Chemical Kinetics

Question1

Consider a general first order reaction,

$A(g) \longrightarrow B(g)+C(g)$

If the initial pressure is 200 mm and after 20 minutes it is 250 mm , then the half-life period of the reaction (in minutes) is ($\log 2=0.30, \log 3=0.48, \log 4=0.60$ **)**

TG EAPCET 2025 (Online) 2nd May Evening Shift

Options:

A.

40.2

B.

50.2

C.

20.5

D.

60.5

Answer: B

Solution:

Given reaction is, $A(g) \longrightarrow B(g)+C(g)$

Let x be the decrease in pressure of A.

The pressure of A at time t is $p_A = p_0 - x$

The total pressure at time t is

$$p_t = p_0 - x + x + x = p_0 + x$$
$$250 = 200 + x; x = 50 \text{ mm}$$
$$p_A = 200 - 50 = 150 \text{ mm}$$

Use the integrated rate law,

$k = \frac{2.303}{t} \log \left(\frac{p_0}{p_A} \right)$

Substitute the value,

$$k = \frac{2.303}{20} \log \left(\frac{200}{150} \right) \Rightarrow k = 0.0138 \text{ min}^{-1}$$
$$t_{1/2} = \frac{0.693}{k}$$
$$t_{1/2} = \frac{0.693}{0.0138} = 50.2 \text{ minutes}$$

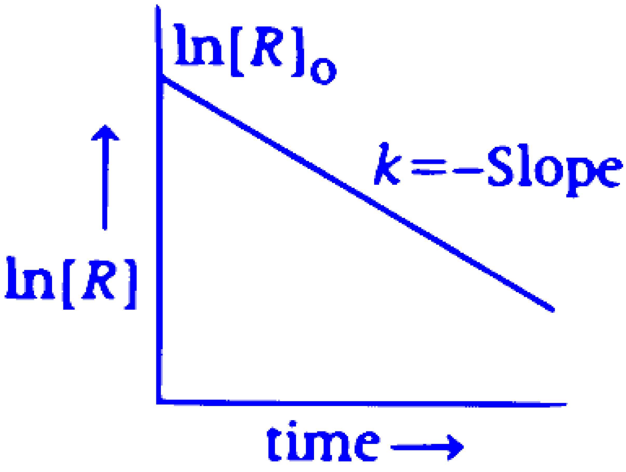
Question2

For the reaction $R \rightarrow P$, half life is independent of initial concentration of the reactant, R. Which one of the following graphs is not correct for the reaction?

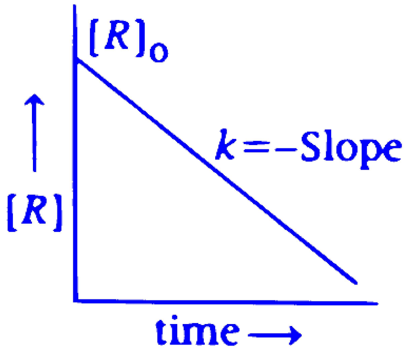
TG EAPCET 2025 (Online) 2nd May Morning Shift

Options:

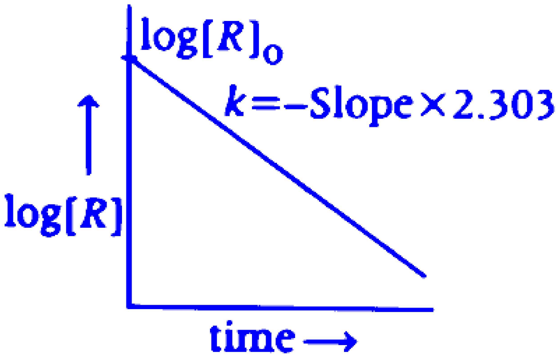
A.



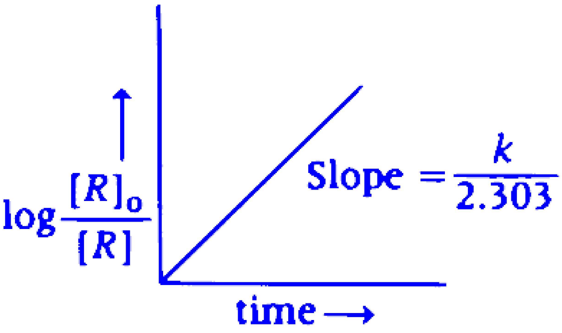
B.



C.



D.



Answer: B

Solution:

$R \longrightarrow P$

Half-life is independent of initial concentration, means it's a first order reaction.

Integrated rate law,

$\ln[R] = -kt + \ln[R_0]$.

Thus, $\ln R vs$ time is a straight line with slope $= -k$.

Option (b) graph is incorrect because it $[R]$ vs t.

So, it should be exponential.

Question3

For the gaseous reaction, $N_2O_5 \longrightarrow 2NO_2 + \frac{1}{2}O_2$

the rate can be expressed as

$$-\frac{d[\ N_2O_5]}{dt} = K_1\ [\ N_2O_5]$$

$$+\frac{d[NO_2]}{dt} = K_2\ [\ N_2O_5]$$

$$+\frac{d[O_2]}{dt} = K_3\ [\ N_2O_5]$$

The correct relation between K_1, K_2 and K_3

TG EAPCET 2024 (Online) 11th May Morning Shift

Options:

A. $K_1 = 2K_2 = 4K_3$

B. $2K_1 = K_2 = 4K_3$

C. $2K_1 = 3K_2 = 4K_3$

D. $4K_1 = 2K_2 = K_3$

Answer: B

Solution:

The given reaction is:

$N_2O_5 \longrightarrow 2NO_2 + \frac{1}{2}O_2$

For this reaction, the rates of change of concentrations are expressed as:

$-\frac{d[N_2O_5]}{dt} = K_1[N_2O_5]$

$\frac{d[NO_2]}{dt} = K_2[N_2O_5]$

$\frac{d[O_2]}{dt} = K_3[N_2O_5]$

From the stoichiometry of the reaction:

$$-\frac{d[\mathrm{N_2O_5}]}{dt} = \frac{1}{2}\frac{d[\mathrm{NO_2}]}{dt} = 2\frac{d[\mathrm{O_2}]}{dt}$$

Therefore, by equating the expressions, we get:

$$K_1[\mathrm{N_2O_5}] = \frac{1}{2}K_2[\mathrm{N_2O_5}] = 2K_3[\mathrm{N_2O_5}]$$

Multiplying through by 2 to eliminate fractions gives:

$$2K_1[\mathrm{N_2O_5}] = K_2[\mathrm{N_2O_5}] = 4K_3[\mathrm{N_2O_5}]$$

Thus, the relationship between the rate constants is:

$$2K_1 = K_2 = 4K_3$$

Question4

For a first order reaction, the graph between $\log \frac{a}{(a-x)}$ (on y-axis) and time (in min, on x-axis) gave a straight line passing through origin. The slope is $2 \times 10^{-3}\ \mathrm{min}^{-1}$. What is the rate constant (in min^{-1})?

TG EAPCET 2024 (Online) 10th May Evening Shift

Options:

A. 2×10^{-3}

B. $\frac{2\times10^{-3}}{2.303}$

C. 4.606×10^{-3}

D. 0.5×10^{-5}

Answer: C

Solution:

The first order reaction equation is

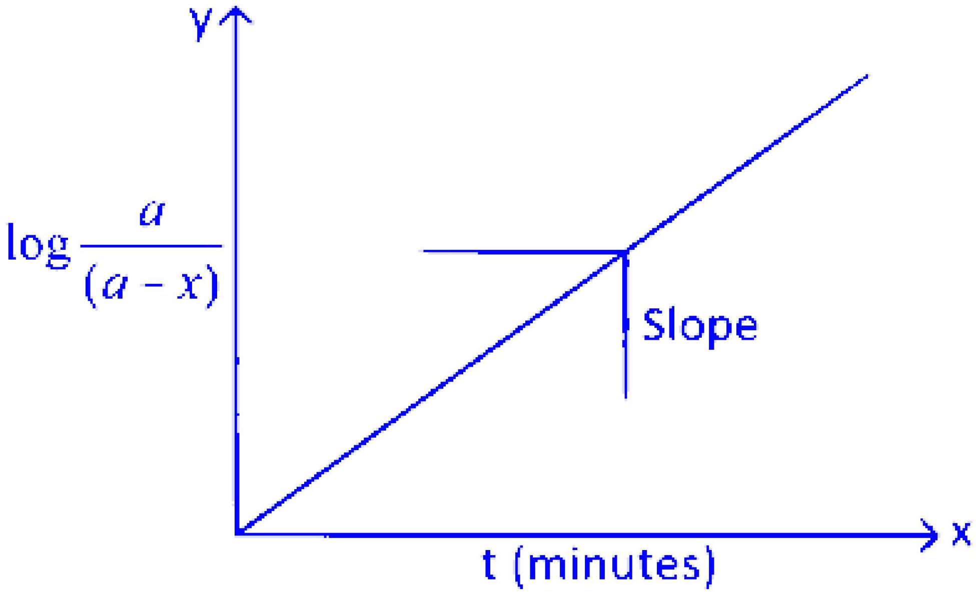
$t = \frac{2.303}{k} \log \frac{(a)}{(a-x)}$

graph is drawn between log $\frac{(a)}{(a-x)}$ on

y-axis and t on x-axis.

Slope from graph is

$\frac{\log \frac{a}{a-x}}{t} = \text{slope}$



From Eq. (i), $\frac{\log \frac{a}{a-x}}{t} = \frac{k}{2303}$

$\text{Slope} = \frac{k}{2303}$

$2 \times 10^{-3} = \frac{k}{2303}$

$k = 4.606 \times 10^{-3}/\text{min}$

Question5

The decomposition of benzene diazonium chloride is a first order reaction. The time taken for its decomposition to $\frac{1}{4}$ and $\frac{1}{10}$ of its initial concentration are $t_{\frac{1}{4}}$ and $t^{10}_{\frac{1}{4}}$ respectively. The value of $\frac{t^{4}_{\frac{1}{4}}}{t_{1}} \times 100$ is (Give: $\log 2 = 0.3$)

TG EAPCET 2024 (Online) 10th May Morning Shift

Options:

A. 60

B. 30

C. 90

D. 45

Answer: A

Solution:

For a first-order reaction, the time required for the reaction to reach a certain fraction of its initial concentration is given by the equation:

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

where:

a is the initial concentration,

$(a-x)$ is the concentration after time t,

k is the rate constant.

Case I: $\frac{1}{4}$ of the initial concentration remains after decomposition.

For this scenario, we have:

$t_{\frac{1}{4}} = \frac{2.303}{k} \log \frac{1}{\frac{1}{4}}$

$\Rightarrow t_{\frac{1}{4}} = \frac{2.303}{k} \times 0.6$

Case II: $\frac{1}{10}$ of the initial concentration remains after decomposition.

For this situation:

$$t_{\frac{1}{10}} = \frac{2.303}{k} \log \frac{1}{\frac{1}{10}}$$

$$\Rightarrow t_{\frac{1}{10}} = \frac{2.303}{k}$$

Calculating the ratio and scaling it by 100:

The task requires finding the ratio $\frac{t_{\frac{1}{4}}}{t_{\frac{1}{10}}}$ and multiplying it by 100.

$$\begin{aligned} \frac{t_{\frac{1}{4}}}{t_{\frac{1}{10}}} \times 100 &= \left(\frac{\frac{2.303}{k} \times 0.6}{\frac{2.303}{k}} \right) \times 100 \\ &= 0.6 \times 100 \\ &= 60 \end{aligned}$$

Thus, the calculated value is 60.

Question6

Consider the gaseous reaction,$A_2 + B_2 \longrightarrow 2AB$The following data was obtained for the above reaction.

$[A_2]_0$	$[B_2]_0$	Initial rate of formation of AB (mol L^{-1} s^{-1})
0.1 M	0.1 M	2.5×10^{-4}
0.2 M	0.1 M	5.0×10^{-4}
0.2 M	0.1 M	1.0×10^{-3}

The value of rate constant for the above reaction is

TG EAPCET 2024 (Online) 9th May Evening Shift

Options:

A. 1.25×10^{-2}

B. 1.25×10^{-3}

C. 2.5×10^{-2}

D. 2.5×10^{-1}

Answer: C

Solution:

$-\frac{dA}{dt} = k[A]^x[B]^y \quad \ldots \text{(i)}$

On doubling [A] rate is doubled

Hence, $2^x = 2$

$x = 1 \quad \ldots \text{(ii)}$

Similarly, $y = 1$

$\therefore \quad -\frac{d[A]}{dt} = k[A][B] \quad \ldots \text{(iii)}$

Substituting the given values,

$$k = \frac{2.5 \times 10^{-4}}{0.1 \times 0.1}$$

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

Question7

For a first order reaction, a plot of $\ln k$ (Y-axis) and $\frac{1}{T}$ (X-axis) gave the straight line with slope equal to -10^3 K and intercept equal to 2.303 (on Y-axis). What is the activation energy (E_a in kJmol^{-1}) of the reaction? (Given, $R = 8314 \mathrm{JK}^{-1} \mathrm{~mol}^{-1}$)

TG EAPCET 2024 (Online) 9th May Morning Shift

Options:

A. 8.314

B. 2.303

C. 2303

D. 83.14

Answer: A

Solution:

The Arrhenius equation describes the temperature dependence of the rate constant k for a reaction, and is given by:

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

Taking the natural logarithm of both sides gives:

$\ln k = \ln A - \frac{E_a}{R} \cdot \frac{1}{T}$

This equation represents a linear plot of $\ln k$ vs. $\frac{1}{T}$, where the slope is $-\frac{E_a}{R}$ and the intercept is $\ln A$.

Given:

Slope $= -10^3\,\mathrm{K}$

Intercept $= 2.303$

$R = 8314\,\mathrm{J\,K^{-1}\,mol^{-1}}$

The slope is related to the activation energy E_a by:

$-\frac{E_a}{R} = -10^3$

Solving for E_a, we have:

$E_a = 10^3 \cdot R$

Substituting the value of R:

$E_a = 10^3 \cdot 8314\,\mathrm{J\,mol^{-1}}$

To convert to kilojoules per mole, divide by 1000:

$E_a = \frac{8314 \times 10^3}{1000}\,\mathrm{kJ\,mol^{-1}} = 8314\,\mathrm{kJ\,mol^{-1}}$

Therefore, the correct answer is:

Option A: 8.314

Question8

The graph obtained between $\ln k$ ($k =$ rate constant) on y-axis and $1/T$ on x-axis is a straight line. The slope of it is -4×10^4 K. The activation energy of the reaction (in $\mathrm{kJmol^{-1}}$) is

$\left(R = 831\ \mathrm{J\,K^{-1}\ mol^{-1}}\right)$

TS EAMCET 2023 (Online) 12th May Evening Shift

Options:

A. 166

B. 332

C. 765

D. 382

Answer: B

Solution:

The graph between $\ln k$ (where k is the rate constant) on the y-axis and $1/T$ on the x-axis is a straight line. The slope of this line is -4×10^4 K. We need to determine the activation energy of the reaction in kJ mol^{-1} given that $R = 831$ J K^{-1} mol^{-1}.

According to the Arrhenius equation:

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

Taking the natural logarithm:

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

This can be compared to the equation of a straight line $y = mx + c$, where:

The slope (m) is $-\frac{E_a}{R}$.

Given the slope $m = -4 \times 10^4$, we equate and solve for E_a:

$-4 \times 10^4 = -\frac{E_a}{831}$

Solving for E_a, we get:

$E_a = 4 \times 10^4 \times 831$

$E_a = 33.2 \times 10^4$ J/mol

Converting joules to kilojoules:

$E_a = 332$ kJ/mol

Question9

If benzene diazonium chloride undergoes first order decomposition at $T(\text{ K})$ with a rate constant of 6.93×10^{-2} min^{-1}, the time for completion of 90% of the reaction (in min) is (nearest integer)

$(\log 2 = 0.30, \log 3 = 0.477\ )$

TS EAMCET 2023 (Online) 12th May Morning Shift

Options:

A. 15

B. 30

C. 33

D. 43

Answer: C

Solution:

To solve the problem, we have the following information:

The reaction follows first-order kinetics: $n = 1$

The rate constant is $k = 6.93 \times 10^{-2}\,\text{min}^{-1}$

Initial concentration, $a = 100$

Final concentration after 90% completion, $a - x = 10$

We need to find the time t it takes for 90% of the reaction to complete.

Using the formula for a first-order reaction rate constant:

$t = \frac{2.303}{k} \log \frac{[A_0]}{[A]}$

Substituting the given values:

$t = \frac{2.303}{6.93 \times 10^{-2}} \log \left(\frac{100}{10}\right)$

Simplifying:

$t = \frac{2.303}{6.93 \times 10^{-2}} \times \log(10)$

Since $\log(10) = 1$:

$t = \frac{2.303}{6.93 \times 10^{-2}}$

Calculating this gives:

$t \approx 33.23\,\text{min}$

Therefore, rounding to the nearest integer, the time for 90% completion of the reaction is approximately 33 minutes.