^{-1} K^{-1}, \ln\left(\frac{2}{3}\right) = 0.4, e^{-3} = 4.0}$)

[NV, Jan. 09, 2020(II)]

Answer: -3.98

Solution:

For a first order reaction, $kt = \ln \frac{[A]}{[A_0]}$

At $300K$, $k_1 \times 60 = \ln \frac{[A]}{[A_0]}$ . . . . (i)

At $400K$, $k_2 \times 40 = \ln \frac{[A]}{[A_0]}$ . . . . (ii)

From equation (1) and (2),

$$\frac{k_2}{k_1} = \frac{60}{40}$$

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

$$\ln\left(\frac{60}{40}\right) = \frac{E_a}{8.3} \times \frac{100}{400 \times 300}$$

$$\ln\left(\frac{3}{2}\right) \times 8.3 \times 1200 = E_a$$

Given : $\ln \frac{2}{3} = 0.4$

Then, $\ln \frac{3}{2} = -0.4$

$\Rightarrow E_a = -0.4 \times 8.3 \times 1200$

$\Rightarrow E_a = -3984 J / mol$

$\Rightarrow E_a = -3.984 kJ / mol$.

Question88

The rate of a certain biochemical reaction at physiological temperature (T) occurs $\mathbf{10^6}$ times faster with enzyme than without. The change in the activation energy upon adding enzyme is:

[Jan .08,2020(I)]

Options:

A. $-6(2.303)RT$

B. $-6RT$

C. $+6(2.303)RT$

D. $+6RT$

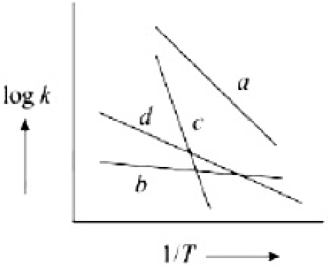
Answer: A

Solution:

Solution:
The rate constant of a reaction is given by
$k = Ae^{-E_a/RT}$
The rate constant in tpresence of catalyst is given by
$k^{'} = Ae^{-E_a^{'}/RT}$
$\frac{k^{'}}{k} = e^{-(E_a^{'} - E_a)/RT}$
$10^6 = e^{-(E_a^{'} - E_a)/RT}$
$\ln 10^6 = -\frac{(E_a^{'} - E_a)}{RT}$
$E_a^{'} - E_a = -6(2.303)RT$

Question89

Consider the following plots of rate constant versus $\frac{1}{T}$ for four different reactions. Which of the following orders is correct for the activation energies of these reactions?



[Jan. 08, 2020 (II)]

Options:

A. $E_b > E_a > E_d > E_c$

B. $E_a > E_c > E_d > E_b$

C. $E_c > E_a > E_d > E_b$

D. $E_b > E_d > E_c > E_a$

Answer: C

Solution:

Arrhenius equation, $k = Ae^{-EaRT}$
$\log k = \log A - \frac{E_a}{2.303RT}$
slope $= -\frac{E_a}{2.303R}$
∴ More negative the slope greater will be the E_a.
So correct order is $E_c > E_a > E_d > E_b$

Question90

For the reaction 2A + 3B + $\frac{3}{2}$C → 3P, which statement is correct? [Sep. 03,2020 (II)]

Options:

A. $\frac{d\,n_A}{d\,t} = \frac{3}{2}\frac{d\,n_B}{d\,t} = \frac{3}{4}\frac{d\,n_C}{d\,t}$

B. $\frac{d\,n_A}{d\,t} = \frac{d\,n_B}{d\,t} = \frac{d\,n_C}{d\,t}$

C. $\frac{d\,n_A}{d\,t} = \frac{2}{3}\frac{d\,n_B}{d\,t} = \frac{4}{3}\frac{d\,n_C}{d\,t}$

D. $\frac{d\,n_A}{d\,t} = \frac{2}{3}\frac{d\,n_B}{d\,t} = \frac{3}{4}\frac{d\,n_C}{d\,t}$

Answer: C

Solution:

Solution:

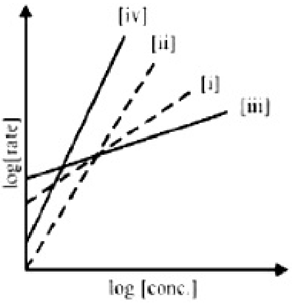
For a given reaction,

rate $= -\frac{1}{2}\frac{d\,n_A}{d\,t} = -\frac{1}{3}\frac{d\,n_B}{d\,t} = -\frac{2}{3}\frac{d\,n_C}{d\,t}$

rate $= \frac{d\,n_A}{d\,t} = \frac{2}{3}\frac{d\,n_B}{d\,t} = \frac{4}{3}\frac{d\,n_C}{d\,t}$

Question91

Consider the following reactions
A → P1; B → P2; C → P3; D → P4
The order of the above reactions are (i), (ii), (iii), and (iv), respectively. The following graph is obtained when log [rate] vs. log[conc.] are plotted:



Among the following, the correct sequence for the order of the reactions is:
[Sep. 06, 2020 (I)]

Options:

A. (iv) > (i) > (ii) > (iii)

B. (i) > (ii) > (iii) > (iv)

C. (iii) > (i) > (ii) > (iv)

D. (iv) > (ii) > (i) > (iii)

Answer: D

Solution:

Solution:
Rate $= k[A]^n$
$\log[\text{ Rate }] = \log k + n\log[A]$
Slope $= n[$ n is order of the reaction $]$
∴ Correct sequence for the order of the reaction is
$(iv) > (ii) > (i) > (iii)$

Question92

A flask contains a mixture of compounds A and B. Both compounds decompose by first-order kinetics. The halflives for A and B are 300s and 180s, respectively. If the concentrations of A and B are equal initially, the time required for the concentration of A to be four times that of B (in s) is: (Use $\ln 2 = 0.693$)

[Sep. 05 , 2020 (I)]

Options:

A. 180

B. 900

C. 300

D. 120

Answer: B

Solution:

$$C_t = C_0 e^{-kt};\ k = \frac{\ln 2}{t_{1/2}}$$

$$(C_t)_A = (C_0)_A e^{-k_A t};\ k_A = \frac{\ln 2}{300}$$

$$(C_t)_B = (C_0)_B e^{-k_B t};\ k_B = \frac{\ln 2}{180}$$

$$\frac{(C_t)_B}{(C_t)_A} = \frac{(C_0)_B}{(C_0)_A} \times e^{(k_B - k_A)t}$$

$$\Rightarrow 4 = e^{(k_B - k_A)t}$$

$$\Rightarrow 2\ln 2 = \left[\frac{\ln 2}{180} - \frac{\ln 2}{300}\right]t$$

$$\Rightarrow 2\ln 2 = \ln 2\left[\frac{1}{180} - \frac{1}{300}\right]t$$

$$\Rightarrow 2 = \left(\frac{120}{180 \times 300}\right)t \Rightarrow t = \frac{2 \times 180 \times 300}{120} = 900\text{sec}$$

Question93

If 75% of a first order reaction was completed in 90 minutes, 60% of the same reaction would be completed in approximately (in minutes) _______ .

(Take :$\log 2 = 0.30$; $\log 2.5 = 0.40$)

[NV, Sep. 04, 2020 (I)]

Answer: 60

Solution:

$$t = \frac{2.303}{k} \log\left[\frac{100}{100 - x\%}\right]$$

$$t_{75\%} = \frac{2.303}{k} \log\left[\frac{100}{25}\right] = 90$$

$$t_{60\%} = \frac{2.303}{k} \log\left[\frac{100}{40}\right]$$

$$\frac{t_{75\%}}{t_{60\%}} = \frac{2 \log 2}{\log 2.5} \Rightarrow \frac{90}{t_{60\%}} = \frac{2 \times 0.3}{0.4}$$

$$\Rightarrow t_{60\%} = \frac{90 \times 4}{6} = 60 \text{min}.$$

Question94

It is true that :

[Sep. 03, 2020 (I)]

Options:

A. A second order reaction is always a multistep reaction

B. A zero order reaction is a multistep reaction

C. A first order reaction is always a single step reaction

D. A zero order reaction is a single step reaction

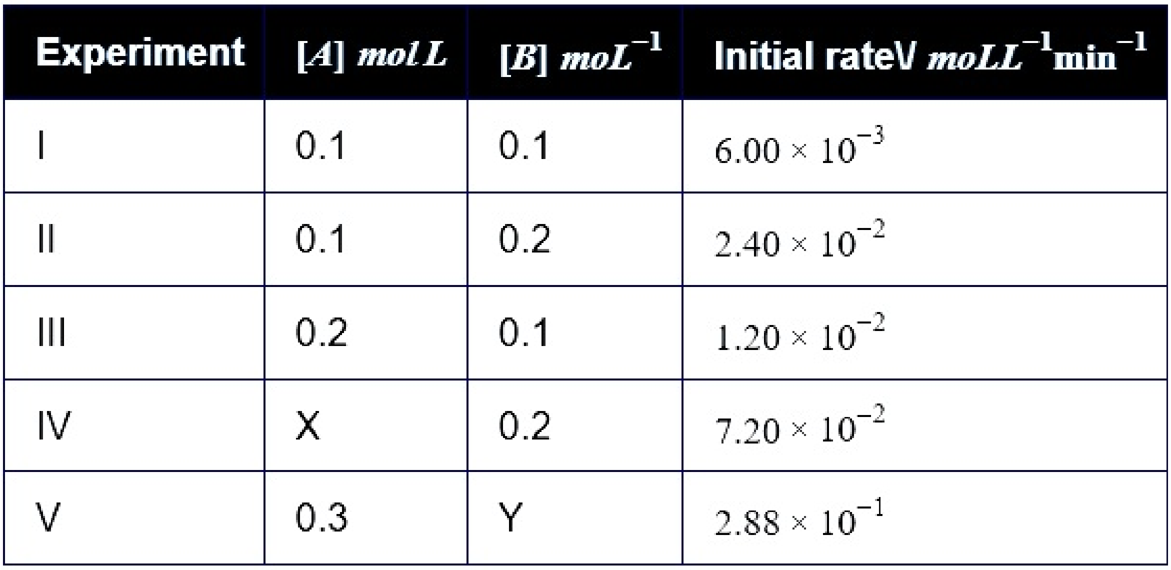
Answer: B

Solution:

Solution:
zero order reaction is a multistep reaction .

Question95

The results given in the below table were obtained during kinetic studies of the following reaction:
$2A + B \rightarrow C + D$



Experiment	$[A]\ mol L$	$[B]\ mol L^{-1}$	Initial rateV $mol L L^{-1} min^{-1}$
I	0.1	0.1	6.00×10^{-3}
II	0.1	0.2	2.40×10^{-2}
III	0.2	0.1	1.20×10^{-2}
IV	X	0.2	7.20×10^{-2}
V	0.3	Y	2.88×10^{-1}

X and Y in the given table are respectively:
[Sep. 02 , 2020(II)]

Options:

A. 0.4, 0.4

B. 0.4, 0.3

C. 0.3, 0.4

D. 0.3, 0.3

Answer: C

Solution:

Rate (R) = $k[A]^a[B]^b$
Exp I $\Rightarrow 6.0 \times 10^{-3} = k[0.1]^a[0.1]^b$
Exp II $\Rightarrow 24.0 \times 10^{-3} = k[0.1]^a[0.2]^b$
Exp III $\Rightarrow 12.0 \times 10^{-3} = k[0.2]^a[0.1]^b$
Exp IV $\Rightarrow 72 \times 10^{-3} = k[X]^a[0.2]^b$
Exp V $\Rightarrow 288 \times 10^{-3} = k[0.3]^a[Y]^b$
From Exp I & II,

$$\Rightarrow \frac{6.0 \times 10^{-3}}{24 \times 10^{-3}} = \frac{k[0.1]^a[0.1]^b}{k[0.1]^a[0.2]^b}$$

$$\Rightarrow \frac{1}{4} = \left(\frac{0.1}{0.2}\right)^b \Rightarrow \left(\frac{1}{2}\right)^2 = \left(\frac{1}{2}\right)^b$$

$\therefore b = 2$
Similarly, from exp I & III we get

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

From Exp. II & IV,

$$\Rightarrow \frac{24 \times 10^{-3}}{72 \times 10^{-3}} = \frac{k[0.1]^a[0.2]^b}{k[X]^a[0.2]^b}$$

$$\Rightarrow \frac{1}{3} = \left(\frac{0.1}{X}\right)^a \Rightarrow \frac{1}{3} = \left(\frac{0.1}{X}\right)^1$$

$\therefore X = 0.3$

From Exp.I & V,

$$\Rightarrow \frac{6.0 \times 10^{-3}}{288 \times 10^{-3}} = \left(\frac{0.1}{0.3}\right)^{a}\left(\frac{0.1}{Y}\right)^{b}$$

$$\Rightarrow \frac{1}{48} = \left(\frac{1}{3}\right)^{1}\left(\frac{0.1}{Y}\right)^{2}$$

$$\Rightarrow \frac{3}{48} = \left(\frac{0.1}{Y}\right)^{2} \Rightarrow \frac{1}{16} = \left(\frac{0.1}{Y}\right)^{2}$$

$$\Rightarrow \left(\frac{1}{4}\right)^{2} = \left(\frac{0.1}{Y}\right)^{2} \Rightarrow \frac{1}{4} = \frac{0.1}{Y}$$

$\therefore Y = 0.4$

Question96

The rate of a reaction decreased by 3.555 times when the temperature was changed from 40˚C to 30˚C. The activation energy (in $kJ\,mol^{-1}$) of the reaction is _______.. Take;

$R = 8.314 J\,mol^{-1}K^{-1} I\,n 3.555 = 1.268$

[NV,Sep. 06 , 2020 (II)]

Answer: 100

Solution:

The Arr henices equation is

$$k = Ae^{\frac{E_a}{RT}}$$

Assuming A and E_a to be independent of temperature

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

$$\ln 3.555 = \frac{E_a}{8.314}\left(\frac{1}{303} - \frac{1}{313}\right)$$

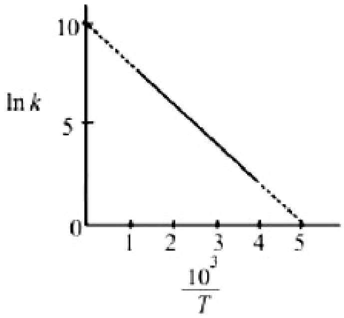
$$\Rightarrow E_a = \frac{1.268 \times 8.314 \times 303 \times 313}{10}$$

$$= 99980.7 = 99.98 kJ / mol$$

Question97

The rate constant (k) of a reaction is measured at different temperatures (T), and the data are plotted in the given figure. The activation energy of the reaction in $kJ\,mol^{-1}$ is:

(R is gas constant)



[Sep. 05, 2020 (II)]

Options:

A. 2 / R

B. 1 / R

C. R

D. 2R

Answer: D

Solution:

Arrhenius equation:
$k = Ae^{-E_a/RT}$
$\ln k = \ln A - \left(\frac{E_a}{R}\right)\frac{1}{T}$
$\ln k = \ln A - \left(\frac{E_a}{R \times 10^3}\right) \times \frac{10^3}{T}$
Slope of graph $= \frac{-E_a}{R \times 10^3} = \frac{-10}{5}$
$E_a = 2R \times 10^3 \text{J} = 2R\text{kJ}$

Question98

The number of molecules with energy greater than the threshold energy for a reaction increases five fold by a rise of temperature from 27˚C to 42˚C. Its energy of activation in J / mol is _______ .
(Take $\ln 5 = 1.6094$; $R = 8.314 \text{J mol}^{-1}\text{K}^{-1}$)
[NV, Sep. 04, 2020 (II)]

Answer: 84297.48

Solution:

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

$\ln \frac{k_2}{k_1} = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) \Rightarrow \ln 5 = \frac{E_a}{R}\left(\frac{1}{300} - \frac{1}{315}\right)$

$\Rightarrow E_a = \frac{1.6094 \times 8.314 \times 300 \times 315}{15} = 84297.48 J / mol$.

Question99

the expression for $\frac{d[A]}{dt}$ is:
[Jan. 10,2019 (II)]

Options:

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

B. $2k_1[A_2] - k_{-1}[A]^2$

C. $k_1[A_2] + k_{-1}[A]^2$

D. $2k_1[A_2] - 2k_{-1}[A]^2$

Answer: D

Solution:

Solution:

Given: $A_2 \frac{k_1}{\text{underset}\, k_{11}} 2A$

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

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

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

Question100

Decomposition of X exhibits a rate constant of 0.05μg/ year. How many years are required for the decomposition of 5μg of X into 2.5μg ?
[Jan. 12, 2019 (I)]

Options:

A. 50

B. 25

C. 20

D. 40

Answer: A

Solution:

Rate constant of decomposition of $X = 0.05\mu g/$ 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 years

Question101

The reaction $2X \rightarrow B$ is a zeroth order reaction. If the initial concentration of X is $0.2M$, the half-life is 6h. When the initial concentration of X is $0.5M$, the time required to reach its final concentration of $0.2M$ will be :
[Jan. 11, 2019 (II)]

Options:

A. 9.0h

B. 12.0h

C. 18.0h

D. 7.2h

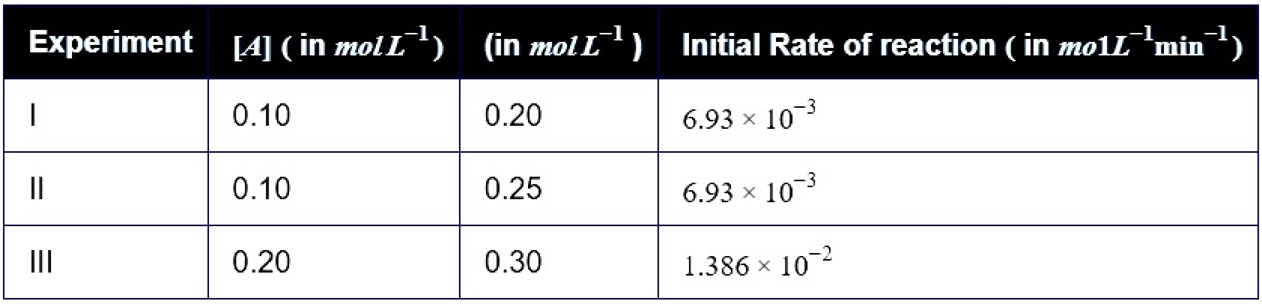
Answer: C

Solution:

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 $[A] = 0.1$
$Kt_{1/2} = 0.2 - 0.1$
$\frac{0.2 - 0.1}{6} = \frac{0.1}{6} M\,hr^{-1}$
∴ Time required to reachfrom $0.5M$ to $0.2M$
$Kt = [A]_0 - [A]$
$\frac{0.1}{6} \times t = (0.5 - 0.2)$; t = 18 hour

Question102

The following results were obtained during kinetic studies of the reaction;
$2A + B \rightarrow$ Products



Experiment	$[A]$ (in $mol L^{-1}$)	(in $mol L^{-1}$)	Initial Rate of reaction (in $mo1L^{-1}min^{-1}$)
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:
[Jan. 9, 2019 (I)]

Options:

A. 5

B. 10

C. 1

D. 100

Answer: A

Solution:

From experiment I and II, it is observed that order of reaction w.r.t. (B) is zero.
From experiment II and III, x can be calculated as:

$$\frac{6.93 \times 10^{-3}}{13.86 \times 10^{-3}} = \left(\frac{0.1}{0.2}\right)^{x}\left(\frac{0.25}{0.3}\right)^{y}$$

$\because$ Order of reaction w.r.t.B = 0,
so $y = 0$

$$\therefore \frac{6.93 \times 10^{-3}}{13.86 \times 10^{-3}} = \left(\frac{0.1}{0.2}\right)^{x}$$

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

$x = 1$

Now, Rate $= K[A]^{1}[B]^{\circ}$

or, $6.93 \times 10^{-3} = K[0.1]^{1}[0.2]^{\circ}$

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

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

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

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

Question103

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?
[Jan. 9, 2019 (II)]

Options:

A. Total order of the reaction is 4

B. Order of the reaction with respect to B is 2

C. Order of the reaction with respect to B is 1

D. Order of the reaction with respect to A is 2

Answer: B

Solution:

$r = K[A]^{\mu}[B]^{y}$

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

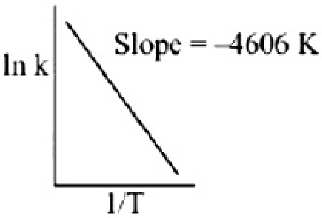
$\frac{r_3}{r_1} = 2x = 2 \Rightarrow x = 1$

$\therefore \; y = 2$

$t_{1/2} = 5$

Question104

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



[Jan. 12, 2019 (II)]

Options:

A. $10^{-6}s^{-1}$

B. $2 \times 10^{-4}s^{-1}$

C. $10^{-4}s^{-1}$

D. $4 \times 10^{-4}s^{-2}$

Answer: C

Solution:

Solution:
From Arrhenius equation,

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

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

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

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

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

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

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

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

Question105

If a reaction follows the Arrhenius equation, the plot $\ln k$ vs $\frac{1}{(RT)}$ gives straight line with a gradient $(-y)$ unit.
The energy required to activate the reactant is:
[Jan. 11,2019(I)]

Options:

A. y / R unit

B. y unit

C. yR unit

D. −y unit

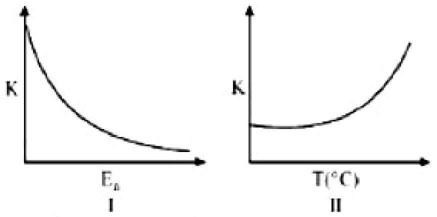
Answer: B

Solution:

From Arrhenius equation,
$k = Ae^{-E_a/RT}$
$\ln k = \ln A - E_a / RT$
slope $= -y$ (given)
$-y = -E_a$
$\Rightarrow E_a = y$

Question106

Consider the given plots for a reaction obeying Arrhenius equation $(0^\circ C < T < 300^\circ C)$: (k and E _ are rate constant and activation energy, respectively)



Choose the correct option:
[Jan. 10,2019 (I)]

Options:

A. I is right but II is wrong

B. Both I and II are correct

C. I is wrong but II is right

D. Both I and II are wrong

Answer: B

Solution:

Solution:
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.

Question107

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:
[April 12, 2019 (I)]

Options:

A. n -Butane and Iso-butane

B. C_2H_2 and C_6H_6

C. C_2H_4 and C_4H_8

D. N_2O_4 and NO_2

Answer: C

Solution:

$xA \rightarrow yB$
$\therefore \ \frac{-dA}{xdt} = \frac{1}{y}\frac{dB}{dt}$
$\frac{-dA}{dt} = \frac{dB}{dt} \times \frac{x}{y}$
$\log\left[\frac{-dA}{dt}\right] = \log\left[\frac{dB}{dt}\right] + \log\left(\frac{x}{y}\right)$
Comparing this equation withthe equation given in question. We get,

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

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

∴ The reaction is of type 2A → B.
Hence, option (3) is correct.

Question108

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:
[April 12,2019 (II)]

Options:

A. $4.167 \times 10^{-3} mol\,L^{-1} min^{-1}$

B. $1.667 \times 10^{-2} mol\,L^{-1} min^{-1}$

C. $8.333 \times 10^{-3} mol\,L^{-1} min^{-1}$

D. $2.083 \times 10^{-3} mol\,L^{-1} min^{-1}$

Answer: B

Solution:

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

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

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

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

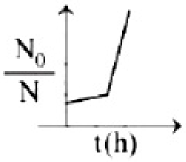
$= 1.67 \times 10^{-2-} M\,min^{-1}$

Question109

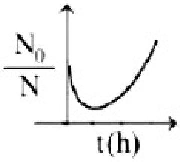
A bacterial infection in an internal wound grows as $N^{'}(t) = N_0 \exp(t)$, where the time tis in hours. A dose of antibiotic, taken orally, needs 1 hour to reach the wound. Once it reaches there, the bacterial population goes down as $\frac{dN}{dt} = -5N^2$. What will be the plot of $\frac{N_0}{N}$ vs .t after 1 hour?
[April 10,2019 (I)]

Options:

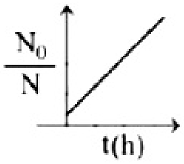
A.



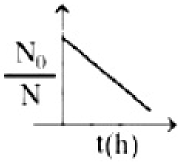
B.



C.



D.



Answer: C

Solution:

Solution:

From 0 to 1h, $N' = N_0 e^t$

When drug is administered bacterial growth is given by

$$\frac{dN}{dt} = -5N^2$$

So, at t = 1h, $N' = eN_0$

Integrating the following equation,

$$\frac{dN}{dt} = -5N^2$$

$$\int_{eN_0}^{N} \frac{dN}{N^2} = -5\int_1^t dt$$

$$\frac{1}{N} - \frac{1}{eN_0} = 5(t-1)$$

$$\frac{N_0}{N} - \frac{1}{e} = 5N_0(t-1)$$

$$\frac{N_0}{N} = 5N_0(t-1) + \frac{1}{e}$$

$$\frac{N_0}{N} = 5N_0 t + \left(\frac{1}{e} - 5N_0\right)$$

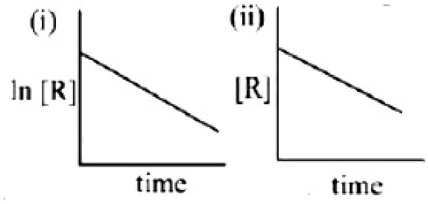
The above equation is similar to straight line equation with positive slope.

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

Question110

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:



[April 9, 2019 (I)]

Options:

A. 1,0

B. 1.1

C. 0,1

D. 0,2

Answer: A

Solution:

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

Question111

For the reaction 2A + B → 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

[April 8, 2019 (I)]

Options:

A. Rate $= k[A][B]^2$

B. Rate $= k[A]^2[B]^2$

C. Rate $= k[A][B]$

D. Rate $= k[A]^2[B]$

Answer: A

Solution:

$2A + B \longrightarrow C$
Rate $= k[A]^x[B]^y$
Exp-1, $0.045 = k[0.05]^x[0.05]$ . . . (i)
$E\,xp - 2,\ 0.090 = k[0.1]^x[0.05]^y$ . . . (ii)
$E\,xp - 3,\ 0.72 = k[0.2]^x[0.1]^y$ . . . (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 $= k[A]^1[B]^2$.

Question112

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:
[April 8, 2019 (II)]

Options:

A. $(k_1 - k_2)[A]$

B. $k_1k_2[A]$

C. $(k_1 + k_2)[A]$

D. $\left(\frac{k_1}{k_2}\right)[A]$

Answer: D

Solution:

$$A \xrightarrow{k_1} B \xrightarrow{k_2} C$$

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

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

Question113

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

[April 10, 2019 (II)]

Options:

A. 166

B. 150

C. 72

D. 59

Answer: A

Solution:

$$K = e^{-\frac{E_a}{RT}} \text{ or } \log K = \frac{-E_a}{2.303RT}$$

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

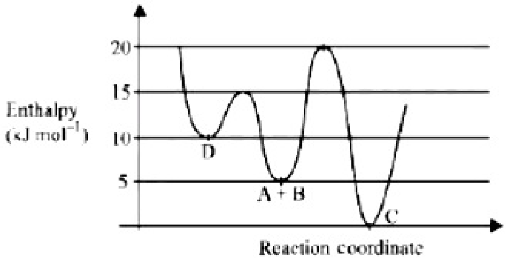
$$\log \frac{1}{2.5 \times 10^{-4}} = \frac{E_a}{8.314 \times 2.303}\left(\frac{1}{600} - \frac{1}{800}\right)$$

$$3.6 = \frac{E_a}{8.314 \times 2.303} \times \frac{200}{600 \times 800}$$

$E_a = 165.4 kJ / mol$ H "166kJ / mol

Question114

Consider the given plot of enthalpy of the following reaction between A and B, A + B → C + D.
Identify the incorrect statement.



[April 9, 2019 (II)]

Options:

A. Activation enthalpy to form C is 5kJ mol $^{-1}$ less than that to form D.

B. C is the thermodynamically stable product.

C. D is kinetically stable product.

D. Formation of A and B from C has highest enthalpy of activation.

Answer: A

Solution:

As we can see from the graph that activation enthalpy to form D from $A+B$ is $15-5=10\text{kJ mol}^{-1}$, whereas, to form C from $A+B$ is $20-5=15\text{kJ mol}^{-1}$. Therefore, activation enthalpy to form C is 5kJ morethan that to form D.

Question115

At 518°C, the rate of decomposition of a sample of gaseous acetaldehyde, initially at a pressure of 363 Torr, was 1.00 Torr s^{-1} when 5% had reacted and 0.5 Torr s^{-1} when 33% had reacted. The order of the reaction is:
[2018]

Options:

A. 2

B. 3

C. 1

D. 0

Answer: A

Solution:

Solution:
$CH_3CHO \longrightarrow CH_4 + CO$
$r \propto (a-x)^m$
Generally
$m =$ order of reaction
$a-x =$ unreacted
$r_1 = 1$ torr s^{-1}, when 5% reacted
$r_2 = 0.5$ torr s^{-1}, when 33% reacted
$(a-x_1) = 0.95$(unreacted)
$(a-x_2) = 0.67$ (unreacted)
$\frac{r_1}{r_2} = \left[\frac{(a-x_1)}{(a-x_2)}\right]^m;\ \frac{1}{0.5} = \left(\frac{0.95}{0.67}\right)^m$
$2 = (1.41)^m \Rightarrow 2 = (\sqrt{2})^m$
$\Rightarrow m = 2$

Question116

If 50% of a reaction occurs in 100 seconds and 75% of the reaction occurs in 200 seconds, the order of this reaction is:
[Online April 16, 2018]

Options:

A. 2

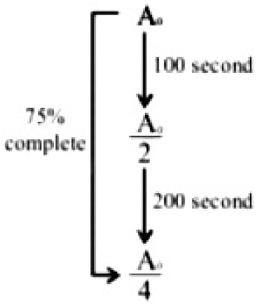
B. 3

C. Zero

D. 1

Answer: D

Solution:



First order reaction as half life is constant.

Question117

N_2O_5 decomposes to NO_2 and O_2 and follows first order kinetics. After 50 minutes, the pressure inside the vessel increases from 50mmHg to 87.5mmHg. The pressure of the gaseous mixture after 100 minutes at constant temperature will be ______.

[Online April 15, 2018 (I)]

Options:

A. 136.25mmHg

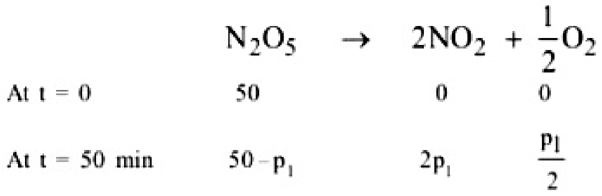
B. 106.25mmHg

C. 175.0mmHg

D. 116.25mmHg

Answer: B

Solution:



	N_2O_5 →	$2NO_2$ +	$\frac{1}{2}O_2$
At t = 0	50	0	0
At t = 50 min	$50-p_1$	$2p_1$	$\frac{p_1}{2}$

Total pressure at 50 minutes

$$= 50 - p_1 + 2p_1 + \frac{p_1}{2} = 87.5$$

$$50 + \frac{3p_1}{2} = 87.5$$

$$\frac{3p_1}{2} = 37.5$$

$$\therefore\ p_1 = \frac{37.5 \times 2}{3} = 25$$

At, t = 100min

$$50 - p_2 \quad 2p_2 \quad \frac{p_2}{2}$$

50 minutes is half life period

For 100 minutes i.e. for 2 half lives $50 - p_2 = 12.5$

$\therefore p_2 = 37.5 \mathrm{mm\ of\ Hg}$

Total pressure at 100 minutes

$= 50 - p_2 + 2p_2 + \frac{p_2}{2}$

$= 50 + \frac{3p_2}{2} = 50 + \frac{3}{2} \times 37.5$

$= 50 + 56.25$

$= 106.25 \mathrm{mm of\ Hg}$

Question118

For a first order reaction, $A \rightarrow P$, $t_{1/2}$ (half-life) is 10 days. The time required for $\frac{1}{4}$ conversion of A(in days) is: ($\ln 2 = 0.693$, $\ln 3 = 1.1$).

[Online April 15,2018(II)]

Options:

A. 3.2

B. 2.5

C. 4.1

D. 5

Answer: C

Solution:

The half life $t_{1/2} = 10$ days

The decay constant

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{10 \mathrm{days}} = 0.0693 \mathrm{days}^{-1}$$

The time required for one fourth conversion

$$t = \frac{2.303}{k} \log_{10} \frac{a}{a - x}$$

$$= \frac{2.303}{0.0693 \ \mathrm{day}^{-1}} \log_{10} \frac{1}{1 - (1/4)} = 4.1 \text{ days}$$

Question119

Two reactions R_1 and R_2 have identical pre-exponential factors. Activation energy of R_1 exceeds that of R_2 by $10 \mathrm{kJ\ mol}^{-1}$. If k_1 and k_2 are rate constants for reactions R_1 and R_2 respectively at $300 \mathrm{K}$, then $\ln(k_2/k_1)$ is equal to: ($R = 8.314 \mathrm{J\ mol}^{-1} \mathrm{K}^{-1}$)

[2017]

Options:

A. 8

B. 12

C. 6

D. 4

Answer: D

Solution:

From Arrhenius equation,

$k = A \cdot e^{\frac{-E a}{RT}}$

so, $k_1 = A \cdot e^{-E_{a_1}/RT}$ . . . . (i)

$k_2 = A \cdot e^{-E_{a_2}/RT}$ . . . . (ii)

On dividing equation (ii) by (i)

$\Rightarrow \frac{k_2}{k_1} = e^{\frac{\left(E_{a_1} - E_{a_2}\right)}{RT}}$

$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_{a_1} - E_{a_2}}{RT} = \frac{10,000}{8.314 \times 300} = 4$

Question120

The rate of a reaction A doubles on increasing the temperature from 300 to 310K. By how much, the temperature of reaction B should be increased from 300K so that rate doubles if activation energy of the reaction B is twice to that of reaction A.

[Online April 8, 2017]

Options:

A. 9.84K

B. 4.92K

C. 2.45K

D. 19.67K

Answer: B

Solution:

$\log \frac{k_2}{k_1} = \frac{E_a}{2.303R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$

For reaction A -

Given, $\frac{k_2}{k_1} = 2, T_1 = 300K, T_2 = 310K$

$\log 2 = \frac{E_a}{2.303R}\left[\frac{1}{300} - \frac{1}{310}\right]$ . . . (i)

For reaction B -

Given, $\frac{k_2}{k_1} = 2, E_a = 2E_a, T_1 = 300K, T_2 = ?$

$\log 2 = \frac{E_a}{2.303R}\left[\frac{1}{300} - \frac{1}{T_2}\right]$ . . . (ii)

From equation (i) and (ii),

$\frac{2E_a}{2.303R}\left[\frac{1}{300} - \frac{1}{T_2}\right] = \frac{E_a}{2.303R}\left[\frac{1}{300} - \frac{1}{310}\right]$

$\Rightarrow 2\left[\frac{1}{300} - \frac{1}{T_2}\right] = \frac{310 - 300}{300 \times 310}$

$\Rightarrow T_2 = 304.92K$

$T_1 = 300K, T_2 = 304.92K$

$\Delta T = T_2 - T_1 = 4.92K$

Question121

The rate of a reaction quadruples when the temperature changes from 300 to 310K. The activation energy of this reaction is : (Assume activation energy and pre-exponential factor are independent of temperature; $\ln 2 = 0.693$; $R = 8.314 J\ mol^{-1} K^{-1}$)
[Online April 9, 2017]

Options:

A. $107.2 kJ\ mol^{-1}$

B. $53.6 kJ\ mol^{-1}$

C. $26.8 kJ\ mol^{-1}$

D. $214.4 kJ\ mol^{-1}$

Answer: A

Solution:

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

$$\ln 4 = \frac{E_a}{8.314}\left(\frac{310 - 300}{310 \times 300}\right)$$

$$2\ln 2 = \frac{E_a}{8.314}\left(\frac{310 - 300}{310 \times 300}\right)$$

$$E_a = \frac{0.693 \times 2 \times 8.314 \times 300 \times 310}{10} = 107.2 kJ\ / mol$$

Question122

The reaction of ozone with oxygen atoms in the presence of chlorine atoms can occur by a two step process shown below:
$O_3(g) + Cl^*(g) \rightarrow O_2(g) + ClO^*(g)$ . . . (i)

$[K_i = 5.2 \times 10^9 \mathrm{Lmol}^{-1} \mathrm{s}^{-1}]$
$\mathrm{ClO}^*(g) + O^*(g) \rightarrow O_2(g) + \mathrm{Cl}^*(g)$ . . . (ii)
$[K_{ai} = 2.6 \times 10^{10} \mathrm{Lmol}^{-1} \mathrm{s}^{-1}]$
The closest rate constant for the overall reaction
$O_3(g) + O^*(g) \rightarrow 2O_2(g)$ is:
[Online April 9, 2016]

Options:

A. $1.4 \times 10^{20} \mathrm{Lmol}^{-1} \mathrm{s}^{-1}$

B. $3.1 \times 10^{10} \mathrm{Lmol}^{-1} \mathrm{S}^{-1}$

C. $5.2 \times 10^{9} \mathrm{Lmol}^{-1} \mathrm{s}^{-1}$

D. $2.6 \times 10^{10} \mathrm{Lmol}^{-1} \mathrm{s}^{-1}$

Answer: A

Solution:

On adding eq. (i) and eq. (ii) we get $O_3(g) + O(g) \longrightarrow 2O_2(g)$
Hence overall rate constant $= k_i \times k_{ii}$
$= 5.2 \times 10^9 \times 2.6 \times 10^{10} \approx 1.4 \times 10^{20} \mathrm{mol}^{-1} \mathrm{Ls}^{-1}$
$O_3(g) + \mathrm{Cl}^*(g) \longrightarrow O_2(g) + \mathrm{ClO}^*(g), k_i$
$\mathrm{ClO}^*(g) + O^*(g) \longrightarrow O_2(g) + \mathrm{Cl}^*(g), k_{ii}$ $O_3(g) + O^*(g) \longrightarrow 2O_2(g), k_{\mathrm{Rate}} = k_i \times k_i$

Question123

The rate law for the reaction below is given by the expression k[A][B]
A + B → Product
If the concentration of B is increased from 0.1 to 0.3mol e, keeping the value of A at 0.1mol e, the rate constant will be:
[Online April 10, 2016]

Options:

A. 3k

B. 9k

C. k / 3

D. k

Answer: D

Solution:

Rate constant is independent of concentration.

Question124

Decomposition of H_2O_2 follows a first order reaction. ln fifty minutes the concentration of H_2O_2 decreases from 0.5 to 0.125M in one such decomposition. When the concentration of H_2O_2 reaches 0.05M, the rate of formation of O_2 will be:
[2016]

Options:

A. $2.66 \mathrm{L\,min}^{-1}$ at STP

B. $1.34 \times 10^{-2} \mathrm{mol\,min}^{-1}$

C. $6.96 \times 10^{-2} \mathrm{mol\,min}^{-1}$

D. $6.93 \times 10^{-4} \mathrm{mol\,min}^{-1}$

Answer: D

Solution:

$H_2O_2(aq) \rightarrow H_2O(aq) + \frac{1}{2}O_2(g)$
For a first order reaction
$k = \frac{2.303}{t}\log\frac{a}{(a-x)}$
Given $a = 0.5$, $(a-x) = 0.125$, $t = 50\mathrm{min}$
$\therefore k = \frac{2.303}{50}\log\frac{0.5}{0.125}$
$r = k[H_2O_2] = 2.78 \times 10^{-2} \mathrm{min}^{-1}$
Now $-\frac{d[H_2O_2]}{dt} = \frac{d[H_2O]}{dt} = \frac{2d[O_2]}{dt}$
$\therefore \frac{2d[O_2]}{dt} = -\frac{d[H_2O_2]}{dt}$
$\therefore \frac{d[O_2]}{dt} = -\frac{1}{2} \times \frac{d[H_2O_2]}{dt}$
$= \frac{1.386 \times 10^{-3}}{2} = 6.93 \times 10^{-4} \mathrm{mol\,min}^{-1}$

Question125

$A + 2B \rightarrow C$, the rate equation for this reaction is given as Rate $= k[A][B]$.
If the concentration of A is kept the same but that of B is doubled what will happen to the rate itself?
[Online April 11,2015]

Options:

A. halved

B. the same

C. doubled

D. quadrupled

Answer: C

Solution:

Rate $= k[A][B] = R$
$R' = k[A][2B]$
$\frac{R}{R'} = \frac{k[A][B]}{k[A][2B]} = \frac{k[A][B]}{2k[A][B]}$
$\Rightarrow 2R = R'$ i.e., rate become doubles.

Question126

Higher order (>3) reactions are rare due to :
[2015]

Options:

A. shifting of equilibrium towards reactants due to elastic collisions

B. loss of active species on collision

C. low probability of simultaneous collision of all the reacting species

D. increase in entropy and activation energy as more molecules are involved

Answer: C

Solution:

Solution:
Reactions of higher order (>3) are very rare due to very less chances of many molecules to undergo effective collisions.

Question127

The reaction $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$ follows first order kinetics. The pressure of a vessel containing only N_2O_5 was found to increase from 50mm Hg to 87.5mmHg in 30min. The pressureexerted by the gases after 60min. will be (assume temperature remains constant):
[Online April 10,2015]

Options:

A. 106.25mmHg

B. 150mmHg

C. 125mmHg

D. 116.25mmHg

Answer: A

Solution:

$$\begin{array}{llll} & 2N_2O_5(g) \rightarrow & 4NO_2(g) + & O_2(g) \\ t = 0 & 50 & 0 & 0 \\ t = 30 & 50 - 2x & 4x & x \end{array}$$

$\Rightarrow 87.5 = 50 + 3x$

$\Rightarrow 3x = 37.5 \quad \Rightarrow x = 12.5$

$\Rightarrow P_{N_2O_5}$ after 30 mins $= 50 - 25 = 25$

$\Rightarrow t_{\frac{1}{2}} = 30\text{min}$

Hence after 60 min, (two half lives),$P_{N_2O_5}$ remaining $= \frac{50}{4} = 12.5\text{torr}$

$\Rightarrow$ Hence decrease in $P_{N_2O_5} = 50 - 12.5 = 37.5$ torr

$\Rightarrow P_{NO_2} = 2 \times 37.5 = 75$ torr

$P_{O_2} = \frac{37.5}{2} = 18.75$ torr

$\Rightarrow P_{total} = 12.5 + 75 + 18.75$

$= 106.25$ torr

Question128

For the equilibrium, A(g) ⇌ B(g), ΔH is $-40\text{kJ}/\text{mol}$. If the ratio of the activation energies of the forward (E_f) and reverse (E_b) reactions is $\frac{2}{3}$ then :

[Online April 11, 2015]

Options:

A. $E_f = 80\text{kJ}/\text{mol}; E_b = 120\text{kJ}/\text{mol}$

B. $E_f = 60\text{kJ}/\text{mol}; E_b = 100\text{kJ}/\text{mol}$

C. $E_f = 30\text{kJ}/\text{mol}; E_b = 70\text{kJ}/\text{mol}$

D. $E_f = 70\text{kJ}/\text{mol}; E_b = 30\text{kJ}/\text{mol}$

Answer: A

Solution:

Solution:

$A(g) \rightleftharpoons B(g)\ \Delta H = -40\text{kJ}$

If activation energy of the reaction $= x$

since, $\frac{E_f}{E_b} = \frac{2}{3}$, therefore, $E_f = \frac{2x}{5}$ and $E_b = \frac{3x}{5}$

$E_b - E_f = +40$

$= \frac{3x}{5} - \frac{2x}{5} = 40; \ \frac{x}{5} = 40; x = 200$

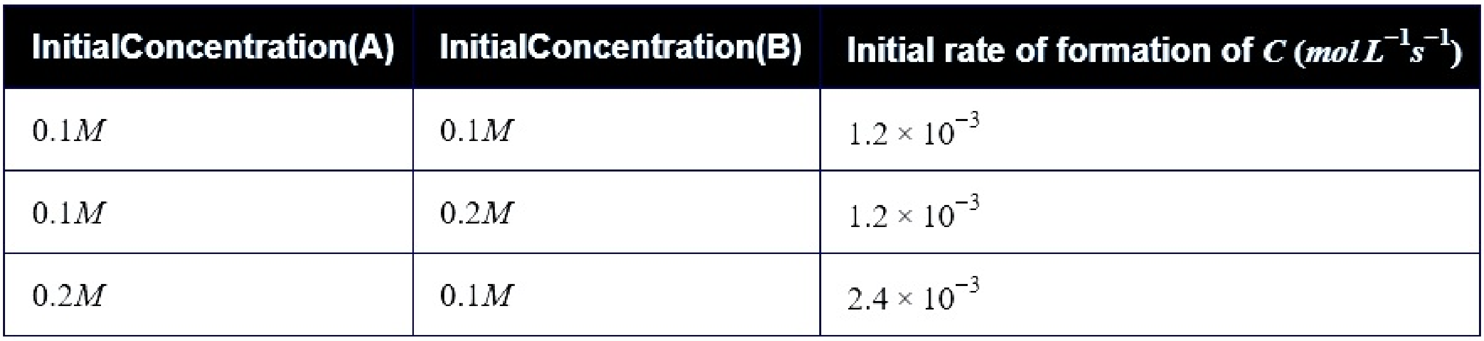
Therefore

$E_b = \frac{3x}{5} = \frac{3 \times 200}{5} = 120 \text{kJ mol}^{-1}$

$E_f = \frac{2x}{5} = \frac{2 \times 200}{5} = 80 \text{kJ mol}^{-1}$

Question129

For the non − stoichiometric reaction 2A + B → C + D, the following kinetic data were obtained in three separate experiments, all at 298K.



InitialConcentration(A)	InitialConcentration(B)	Initial rate of formation of $C\ (mol L^{-1}s^{-1})$
$0.1M$	$0.1M$	1.2×10^{-3}
$0.1M$	$0.2M$	1.2×10^{-3}
$0.2M$	$0.1M$	2.4×10^{-3}

The rate law for the formation of C is:
[2014]

Options:

A. $\frac{dc}{dt} = k[A][B]$

B. $\frac{dc}{dt} = k[A]^2[B]$

C. $\frac{dc}{dt} = k[A][B]^2$

D. $\frac{dc}{dt} = k[A]$

Answer: D

Solution:

Solution:

Let rate of reaction $= \frac{d[C]}{t} = k[A]^x[B]^y$

Now from the given data

$1.2 \times 10^{-3} = k[0.1]^x[0.1]^y \ldots$ (i)

$1.2 \times 10^{-3} = k[0.1]^x[0.2]^y \ldots$ (ii)

$2.4 \times 10^{-3} = k[0.2]^x[0.1]^y \ldots$ (iii)

Dividing equation (i) by (ii)

$\Rightarrow \frac{1.2 \times 10^{-3}}{1.2 \times 10^{-3}} = \frac{k[0.1]^x[0.1]^y}{k[0.1]^x[0.2]^y}$

We find, $y = 0$

Now dividing equation (i) by (iii)

$\Rightarrow \frac{1.2 \times 10^{-3}}{2.4 \times 10^{-3}} = \frac{k[0.1]^x[0.1]^y}{k[0.2]^x[0.1]^y}$

We find, $x = 1$

Hence, $\frac{d[C]}{dt} = k[A]^1[B]^0$

Question130

In the reaction of formation of sulphur trioxide by contact process $2SO_2 + O_2 \rightleftharpoons 2SO_3$ the rate of reaction was measured as $\frac{d[O_2]}{dt} = -2.5 \times 10^{-4} \text{mol L}^{-1}\text{s}^{-1}$. The rate of reaction is terms of $[SO_2]$in $\text{mol L}^{-1}\text{s}^{-1}$ will be:
[Online April 11, 2014]

Options:

A. -1.25×10^{-4}

B. -2.50×10^{-4}

C. -3.75×10^{-4}

D. -5.00×10^{-4}

Answer: D

Solution:

Solution:
From rate law

$$-\frac{1}{2}\frac{d[SO_2]}{dt} = -\frac{d[O_2]}{dt} = \frac{1}{2}\frac{d[SO_3]}{dt}$$

$$\therefore -\frac{d[SO_2]}{dt} = -2 \times \frac{d[O_2]}{dt}$$

$$= -2 \times 2.5 \times 10^{-4}$$

$$= -5 \times 10^{-4} \text{mol L}^{-1}\text{s}^{-1}$$

Question131

For the reaction, $2N_2O_5 \rightarrow 4NO_2 + O_2$, the rate equation can be expressed in two ways $-\frac{d[N_2O_5]}{dt} = k[N_2O_5]$ and $+\frac{d[NO_2]}{dt} = k'[N_2O_5]$
k and k' are related as:
[Online April 11, 2014]

Options:

A. $k = k'$

B. $2k = k'$

C. $k = 2k'$

D. $k = 4k'$

Answer: B

Solution:

Rate of disappearance of reactant = Rate of appearance of products

$$-\frac{1}{2}\frac{d[N_2O_5]}{dt}=\frac{1}{4}\frac{d[NO_2]}{dt}$$

$$\frac{1}{2}k[N_2O_5]=\frac{1}{4}k'[N_2O_5]$$

$$\frac{k}{2}=\frac{k'}{4}$$

$$\therefore k'=2k$$

Question132

For the reaction, 3A + 2B → C + D the differential rate law can be written as:
[Online April 19, 2014]

Options:

A. $\frac{1}{3}\frac{d[A]}{dt}=\frac{d[C]}{dt}=k[A]^n[B]^m$

B. $-\frac{d[A]}{dt}=\frac{d[C]}{dt}=k[A]^n[B]^m$

C. $+\frac{1}{3}\frac{d[A]}{dt}=-\frac{d[C]}{dt}=k[A]^n[B]^m$

D. $-\frac{1}{3}\frac{d[A]}{dt}=\frac{d[C]}{dt}=k[A]^n[B]^m$

Answer: D

Solution:

For the reaction

$3A+2B\longrightarrow C+D$

Rate of disappearance of A = Rate of appearance of C reaction

$$=-\frac{1}{3}\frac{d[A]}{dt}=\frac{d[C]}{dt}=k[A]^n[B]^m$$

Question133

The half-life period of a first order reaction is 15 minutes. The amount of substance left after one hour will be:
[Online April 9, 2014]

Options:

A. $\frac{1}{4}$ of the original amount

B. $\frac{1}{8}$ of the original amount

C. $\frac{1}{16}$ of the original amount

D. $\frac{1}{32}$ of the original amount

Answer: C

Solution:

Given $t_{1/2} = 15$ minutes
Total time $(T) = 1\text{hr} = 60\text{min}$
From $T = n \times t_{1/2}$
$n = \frac{60}{15} = 4$
Now from the formula $\frac{N}{N_0} = \left(\frac{1}{2}\right)^n = \left(\frac{1}{2}\right)^4 = \frac{1}{16}$
Where $N_0 =$ initial amout
$N =$ amount left after time t
Hence the amount of substance left after 1 hour will be $\frac{1}{16}$

Question134

The rate coefficient (k) for a particular reactions is $1.3 \times 10^{-4} M^{-1} s^{-1}$ at $100^{\circ}C$, and $1.3 \times 10^{-3} M^{-1} s^{-1}$ at $150^{\circ}C$. What is the energy of activation (E_0)(inkJ) for this reaction? (R = molargas constant $= 8.314 J K^{-1} mol^{-1}$)
[Online April 12, 2014]

Options:

A. 16

B. 60

C. 99

D. 132

Answer: B

Solution:

Solution:
According to Arrhenius equation $\log \frac{k_2}{k_1} = \frac{E_a}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$
$\log \frac{1.3 \times 10^{-3}}{1.3 \times 10^{-4}} = \frac{E_a}{2.303 \times 8.314}\left[\frac{1}{373} - \frac{1}{423}\right]$
$1 = \frac{E_a}{2.303 \times 8.314}\left[\frac{1}{373} - \frac{1}{423}\right]$
$E_a = 60\text{kJ} / \text{mol}$

Question135

The instantaneous rate of disappearance of MnO_4^- ion in the following reaction is $4.56 \times 10^{-3} M s^{-1}$
$2MnO_4^- + 10I^- + 16H^+ \rightarrow 2Mn^{2+} + 5I_2 + 8H_2O$
The rate of appearance I_2 is:
[Online April 9, 2013]

Options:

A. $4.56 \times 10^{-4} M s^{-1}$

B. $1.14 \times 10^{-2} M s^{-1}$

C. $1.14 \times 10^{-3} M s^{-1}$

D. $5.7 \times 10^{-3} M s^{-1}$

Answer: B

Solution:

Given $-\frac{d[MnO_4^-]}{dt} = 4.56 \times 10^{-3} M s^{-1}$
From the reaction given,
$-\frac{1}{2}\frac{d[MnO_4^-]}{dt} = \frac{4.56 \times 10^{-3}}{2} M s^{-1}$
$-\frac{1}{2}\frac{d[MnO_4^-]}{dt} = \frac{1}{5}\frac{dI_2}{dt}$
$\therefore -\frac{5}{2}\frac{d[MnO_4^-]}{dt} = \frac{dI_2}{dt}$
On substituting the given value
$\therefore \frac{dI_2}{dt} = \frac{4.56 \times 10^{-3} \times 5}{2} = 1.14 \times 10^{-2} M / s$

Question136

The rate constant of a zero order reaction is $2.0 \times 10^{-2} mol\ L^{-1} s^{-1}$. If the concentration of the reactant after 25 seconds is $0.5M$. What is the initial concentration?
[Online April 23, 2013]

Options:

A. $0.5M$

B. $1.25M$

C. $12.5M$

D. $1.0M$

Answer: D

Solution:

For a zero order reaction Rate constant

$= k = \frac{a - x}{t}$

$2 \times 10^{-2} = \frac{a - 0.5}{25}$

$a - 0.5 = 0.5$

$a = 1.0M$

Question137

A radioactive isotope having a half - life period of 3 days was received after 12 days. If 3g of the isotope is left in the container, what would be the initial mass of the isotope?
[Online April 25, 2013]

Options:

A. 12g

B. 36g

C. 48g

D. 24g

Answer: C

Solution:

Given $t_{1/2} = 3$

Total time $T = 12$

No. of half lives $(n) = \frac{12}{3} = 4$

$\left(\frac{1}{2}\right)^n = \frac{N}{N_o}$

$\therefore \left(\frac{1}{2}\right)^4 = \frac{3}{N}$

$\frac{3}{N} = \frac{1}{16}$

$N = 48g$

Question138

The rate of a reaction doubles when its temperature changes from 300K to 310K. Activation energy of the reaction will be : ($R = 8.314 J\ K^{-1} mol^{-1}$ and $\log 2 = 0.301$)
[2013]

Options:

A. $53.6 kJ\ mol^{-1}$

B. 48.6kJ mol^{-1}

C. 58.5kJ mol^{-1}

D. 60.5kJ mol^{-1}

Answer: A

Solution:

Activation energy can be calculated from the equation

$$\frac{\log k_2}{\log k_1} = \frac{E_a}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Given $\frac{k_2}{k_1} = 2; T_2 = 310\text{K}; T_1 = 300\text{K}$

$$\therefore \log 2 = \frac{E_a}{2.303 \times 8.314}\left(\frac{1}{300} - \frac{1}{310}\right)$$

$E_a = 53598.6\text{J}/\text{mol} = 53.6\text{kJ}/\text{mol}$

Question139

The reaction $X \rightarrow Y$ is an exothermic reaction. Activation energy of the reaction for X into Y is 150 kJ mol^{-1}. Enthalpy of reaction is 135kJ mol^{-1}. The activation energy for the reverse reaction, $Y \rightarrow X$ will be:

[Online April 22, 2013]

Options:

A. 280kJ mol^{-1}

B. 285kJ mol^{-1}

C. 270kJ mol^{-1}

D. 15kJ mol^{-1}

Answer: B

Solution:

$X \longrightarrow Y; \Delta H = -135\text{kJ}/\text{mol},$

$E_a = 150\text{kJ}/\text{mol}$

For an exothermic reaction

$E_{a(F.R.)} = \Delta H + E_{a(B.R)}'$

$150 = -135 + E_{a(B.R.)}'$

$E_{a(B.R)}' = 285\text{kJ}/\text{mol}$

Question140

For a first order reaction (A)→ products, the concentration of A changes from 0.1M to 0.025M in 40 minutes. The rate of reaction when the concentration of A is 0.01M is :
[2012]

Options:

A. $1.73 \times 10^{-5}\mathrm{M}$ / min

B. $3.47 \times 10^{-4}\mathrm{M}$ / min

C. $3.47 \times 10^{-5}\mathrm{M}$ / min

D. $1.73 \times 10^{-4}\mathrm{M}$ / min

Answer: B

Solution:

For a first order reaction

$$k = \frac{2.303}{t}\log\frac{a}{a-x} = \frac{2.303}{40}\log\frac{0.1}{0.025}$$

$$= \frac{2.303}{40}\log 4 = \frac{2.303 \times 0.6020}{40}$$

$$= 3.47 \times 10^{-2}$$

$$R = k(A)^1 = 3.47 \times 10^{-2} \times 0.01$$

$$= 3.47 \times 10^{-4}$$

Question141

In a chemical reaction A is converted into B. The rates of reaction, starting with initial concentrations of A as $2 \times 10^{-3}\mathrm{M}$ and $1 \times 10^{-3}\mathrm{M}$, are equal to $2.40 \times 10^{-4}\mathrm{M\,s^{-1}}$ and $0.60 \times 10^{-4}\mathrm{M\,s^{-1}}$ respectively. The orderof reaction with respect to reactant A will be
[Online May 12,2012]

Options:

A. 0

B. 1.5

C. 1

D. 2

Answer: D

Solution:

$A \longrightarrow B$

Initial concentration

$2 \times 10^{-3}\mathrm{M}$

$1 \times 10^{-3}\mathrm{M}$

Rate of reaction
$2.40 \times 10^{-4} \mathrm{M\ s^{-1}}$
$0.60 \times 10^{-4} \mathrm{M\ s^{-1}}$
rate of reaction
$r = k[A]^{x}$
where $x =$ order of reaction
hence
$2.40 \times 10^{-4} = k[2 \times 10^{-3x}]$....... (i)
$0.60 \times 10^{-4} = k[1 \times 10^{-3}]^{x}$ ....(ii)
On dividing eqn.(i) from eqn. (ii) we get
$4 = (2)^{x}$
$\therefore x = 2$
i.e. order of reaction = 2

Question142

For a reaction A→ Products, a plot of $\log t_{1/2}$ versus log a_0 is shown in the figure. If the initial concentration of A is represented by a_0, the order of the reaction is
[Online May 19, 2012]

Options:

A. one

B. zero

C. two

D. three

Answer: B

Solution:

Solution:
Plot given is for zero order reaction.

Question143

The activation energy for a reaction which doubles the rate when the temperature is raised from 298K to 308K is
[Online May 26, 2012]

Options:

A. $59.2 \mathrm{kJ\ mol^{-1}}$

B. $39.2 \mathrm{kJ\ mol}$

C. $52.9 \mathrm{kJ\ mol^{-1}}$

D. $29.5 \mathrm{kJ\ mol^{-1}}$

Answer: C

Solution:

Activation energy can be calculated from the equation

$$\frac{\log k_2}{\log k_1} = \frac{E_a}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Given, $\frac{\log k_2}{\log k_1} = 2$ $T_2 = 308$; $T_1 = 298$

$$\therefore \log 2 = \frac{E_a}{2.303 \times 8.314}\left(\frac{1}{298} - \frac{1}{308}\right)$$

$E_a = 52.9 \text{kJ mol}^{-1}$

Question144

Reaction rate between two substance A and B is expressed as following:
rate $= k[A]^n[B]^m$
If the concentration of A is doubled and concentration of B is made half of initial concentration, the ratio of the new rate to the earlierrate will be:
[Online May 7, 2012, Offline 2003]

Options:

A. $m + n$

B. $n - m$

C. $\frac{1}{2^{(m+n)}}$

D. $2^{(n-m)}$

Answer: D

Solution:

Rate $_1 = k[A]^n[B]^m$

Rate $_2 = k[2A]^n\left[\frac{1}{2}B\right]^m$

$$\therefore \frac{\text{Rate}_2}{\text{Rate}_1} = \frac{k[2A]^n\left[\frac{1}{2}B\right]^m}{k[A]^n[B]^m} = (2)^n\left(\frac{1}{2}\right)^m$$

$$= 2^n \cdot (2)^{-m} = 2^{n-m}$$

Question145

A reactant (A) from two products :

$A \xrightarrow{k_1} B$, Activation energy Ea_1

$A \xrightarrow{k_2} C$, Activation energy Ea_2

If $Ea_2 = 2Ea_1$, then k_1 and k_2 are related as :
[2011RS]

Options:

A. $k_2 = k_1 e^{Ea_1/RT}$

B. $k_2 = k_1 e^{Ea_2/RT}$

C. $k_1 = Ak_2 e^{Ea_1/RT}$

D. $k_1 = 2k_2 e^{Ea_2/RT}$

Answer: C

Solution:

$k_1 = A_1 e^{-Ea_1/RT}$
$k_2 = A_2 e^{-Ea_2/RT}$ .....(ii)
On dividing eqn. (i) by eqn. .....(ii)
$\frac{k_1}{k_2} = \frac{A_1}{A_2}\left(E_{a_2} - E_{a_1}\right) / RT$ ......(iii)
Given $Ea_2 = 2Ea_1$
On substituting this value in eqn. (iii)
$k_1 = k_2 A \times e^{Ea_1/RT}$

Question146

Consider the reaction:
$Cl_2(aq) + H_2S(aq) \rightarrow S(s) + 2H^+(aq) + 2Cl^-(aq)$
The rate equation for this reaction is rate $= k[Cl_2][H_2S]$
Which of these mechanisms is/are consistent with this rate equation?
A. $Cl_2 + H_2S \rightarrow H^+ + Cl^- + Cl^+ + HS^-$ (slow)
$Cl^+ + HS^- \rightarrow H^+ + Cl^- + S$ (fast)
B $H_2S \rightleftharpoons H^+ + HS^-$ (fast equilibrium)
$Cl_2 + HS^- \rightarrow 2Cl^- + H^+ + S$(Slow)
[2010]

Options:

A. B only

B. Both A and B

C. Neither A nor B

D. A only

Answer: D

Solution:

Since the slow step is the rate determining step, hence if we consider mechanism A. We find

Rate $= k[Cl_2][H_2S]$

Now, if we consider mechanism B. We find

Rate $= k[Cl_2][HS^-]$ . . . (i)

From step 1 (fast equilibrium) of mechanism B.

$$k = \frac{[H^+][HS^-]}{H_2S}$$

or $[HS^-] = \frac{k[H_2S]}{H^+}$

Substituting this value in equation (i) we find

Rate $= k[Cl_2]k\frac{[H_2S]}{H^+} = k'\frac{[Cl_2][H_2S]}{[H^+]}$

Hence, onlymechanism A is consistent with the given rate equation.

Question147

The time for half life period of a certain reaction A$\longrightarrow$ Products is 1 hour. When the initial concentration of the reactant ' A ' is 2.0mol L^{-1}, how much time does it take for its concentration to come from 0.50 to 0.25mol L^{-1}if it is a zero order reaction?

[2010]

Options:

A. 4h

B. 0.5h

C. 0.25h

D. 1h

Answer: C

Solution:

For the reaction

A→ Product

Given $t_{1/2} = 1$ hour

For a zero order reaction

$$t_{\text{completion}} = \frac{[A_0]}{k} = \frac{\text{Initial conc.}}{\text{Rate constant}}$$

$$\therefore t_{1/2} = \frac{[A_0]}{2k}$$

or $k = \frac{[A_0]}{2t_{1/2}} = \frac{2}{2 \times 1} = 1\text{mol lit}^{-1}\text{hr}^{-1}$

Further for a zero order reaction

$$k = \frac{dx}{dt} = \frac{\text{change in concentration}}{\text{time}}$$

$$1 = \frac{0.50 - 0.25}{\text{time}}$$

$\therefore$ time $= 0.25$hr.

Question148

The half life period of a first order chemical reaction is 6.93 minutes. The time required for the completion of 99% of the chemical reaction will be ($\log 2 = 0.301$)
[2009]

Options:

A. 23.03 minutes

B. 46.06 minutes

C. 460.6 minutes

D. 230.03 minutes

Answer: B

Solution:

For first order reaction,

$$k = \frac{2.303}{t} \log \frac{100}{100-99}$$

$$\frac{0.693}{6.93} = \frac{2.303}{t} \log \frac{100}{1}$$

$$\frac{0.693}{6.93} = \frac{2.303 \times 2}{t} \, 6.93$$

$$t = 46.06 \mathrm{min}$$

Question149

For a reaction $\frac{1}{2}A \rightarrow 2B$, rate of disappearance of ^{c}A ' is related to the rate of appearance of ' B ' by the expression
[2008]

Options:

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

B. $-\frac{d[A]}{dt} = \frac{1}{4}\frac{d[B]}{dt}$

C. $-\frac{d[A]}{dt} = \frac{d[B]}{dt}$

D. $-\frac{d[A]}{dt} = 4\frac{d[B]}{dt}$

Answer: B

Solution:

The rate of reaction for the reaction

$\frac{1}{2}\mathrm{A} \longrightarrow 2\mathrm{B}$

can be written either as

$-2\frac{\mathrm{d}}{\mathrm{d\,t}}[\mathrm{A}]$ with respect to ' A ' or $\frac{1}{2}\frac{\mathrm{d}}{\mathrm{d\,t}}[\mathrm{B}]$ with respect to 'B'

From the above, we have

$-2\frac{\mathrm{d}}{\mathrm{d\,t}}[\mathrm{A}] = \frac{1}{2}\frac{\mathrm{d}}{\mathrm{d\,t}}[\mathrm{B}]$

or $-\frac{\mathrm{d}}{\mathrm{d\,t}}[\mathrm{A}] = \frac{1}{4}\frac{\mathrm{d}}{\mathrm{d\,t}}[\mathrm{B}]$

Question150

Consider the reaction, 2A + B→ products. When concentration of B alone was doubled, the half-life did not change. When the concentration of A alone was doubled, the rate increased by two times. The unit of rate constant for this reaction is
[2007]

Options:

A. s^{-1}

B. $\mathrm{Lmol}^{-1}\mathrm{s}^{-1}$

C. no unit

D. $\mathrm{mol\,L}^{-1}\mathrm{S}^{-1}$.

Answer: B

Solution:

Solution:

For a first order reaction, $\mathrm{t}_{1/2} = \frac{0.693}{\mathrm{K}}$ i.e. for a first order reaction $\mathrm{t}_{1/2}$ does not depend up on the concentration. From the given data, we can say that order of reaction with respect to B = 1 because change in concentrationof B does not change half life. Order of reaction with respect to A = 1 because rate of reaction doubles when concentration of B is doubled keeping concentration of A constant.

∴ Order of reaction $= 1 + 1 = 2$ and units of second order reactionare $\mathrm{Lmol}^{-1}\mathrm{sec}^{-1}$.

Question151

A radioactive element gets spilled over the floor of a room. Its half-life period is 30 days. If the initial velocity is ten times the permissible value, after how many days will it be safe to enter the room?
[2007]

Options:

A. 100 days

B. 1000 days

C. 300 days

D. 10 days.

Answer: A

Solution:

Suppose activity of safe working $= A$
Given $A_0 = 10A$

$$\lambda = \frac{0.693}{t_{1/2}} = \frac{0.693}{30}$$

$$t_{1/2} = \frac{2.303}{\lambda} \log \frac{A_0}{A} = \frac{2.303}{0.693/30} \log \frac{10A}{A}$$

$$= \frac{2.303 \times 30}{0.693} \times \log 10 = 100 \text{ days.}$$

Question152

The energies of activation for forward and reverse reactions for $A_2 + B_2 \rightleftharpoons 2AB$ are 180kJ mol^{-1} and 200kJ mol^{-1} respectively. The presence of a catalyst lowers the activation energy of both (forward and reverse) reactions by 100kJ mol^{-1}. The enthalpy change of the reaction ($A_2 + B_2 \rightarrow 2AB$) in the presence of a catalyst will be (in kJ mol^{-1}) [2007]

Options:

A. 20

B. 300

C. 120

D. 280

Answer: A

Solution:

Solution:
$\Delta H_R = E_f - E_b = 180 - 200 = -20 \text{kJ} / \text{mol}$
The nearest correct answer given in choices may be obtained by neglecting sign.

Question153

The following mechanism has been proposed for the reaction of NO with Br_2 to form NOBr :

$NO(g) + Br_2(g) \rightleftharpoons NOBr_2(g)$

$\mathrm{NOBr_2(g) + NO(g) \longrightarrow 2NOBr(g)}$

If the second step is the rate determining step, the order of the reaction with respect to $\mathrm{NO(g)}$ is
[2006]

Options:

A. 3

B. 2

C. 1

D. 0

Answer: B

Solution:

Solution:
(i) $\mathrm{NO(g) + Br_2(g) \rightleftharpoons NOBr_2(g)}$
(ii) $\mathrm{NOBr_2(g) + NO(g) \longrightarrow 2NOBr(g)}$
Rate law equation $= \mathrm{k[NOBr_2][NO]}$
But $\mathrm{NOBr_2}$ is intermediate and must not appear in the rate law equation.
From Ist step,
$\mathrm{K_c} = \frac{\mathrm{[NOBr_2]}}{\mathrm{[NO][Br_2]}}$
$\therefore \mathrm{[NOBr_2] = K_c[NO][Br_2]}$
$\therefore$ Rate law equation $= \mathrm{k \cdot K_c[NO]^2[Br_2]}$
Hence order of reaction w.r.t. NO is 2 .

Question154

A reaction was found to be second order with respect to the concentration of carbon monoxide. If the concentration of carbon monoxide is doubled, with everything else kept the same, the rate of reaction will
[2006]

Options:

A. increase by a factor of 4

B. double

C. remain unchanged

D. triple

Answer: A

Solution:

Since the reaction is 2nd order w.r.t CO. Thus, rate law is given as.
$r = k[CO]^2$
Let initial concentration of CO is a
i.e. [CO] = a
$\therefore\ r_1 = k(a)^2 = ka^2$
When concentration becomes doubled, i.e. [CO] = 2a
$\therefore\ r_2 = k(2a)^2 = 4ka^2 \quad \therefore r_2 = 4r_1$
So, the rate of reaction becomes 4 times.

Question155

Rate of a reaction can be expressed by Arrhenius equation as: $k = Ae^{-E a / RT}$ In this equation, E a represents [2006]

Options:

A. the total energy of the reacting molecules at a temperature, I

B. the fraction of molecules with energy greater than the activation energy of the reaction

C. the energy above which all the colliding molecules will react

D. the energy below whichall the colliding molecules will react

Answer: C

Solution:

Solution:
In Arrhenius equation, $k = Ae^{-E aRT}$, E_a is the energy of activation, which is required by the colliding molecules to react resulting in the formation of products.

Question156

A reaction involving two different reactants can never be [2005]

Options:

A. bimolecular reaction

B. second order reaction

C. first order reaction

D. unimolecular reaction

Answer: D

Solution:

The molecularity of a reaction is the number of reactant molecules taking part in a single step of the reaction. Note: The reaction involving two different reactants can never be unimolecular.

Question157

$\frac{t_{1}}{4}$ can be taken as the time taken for the concentration of a reactant to drop to $\frac{3}{4}$ of its initial value. If the rate constant for a first order reaction is k, the $\frac{t_{1}}{4}$ can be written as
[2005]

Options:

A. 0.75 / k

B. 0.69 / k

C. 0.29 / k

D. 0.10 / k

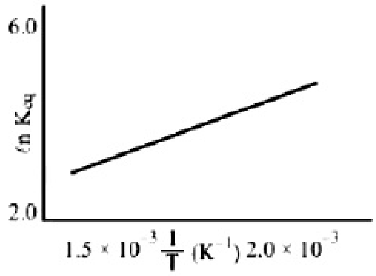
Answer: C

Solution:

$$t_{1/4} = \frac{2.303}{K}\log\frac{1}{3/4} = \frac{2.303}{K}\log\frac{4}{3}$$
$$= \frac{2.303}{K}(\log 4 - \log 3)$$
$$= \frac{2.303}{K}(2\log 2 - \log 3)$$
$$= \frac{2.303}{K}(2 \times 0.301 - 0.4771) = \frac{0.29}{K}$$

Question158

A schematic plot of $\ln K_{eg}$ versus inverse of temperature for a reaction is shown below



The reaction must be
[2005]

Options:

A. highly spontaneous at ordinary temperature

B. one with negligible enthalpy change

C. endothermic

D. exothermic

Answer: D

Solution:

The graph shows that reaction is exothermic.

$\log k = \frac{-\Delta H}{RT} + 1$

For exothermic reaction $\Delta H < 0$

$\therefore \log k V s \frac{1}{T}$ would be negative straight line with positive slope.

Question159

Consider an endothermic reaction $X \rightarrow Y$ with the activation energies E_b and E_f for the backward and forward reactions, respectively. In general [2005]

Options:

A. there is no definite relation between E_b and E_f

B. $E_b = E_f$

C. $E_b > E_f$

D. $E_b < E_f$

Answer: D

Solution:

Solution:

Enthalpy of reaction $(\Delta H) = E_f - E_b$

For an endothermic reaction, $\Delta H = +$ ve hence for ΔH to be negative

$E_b < E_f$

Question160

The rate equation for the reaction $2A + B \rightarrow C$ is found to be: rate $= k[A][B]$. The correct statement in relation to this reaction is that the [2004]

Options:

A. rate of formation of C is twice the rate of disappearance of A

B. $t_{1/2}$ is a constant

C. unit of k must be s^{-1}

D. value of k is independent of the initial concentrations of A and B

Answer: D

Solution:

Solution:
The velocity constant depends on temperature only. It is independent of concentration of reactants.

Question161

In a first order reaction, the concentration of the reactant, decreases from 0.8M to 0.4M in 15 minutes. The time taken for the concentration to change from 0.1M to 0.025 M is [2004]

Options:

A. 7.5 minutes

B. 15 minutes

C. 30 minutes

D. 60 minutes

Answer: C

Solution:

Solution:
As the concentration of reactant decreases from 0.8 to 0.4 in 15 minutes hence the $t_{1/2}$ is 15 minutes. To fall the concentration from 0.1 to 0.025 we need two halflives i.e., 30 minutes.

Question162

The half-life of a radioisotope is four hours. If the initial mass of the isotope was 200g, the mass remaining after 24 hours undecayed is [2004]

Options:

A. 3.125g

B. 2.084g

C. 1.042g

D. 4.167g

Answer: A

Solution:

$N_t = N_0 \left(\frac{1}{2}\right)^n$ where n is number of half life periods.

$n = \frac{\text{Total time}}{\text{half life}} = \frac{24}{4} = 6$

$\therefore N_t = 200\left(\frac{1}{2}\right)^6 = 3.125g$

Question163

For the reaction system:
$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
volume is suddenly reduced to half its value by increasing the pressure on it. If the reaction is of first order with respect to O_2 and second order with respect to NO, the rate of reaction will
[2003]

Options:

A. diminish to one-eighth of its initial value

B. increase to eight times of its initial value

C. increase to four times of its initial value

D. diminish to one-fourth of its initial value

Answer: B

Solution:

Solution:
$r = k[O_2][NO]^2$. When the volume is reduced to 1 / 2, the conc. will double
$\therefore$ New rate $= k[2O_2][2NO]^2 = 8k[O_2][NO]^2$
The new rate increases to eight times of its initial.

Question164

The half-life of a radioactive isotope is three hours. If the initial mass of the isotope were 256g, the mass of it remaining undecayed after 18 hours would be
[2003]

Options:

A. 8.0g

B. 12.0g

C. 16.0g

D. 4.0g

Answer: D

Solution:

$t_{1/2} = 3h, T = 18h$
$\because T = n \times t_{1/2}$
$\therefore n = \frac{18}{3} = 6$
Initial mass $(C_0) = 256g$
$\therefore C_n = \frac{C_0}{2^n} = \frac{256}{(2)^6} = \frac{256}{64} = 4g$

Question165

In respect of the equation $k = Ae^{-E_a/RT}$ in chemical kinetics, which one of the following statements is correct? [2003]

Options:

A. A is adsorption factor

B. E_a is energy of activation

C. R is Rydberg's constant

D. k is equilibrium constant

Answer: B

Solution:

In equation $k = Ae^{-E_a/RT}$;
$A =$ Frequency factor
$k =$ velocity constant, $R =$ gas constant and $E_a =$ energy of activation

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Question166

The differential rate law for the reaction

$H_2 + I_2 \rightarrow 2HI$ is
[2002]

Options:

A. $-\frac{d[H_2]}{dt} = -\frac{d[I_2]}{dt} = -\frac{d[HI]}{dt}$

B. $\frac{d[H_2]}{dt} = \frac{d[I_2]}{dt} = \frac{1}{2}\frac{d[HI]}{dt}$

C. $\frac{1}{2}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[I_2]}{dt} = -\frac{d[HI]}{dt}$

D. $-2\frac{d[H_2]}{dt} = -2\frac{d[I_2]}{dt} = \frac{d[HI]}{dt}$

Answer: D

Solution:

Rate of appearance of HI $= \frac{1}{2}\frac{d[HI]}{dt}$

Rate of disappearance of $H_2 = \frac{-d[H_2]}{dt}$

Rate of disappearance of $I_2 = \frac{-d[I_2]}{dt}$

hence $-\frac{d[H_2]}{dt} = -\frac{d[I_2]}{dt} = \frac{1}{2}\frac{d[HI]}{dt}$

or $-2\frac{d[H_2]}{dt} = -2\frac{d[I_2]}{dt} = \frac{d[HI]}{dt}$

Question167

The integrated rate equation is
$Rt = \log C_0 - \log C_i$
The straight line graph is obtained by plotting
[2002]

Options:

A. time vs $\log C,$

B. $\frac{1}{\text{time}}$ vs C_t

C. time vs C_t

D. $\frac{1}{\text{time}}$ vs $\frac{1}{C_t}$

Answer: A

Solution:

Solution:
$Rt = \log C_0 - \log C_t$
It is clear from the equation that if we plot a graph between $\log C_t$ and time, a straight line with a slope equal to $-\frac{k}{2.303}$ and intercept equal to $\log[C_0]$ will be obtained.

Question168

Units of rate constant of first and zero order reactions in terms of molarity M unit are respectively
[2002]

Options:

A. $\sec^{-1}$, $\mathrm{M} \sec^{-1}$

B. $\sec^{-1}$, M

C. $\mathrm{M} \sec^{-1}$, $\sec^{-1}$

D. M , $\sec^{-1}$.

Answer: A

Solution:

For a zero order reaction.
rate $= \mathrm{k}[\mathrm{A}]^{\circ}$ i.e. rate $= \mathrm{k}$
hence unit of $\mathrm{k} = \mathrm{M} \cdot \sec^{-1}$
For a first order reaction. rate $= \mathrm{k}[\mathrm{A}]$
$\therefore \mathrm{k} = \mathrm{M} \cdot \sec^{-1} / \mathrm{M} = \sec^{-1}$

Question169

For the reaction A + 2B → C, rate is given by R = $[A][B]^2$ then the order of the reaction is
[2002]

Options:

A. 3

B. 6

C. 5

D. 7

Answer: A

Solution:

Solution:
Note : Order is the sum of the power of the concentration terms in rate law expression. Hence the order of reaction is $= 1 + 2 = 3$

Question170

If half-life of a substance is 5 yrs, then the total amount of substance left after 15 years, when initial amount is 64 grams is
[2002]

Options:

A. 16 grams

B. 2 grams

C. 32 grams

D. 8 grams

Answer: D

Solution:

$t_{1/2} = 5$ years, $T = 15$ years.

Hence total number of half life periods $= \frac{15}{5} = 3$

$\therefore$ Amount left $= \frac{64}{(2)^3} = 8\text{g}$
