Chemical Kinetics

Question1

Which of the following statements is NOT correct regarding order of reaction?

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Options:

A.

It is determined experimentally.

B.

It is not influenced by stoichiometric coefficient of reactants.

C.

It is sum of power to the concentration terms of reactants in rate law equation.

D.

It is always whole number.

Answer: D

Solution:

- **Option A:** *It is determined experimentally.*

  ✅ Correct — the order of a reaction cannot be determined from the balanced equation alone; it must be determined experimentally.

- **Option B:** *It is not influenced by stoichiometric coefficient of reactants.*

  ✅ Correct — reaction order is not directly tied to stoichiometric coefficients, only to the experimental rate law.

- **Option C:** *It is sum of power to the concentration terms of reactants in rate law equation.*

  ✅ Correct — order of reaction is indeed the sum of the exponents of concentration terms in the experimentally determined rate law.

- **Option D:** *It is always whole number.*

  ❌ Incorrect — order can be fractional (e.g., 1/2) or even zero, not necessarily a whole number.

✅ Final Answer:

Option D: It is always whole number is NOT correct.

Question2

What is half life of a first order reaction if time required to decrease concentration of reactant from 0.4 M to 0.1 M is x hour?

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Options:

A.

x hour

B.

$2x$ hour

C.

$\frac{x}{2}$ hour

D.

$\frac{3x}{2}$ hour

Answer: C

Solution:

Step 1: General formula for first-order kinetics

For a first-order reaction,

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

or equivalently,

$\ln \frac{[A]_0}{[A]} = kt$

Step 2: Information given

- Initial concentration $[A]_0 = 0.4\,M$
- Final concentration $[A] = 0.1\,M$
- Time taken = x hours

So,

$k = \frac{1}{x} \ln \frac{0.4}{0.1} = \frac{1}{x} \ln(4) = \frac{1}{x} \cdot (1.386) = \frac{1.386}{x}$

Step 3: Half-life for a first-order reaction

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

Substitute value of k:

$t_{1/2} = \frac{0.693}{1.386/x} = \frac{0.693 \cdot x}{1.386} = \frac{x}{2}$

✅ Final Answer

The half-life of the reaction is:

$\boxed{\frac{x}{2}\text{ hour}}$

Correct option: C

Question3

For the reaction,

$$2\,N_2O_{5(g)} \longrightarrow 4NO_{2(g)} + O_{2(g)}$$

N_2O_5 disappears at a rate of x $\text{moldm}^{-3}\,\text{s}^{-1}$

Find the rate of formation of O_2 ?

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Options:

A.

$x \text{ moldm}^{-3}\text{ s}^{-1}$

B.

$2x \text{ moldm}^{-3}\text{ s}^{-1}$

C.

$\frac{x}{2} \text{ moldm}^{-3}\text{ s}^{-1}$

D.

$\frac{3x}{2} \text{ moldm}^{-3}\text{ s}^{-1}$

Answer: C

Solution:

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

Rate of disappearance of N_2O_5

$$= -\frac{d[N_2O_5]}{dt} = x \text{ moldm}^{-3}\text{ s}^{-1}$$

Rate of formation of O_2 is given by,

$$\frac{d[O_2]}{dt} = \frac{-1}{2}\frac{d[N_2O_5]}{dt}$$
$$\therefore \quad \frac{d[O_2]}{dt} = \frac{x}{2} \text{ moldm}^{-3}\text{ s}^{-1}$$

Question4

A complex reaction takes place in following steps.

$$NO_2Cl_{(g)} \longrightarrow NO_{2(g)} + Cl_{(g)} \text{ (slow)}$$
$$NO_2Cl_{(g)} + Cl_{(g)} \longrightarrow NO_{2(g)} + Cl_{2(g)} \text{ (fast)}$$

Identify rate law equation for this reaction.

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Options:

A.

$r = k[NO_2Cl]$

B.

$r = k[NO_2][Cl]$

C.

$r = k[NO_2]^2$

D.

$r = k[NO_2Cl]^2$

Answer: A

Solution:

The first step being slower than the second it is the rate determining step.

$\therefore$ The rate law is $r = k[NO_2Cl]$

For a complex reaction, rate law can be written as law of mass action for the rate determining step.

Question5

What is the average rate of reaction when change in concentration of product is 0.05 M in 20 seconds?

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Options:

A.

$0.0025 \mathrm{Ms}^{-1}$

B.

$0.05 \mathrm{Ms}^{-1}$

C.

$1.0 \mathrm{Ms}^{-1}$

D.

$4.0 \mathrm{Ms}^{-1}$

Answer: A

Solution:

The average rate of reaction tells us how fast the amount of product changes over time.

To find the average rate, use this formula:

Average rate $= \frac{\Delta[\text{Product}]}{\Delta t}$

Here, $\Delta[\text{Product}]$ is the change in concentration of the product, which is $0.05\,\mathrm{M}$, and Δt is the time taken, which is $20\,\mathrm{s}$.

Now, substitute the values into the formula:

Average rate $= \frac{0.05\ \mathrm{M}}{20\ \mathrm{s}}$

Do the division:

Average rate $= 0.0025\,\mathrm{M\,s^{-1}}$

So, the average rate of reaction is $0.0025\,\mathrm{M\,s^{-1}}$.

Question6

Identify role of glycerol in following reaction.

$$2\mathrm{H_2O_{2(\ell)}} \xrightarrow{\text{glycerol}} 2\mathrm{H_2O_{(\ell)}} + \mathrm{O_{2(g)}}$$

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Options:

A.

Homogeneous catalyst

B.

Inhibitor

C.

Promoter

D.

Heterogeneous catalyst

Answer: B

Solution:

We are given the reaction:

$$2H_2O_2 \xrightarrow{\text{glycerol}} 2H_2O + O_2$$

Step 1: Recall what happens here

Hydrogen peroxide decomposes to water and oxygen. This decomposition can be catalyzed by various substances.

- **Catalysts** speed up the decomposition.
- **Inhibitors** slow down or prevent it.
- **Promoters** increase the efficiency of a catalyst (but do not act as catalyst themselves).
- **Homogeneous catalyst**: Catalyst is in the same phase as the reactant.
- **Heterogeneous catalyst**: Catalyst is in a different phase than the reactant.

Step 2: Glycerol’s actual role

- Glycerol is not a catalyst in this decomposition. Instead, **glycerol acts as an inhibitor** of hydrogen peroxide decomposition.
- Substances like glycerol, alcohol, acetanilide, etc., are commonly added to hydrogen peroxide solutions to stabilize them and prevent rapid decomposition.

Correct Answer:

Option B: Inhibitor ✅

Question7

The slope of a graph $\log[A]_1$ versus ' t ' for first order reaction is $-2.5 \times 10^{-3}\ \mathrm{s}^{-1}$. Find rate constant of the reaction?

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Options:

A.

$1.263 \times 10^{-3}\ \mathrm{s}^{-1}$

B.

$3.471 \times 10^{-3}\ \mathrm{s}^{-1}$

C.

$5.757 \times 10^{-3}\ \mathrm{s}^{-1}$

D.

$8.125 \times 10^{-3}\ \mathrm{s}^{-1}$

Answer: C

Solution:

The integrated rate law for the first reaction is

$k=\frac{2.303}{t} \log_{10} \frac{[A]_0}{[A]_t}$

On rearrangement, the equation becomes

$$\underset{y}{\underset{\updownarrow}{\log_{10}[A]_t}} = -\underset{m}{\underset{\updownarrow}{\frac{k}{2.303}}}\,\underset{x}{\underset{\updownarrow}{t}} + \underset{c}{\underset{\updownarrow}{\log_{10}[A]_0}}$$

A graph of $\log_{10}[\mathrm{A}]_1$ versus t yields a straight line with slope $= -\frac{\mathrm{k}}{2.303} = -2.5 \times 10^{-3}\ \mathrm{s}^{-1}$

$\mathrm{k} = (-2.303) \times (-2.5 \times 10^{-3})\mathrm{s}^{-1}$

$= 5.7575 \times 10^{-3}\ \mathrm{s}^{-1}$

Question8

Identify the order of following reaction?

$$2\mathrm{NO}_{2(\mathrm{g})} \rightarrow 2\mathrm{NO}_{(\mathrm{g})} + \mathrm{O}_{2(\mathrm{g})}$$

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Options:

A. 1

B. 1.5

C. 2

D. 3

Answer: C

Solution:

The rate law for the elementary reaction $2\mathrm{NO}_{2(\mathrm{g})} \longrightarrow 2\mathrm{NO}_{(\mathrm{g})} + \mathrm{O}_{2(\mathrm{g})}$ is found to be rate $= \mathrm{k}[\mathrm{NO}_2]^2$.

$\therefore$ The reaction follows second order.

Question9

Rate of the reaction $\mathrm{A} + \mathrm{B} \longrightarrow$ product is $3.6 \times 10^{-2}\ \mathrm{moldm}^{-3}\ \mathrm{s}^{-1}$ and rate law is $\mathrm{r} = \mathrm{k}[\mathrm{A}][\mathrm{B}]^2$. What is rate constant of the reaction if $[\mathrm{A}] = 0.2\mathrm{M}$ and $[\mathrm{B}] = 0.1\mathrm{M}$?

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Options:

A. $10 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

B. $18 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

C. $24 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

D. $4.8 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

Answer: B

Solution:

$$
\begin{aligned}
& \mathrm{r}=\mathrm{k}[\mathrm{~A}][\mathrm{B}]^2 \\
& 3.6 \times 10^{-2}=\mathrm{k} \times(0.2) \times(0.1)^2 \\
& 3.6 \times 10^{-2}=\mathrm{k} \times 2 \times 10^{-3} \\
& \mathrm{k}=\frac{3.6 \times 10^{-2}}{2 \times 10^{-3}}=18 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}
\end{aligned}
$$

Question10

Calculate rate constant of a first order reaction having pre exponential factor $1.6 \times 10^{13} \mathrm{~s}^{-1}$.

$\left(\mathrm{E}_{\mathrm{a}} / 2.303 \mathrm{RT}=21\right)$

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Options:

A. 1.6×10^{-13}

B. 3.2×10^{-13}

C. 3.2×10^{-8}

D. 1.6×10^{-8}

Answer: D

Solution:

$$\begin{aligned}
&\text{Arrehenius equation is } \mathrm{k}=\mathrm{Ae}^{-\mathrm{Ea}/\mathrm{RT}} \\
&\log_{10}\mathrm{k}=\log_{10}\ \mathrm{A}-\frac{\mathrm{E_a}}{2.303\mathrm{RT}} \\
&\log_{10}\mathrm{k}=\log\left(1.6\times10^{13}\right)-21 \\
&\qquad=\log\left(16\times10^{12}\right)-21 \\
&\qquad=\log 16+\log 10^{12}-21 \\
&\qquad=\log 2^4+12-21 \\
&\qquad=4\times0.3010-9=-7.796 \\
&\therefore\quad \mathrm{k}=\ \text{Antilog}\,(-7.796) \\
&\qquad=1.5995\times10^{-8}=1.6\times10^{-8}
\end{aligned}$$

Question11

If Instantaneous rate of reaction is stated as $-\frac{1}{2}\frac{\mathrm{d}[x]}{\mathrm{dt}}=-\frac{\mathrm{d[y]}}{\mathrm{dt}}=\frac{1}{2}\frac{\mathrm{d[z]}}{\mathrm{dt}}$, identify the reaction.

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Options:

A. $x-2y\longrightarrow 2z$

B. $2x+y\longrightarrow 2z$

C. $x+y\longrightarrow z$

D. $2x-2y\longrightarrow z$

Answer: B

Solution:

We are given the relation for the instantaneous rate of reaction:

$$-\frac{1}{2}\frac{d[x]}{dt}=-\frac{d[y]}{dt}=\frac{1}{2}\frac{d[z]}{dt}.$$

Recall:

For a general reaction

$aA + bB \longrightarrow cC,$

the rate expression is

$\text{Rate} = -\frac{1}{a}\frac{d[A]}{dt} = -\frac{1}{b}\frac{d[B]}{dt} = \frac{1}{c}\frac{d[C]}{dt}.$

Step 1: Compare coefficients

From the given expression:

- For species **x**, the rate has factor $1/2 \rightarrow$ stoichiometric coefficient of x is **2**.
- For species **y**, the rate has factor $1 \rightarrow$ stoichiometric coefficient of y is **1**.
- For species **z**, the rate has factor $1/2 \rightarrow$ stoichiometric coefficient of z is **2**.

Step 2: Write the reaction

So the balanced stoichiometric reaction should be:

$2x + y \longrightarrow 2z.$

Step 3: Match option

That corresponds to **Option B**.

✅ **Final Answer: Option B:** $2x + y \rightarrow 2z$

Question12

Which of the following is an example of second order reaction?

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Options:

A.

$2H_2O_{2(g)} \longrightarrow 2H_2P_{(\ell)} + O_{2(g)}$

B.

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

C.

$CH_3CHO_{(g)} \longrightarrow CH_{4(g)} + CO_{(g)}$

D.

$2NO_{(g)} + 2H_{2(g)} \longrightarrow N_{2(g)} + 2H_2O_{(g)}$

Answer: B

Solution:

Step 1: Recall definition

- **Order of a reaction** is defined by the dependence of the rate of reaction on the reactant concentration(s).
- A **second-order reaction** has a rate law of either:
- Rate = $k[A]^2$, **or**
- Rate = $k[A][B]$.

So, it involves **two molecules colliding** (bimolecular reaction).

Step 2: Analyze each option

Option A:

$2H_2O_2 \rightarrow 2H_2O + O_2$

This is the decomposition of hydrogen peroxide. In reality, in the gas phase/decomposition (or catalyzed), it is **first-order** (unimolecular decomposition kinetically). Not second order.

Option B:

$H_2 + I_2 \rightarrow 2HI$

Here, the molecularity is **bimolecular** (collision between H_2 and I_2).

Rate law: $r = k[H_2][I_2]$.

This is a **second-order reaction** (order = 1+1 = 2). ✅

Option C:

$CH_3CHO \rightarrow CH_4 + CO$

This is the decomposition of acetaldehyde, which is a **first-order reaction**. ❌

Option D:

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

The experimentally found rate law is $r = k[NO]^2[H_2]$.

That means order = 3 (two in NO, one in H_2). Not second order. ❌

✅ **Correct Answer:**

Option B: $H_2 + I_2 \longrightarrow 2HI$

This is a classic **second-order reaction**.

Question13

What is the time required for 99% completion of a first order reaction if rate constant is $23.03\ \mathrm{min}^{-1}$?

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Options:

A. 0.2 minute

B. 0.4 minute

C. 6.2 minute

D. 8.1 minute

Answer: A

Solution:

For a first order reaction,

$$t = \frac{2.303}{k} \log_{10} \frac{[A]_0}{[A]_t}$$

99% of reaction is complete, so if $[A]_0 = 100$,

$$\begin{aligned} & [A]_t = 100 - 99 = 1 \\ \therefore \quad & t = \frac{2.303}{k} \log_{10} \frac{(100)}{(1)} \\ & = \frac{2.303}{k} \times 2 = \frac{2.303}{23.03} \times 2 = 0.2 \text{ minutes} \end{aligned}$$

Question14

For a reaction A → Product, rate constant is 6.93×10^{-3} hour $^{-1}$. What is order of reaction?

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Options:

A. Zero

B. 1

C. 1.5

D. 2

Answer: B

Solution:

For first order reaction, the unit of rate constant (k) is (time) $^{-1}$.

Question15

If $r=k[A]^2[B]$ is rate law equation for reaction $\mathrm{A}+\mathrm{B} \rightarrow \mathrm{C}$, at $[\mathrm{A}]=1 \mathrm{M}$ and $[\mathrm{B}]=0.2 \mathrm{M}$, Calculate rate of reaction if rate constant is $6.25 \mathrm{M}^{-2} \mathrm{~s}^{-1}$.

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Options:

A. $1.25 \mathrm{~moldm}^{-3} \mathrm{~s}^{-1}$

B. $3.40 \mathrm{~moldm}^{-3} \mathrm{~s}^{-1}$

C. $5.88 \mathrm{~moldm}^{-3} \mathrm{~s}^{-1}$

D. $8.58 \mathrm{~moldm}^{-3} \mathrm{~s}^{-1}$

Answer: A

Solution:

Given data:

- $k = 6.25\,\mathrm{M}^{-2}\,\mathrm{s}^{-1}$
- $[A] = 1.0\,\mathrm{M}$
- $[B] = 0.20\,\mathrm{M}$

Step 1: Substitute values

$r = (6.25)(1.0)^2(0.20)$

$r = 6.25 \times 0.20$

$r = 1.25$

Step 2: Units

Units: $(\mathrm{M}^{-2}\,\mathrm{s}^{-1})(\mathrm{M}^2)(\mathrm{M}) = \mathrm{M\,s}^{-1} = \mathrm{mol\,dm}^{-3}\,\mathrm{s}^{-1}$

✅ **Final Answer:**

$\boxed{1.25\,\mathrm{mol\,dm}^{-3}\,\mathrm{s}^{-1}}$

Correct option: **A**

Question16

The rate constant is doubled when temperature increases from $27^\circ\mathrm{C}$ to $37^\circ\mathrm{C}$. What is activation energy in kJ ?

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Options:

A. 21.32

B. 34.12

C. 53.60

D. 43.54

Answer: C

Solution:

$$
\begin{aligned}
\log_{10} \frac{k_2}{k_1} &= \frac{E_s}{2.303R}\left[\frac{T_2 - T_1}{T_1 \times T_2}\right] \\
\log_{10} 2 &= \frac{E_a}{2.303 \times 8.314}\left[\frac{310 - 300}{300 \times 310}\right]
\end{aligned}
$$

$$
\begin{aligned}
\therefore E_a &= \frac{\log_{10} 2 \times 2.303 \times 8.314 \times 300 \times 310}{10} \\
&= \frac{0.3010 \times 2.303 \times 8.314 \times 300 \times 310}{10} \\
&= 53,598 \text{ J} = 53.6 \text{ kJ}
\end{aligned}
$$

Question17

Rate law for the reaction,

$$\mathrm{NO}_{2(\mathrm{~g})} + \mathrm{CO}_{(\mathrm{g})} \longrightarrow \mathrm{NO}_{(\mathrm{g})} + \mathrm{CO}_{2(\mathrm{~g})}$$

is as $\mathrm{R} = \mathrm{k}\left[\mathrm{NO}_2\right]^2$. What is the order of reaction w.r.t. CO ?

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Options:

A. Zero

B. 1

C. 1.5

D. 2

Answer: A

Solution:

We are given the reaction:

$$\mathrm{NO}_{2(g)} + \mathrm{CO}_{(g)} \longrightarrow \mathrm{NO}_{(g)} + \mathrm{CO}_{2(g)}$$

and the experimental rate law is:

$R = k[\mathrm{NO_2}]^2$

Step 1: Compare the rate law with the general form

For a general reaction $aA + bB \rightarrow \ldots$, the rate law can be written as:

$R = k[A]^m[B]^n$

Here, m is order with respect to A, n is order with respect to B.

Step 2: Match given rate law

We are told:

$R = k[\mathrm{NO_2}]^2$

So:

- Order with respect to NO_2 = 2
- There is **no term** involving $[\mathrm{CO}]$ in the rate law.

Step 3: Interpret

The absence of $[\mathrm{CO}]$ means the reaction rate does not depend on CO concentration. That means the order with respect to CO = 0.

✅ Final Answer:

Order of reaction with respect to CO = **Zero**

Correct Option: A

Question18

A first order reaction is 50% completed in 16 minutes. Find the percentage of reactant reacting in 32 minutes.

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Options:

A. 25%

B. 40%

C. 50%

D. 75%

Answer: D

Solution:

For a first order reaction,

$$\underset{[A]_o}{100\ \%} \xrightarrow{t_{1/2}} \underset{[A]_t}{50\ \%} \xrightarrow{t_{1/2}} \underset{[A]_t}{25\ \%}$$

To complete 50% of the reaction, time required is 16 minutes. i.e., $t_{1/2} = 16$ minutes.

$\therefore$ % of reactant reacting in 32 minutes $\left(2 \times t_{1/2}\right)$ will be 75%.

Question19

Find the rate of following reaction.

$$2\ N_2O_{5(g)} \longrightarrow 4NO_{2(g)} + O_{2(g)}$$

if concentration of NO_2 increases to 5.2×10^{-3}M in 100 seconds.

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Options:

A. $1.3 \times 10^{-3}\ moldm^{-3}\ s^{-1}$

B. $1.3 \times 10^{-5}\ moldm^{-3}\ s^{-1}$

C. $4.0 \times 10^{-5}\ moldm^{-3}\ s^{-1}$

D. $5.2 \times 10^{-5}\ moldm^{-3}\ s^{-1}$

Answer: B

Solution:

For the reaction, $2\,\mathrm{N_2O_{5(g)}} \longrightarrow 4\mathrm{NO_{2(g)}} + \mathrm{O_{2(g)}}$

$$\begin{aligned}\text{Rate } &= \frac{1}{4}\frac{\mathrm{d}[\mathrm{NO_2}]}{\mathrm{dt}} = \frac{1 \times 5.2 \times 10^{-3}}{4 \times 100}\\ &= 1.3 \times 10^{-5}\ \mathrm{moldm^{-3}\,s^{-1}}\end{aligned}$$

Question20

The rate constant of a first order reaction is $1.15 \times 10^{-3}\ \mathrm{s}^{-1}$. How long will 5 g of reactant take to reduce to 3 g ?

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Options:

A. 314 seconds

B. 240 seconds

C. 404 seconds

D. 444 seconds

Answer: D

Solution:

We are given:

- First-order reaction
- Rate constant: $k = 1.15 \times 10^{-3}\ \mathrm{s}^{-1}$
- Initial amount: $a = 5$ g
- Left after time t: $x = 3$ g

Step 1: First-order rate law

For first order kinetics:

$\ln \frac{[A]_0}{[A]} = kt$

Here,

$[A]_0 = 5, [A] = 3.$

$\ln \frac{5}{3} = kt$

Step 2: Solve for t

$t = \frac{1}{k} \ln \left(\frac{5}{3}\right)$

First evaluate ratio:

$\frac{5}{3} = 1.666\ldots$

$\ln(1.666) \approx 0.5108$

Now divide:

$t = \frac{0.5108}{1.15 \times 10^{-3}}$

$t \approx 444.2 \mathrm{~s}$

✅ **Final Answer:**

$\boxed{444 \mathrm{~s}}$

Correct option: D (444 seconds)

Question21

Which of the following is an example of first order reaction?

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Options:

A.

$\mathrm{CH_3CHO_{(g)}} \longrightarrow \mathrm{CH_{4(g)}} + \mathrm{CO_{(g)}}$

B.

$\mathrm{2NO_{2(g)}} + \mathrm{F_{2(g)}} \longrightarrow \mathrm{2NO_2F_{(g)}}$

C.

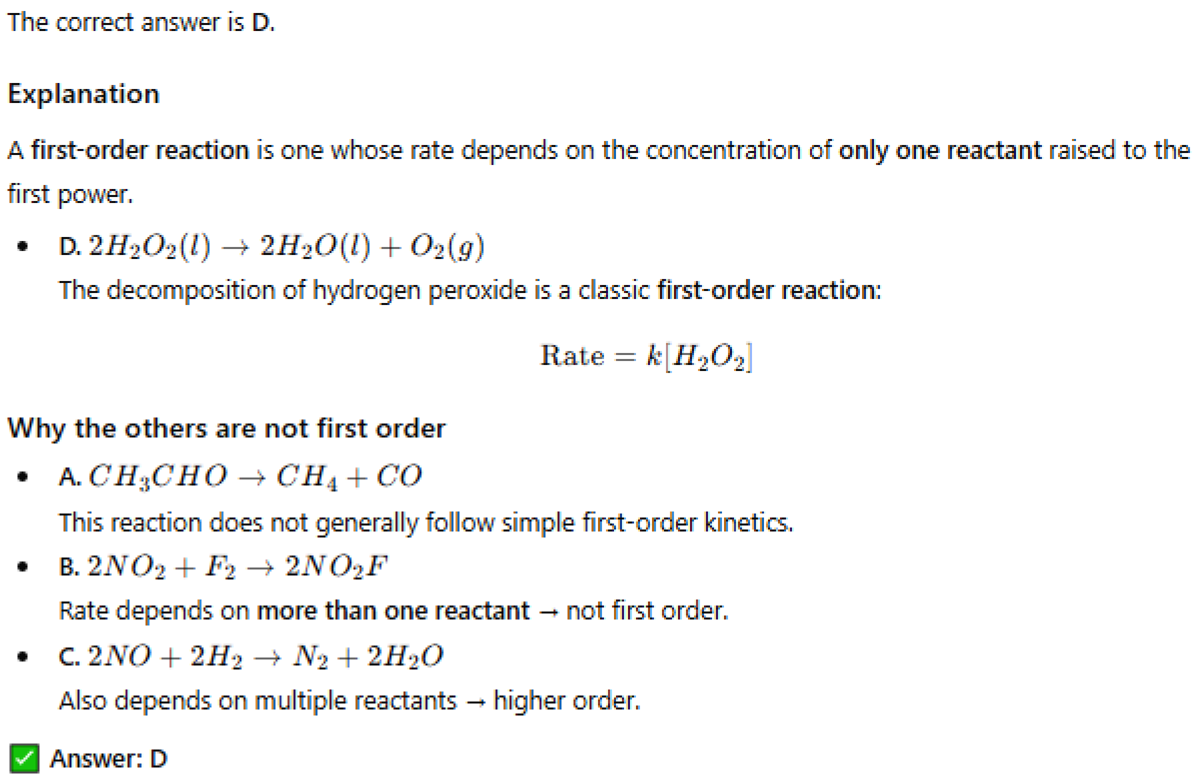
$\mathrm{2NO_{(g)}} + \mathrm{2H_{2(g)}} \longrightarrow \mathrm{N_{2(g)}} + \mathrm{2H_2O_{(g)}}$

D.

$2H_2O_{2(l)} \longrightarrow 2H_2O_{(l)} + O_{2(g)}$

Answer: D

Solution:



The correct answer is **D**.

Explanation

A **first-order reaction** is one whose rate depends on the concentration of **only one reactant** raised to the first power.

- D. $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$
  The decomposition of hydrogen peroxide is a classic **first-order reaction**:

$$\text{Rate} = k[H_2O_2]$$

Why the others are not first order

- A. $CH_3CHO \rightarrow CH_4 + CO$
  This reaction does not generally follow simple first-order kinetics.
- B. $2NO_2 + F_2 \rightarrow 2NO_2F$
  Rate depends on **more than one reactant** → not first order.
- C. $2NO + 2H_2 \rightarrow N_2 + 2H_2O$
  Also depends on multiple reactants → higher order.

✅ **Answer: D**

Question22

Nitric oxide reacts with H_2 according to reaction. $2NO_{(g)} + 2H_{2(g)} \rightarrow N_{2(g)} + 2H_2O_{(g)}$.

Identify the correct relationship for consumption of reactant and formation of product.

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Options:

A.

$\frac{1}{2}\frac{d[NO]}{dt} = -\frac{d[H_2]}{dt}$

B.

$\frac{d[N_2]}{dt} = -\frac{1}{2}\frac{d[H_2]}{dt}$

C.

$\frac{d[H_2O]}{dt} = \frac{d[N_2]}{dt}$

D.

$\frac{d[H_2O]}{dt} = \frac{1}{2}\frac{d[N_2]}{dt}$

Answer: B

Solution:

For the given reaction,

$2NO_{(g)} + 2H_{2(g)} \longrightarrow N_{2(g)} + 2H_2O_{(g)}$

Rate of reaction

$= -\frac{1}{2}\frac{d[NO]}{dt} = -\frac{1}{2}\frac{d[H_2]}{dt} = \frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[H_2O]}{dt}$

Question23

For a reaction,

$A + B \longrightarrow$ product, it is found that rate law is $r = k[A]^{1.5}[B]^{2.5}$. What is the order of reaction?

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Options:

A. 1.5

B. 2.5

C. +1

D. 4

Answer: D

Solution:

For the reaction, $\mathrm{A}+\mathrm{B} \longrightarrow$ product

Rate of the reaction $=\mathrm{r}=\mathrm{k}[\mathrm{A}]^{1.5}[\mathrm{~B}]^{2.5}$

Overall order of reaction $=1.5+2.5=4$

Question24

The half life values for two different first order reaction A and B are 75 minute and 2.5 hour respectively. What is the $\frac{r_B}{r_A}$ ratio of rate constants?

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Options:

A. 2.0

B. 0.5

C. 14.2

D. 22.0

Answer: B

Solution:

Step 1. Recall relation between half life and rate constant for first order reaction:

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

i.e.,

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

Step 2. Calculate rate constants:

For A:

$t_{1/2,A} = 75\,\text{min} = 1.25\,\text{h}$

So,

$k_A = \frac{0.693}{1.25} = 0.554\,\text{h}^{-1}$

For B:

$t_{1/2,B} = 2.5\,\text{h}$

So,

$k_B = 0.693/2.5 = 0.277\,\text{h}^{-1}$

Step 3. Ratio:

$\frac{k_B}{k_A} = \frac{0.277}{0.554} \approx 0.5$

✅ **Answer: Option B (0.5)**

Question25

Nitric oxide reacts with H_2 according to reaction. $2NO_{(g)} + 2H_{2(g)} \rightarrow N_{2(g)} + 2H_2O_{(g)}$, identify correct relation for disappearance of reactant and appearance of product.

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Options:

A.

$\frac{1}{2}\frac{d[NO]}{dt} = -\frac{d[H_2]}{dt}$

B.

$\frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[H_2O]}{dt}$

C.

$-\frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[H_2O]}{dt}$

D.

$\frac{d[H_2]}{dt} = -\frac{1}{2}\frac{d[N_2]}{dt}$

Answer: B

Solution:

$$2NO_{(g)} + 2H_{2(g)} \longrightarrow N_2 + 2H_2O_{(g)}$$

$$\text{rate} = -\frac{1}{2}\frac{d[NO]}{dt} = -\frac{1}{2}\frac{d[H_2]}{dt} = \frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[H_2O]}{dt}$$

Question26

Rate law for the reaction $aA + bB \rightarrow cC + dD$ is $r = k[A][B]$, the rate of reaction doubles if

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Options:

A. Concentration of both A and B are doubled.

B. Concentration of A is doubled and concentration of B is kept constant.

C. Concentration of B is doubled and concentration of A is halved.

D. Concentration of A is kept constant and concentration of B is halved.

Answer: B

Solution:

We are given:

Reaction:

$aA + bB \rightarrow cC + dD$

Rate law:

$r = k[A][B]$

Step 1: Initial rate

$r = k[A][B]$

Step 2: Analyze each option

Option A: If $[A]$ and $[B]$ are both doubled:

$r' = k(2[A])(2[B]) = 4k[A][B] = 4r$

$\rightarrow$ rate becomes **4 times**, not double.

Option B: If $[A]$ is doubled and $[B]$ constant:

$r' = k(2[A])[B] = 2k[A][B] = 2r$

$\rightarrow$ rate becomes **double**.

Option C: If $[B]$ is doubled and $[A]$ halved:

$r' = k\left(\frac{1}{2}[A]\right)(2[B]) = k[A][B] = r$

$\rightarrow$ rate remains **the same**.

Option D: If $[A]$ constant and $[B]$ halved:

$r' = k[A]\left(\frac{1}{2}[B]\right) = \frac{1}{2}r$

$\rightarrow$ rate becomes **half**, not double.

✅ Correct Answer:

Option B: Concentration of A is doubled and concentration of B is kept constant.

Question27

Half life of a first order reaction is 20 minutes. The time taken to reduce the initial concentration of reactant to $(1/10)^{\text{th}}$ is __________

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Options:

A. 46.60 min

B. 66.46 min

C. 79.68 min

D. 88.00 min

Answer: B

Solution:

Step 1: Relation of half-life and rate constant

For first-order:

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

$k = \frac{0.693}{20} = 0.03465\,\text{min}^{-1}$

Step 2: Integrated rate law

For first-order:

$t = \frac{2.303}{k} \log\left(\frac{[A]_0}{[A]}\right)$

Here, $\frac{[A]_0}{[A]} = \frac{[A]_0}{[A]_0/10} = 10$.

Step 3: Substitute values

$t = \frac{2.303}{0.03465} \log(10)$

Since $\log(10) = 1$:

$t = \frac{2.303}{0.03465}$

$t \approx 66.46\,\text{min}$

✅ Final Answer:

Option B: 66.46 min

Question28

Rate law for the reaction $aA + bB \rightarrow cC + dD$ is $\mathrm{r} = \mathrm{k}[\mathrm{A}][\mathrm{B}]$. Which from following conditions does NOT affect the rate of reaction?

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Options:

A. Concentration of A is doubled and concentration of B is kept constant.

B. Concentration of B is doubled and concentration of A is kept constant.

C. Concentration of B is doubled and concentration of A is halved.

D. Concentration of A is kept constant and concentration of B is halved.

Answer: C

Solution:

When concentration of A is halved and the concentration of B is doubled, the rate does not change. The rate remains the same because the net effect is neutral as order with respect to both A and B is 1 . This can be represented as,

$$\mathrm{r}=\mathrm{k}[\mathrm{~A}][\mathrm{B}]$$

$$\text { Rate }=\mathrm{k} \times \frac{[\mathrm{A}]}{2} \times[2 \mathrm{~B}]=\mathrm{k}[\mathrm{~A}][\mathrm{B}]$$

Question29

Nitric oxide reacts with H_2 according to reaction, $2 \mathrm{NO}_{(\mathrm{g})}+2 \mathrm{H}_{2(\mathrm{~g})} \longrightarrow \mathrm{N}_{2(\mathrm{~g})}+2 \mathrm{H}_2 \mathrm{O}_{(\mathrm{g})}$, identify the correct relationship among the following.

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Options:

A.

$-\frac{1}{2} \frac{\mathrm{d}[\mathrm{NO}]}{\mathrm{dt}}=\frac{\mathrm{d}\left[\mathrm{H}_2 \mathrm{O}\right]}{\mathrm{dt}}$

B.

$-\frac{\mathrm{d}[\mathrm{NO}]}{\mathrm{dt}}=\frac{\mathrm{d}\left[\mathrm{H}_2 \mathrm{O}\right]}{\mathrm{dt}}$

C.

$-\frac{3}{2} \frac{\mathrm{d}[\mathrm{NO}]}{\mathrm{dt}}=\frac{\mathrm{d}\left[\mathrm{H}_2 \mathrm{O}\right]}{\mathrm{dt}}$

D.

$2\frac{\mathrm{d[H_2]}}{\mathrm{dt}} = \frac{\mathrm{d[N_2]}}{\mathrm{dt}}$

Answer: B

Solution:

Rate can be expressed as,

$$= -\frac{1}{2}\frac{\mathrm{d[NO]}}{\mathrm{dt}} = -\frac{1}{2}\frac{\mathrm{d[H_2]}}{\mathrm{dt}} = \frac{\mathrm{d[\ N_2]}}{\mathrm{dt}} = \frac{1}{2}\frac{\mathrm{d[H_2O]}}{\mathrm{dt}}$$

$\therefore$ Correct relationship amongst given options will be $-\frac{\mathrm{d[NO]}}{\mathrm{dt}} = \frac{\mathrm{d[H_2O]}}{\mathrm{dt}}$

Question30

For a reaction, $\mathrm{A} \longrightarrow \mathrm{B}$, rate equation is $\mathrm{r} = \mathrm{k[A]^{\circ}}$. If initial concentration of reactant is 'a'mol dm $^{-3}$ find half life time of reaction.

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Options:

A. $\frac{\mathrm{a}}{4\mathrm{k}}$

B. $\frac{\mathrm{a}}{\mathrm{k}}$

C. $\frac{5\mathrm{a}}{2\mathrm{k}}$

D. $\frac{\mathrm{a}}{2\mathrm{k}}$

Answer: D

Solution:

We are tasked with solving:

Reaction:

$\mathrm{A} \longrightarrow \mathrm{B}$

Rate law:

$r = k[A]^0$

Step 1: Identify order of reaction

Since the exponent is zero, this is a **zero-order reaction**.

So,

$r = -\frac{d[A]}{dt} = k$

Step 2: Integrated rate law for zero-order reaction

$[A]_t = [A]_0 - kt$

Here,

- Initial concentration: $[A]_0 = a$
- At time t, concentration is $[A]_t$.

Step 3: Condition for half-life

Half-life ($t_{1/2}$) is the time when:

$[A]_{t_{1/2}} = \frac{a}{2}$

Using integrated form:

$\frac{a}{2} = a - kt_{1/2}$

Step 4: Solve for $t_{1/2}$

$kt_{1/2} = a - \frac{a}{2} = \frac{a}{2}$

$t_{1/2} = \frac{a}{2k}$

✅ Final Answer:

$t_{1/2} = \frac{a}{2k}$

Correct option: D

Question31

What is the order and molecularity respectively for the elementary reaction given below?

$$\mathrm{O_{3(g)} + O_{(g)} \longrightarrow 2O_{2(g)}}\ \text{if}\ \mathrm{r = k\,[O_3][O]}$$

Options:

A. 2^{nd} and 1

B. 2^{nd} and 2

C. 1^{st} and 2

D. 1^{st} and 1

Answer: B

Solution:

Given reaction:

$O_{3(g)} + O_{(g)} \longrightarrow 2O_{2(g)}$

Given rate law:

$r = k[O_3][O]$

Step 1: Order of the reaction

Order is the sum of the powers of concentration terms in the rate law.

Here, rate law is

$r = k[O_3]^1[O]^1$

So, order $= 1 + 1 = 2$.

Step 2: Molecularity of the reaction

For an elementary reaction, molecularity is the total number of reactant molecules (or atoms/ions) taking part in one elementary reaction step.

Here,

$O_3 + O \Rightarrow 2$ reactant species

So, molecularity $= 2$.

Final Answer:

- Order $= 2$
- Molecularity $= 2$

So, the correct option is:

Option B: 2^{nd} and 2

Question32

Calculate the time in seconds required to reduce the concentration of reactant to half of initial concentration for first order reaction if rate constant is $1.386 \times 10^{-3}\ \mathrm{s}^{-1}$.

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Options:

A. 140 s

B. 290 s

C. 390 s

D. 500 s

Answer: D

Solution:

For a first order reaction, the half-life ($t_{1/2}$), which is the time required for the concentration of the reactant to become half of its initial value, is given by:

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

Given:

- Rate constant, $k = 1.386 \times 10^{-3}\ \mathrm{s}^{-1}$

Substitute the value of k into the formula:

$t_{1/2} = \frac{0.693}{1.386 \times 10^{-3}}$

$t_{1/2} = \frac{0.693}{0.001386}$

$t_{1/2} = 500\ \mathrm{s}$

Correct answer: Option D (500 s)

Question33

If concentration of reactant ' A ' is increased by 10 times, rate of reaction increases 100 times. What is order of reaction if rate law is $r=k[A]^r$?

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Options:

A. 1

B. 1.5

C. 2

D. 3

Answer: C

Solution:

Given rate law:

$r=k[A]^n$

Let the initial concentration of A be $[A]$ and initial rate be r_1:

$r_1=k[A]^n$

When $[A]$ is increased 10 times, new concentration = $10[A]$, new rate = r_2:

$r_2=k(10[A])^n=k\times 10^n\times[A]^n$

According to question:

$\frac{r_2}{r_1}=100$

Substitute the values:

$\frac{k\times 10^n\times[A]^n}{k[A]^n}=100$

$10^n=100$

We know $100=10^2$,

Thus,

$n=2$

Order of reaction $=2$

Correct Option: C (2)

Question34

What is order and molecularity of following reaction?

$$\mathrm{NO}_{2(\mathrm{~g})}+\mathrm{NO}_{2(\mathrm{~g})} \longrightarrow 2 \mathrm{NO}_{(\mathrm{g})}+\mathrm{O}_{2(\mathrm{~g})}$$

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Options:

A. order 1 , molecularity 2

B. order 2, molecularity 2

C. order 2 , molecularity 3

D. order zero, molecularity 2

Answer: B

Solution:

Given reaction:

$\mathrm{NO}_2(g)+\mathrm{NO}_2(g) \longrightarrow 2\ \mathrm{NO}(g)+\mathrm{O}_2(g)$

Step 1: Molecularity

Molecularity is the number of reactant molecules taking part in an elementary step.

Here, **2 molecules of** NO_2 are reacting in one step, so:

- **Molecularity = 2** (Bimolecular reaction)

Step 2: Order of Reaction

Order is the sum of the powers of concentration terms in the **rate law**.

For this reaction (from NCERT examples), the experimentally determined rate law is:

$\text{Rate} = k[\mathrm{NO}_2]^2$

Here, the power of concentration of NO_2 is 2.

- **Order = 2**

Step 3: Correct Option

- **Order = 2**
- **Molecularity = 2**

Final Answer:

Option **B**

order 2, molecularity 2

Question35

Half life of a first order reaction is 3 minute. What is the time required to reduce the concentration of reactant by 90% of its initial concentration?

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Options:

A. 4.12 minute

B. 7.32 minute

C. 9.97 minute

D. 12.05 minute

Answer: C

Solution:

Given:

- Half-life $(t_{1/2})$ of a first order reaction $= 3$ min
- We need the time required to reduce reactant concentration by 90%, i.e., to 10% of initial concentration.

Let's solve step-by-step:

Step 1: First Order Kinetics Formula

For a first order reaction:

$$k = \frac{2.303}{t} \log\left(\frac{[R]_0}{[R]}\right)$$

where:

- $[R]_0$ = Initial concentration
- $[R]$ = Final concentration after time t

Step 2: Use Half-Life to find Rate Constant (k)

First order half-life:

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

Given $t_{1/2} = 3$ min,

$$3 = \frac{0.693}{k} \implies k = \frac{0.693}{3} = 0.231\,\text{min}^{-1}$$

Step 3: Calculate Time for 90% Reduction

After 90% reaction, $[R] = 10\%$ of $[R]_0$

$$[R] = \frac{[R]_0}{10}$$

Plug into the equation:

$$k = \frac{2.303}{t} \log\left(\frac{[R]_0}{[R]}\right)$$

$$0.231 = \frac{2.303}{t} \log\left(\frac{[R]_0}{[R]_0/10}\right)$$

$$0.231 = \frac{2.303}{t} \log(10)$$

$$\log(10) = 1$$

So,

$$0.231 = \frac{2.303}{t} \times 1$$

$$t = \frac{2.303}{0.231}$$

$$t \approx 9.97\text{ min}$$

Final Answer

The time required is 9.97 minutes.

Option C is correct.

Question36

The rate constant for a first order reaction is $0.58\ \mathrm{s}^{-1}$ at 300 K and $0.026\ \mathrm{s}^{-1}$ at 290 K .

What is the energy of activation?

$$\left(\mathrm{R}=8.314\ \mathrm{J\ K}^{-1}\ \mathrm{mol}^{-1}\right)$$

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Options:

A. 124.48 kJ

B. 224.55 kJ

C. 348.18 kJ

D. 513.21 kJ

Answer: B

Solution:

$$\log_{10}\frac{\mathrm{k}_2}{\mathrm{k}_1}=\frac{\mathrm{E}_\mathrm{a}}{2.303\mathrm{R}}\left(\frac{\mathrm{T}_2-\mathrm{T}_1}{\mathrm{T}_1\ \mathrm{T}_2}\right)$$

Substituting the given values

$$\log_{10}\frac{0.58}{0.026}=\frac{\mathrm{E}_\mathrm{a}}{2.303\times 8.314\ \mathrm{J\ K}^{-1}\ \mathrm{mol}^{-1}}\times\left[\frac{300\ \mathrm{K}-290\ \mathrm{K}}{290\ \mathrm{K}\times 300\ \mathrm{K}}\right]$$

$$\log_{10}22.307=\frac{\mathrm{E}_\mathrm{a}}{2.303\times 8.314\ \mathrm{J\ mol}^{-1}}\times\frac{10}{290\times 300}$$

$$1.3484=\frac{\mathrm{E}_\mathrm{a}}{19.15\ \mathrm{J\ mol}^{-1}}\times 1.15\times 10^{-4}$$

$$\begin{aligned}\mathrm{E}_\mathrm{a}&=\frac{1.3484\times 19.15}{1.15\times 10^{-4}}\ \mathrm{J\ mol}^{-1}\\&=224538\ \mathrm{J\ mol}^{-1}\\&\approx 224.55\ \mathrm{kJ\ mol}^{-1}\end{aligned}$$

Question37

The rate constant for the reaction, $2\ \mathrm{N_2O_{5(\,g)}} \rightarrow 2\ \mathrm{N_2O_{4(\,g)}} + \mathrm{O_{2(\,g)}}$ is $4.98 \times 10^{-4}\ \mathrm{s}^{-1}$. What is the order of reaction?

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Options:

A. 2

B. 1

C. Zero

D. 3

Answer: B

Solution:

Given:

Rate constant for the reaction, $2\ \mathrm{N_2O_5(g)} \rightarrow 2\ \mathrm{N_2O_4(g)} + \mathrm{O_2(g)}$ is $4.98 \times 10^{-4}\ \mathrm{s}^{-1}$.

Step 1: **Recall the unit of rate constant for different order reactions** (NCERT Table):

- For **zero order**, unit is $\mathrm{mol\ L^{-1}\ s^{-1}}$
- For **first order**, unit is $\mathrm{s^{-1}}$
- For **second order**, unit is $\mathrm{L\ mol^{-1}\ s^{-1}}$
- For **third order**, unit is $\mathrm{L^2\ mol^{-2}\ s^{-1}}$

Step 2: **Compare given unit**

The rate constant's unit is $\mathrm{s^{-1}}$.

Step 3: **Matching the unit**

Only a **first order** reaction has rate constant with the unit $\mathrm{s^{-1}}$.

Correct answer:

Option B

1

Question38

Which of the following equations is correct regarding rate of disappearance of reactant and appearance of product for

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

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Options:

A.

$$3\frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[N_2]}{dt}$$

B.

$$\frac{1}{2}\frac{d[N_2]}{dt} = \frac{1}{3}\frac{d[H_2]}{dt}$$

C.

$$2\frac{d[NH_3]}{dt} = 3\frac{d[H_2]}{dt}$$

D.

$$3\frac{d[NH_3]}{dt} = -2\frac{d[H_2]}{dt}$$

Answer: D

Solution:

Given reaction:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

According to NCERT, the rate of reaction is written as:

$$\text{Rate} = -\frac{1}{1}\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$$

Let us check each option:

Option A:

$$3\frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[N_2]}{dt}$$

This is incorrect. There should be different species on either side.

Option B:

$$\frac{1}{2} \frac{d[N_2]}{dt} = \frac{1}{3} \frac{d[H_2]}{dt}$$

This is incorrect. Coefficients and species associated do not match the actual stoichiometry.

Option C:

$$2 \frac{d[NH_3]}{dt} = 3 \frac{d[H_2]}{dt}$$

This is incorrect. According to the rate law,

$$\frac{1}{2} \frac{d[NH_3]}{dt} = -\frac{1}{3} \frac{d[H_2]}{dt}$$

Option D:

$$3 \frac{d[NH_3]}{dt} = -2 \frac{d[H_2]}{dt}$$

Let's check:

If we multiply both sides of the rate equation by 6:

$$\frac{1}{2} \frac{d[NH_3]}{dt} = -\frac{1}{3} \frac{d[H_2]}{dt}$$

Multiply both sides by 6:

$$6 \times \frac{1}{2} \frac{d[NH_3]}{dt} = 6 \times \left(-\frac{1}{3} \frac{d[H_2]}{dt}\right)$$

$$3 \frac{d[NH_3]}{dt} = -2 \frac{d[H_2]}{dt}$$

This matches **Option D**.

Correct Answer:

Option D

$$3 \frac{d[NH_3]}{dt} = -2 \frac{d[H_2]}{dt}$$

Question39

Which from following is a correct representation of reaction rate for reaction stated below?

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

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Options:

A.

$$\frac{d[N_2]}{dt} = -3\frac{d[H_2]}{dt} = 2\frac{d[NH_3]}{dt}$$

B.

$$\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$$

C.

$$-\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$$

D. $\frac{d[N_2]}{dt} = \frac{1}{3}\frac{d[H_2]}{dt} = -\frac{1}{2}\frac{d[NH_3]}{dt}$

Answer: C

Solution:

For the reaction

$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$

The general formula for rate of reaction is:

If a reaction is of the form

$aA + bB \longrightarrow cC + dD$

then the rate of reaction is written as

$$\text{Rate} = -\frac{1}{a}\frac{d[A]}{dt} = -\frac{1}{b}\frac{d[B]}{dt} = \frac{1}{c}\frac{d[C]}{dt} = \frac{1}{d}\frac{d[D]}{dt}$$

Step-by-step for given reaction:

1. Compare coefficients:

- For N_2, coefficient $a = 1$
- For H_2, coefficient $b = 3$
- For NH_3, coefficient $c = 2$

1. Write expression for rate:

$$\text{Rate} = -\frac{1}{1}\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$$

3. Observe that the negative sign comes with reactants (since their concentration decreases with time), and positive sign with product (since concentration increases).

Correct representation:

$-\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$

✅ So the correct answer is **Option C**.

Question40

Half life of a first order reaction is 900 minute at 400 K , find its half life at 300 K ?

$$\left(\frac{\mathrm{E_a}}{2.303\mathrm{R}} = 1.3056 \times 10^3\right)$$

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Options:

A. 5512.5 minute

B. 11025.0 minute

C. 8314.3 minute

D. 2303.1 minute

Answer: B

Solution:

Given data:

- Half-life at 400 K, $t_{1/2,1} = 900\,\text{min}$
- $E_a/(2.303R) = 1.3056 \times 10^3$
- Temperature 1: $T_1 = 400\,\text{K}$
- Temperature 2: $T_2 = 300\,\text{K}$
- We are to find $t_{1/2,2}$ at $300\,\text{K}$

Step 1: Recall the half-life formula for a first-order reaction:

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

Step 2: Use the Arrhenius equation in its logarithmic form:

$$\log\left(\frac{k_2}{k_1}\right) = \frac{E_a}{2.303R}\left(\frac{T_1-T_2}{T_1T_2}\right)$$

Step 3: Relate half-lives to rate constants:

$$\frac{t_{1/2,2}}{t_{1/2,1}} = \frac{k_1}{k_2}$$

Step 4: Substitute values into the Arrhenius expression:

Given,

$$\log\left(\frac{k_2}{k_1}\right) = \frac{E_a}{2.303R}\left(\frac{400-300}{400\times300}\right)$$

$$\log\left(\frac{k_2}{k_1}\right) = 1.3056\times10^3\times\left(\frac{100}{120000}\right)$$

$$\log\left(\frac{k_2}{k_1}\right) = 1.3056\times10^3\times\frac{1}{1200}$$

$$\log\left(\frac{k_2}{k_1}\right) = \frac{1.3056\times10^3}{1200}$$

$$\log\left(\frac{k_2}{k_1}\right) = 1.088$$

Step 5: Write k_1/k_2:

$$\log\left(\frac{k_1}{k_2}\right) = -1.088$$

So,

$$\frac{k_1}{k_2} = 10^{-1.088}$$

Step 6: Relate half-lives:

$$\frac{t_{1/2,2}}{t_{1/2,1}} = \frac{k_1}{k_2} = 10^{-1.088}$$

So,

$$t_{1/2,2} = t_{1/2,1}\times10^{1.088}$$

Step 7: Calculate $10^{1.088}$

First,

$$10^{1.088}\approx12.25$$

Step 8: Calculate $t_{1/2,2}$

$$t_{1/2,2} = 900\times12.25 = 11025\,\text{min}$$

Final Answer:

Option **B** 11025.0 minute

Question41

Rate law for the reaction,

$C_2H_5I_{(g)} \longrightarrow C_2H_{4(\ g)} + HI_{(g)}$ is $r = k\left[C_2H_5I\right]$

What is the order and molecularity of this reaction?

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Options:

A. order and molecularity both are 1

B. order is 1 and molecularity is 2

C. order and molecularity both are 2

D. order is 2 and molecularity is 1

Answer: A

Solution:

Given rate law:

$r = k[C_2H_5I]$

Step 1: Determining Order of the Reaction

- The *order of a reaction* is the sum of powers of concentration terms in the rate law.
- Here, only $[C_2H_5I]$ is present with power 1.

So, **order = 1**.

Step 2: Determining Molecularity of the Reaction

- *Molecularity* is the number of reacting species involved in an elementary reaction (single step).
- The given reaction is:

  $C_2H_5I \rightarrow C_2H_4 + HI$
- Only **one molecule decomposes in one step**.

So, **molecularity = 1**.

Step 3: Choosing the Correct Option

From above,

- order = 1
- molecularity = 1

Correct answer: Option A

Order and molecularity both are 1.

Question42

For the reaction, $\mathrm{NO}_{2(\ g)} + \mathrm{CO}_{(\mathrm{g})} \longrightarrow \mathrm{NO}_{(\mathrm{g})} + \mathrm{CO}_{2(\ g)}$, rate of formation of $\mathrm{NO}_{(\mathrm{g})}$ is $\mathrm{Ymoldmds}^{-1}$. Find the rate of disappearance of $\mathrm{CO}_{(\mathrm{g})}$.

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Options:

A. $\mathrm{Y\ mol\ dm}^{-3}\ \mathrm{s}^{-1}$

B. $\mathrm{2Y\ mol\ dm}^{-3}\ \mathrm{s}^{-1}$

C. $\frac{\mathrm{Y}}{2}\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

D. $\frac{3}{2}\mathrm{Y\ mol\ dm}^{-3}\ \mathrm{s}^{-1}$

Answer: A

Solution:

Given reaction:

$\mathrm{NO}_{2\,(g)} + \mathrm{CO}_{(g)} \longrightarrow \mathrm{NO}_{(g)} + \mathrm{CO}_{2\,(g)}$

According to NCERT, the rate of a reaction in terms of reactants and products is given by:

$$-\frac{d[\mathrm{NO}_2]}{dt} = -\frac{d[\mathrm{CO}]}{dt} = \frac{d[\mathrm{NO}]}{dt} = \frac{d[\mathrm{CO}_2]}{dt}$$

If the rate of formation of NO is Y mol $\mathrm{dm}^{-3}\ \mathrm{s}^{-1}$:

$$\frac{d[\mathrm{NO}]}{dt} = Y$$

So,

$-\frac{d[\mathrm{CO}]}{dt} = Y$

This implies,

Rate of disappearance of CO $= Y$ mol dm^{-3} s^{-1}

Correct option:

$\boxed{\text{Option A: } Y \text{ mol dm}^{-3}\text{ s}^{-1}}$

Question43

In a first order reaction concentration of reactant decreases from 20 milli moldm^{-3} to 8 milli moldm^{-} in 40 minutes, find rate constant of reaction?

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Options:

A. 0.011 minute $^{-1}$

B. 0.023 minute $^{-1}$

C. 0.032 minute $^{-1}$

D. 0.041 minute $^{-1}$

Answer: B

Solution:

Let the initial concentration of reactant, $[R]_0 = 20$ milli mol dm^{-3}

Final concentration after time $t = 40$ min, $[R] = 8$ milli mol dm^{-3}

For a first order reaction, the rate law is:

$k = \frac{1}{t}\ln\left(\frac{[R]_0}{[R]}\right)$

Substitute the values:

1. $[R]_0 = 20$
2. $[R] = 8$
3. $t = 40$ min

$k = \frac{1}{40} \ln\left(\frac{20}{8}\right)$

Calculate $\frac{20}{8}$:

$\frac{20}{8} = 2.5$

So,

$k = \frac{1}{40} \ln(2.5)$

Now, calculate $\ln(2.5)$:

$\ln(2.5) \approx 0.916$

Now calculate k:

$k = \frac{0.916}{40}$

$k = 0.0229 \text{ min}^{-1}$

Rounding off, $k \approx 0.023 \text{ min}^{-1}$

Correct Option:

$\boxed{\text{Option B } 0.023 \text{ minute}^{-1}}$

Question44

Which from following is correct regarding $t_{1/2}$ of reaction if we double the initial concentration of a reactant in first order reaction?

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Options:

A. $t_{1/2}$ will increase by two times

B. $t_{1/2}$ will decrease by four times

C. $t_{1/2}$ remains the same

D. $t_{1/2}$ will decrease by two times

Answer: C

Solution:

For a first order reaction, the half-life $(t_{1/2})$ is given by:

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

Here, k is the rate constant.

Notice that $t_{1/2}$ does **not** depend on the initial concentration of the reactant.

So, even if we double the initial concentration, $t_{1/2}$ remains the **same**.

Correct option:

Option C

$t_{1/2}$ remains the same

Question45

In a first order reaction 60% of the reactant converts into product in 45 minute. Calculate rate constant of the reaction.

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Options:

A. 0.0102 minute $^{-1}$

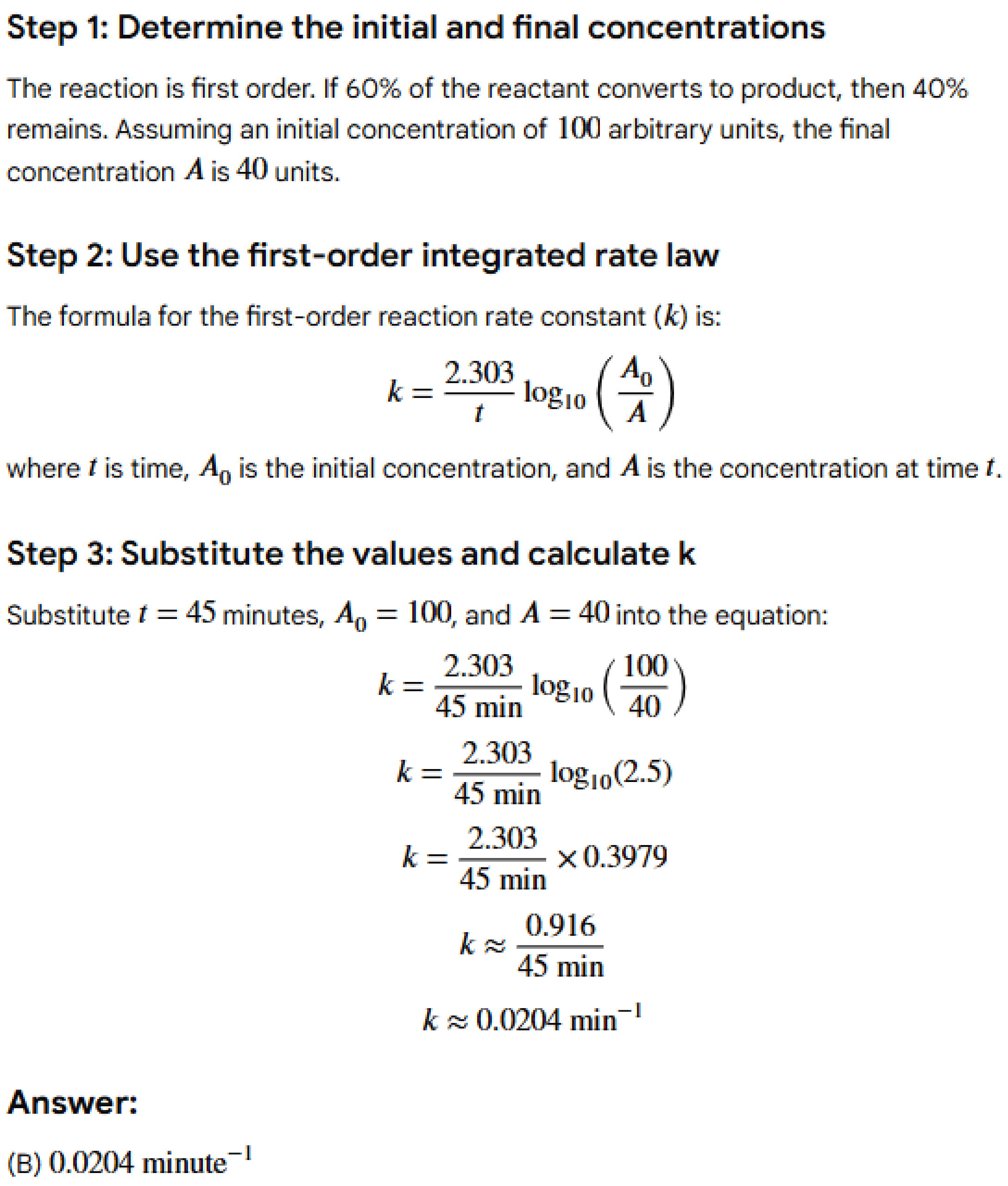
B. 0.0204 minute^{-1}

C. 0.0306 minute $^{-1}$

D. 0.0408 minute $^{-1}$

Answer: B

Solution:



Step 1: Determine the initial and final concentrations

The reaction is first order. If 60% of the reactant converts to product, then 40% remains. Assuming an initial concentration of 100 arbitrary units, the final concentration A is 40 units.

Step 2: Use the first-order integrated rate law

The formula for the first-order reaction rate constant (k) is:

$$k = \frac{2.303}{t} \log_{10}\left(\frac{A_0}{A}\right)$$

where t is time, A_0 is the initial concentration, and A is the concentration at time t.

Step 3: Substitute the values and calculate k

Substitute $t = 45$ minutes, $A_0 = 100$, and $A = 40$ into the equation:

$$k = \frac{2.303}{45 \text{ min}} \log_{10}\left(\frac{100}{40}\right)$$

$$k = \frac{2.303}{45 \text{ min}} \log_{10}(2.5)$$

$$k = \frac{2.303}{45 \text{ min}} \times 0.3979$$

$$k \approx \frac{0.916}{45 \text{ min}}$$

$$k \approx 0.0204 \text{ min}^{-1}$$

Answer:

(B) $0.0204 \text{ minute}^{-1}$

Question46

For the reaction,

$$\mathrm{CH_3Br_{(aq)} + OH^-_{(aq)} \longrightarrow CH_3OH_{(aq)} + Br^-_{(aq)}}$$

rate of consumption of $\mathrm{OH^-_{(aq)}}$ is $\mathrm{x mol dm^{-3}\ s^{-1}}$What is the rate of formation of $\mathrm{Br^-_{(9q)}}$?

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Options:

A. $0.5\mathrm{x}\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

B. $\mathrm{x}\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

C. $2x\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

D. $\frac{3}{2}x\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

Answer: B

Solution:

Rate of consumption of $\left[\mathrm{OH}^{-}\right]=-\frac{\mathrm{d}\left[\mathrm{OH}^{-}\right]}{\mathrm{dt}}$

$=\mathrm{x}\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

Rate of formation of $\left[\mathrm{Br}^{-}\right]$

$=\frac{\mathrm{d}\left[\mathrm{Br}^{-}\right]}{\mathrm{dt}}=-\frac{\mathrm{d}\left[\mathrm{OH}^{-}\right]}{\mathrm{dt}}=\mathrm{x}\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

Question47

For the reaction,

$$\mathrm{H}_{2(g)}+\mathrm{Br}_{2(8)} \longrightarrow 2\mathrm{HBr}_{(\mathrm{g})}, \mathrm{r}=\mathrm{k}\left[\mathrm{H}_2\right]\left[\mathrm{Br}_2\right]^{\frac{1}{2}}$$

What is the molecularity and order of reaction respectively?

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Options:

A. Monomolecular and $\frac{1}{2}$

B. Monomolecular and 1

C. Monomolecular and $\frac{3}{2}$

D. Bimolecular and $\frac{3}{2}$

Answer: D

Solution:

For the reaction $H_{2(g)} + Br_{2(g)} \longrightarrow 2HBr_{(g)}$, the rate law is given by $r = k[H_2][Br_2]^{\frac{1}{2}}$.

Molecularity: This refers to the number of molecules involved in an elementary step of the reaction. In this given reaction equation, one molecule of H_2 and one molecule of Br_2 are involved in the reaction. Therefore, the reaction is **bimolecular**.

Order of Reaction: This is determined by summing the exponents in the rate law:

The exponent of $[H_2]$ is 1.

The exponent of $[Br_2]$ is $\frac{1}{2}$.

Therefore, the order of the reaction is the sum of these exponents:

$1 + \frac{1}{2} = \frac{3}{2}$.

Hence, the molecularity is bimolecular and the order of reaction is $\frac{3}{2}$. Therefore, the correct answer is **Option D: Bimolecular and $\frac{3}{2}$.**

Question48

Which of the following is true for a reaction as per coliision theory?

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Options:

A. Every collision between reactants leads to chemical reaction.

B. It may be expected that rate of reaction is equal to rate of collision.

C. For gas phase reactions, the number of collisions is far less compared to observed rate.

D. The colliding molecules do not need proper orientation.

Answer: B

Solution:

Not every collision between reactants leads to a chemical reaction.

For gas phase reactions, the number of collisions is more as compared to the observed rate.

The colliding molecules need proper orientation.

Question49

Rate of reaction, $\mathrm{A}+\mathrm{B} \rightarrow$ product, is $7.2 \times 10^{-2} \mathrm{moldm}^{-3} \mathrm{~s}^{-1}$ at $[\mathrm{A}]=0.4 \mathrm{~moldm}^{-3}$ and $[B]=0.1 \mathrm{~moldm}^{-3}$. The reaction is first order in A and second order in B. Calculate rate constant.

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Options:

A. $14 \mathrm{~mol}^{-2} \mathrm{dm}^{6} \mathrm{~s}^{-1}$

B. $12 \mathrm{~mol}^{-2} \mathrm{dm}^{6} \mathrm{~s}^{-1}$

C. $18 \mathrm{~mol}^{-2} \mathrm{dm}^{6} \mathrm{~s}^{-1}$

D. $20 \mathrm{~mol}^{-2} \mathrm{dm}^{6} \mathrm{~s}^{-1}$

Answer: C

Solution:

$\text{Rate} = k[A][B]^2$

$\mathrm{k}=\frac{\text{Rate}}{[\mathrm{A}][\mathrm{B}]^{2}}=\frac{0.072 \mathrm{~moldm}^{-3} \mathrm{~s}^{-1}}{0.4 \mathrm{~moldm}^{-3} \times\left(0.1 \mathrm{~moldm}^{-3}\right)^{2}}$

$=18 \mathrm{~mol}^{-2} \mathrm{dm}^{6} \mathrm{~s}^{-1}$

Question50

A zero order reaction has half life time of 0.2 minute. If initial concentration of reactant is 0.2 moldm^{-3}. Find rate constant.

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Options:

A. $0.2 \text{ mol}^{-1} \text{dm}^3 \text{ min}^{-1}$

B. $0.5 \text{ mol}^{-1} \text{dm}^3 \text{ min}^{-1}$

C. $1.4 \text{ mol}^{-1} \text{dm}^3 \text{ min}^{-1}$

D. $6.0 \text{ mol}^{-1} \text{dm}^3 \text{ min}^{-1}$

Answer: B

Solution:

For zero order reaction, $t_{1/2} = \frac{[A]_0}{2k}$

$$0.2 \text{ min} = \frac{0.2 \text{ moldm}^{-3}}{2\text{k}}$$

$$\therefore \quad \text{k} = 0.5 \text{ mol dm}^{-3} \text{ min}^{-1}$$

Question51

The reaction given below $2NH_{3(g)} \xrightarrow{Pt} N_{2(g)} + 3H_{2(g)}$ has rate of reaction $2.5 \times 10^{-6} \text{ moldm}^{-3}\text{sec}^{-1}$ formation of $H_{2(g)}$?

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Options:

A. $2.5 \times 10^{-6} \text{ mol dm}^{-3}\text{sec}^{-1}$

B. 3.75×10^{-6} mol dm^{-3}sec^{-1}

C. 5.0×10^{-6} mol dm^{-3}sec^{-1}

D. 7.5×10^{-6} mol dm^{-3}sec^{-1}

Answer: D

Solution:

For the reaction, $2NH_{3(g)} \xrightarrow{Pt} N_{2(g)} + 3H_{2(g)}$

Rate of reaction $= -\frac{1}{2}\frac{d[NH_3]}{dt} = \frac{d[N_2]}{dt} = \frac{1}{3}\frac{d[H_2]}{dt}$

$\therefore \quad 2.5 \times 10^{-6} = \frac{1}{3}\frac{d[H_2]}{dt}$

$\therefore \quad \frac{d[H_2]}{dt} = 3 \times 2.5 \times 10^{-6} = 7.5 \times 10^{-6}$ moldm^{-3}sec^{-1}

Question52

For reaction $A + B \rightarrow$ product, rate of reaction is 3.6×10^{-2}sec^{-1}. When $[A] = 0.2$moldm^{-3} and $[B] = 0.1$moldm^{-3}, calculate rate constant of reaction if reaction is first order in A and second order is B ?

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Options:

A. 3.6 mol^{-2}dm^{6}sec^{-1}

B. 1.8 mol^{-2}dm^{-6}sec^{-1}

C. 18 mol^{-2}dm^{6}sec^{-1}

D. 36 mol^{-2}dm^{-6}sec^{-1}

Answer: C

Solution:

For the reaction

$\mathrm{A}+\mathrm{B} \rightarrow$ product,

the rate law can be expressed using the orders of reaction:

Rate $= k[\mathrm{A}]^1[\mathrm{B}]^2$,

where k is the rate constant, [A] is the concentration of A, and [B] is the concentration of B. Given the reaction is first-order in A and second-order in B, the expression becomes:

$3.6 \times 10^{-2} = k(0.2)^1(0.1)^2$.

Substitute the concentrations into the equation:

$3.6 \times 10^{-2} = k \times 0.2 \times 0.01$.

Combine and simplify:

$3.6 \times 10^{-2} = k \times 0.002$.

Solve for k by dividing both sides by 0.002:

$k = \frac{3.6 \times 10^{-2}}{0.002}$.

Calculate the value:

$k = 18 \mathrm{~mol}^{-2} \mathrm{~dm}^6 \mathrm{~sec}^{-1}$.

Thus, the rate constant of the reaction is given by option C:

$18 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{sec}^{-1}$.

Question53

For a reaction $\mathrm{r} = \mathrm{k}[\mathrm{A}][\mathrm{B}]^2$, if concentration of A is doubled the rate of reaction

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Options:

A. increases by 2

B. increases by 4

C. decrease by $\frac{1}{2}$

D. decreases by 2

Answer: A

Solution:

In the reaction given by the rate equation:

$r=k[A][B]^2$

where $[A]$ and $[B]$ are the concentrations of reactants A and B, respectively, k is the rate constant, and r is the rate of the reaction.

If the concentration of A is doubled, that means new concentration of A becomes $2[A]$. Substituting this into the rate equation:

$r'=k(2[A])[B]^2$

The original rate of the reaction is:

$r=k[A][B]^2$

The new rate of the reaction after doubling $[A]$ is:

$r'=2k[A][B]^2$

The factor by which the rate of reaction increases can be calculated by taking the ratio of the new rate r' to the original rate r:

Increasing Factor $=\frac{r'}{r}=\frac{2k[A][B]^2}{k[A][B]^2}=2$

Therefore, when the concentration of A is doubled, the rate of the reaction increases by a factor of 2.

So, **Option A: increases by 2** is the correct choice.

Question54

For the reaction

$$2NO_{(s)}+2H_{2(g)}\longrightarrow N_{2(g)}+2H_2O_{(s)}$$

$$\text{rate }=k[NO]^2[H_2].$$

What is the order of reaction with respect to H_2 and overall order of reaction respectively?

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Options:

A. $0, 2$

B. $1, 3$

C. $2, 1$

D. $3, 1$

Answer: B

Solution:

The given rate law for the reaction is:

$\text{rate} = k[\mathrm{NO}]^2[\mathrm{H}_2]$

The order of the reaction with respect to each reactant is determined by the exponents in the rate law:

The order with respect to NO is 2, as indicated by the exponent on $[\mathrm{NO}]$.

The order with respect to H_2 is 1, as indicated by the exponent on $[\mathrm{H}_2]$.

The overall order of the reaction is the sum of the orders with respect to each reactant:

$\text{Overall order} = 2 + 1 = 3$

Thus, with respect to H_2, the order is 1, and the overall order of reaction is 3. Therefore, the correct answer is:

Option B

$1, 3$

Question55

Which among the following statements is NOT true about rate constant?

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Options:

A. It is independent of concentration.

B. It varies with temperature.

C. For unit concentration it is equal to rate of reaction.

D. It is directly proportional to the product of concentrations of reacting species.

Answer: D

Solution:

Option D: It is directly proportional to the product of concentrations of reacting species.

The rate constant, often denoted as k, is a crucial factor in the rate law for chemical reactions. It depends on factors such as temperature and the presence of a catalyst but not directly on the concentration of reactants.

Here's a clear explanation:

Independence from Concentration: The rate constant is indeed independent of the concentration of the reacting species for a given reaction at a particular temperature (Option A).

Variation with Temperature: The rate constant varies with temperature. This relationship is often modeled by the Arrhenius equation, $k = Ae^{-\frac{E_a}{RT}}$ where A is the pre-exponential factor, E_a is the activation energy, R is the gas constant, and T is the temperature in Kelvin (Option B).

Unit Concentration: For a first-order reaction, the rate of reaction is equal to the rate constant when the concentration of the reactant is unit concentration (Option C).

Dependence on Product of Concentrations: The rate of reaction, not the rate constant itself, is proportional to the product of the concentrations of the reacting species raised to the power of their stoichiometric coefficients as described by the rate law expression. Thus, Option D is incorrect regarding the nature of the rate constant.

In summary, the rate constant (k) is not directly proportional to the concentrations of the reactants, but it is an intrinsic property of the reaction at a specific temperature and must be calculated or determined experimentally.

Question56

A first order reaction takes 40 minute for 20% decomposition. Calculate its rate constant.

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Options:

A. 5.6×10^{-3} minute $^{-1}$

B. 4.5×10^{-3} minute $^{-1}$

C. 6.5×10^{-3} minute $^{-1}$

D. 7.2×10^{-3} minute $^{-1}$

Answer: A

Solution:

If $[\mathrm{A}]_0=100$, then $[\mathrm{A}]_{\mathrm{t}}=100-20=80$

For a first order reaction,

$$
\begin{aligned}
\mathrm{k} & =\frac{2.303}{\mathrm{t}} \log _{10} \frac{[\mathrm{~A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}=\frac{2.303}{40 \mathrm{~min}} \log _{10} \frac{100}{80} \\
& =\frac{2.303}{40 \mathrm{~min}} \log _{10} 1.25=\frac{2.303}{40 \mathrm{~min}} \times 0.097 \\
& =5.57 \times 10^{-3} \text {minute}^{-1} \approx 5.6 \times 10^{-3} \text {minute}^{-1}
\end{aligned}
$$

Question57

If instantaneous rate of reaction is given as $-\frac{1}{\mathrm{a}} \frac{\mathrm{d}[\mathrm{~A}]}{\mathrm{dt}}=-\frac{1}{\mathrm{b}} \frac{\mathrm{d}[\mathrm{~B}]}{\mathrm{dt}}=\frac{1 \mathrm{~d}[\mathrm{C}]}{\mathrm{c}]}=\frac{1 \mathrm{~d}[\mathrm{D}]}{\mathrm{d}]}$ the reaction is represented as

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Options:

A. $\mathrm{A}+\mathrm{B} \longrightarrow \mathrm{C}+\mathrm{D}$

B. $\mathrm{aA}+\mathrm{B} \longrightarrow \mathrm{cC}+\mathrm{dD}$

C. $\mathrm{aA}+\mathrm{bB} \longrightarrow \mathrm{cC}+\mathrm{dD}$

D. $\mathrm{aA}+\mathrm{bB} \longrightarrow \mathrm{C}+\mathrm{D}$

Answer: C

Solution:

The instantaneous rate of reaction can be represented using the stoichiometric coefficients in a balanced chemical equation. The given expression for the instantaneous rate is :

$-\frac{1}{a}\frac{d[\mathrm{A}]}{dt}=-\frac{1}{b}\frac{d[\mathrm{B}]}{dt}=\frac{1}{c}\frac{d[\mathrm{C}]}{dt}=\frac{1}{d}\frac{d[\mathrm{D}]}{dt}$

This expression suggests that the reaction involves reactants A and B, and products C and D, with their rates of consumption and formation scaled by the inverse of their stoichiometric coefficients. Therefore, the corresponding balanced chemical equation for the reaction would be :

$a\mathrm{A}+b\mathrm{B}\rightarrow c\mathrm{C}+d\mathrm{D}$

Thus, the correct option is **Option C**.

Question58

Rate law for a reaction is $r=k[A]^2[B]$. If rate constant is $6.25\ \mathrm{mol}^{-2}\mathrm{dm}^6\ \mathrm{s}^{-1}$, what is the rate of reaction when $[\mathrm{A}]=1\ \mathrm{moldm}^{-3}$ and $[\mathrm{B}]=0.2\ \mathrm{moldm}^{-3}$?

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Options:

A. $1.250\ \mathrm{mol\ dm}^{-3}\mathrm{sec}^{-1}$

B. $2.125\ \mathrm{mol\ dm}^{-3}\mathrm{sec}^{-1}$

C. $3.105\ \mathrm{mol\ dm}^{-3}\mathrm{sec}^{-1}$

D. $2.0\times10^{-2}\ \mathrm{mol\ dm}^{-3}\mathrm{sec}^{-1}$

Answer: A

Solution:

The rate of a reaction can be determined using the rate law expression given by:

$r=k[A]^2[B]$

where:

$k=6.25\,\mathrm{mol}^{-2}\,\mathrm{dm}^6\,\mathrm{s}^{-1}$ is the rate constant,

$[A]=1\,\mathrm{mol\,dm}^{-3}$ is the concentration of reactant A,

$[B] = 0.2\,\mathrm{mol\,dm^{-3}}$ is the concentration of reactant B.

Substitute these values into the rate law equation:

$r = 6.25\,(1\,\mathrm{mol\,dm^{-3}})^2\,(0.2\,\mathrm{mol\,dm^{-3}})$

$r = 6.25 \times 1^2 \times 0.2\,\mathrm{mol\,dm^{-3}\,s^{-1}}$

$r = 6.25 \times 0.2\,\mathrm{mol\,dm^{-3}\,s^{-1}}$

$r = 1.25\,\mathrm{mol\,dm^{-3}\,s^{-1}}$

Thus, the rate of the reaction is $1.250\,\mathrm{mol\,dm^{-3}\,s^{-1}}$.

The correct answer is **Option A**: $1.250\,\mathrm{mol\,dm^{-3}\,sec^{-1}}$.

Question59

What is the time needed to reduce the initial concentration of reactant to 10% in a first order reaction if its half life time is 10 minutes?

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Options:

A. 15 minute

B. 20 minute

C. 25 minute

D. 33 minute

Answer: D

Solution:

For a first-order reaction, the relation between the rate constant k and the half-life $t_{1/2}$ is given by:

$t_{1/2} = \frac{\ln(2)}{k}$

Given the half-life $t_{1/2} = 10$ minutes, we can calculate the rate constant k:

$k = \frac{\ln(2)}{10}$

To find the time required to reduce the concentration of the reactant to 10% of the initial concentration, we use the first-order kinetics formula:

$$\ln\left(\frac{[A]_0}{[A]}\right) = kt$$

where $[A]_0$ is the initial concentration and $[A]$ is the final concentration. Here, $[A] = 0.1[A]_0$, which means the reactant is reduced to 10% of its initial concentration.

Substitute into the equation:

$$\ln\left(\frac{[A]_0}{0.1[A]_0}\right) = kt$$

$$\ln(10) = kt$$

Now substitute the expression for k:

$$\ln(10) = \left(\frac{\ln(2)}{10}\right)t$$

Solve for t:

$$t = \frac{10\ln(10)}{\ln(2)}$$

Using the approximations: $\ln(10) \approx 2.302$ and $\ln(2) \approx 0.693$, calculate t:

$$t \approx \frac{10\times 2.302}{0.693}$$

$$t \approx \frac{23.02}{0.693}$$

$$t \approx 33.22$$

Thus, the time needed to reduce the initial concentration to 10% is approximately **33 minutes**.

Therefore, the correct option is:

Option D: 33 minute

Question60

Which of the following is an elementary reaction?

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Options:

A. $C_2H_5I_{(g)} \longrightarrow C_2H_{4(\ g)} + HI_{(g)}$

B. $2NO_2Cl_{(g)} \longrightarrow 2NO_{2(z)} + Cl_{2(z)}$

C. $2NO_{2(g)} + F_{2(g)} \longrightarrow 2NO_2F_{(g)}$

D. $2NO_{(g)} + Cl_{2(g)} \longrightarrow 2NaCl_{(g)}$

Answer: A

Solution:

Among the given options, only $C_2H_5I \rightarrow C_2H_4 + HI$ occurs in single step involving a single reactant. Hence, it is an elementary reaction.

Question61

Rate constant of a reaction,$2NO_2Cl_{2(g)} \longrightarrow 2NO_{2(g)} + Cl_{2(g)}$ is 4.7672 minute $^{-1}$. Calculate half life of reaction.

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Options:

A. 0.0727 minute

B. 0.1454 minute

C. 0.2181 minute

D. 0.4362 minute

Answer: B

Solution:

For a first-order reaction, the half-life ($t_{1/2}$) is independent of the initial concentration and can be calculated using the formula:

$t_{1/2} = \frac{\ln(2)}{k}$

where:

k is the rate constant of the reaction,

$\ln(2)$ is the natural logarithm of 2, approximately 0.693.

Given:

Rate constant, $k = 4.7672\,\text{minute}^{-1}$

Plugging in the values, we find:

$t_{1/2} = \frac{0.693}{4.7672} \approx 0.1454\,\text{minutes}$

Therefore, the half-life of the reaction is **0.1454 minutes**. Thus, the correct answer is **Option B**.

Question62

Initial concentration of reactant in a first order reaction is $0.08\ \mathrm{mol\ dm}^{-3}$ What concentration would remain after 40 minute?

$$\left(\text{given } \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}} = 5.00\right)$$

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Options:

A. $0.008\ \mathrm{mol\ dm}^{-3}$

B. $0.08\ \mathrm{mol\ dm}^{-3}$

C. $0.016\ \mathrm{mol\ dm}^{-3}$

D. $0.032\ \mathrm{mol\ dm}^{-3}$

Answer: C

Solution:

For a first order reaction,

$$t = \frac{2.303}{k} \log_{10} \frac{[\mathrm{~A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}$$

$$\therefore \quad k = \frac{2.303}{40\ \text{minute}} \log_{10} 5 = 0.04$$

Now,

$\therefore \quad 0.04=\frac{2.303}{40} \log_{10} \frac{0.08}{[\mathrm{~A}]_{\mathrm{t}}}$
$\therefore \quad 0.69=\log_{10}(0.08)-\log_{10}[\mathrm{~A}]_{\mathrm{t}}$
$\therefore \quad \log_{10}[\mathrm{~A}]_{\mathrm{t}}=-1.096-0.69$
$\therefore \quad \log_{10}[\mathrm{~A}]_{\mathrm{t}}=-1.78$
$\therefore \quad [\mathrm{~A}]_{\mathrm{t}}=$ Antilog $\log_{10}(-1.78)=0.016$

Alternate method:

Given $\frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}=5$

$\therefore \quad [\mathrm{A}]_{\mathrm{t}}=\frac{0.08}{5}=0.016$

Question63

For the reaction $\mathrm{A}+\mathrm{B} \longrightarrow$ product, rate law equation is, rate $=k[A]^2[B]$. If rate of reaction is $0.22 \mathrm{~mol} \mathrm{~L}^{-1} \mathrm{~s}^{-1}$, calculate rate constant. ($[\mathrm{A}]=1 \mathrm{~mol} \mathrm{~L}^{-1},[\mathrm{~B}]=0.25 \mathrm{~mol} \mathrm{~L}^{-1}$)

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Options:

A. $0.44 \mathrm{~mol}^{-2} \mathrm{~L}^2 \mathrm{~s}^{-1}$

B. $0.88 \mathrm{~mol}^{-2} \mathrm{~L}^2 \mathrm{~s}^{-1}$

C. $1.136 \mathrm{~mol}^{-2} \mathrm{~L}^2 \mathrm{~s}^{-1}$

D. $3.52 \mathrm{~mol}^{-2} \mathrm{~L}^2 \mathrm{~s}^{-1}$

Answer: B

Solution:

To calculate the rate constant k, use the provided rate law equation:

rate $=k[A]^2[B]$

Substitute the known values into the equation:

Rate $=0.22 \mathrm{~mol} \mathrm{~L}^{-1} \mathrm{~s}^{-1}$

$[A]=1 \mathrm{~mol} \mathrm{~L}^{-1}$

$[B] = 0.25 \text{ mol L}^{-1}$

Plug these values into the rate law:

$0.22 \text{ mol L}^{-1} \text{ s}^{-1} = k(1 \text{ mol L}^{-1})^2(0.25 \text{ mol L}^{-1})$

Simplify the expression:

$0.22 = k \cdot 1 \cdot 0.25$

Solve for k:

$k = \frac{0.22}{0.25}$

Calculate k:

$k = 0.88 \text{ mol}^{-2} \text{ L}^2 \text{ s}^{-1}$

The answer is Option B: $0.88 \text{ mol}^{-2} \text{ L}^2 \text{ s}^{-1}$.

Question64

Rate of a first order reaction is $1.5 \times 10^{-2} \text{ mol L}^{-1}$ minute $^{-1}$ at 0.5 M concentration of reactant, calculate half life of reaction.

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Options:

A. 0.383 minute

B. 7.53 minute

C. 8.73 minute

D. 23.1 minute

Answer: D

Solution:

For a first-order reaction, the rate of the reaction is given by:

$\text{Rate} = k[A]$

where k is the rate constant, and $[A]$ is the concentration of the reactant.

Given the rate:

$1.5 \times 10^{-2}\,\mathrm{mol\,L^{-1}\,min^{-1}}$

and the concentration $[A] = 0.5\,\mathrm{M}$, we can solve for the rate constant k:

$k = \frac{\text{Rate}}{[A]} = \frac{1.5\times10^{-2}}{0.5} = 3.0 \times 10^{-2}\,\mathrm{min}^{-1}$

For a first-order reaction, the half-life $(t_{1/2})$ is given by:

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

Substitute the value of k:

$t_{1/2} = \frac{0.693}{3.0\times10^{-2}} = 23.1\,\mathrm{minutes}$

Thus, the correct option is:

Option D: 23.1 minute

Question65

Which from following decides the rate of multistep reaction?

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Options:

A. rate of slowest step.

B. rate of fastest step.

C. average rate of various steps.

D. instantaneous rate of fast step.

Answer: A

Solution:

The rate of a multistep reaction is determined by the **rate of the slowest step**, often referred to as the **rate-determining step** or **rate-limiting step**. This is because the slowest step in a reaction pathway acts as a bottleneck, restricting the overall speed at which the entire reaction can proceed.

In multistep reactions, each step can have its own reaction rate, and the step with the highest energy barrier (lowest rate) will limit the rate at which products are formed. By analogy, imagine cars moving through a series of traffic lights: even if some lights are green for longer (fast steps), the time it takes to pass through the sequence is restricted by the light that stays red the longest (the slowest step).

Therefore, the correct choice is:

Option A: rate of slowest step.

Question66

For a zero order reaction, $\mathrm{A} \longrightarrow$ product, concentration of A decreases from $1.2 \mathrm{~mol} \mathrm{dm}^{-3}$ to $0.4 \mathrm{~mol} \mathrm{dm}^{-3}$ in 240 second. What is rate constant of the reaction?

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Options:

A. $0.1 \mathrm{~mol} \mathrm{dm}^{-3}$ minute $^{-1}$

B. $0.2 \mathrm{~mol} \mathrm{dm}^{-3}$ minute $^{-1}$

C. $0.3 \mathrm{~mol} \mathrm{dm}^{-3}$ minute $^{-1}$

D. $0.4 \mathrm{~mol} \mathrm{dm}^{-3}$ minute $^{-1}$

Answer: B

Solution:

For a zero-order reaction, the rate of reaction is independent of the concentration of the reactant. The rate law for a zero-order reaction is given by:

$[A]=[A]_0-k t$

where:

$[A]$ is the concentration of the reactant at time t,

$[A]_0$ is the initial concentration of the reactant,

k is the rate constant,

t is the time.

Given:

Initial concentration, $[A]_0 = 1.2\,\mathrm{mol\,dm^{-3}}$,

Final concentration, $[A] = 0.4\,\mathrm{mol\,dm^{-3}}$,

Time, $t = 240\,\mathrm{seconds}$.

Using the zero-order kinetics equation, substitute the known values:

$0.4 = 1.2 - k \times 240$

Rearrange to solve for the rate constant, k:

$0.4 = 1.2 - 240k$

$240k = 1.2 - 0.4$

$240k = 0.8$

$k = \frac{0.8}{240}$

Calculate k:

$k = 0.00333\,\mathrm{mol\,dm^{-3}\,s^{-1}}$

Convert k from seconds to minutes by multiplying by 60 seconds per minute:

$k = 0.00333\,\mathrm{mol\,dm^{-3}\,s^{-1}} \times 60\,\mathrm{s\,min^{-1}}$

$k = 0.2\,\mathrm{mol\,dm^{-3}\,min^{-1}}$

Therefore, the rate constant of the reaction is:

Option B: $0.2\,\mathrm{mol\,dm^{-3}\,minute^{-1}}$.

Question67

Which of the following is NOT true about order of a reaction?

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Options:

A. It is sum of power of concentration terms in rate law equation.

B. It may be integer, fraction or zero.

C. It is theoretical quantity

D. It is experimentally determined quantity.

Answer: C

Solution:

Option C: It is a theoretical quantity.

The order of a reaction is determined experimentally and represents the sum of the powers of concentration terms in the rate law equation. It can indeed be an integer, a fraction, or even zero, depending on the reaction. This experimental determination highlights its nature as an observed, rather than theoretical, property.

Question68

Calculate rate constant of first order reaction if concentration of reactant decreases by 90% in 30 minute?

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Options:

A. $2.16 \times 10^{-2} \text{ min}^{-1}$

B. $3.52 \times 10^{-2} \text{ min}^{-1}$

C. $4.81 \times 10^{-2} \text{ min}^{-1}$

D. $7.67 \times 10^{-2} \text{ min}^{-1}$

Answer: D

Solution:

Concentration decreases by 90%. Hence, 10% reactant is left after 30 minutes.

$$\begin{aligned} k &= \frac{2.303}{t} \log_{10} \frac{[A]_0}{[A]_t} \\ k &= \frac{2.303}{30} \log_{10} \frac{100}{10} \\ &= \frac{2.303}{30 \text{ min}} = 7.67 \times 10^{-2} \text{ min}^{-1} \end{aligned}$$

Question69

The reaction $2\mathrm{~A}+\mathrm{B}+\mathrm{C} \longrightarrow \mathrm{D}+\mathrm{E}$ is found to be first order in A , second order in B and zero order in C . What is the effect of increasing concentration of all reactants twice?

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Options:

A. Rate of reaction increases 8 times.

B. Rate of reaction increases 24 times.

C. Rate of reaction increases 36 times.

D. Rate of reaction remains unaffected.

Answer: A

Solution:

$$\begin{aligned} \text{Rate } & =\mathrm{k}[\mathrm{~A}][\mathrm{B}]^2[\mathrm{C}]^0 \\ (\text{ Rate })_1 & =\mathrm{k} \times[2 \mathrm{~A}] \times[2 \mathrm{~B}]^2 \times[2 \mathrm{C}]^0 \\ & =8 \mathrm{k}[\mathrm{~A}][\mathrm{B}]^2[\mathrm{C}]^0 \end{aligned}$$

$$\therefore \quad \frac{(\text{ Rate })_1}{\text{Rate}}=\frac{8 \mathrm{k}[\mathrm{~A}][\mathrm{B}]^2[\mathrm{C}]^0}{\mathrm{k}[\mathrm{~A}][\mathrm{B}]^2[\mathrm{C}]^0}=8$$

$\therefore \quad (\text{Rate})_1=8 \times \text{Rate}$

i.e., rate of reaction increases 8 times.

Question70

Identify order of following reaction.

$$\mathrm{H}_2 \mathrm{O}_{2(\mathrm{~g})} \longrightarrow 2 \mathrm{H}_2 \mathrm{O}_{(\mathrm{l})}+\mathrm{O}_{2(\mathrm{~g})}$$

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Options:

A. 0

B. 1

C. 1.5

D. 2

Answer: B

Solution:

To determine the order of a reaction, it's essential to analyze the rate law, which is usually determined experimentally. However, without experimental data or a given rate expression, predicting the order solely from the chemical equation is not possible.

For the decomposition of hydrogen peroxide:

$H_2O_{2(g)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$

Experimentally, the decomposition of hydrogen peroxide is typically found to be a first-order reaction with respect to H_2O_2. Thus, the total reaction order is:

Option B: 1

This is based on common experimental findings for the decomposition of hydrogen peroxide in gaseous form.

Question71

For the reaction, $N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}NH_3$ is formed at a rate of $0.088 \text{ mol dm}^{-3}\text{ s}^{-1}$. Calculate consumption rate of $N_{2(g)}$.

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Options:

A. $0.011 \text{ mol dm}^{-3}\text{ s}^{-1}$

B. $0.022 \text{ mol dm}^{-3}\text{ s}^{-1}$

C. $0.033 \text{ mol dm}^{-3}\text{ s}^{-1}$

D. $0.044 \text{ mol dm}^{-3}\text{ s}^{-1}$

Answer: D

Solution:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

$$\text{Rate } = -\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$$

$$\therefore \quad -\frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt} = \frac{1}{2} \times 0.088$$

$$\therefore \quad \text{Rate of consumption of } N_2 = 0.044 \text{ mol dm}^{-3}\text{ s}^{-1}$$

Question72

Consider the reaction $3I^- + S_2O_8^{2-} \longrightarrow I_3^- + 2SO_4^{2-}$, at a particular time t , $\frac{d[SO_4^{2-}]}{dt}$ is 2.2×10^{-2} moldm^{-3} s^{-1}. What is the value of $\frac{d[S_2O_8^{2-}]}{dt}$?

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Options:

A. 1.1×10^{-2} mol dm^{-3} s^{-1}

B. 2.2×10^{-2} mol dm^{-3} s^{-1}

C. 3.3×10^{-2} mol dm^{-3} s^{-1}

D. 6.6×10^{-2} mol dm^{-3} s^{-1}

Answer: A

Solution:

$$3I^- + S_2O_8^{2-} \longrightarrow I_3^- + 2SO_4^{2-}$$

Rate of reaction

$$= -\frac{1}{3}\frac{d[I^-]}{dt} = -\frac{d[S_2O_8^{2-}]}{dt} = \frac{d[I_3^-]}{dt} = \frac{1}{2}\frac{d[SO_4^{2-}]}{dt}$$

Hence, $\frac{d[S_2O_8^{2-}]}{dt} = \frac{1}{2}\frac{d[SO_4^{2-}]}{dt}$

$= \frac{1}{2} \times 2.2 \times 10^{-2} = 1.1 \times 10^{-2} \mathrm{~mol~dm}^{-3} \mathrm{~s}^{-1}$

Question73

For the reaction, $NO_{2(g)} + CO_{(g)} \longrightarrow NO_{(g)} + CO_{2(g)}$ rate of reaction is proportional to square of $[NO_2]$ and independent of [CO].What is the rate law equation?

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Options:

A. $r = k\frac{[NO_2]^{\frac{1}{2}}}{[CO]}$

B. $r = k[NO_2]^2[CO]^0$

C. $r = \frac{1}{2}k[NO_2][CO]$

D. $r = k[NO_2]^2[CO]$

Answer: B

Solution:

The rate law equation describes the relationship between the rate of a chemical reaction and the concentration of the reactants. For the given reaction:

$NO_{2(g)} + CO_{(g)} \longrightarrow NO_{(g)} + CO_{2(g)}$

It is stated that the rate of reaction is proportional to the square of $[NO_2]$ and is independent of $[CO]$. This means that the reaction rate depends solely on the concentration of NO_2 raised to the power of 2, and there is no dependence on the $[CO]$ concentration.

Therefore, the rate law equation for this reaction is:

$r = k[NO_2]^2[CO]^0$

Since any concentration term raised to the power of 0 is equal to 1, the rate law simplifies to:

$r = k[NO_2]^2$

This matches with the expression given in Option B:

$r = k[NO_2]^2[CO]^0$

Option B is the correct rate law equation.

Question74

Find the percentage of unreacted reactant for zero order reaction in 90 second having rate constant $1 \mathrm{~mol} \mathrm{dm}^{-3} \mathrm{~s}^{-1}$.

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Options:

A. 5%

B. 10%

C. 15%

D. 20%

Answer: B

Solution:

For a zero-order reaction, the rate of reaction is given by:

Rate $= \frac{d[A]}{dt} = -k$

where k is the rate constant and $[A]$ is the concentration of the reactant. The integrated rate equation for a zero-order reaction can be expressed as:

$[A] = [A]_0 - kt$

where $[A]_0$ is the initial concentration of the reactant and t is the time.

To find the percentage of unreacted reactant, we can rearrange the equation to solve for the concentration of the reactant that is left:

Assume an initial concentration $[A]_0$.

Substitute the given rate constant ($k = 1 \text{ mol dm}^{-3} \text{ s}^{-1}$) and time ($t = 90$ s) into the equation:

$[A] = [A]_0 - (1 \text{ mol dm}^{-3} \text{ s}^{-1})(90 \text{ s}) = [A]_0 - 90$

If $[A]_0 = 100 \text{ mol dm}^{-3}$ for simplicity (since it will cancel out besides demonstrating the percentage), calculate $[A]$:

$[A] = 100 - 90 = 10 \text{ mol dm}^{-3}$

The percentage of unreacted reactant is then:

$\left(\frac{[A]}{[A]_0}\right) \times 100\% = \left(\frac{10}{100}\right) \times 100\% = 10\%$

Therefore, the percentage of unreacted reactant after 90 seconds is 10%.

Option B: 10%

Question75

What is the half life of a first order reaction if time required to decrease concentration of reactant from 0.8 M to 0.2 M is 12 hour?

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Options:

A. 1.5 hour

B. 3 hour

C. 6 hour

D. 12 hour

Answer: C

Solution:

To determine the half-life of a first-order reaction, we can use the first-order kinetics formula:

$k = \frac{1}{t} \ln\left(\frac{[A_0]}{[A]}\right)$

where

k is the rate constant,

t is the time elapsed,

$[A_0]$ is the initial concentration, and

$[A]$ is the concentration at time t.

Given:

$[A_0] = 0.8\,\mathrm{M},$

$[A] = 0.2\,\mathrm{M},$

$t = 12\,\text{hours}.$

Substituting the values:

$k = \frac{1}{12} \ln\left(\frac{0.8}{0.2}\right) = \frac{1}{12} \ln(4)$

The half-life ($t_{1/2}$) for a first-order reaction is given by:

$t_{1/2} = \frac{\ln(2)}{k}$

Substitute the expression for k:

$t_{1/2} = \frac{\ln(2)}{\frac{1}{12} \ln(4)}$

This simplifies to:

$t_{1/2} = 12 \times \frac{\ln(2)}{\ln(4)}$

Since $\ln(4) = 2 \ln(2)$, we have:

$t_{1/2} = 12 \times \frac{\ln(2)}{2 \ln(2)} = 12 \times \frac{1}{2} = 6\,\text{hours}$

Thus, the half-life of this first-order reaction is **6 hours**.

Answer: Option C: 6 hour

Question76

For the reaction $2\,\mathrm{N_2O_{5(\,g)}} \longrightarrow 4\mathrm{NO_{2(\,g)}} + \mathrm{O_{2(\,g)}}$ rate and rate constant are $1.02 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}$ and $3.4 \times 10^{-5}\ \mathrm{s^{-1}}$.What is the conc. of $\mathrm{N_2O_5}$?

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Options:

A. $1.7\ \mathrm{mol\ L^{-1}}$

B. $3.0\ \mathrm{mol\ L^{-1}}$

C. $3.4\ \mathrm{mol\ L^{-1}}$

D. 5.1 $\mathrm{mol\,L^{-1}}$

Answer: B

Solution:

The rate law for a reaction is expressed as:

Rate $= k \times [\text{Reactants}]^n$

For this specific reaction, it can be written as:

Rate $= k \times [\mathrm{N_2O_5}]$

Here, the rate ($1.02 \times 10^{-4}\,\mathrm{mol\,L^{-1}\,s^{-1}}$) and the rate constant ($k = 3.4 \times 10^{-5}\,\mathrm{s^{-1}}$) are provided. Since the stoichiometry of the reactant 2 $\mathrm{N_2O_5}$ implies a first-order dependence, the concentration of $\mathrm{N_2O_5}$ is the unknown variable.

Rearranging the rate law gives:

$[\mathrm{N_2O_5}] = \frac{\text{Rate}}{k}$

Substituting the given values:

$[\mathrm{N_2O_5}] = \frac{1.02\times10^{-4}\,\mathrm{mol\,L^{-1}\,s^{-1}}}{3.4\times10^{-5}\,\mathrm{s^{-1}}}$

Calculating this gives:

$[\mathrm{N_2O_5}] = 3.0\,\mathrm{mol\,L^{-1}}$

Thus, the concentration of $\mathrm{N_2O_5}$ is **Option B**: $3.0\,\mathrm{mol\,L^{-1}}$.

Question77

What is the order of following reaction

$$2\mathrm{H_2O_{2(\,g)}} \longrightarrow 2\mathrm{H_2O_{(l)}} + \mathrm{O_{2(\,g)}}$$

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Options:

A. 1

B. 0

C. 3

D. 2

Answer: A

Solution:

To determine the order of the reaction $2H_2O_{2(g)} \longrightarrow 2H_2O_{(l)} + O_{2(g)}$, it is essential to review how reaction orders are defined. The order of a reaction with respect to a particular reactant is determined experimentally and indicates the power to which the concentration of that reactant is raised in the rate law expression.

For a reaction:

If the reaction is zero-order, the rate is independent of the concentration of the reactants.

If the reaction is first-order, the rate is directly proportional to the concentration of one reactant.

If the reaction is second-order, the rate depends on the concentration of one reactant squared or the product of the concentrations of two different reactants.

If the reaction is third-order, the rate depends on the concentration of one reactant cubed, or the combination of concentrations for three reactions.

However, a reaction's stoichiometry (i.e., the coefficients in the balanced equation) does not directly give the reaction order; it must be determined through experiments. In practice, the decomposition of hydrogen peroxide is typically first-order with respect to hydrogen peroxide, given its concentration dependence:

$\text{rate} = k[H_2O_2]$

Therefore, the reaction order is often experimentally found to be 1.

So, the answer is **Option A - 1**.

Question78

Half life of a first order reaction is 1 hour. What fraction of it will remain after 3 hour?

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Options:

A. $\frac{1}{8}$

B. $\frac{1}{9}$

C. $\frac{1}{16}$

D. $\frac{1}{64}$

Answer: A

Solution:

For a first orer reaction,

$$\begin{array}{ccccccc} 1 & \xrightarrow{t_{1/2}} & \dfrac{1}{2} & \xrightarrow{t_{1/2}} & \dfrac{1}{4} & \xrightarrow{t_{t/2}} & \dfrac{1}{8} \\ [A]_0 & & [\mathrm{~A}]_t & & [A]_t & & [\mathrm{~A}]_t \end{array}$$

$\because$ Half life $\left(\mathrm{t}_{1/2}\right)=1$ hour

$\therefore$ Fraction available after 3 hours $=3 \times \mathrm{t}_{1/2}=\frac{1}{8}$

Question79

In the Arrhenius plot of logk versus $1/T$ find the value of intercept on y axis

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Options:

A. $\log_{10} \mathrm{A}$

B. $\frac{-\mathrm{Ea}}{\mathrm{R}}$

C. lnk

D. $\mathrm{R}/\mathrm{E}_{\mathrm{a}}$

Answer: A

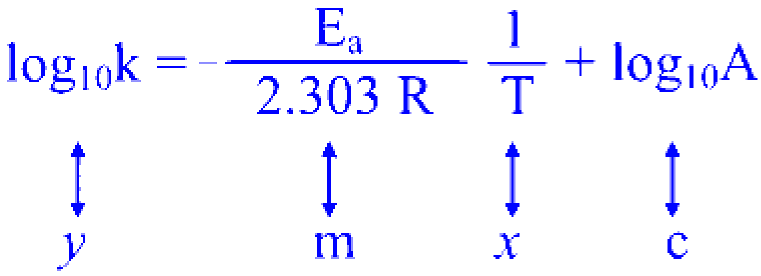
Solution:

The Arrhenius equation is

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

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

$\therefore$



$$\log_{10} k = -\frac{E_a}{2.303\,R}\,\frac{1}{T} + \log_{10} A$$

Thus, the value of intercept on y axis is $\log_{10}$A.

Question80

For the reaction,

$$3I^-_{(aq.)} + S_2O^{2-}_{8(aq.)} \longrightarrow 2SO^{2-}_{4(aq.)} + I^-_{3(aq.)}$$

rate of formation of $SO^{2-}_{4(aq.)}$ **is** $0.044\ \text{mol dm}^{-3}\ \text{s}^{-1}$**.**

Calculate rate of consumption of $I^-_{(aq.)}$**.**

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Options:

A. $0.022\ \text{mol dm}^{-3}\ \text{s}^{-1}$

B. $0.044\ \text{mol dm}^{-3}\ \text{s}^{-1}$

C. $0.066\ \text{mol dm}^{-3}\ \text{s}^{-1}$

D. $0.088\ \text{mol dm}^{-3}\ \text{s}^{-1}$

Answer: C

Solution:

$3I_{(aq)} + S_2O_{8(aq)}^{2-} \longrightarrow 2SO_{4(aq)}^{2-} + I_{3(aq)}^{-}$

Rate of reaction

$= -\frac{1}{3}\frac{d[I^-]}{dt} = -\frac{d[\ S_2O_8^{2-}]}{dt} = \frac{1\ d[SO_4^{2-}]}{2} = \frac{d[I_3^-]}{dt}$

$\therefore$ Rate of consumption of $I^- = -\frac{1}{3}\frac{d[I^-]}{dt} = \frac{1}{2}\frac{d[SO_4^2]}{dt}$

$= \frac{3\times0.044}{2} = 0.066 \text{ mol dm}^{-3}\text{ s}^{-1}$

Question81

For a zero order reaction $A \longrightarrow$ product. Conc. of A decreases from 0.8 mol dm^{-3} to 0.2 mol dm^{-3} in 6 minute. What is rate constant of the reaction?

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Options:

A. $0.01 \text{ mol dm}^{-3} \text{ minute}^{-1}$

B. $1.0 \text{ mol dm}^{-3} \text{ minute}^{-1}$

C. $0.1 \text{ mol dm}^{-3} \text{ minute}^{-1}$

D. $1.66 \text{ mol dm}^{-3} \text{ minute}^{-1}$

Answer: C

Solution:

For a zero-order reaction, the rate of reaction is constant and can be defined by the equation:

$[A] = [A_0] - kt$

where:

$[A]$ is the concentration of the reactant at time t,

$[A_0]$ is the initial concentration of the reactant,

k is the rate constant, and

t is the time.

Given that the concentration of A decreases from $0.8\,\mathrm{mol\,dm^{-3}}$ to $0.2\,\mathrm{mol\,dm^{-3}}$ in 6 minutes, we can use these values to determine k.

Substitute the given values into the equation:

$0.2 = 0.8 - k \times 6$

Rearrange to solve for k:

$k \times 6 = 0.8 - 0.2 = 0.6$

$k = \frac{0.6}{6} = 0.1\,\mathrm{mol\,dm^{-3}\,minute^{-1}}$

Thus, the rate constant of the reaction is:

Option C: $0.1\,\mathrm{mol\,dm^{-3}\,minute^{-1}}$

Question82

What is rate constant of a first order reaction if 60% reactant decompose in 45 minute?

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Options:

A. 0.010 minute $^{-1}$

B. 0.015 minute $^{-1}$

C. 0.020 minute $^{-1}$

D. 0.025 minute $^{-1}$

Answer: C

Solution:

For a first-order reaction, the rate constant (k) can be determined using the formula:

$k = \frac{1}{t}\ln\left(\frac{[R]_0}{[R]}\right)$

where:

t is the time taken for the reaction,

$[R]_0$ is the initial concentration of the reactant,

$[R]$ is the concentration of the reactant at time t,

ln is the natural logarithm.

Given that 60% of the reactant decomposes, 40% remains, thus:

$\frac{[R]}{[R]_0} = 0.40$

Substitute this expression and the given time into the equation:

$k = \frac{1}{45} \ln\left(\frac{1}{0.40}\right)$

Calculate the natural logarithm:

$\ln\left(\frac{1}{0.40}\right) = \ln(2.5) \approx 0.9163$

Now, calculate the rate constant:

$k = \frac{1}{45} \times 0.9163 \approx 0.02036 \text{ minute}^{-1}$

Thus, the rate constant is approximately:

$0.020 \text{ minute}^{-1}$

Therefore, the correct option is **C**.

Question83

Half life of a zero order reaction is directly proportional to ______________

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Options:

A. temperature

B. rate constant

C. amount of product formed

D. initial concentration of reactant

Answer: D

Solution:

The half-life of a zero-order reaction is directly proportional to the initial concentration of the reactant.

To understand this, consider the integrated rate law for a zero-order reaction:

$[A] = [A]_0 - kt$

where:

$[A]$ is the concentration of the reactant at time t,

$[A]_0$ is the initial concentration,

k is the rate constant, and

t is the time.

For a zero-order reaction, the half-life ($t_{1/2}$) is the time required for the concentration of the reactant to decrease to half of its initial concentration ($[A]_0/2$). Setting $[A] = [A]_0/2$ and solving for $t_{1/2}$ gives:

$t_{1/2} = \frac{[A]_0}{2k}$

From this expression, it is evident that the half-life of a zero-order reaction is directly proportional to the initial concentration of the reactant, $[A]_0$. Hence, the correct option is:

Option D: initial concentration of reactant

Question84

What time is required for 100 g of reactant to reduce to 25 g in a first order reaction having half life 5760 year?

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Options:

A. 4760 year

B. 8640.26 year

C. 2880.15 year

D. 11526.48 year

Answer: D

Solution:

For a first order reaction,
$[A_0] = 100\text{ g}; [A_t] = 25\text{ g}; t_{1/2} = 5760\text{ years}$
$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{5760\text{yr}} = 1.203 \times 10^{-1}\text{year}^{-1}$$
$$t = \frac{2.303}{k}\log_{10}\frac{[A_0]}{[A_t]}$$
$$t = \frac{2.303}{1.203 \times 10^{-1}\text{year}^{-1}}\log_{10}\frac{100}{25}$$
$= 19141.8 \times 0.6021 = 11526.5\text{ years}$

Question85

Which of the following statements is correct about zero order reaction?

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Options:

A. Rate of reaction depends on rate constant.

B. Rate of reaction is independent of concentration of reactant.

C. Half life of reaction does not depend on initial concentration.

D. Unit of rate constant is time $^{-1}$.

Answer: B

Solution:

The correct statement about a zero-order reaction is:

Option B

Rate of reaction is independent of concentration of reactant.

Explanation:

In a zero-order reaction, the rate of reaction is given by:

$\text{Rate} = k$

where k is the rate constant. This implies that the rate of reaction is constant and does not depend on the concentration of the reactant.

Option A is misleading because, while the rate does involve the rate constant, it is independent of the concentration of the reactants in a zero-order reaction.

Option C is incorrect for zero-order reactions because the half-life ($t_{1/2}$) for a zero-order reaction is given by:

$t_{1/2} = \frac{[A]_0}{2k}$

where $[A]_0$ is the initial concentration. This clearly shows that the half-life depends on the initial concentration for zero-order reactions.

Option D is incorrect because the unit of the rate constant (k) for a zero-order reaction is not time $^{-1}$. Instead, it is:

$\text{Unit of } k = \text{concentration} \cdot \text{time}^{-1}$

For example, in a reaction where concentration is measured in mol/L and time in seconds, the unit of k would be $\mathrm{mol\ L^{-1}\ s^{-1}}$.

Question86

For a reaction, $2\ \mathrm{N_2O_{5(\ g)}} \longrightarrow 4\mathrm{NO_{2(\ g)}} + \mathrm{O_{2(\ g)}}$

$\mathrm{N_2O_5}$ disappears at a rate of $0.06\ \mathrm{mol\ dm^{-3}\ s^{-1}}$ What is rate of $\mathrm{NO_{2(\ g)}}$ formation?

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Options:

A. $0.06\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

B. $0.12\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

C. $0.18\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

D. $0.24\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

Answer: B

Solution:

$2\,N_2O_5 \longrightarrow 4NO_2 + O_2$

Overall rate of reaction can be expressed as:

$-\frac{1}{2}\frac{d[\,N_2O_5]}{dt} = \frac{1}{4}\frac{d[NO_2]}{dt} = \frac{d[O_2]}{dt}$

Rate of formation of NO_2 :

$\frac{d[NO_2]}{dt} = -\frac{4}{2}\frac{d[\,N_2O_5]}{dt} = \frac{4\times 0.06}{2} = 0.12\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

Question87

What is the value of slope, if $\log_{10}\ \mathrm{K}$ (y-axis) is plotted versus $1/\mathrm{T}$ (x-axis) for Arrhenius equation?

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Options:

A. $\log_{10}\mathrm{A}$

B. $\frac{2\cdot 303R}{E_a}$

C. $\frac{-\mathrm{E_a}}{2\cdot 303\mathrm{R}}$

D. $-\log_{10}\mathrm{A}$

Answer: C

Solution:

Arrhenius equation is $\mathrm{k} = \mathrm{Ae}^{\mathrm{E_G}/\mathrm{RT}}$

$\therefore \quad \ln \mathrm{k} = -\frac{\mathrm{E_a}}{\mathrm{RT}} + \ln \mathrm{A}$

$$\therefore \quad \underset{\substack{\updownarrow \\ y}}{\log_{10}k} = -\underset{\substack{\updownarrow \\ m}}{\frac{E_a}{2.303\ R}}\ \underset{\substack{\updownarrow \\ x}}{\frac{1}{T}} + \underset{\substack{\updownarrow \\ c}}{\log_{10}A}$$

Thus, the slope of the line is $-\frac{E_a}{2.303R}$

Question88

For the reaction, $A+3B \longrightarrow 2C$

rate of consumption of A is $1.4 \mathrm{~mol~dm}^{-3} \mathrm{~sec}^{-1}$. Calculate rate of formation of C ?

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Options:

A. $0.07 \mathrm{~mol~dm}^{-3} \mathrm{~sec}^{-1}$

B. $1.4 \mathrm{~mol~dm}^{-3} \mathrm{~sec}^{-1}$

C. $2.8 \mathrm{~mol~dm}^{-3} \mathrm{~sec}^{-1}$

D. $3.5 \mathrm{~mol~dm}^{-3} \mathrm{~sec}^{-1}$

Answer: C

Solution:

For the reaction,

$A+3B \rightarrow 2C$

Rate of reaction

$=\frac{-\mathrm{d}[\mathrm{~A}]}{\mathrm{dt}}=-\frac{1}{3} \frac{\mathrm{d}[\mathrm{~B}]}{\mathrm{dt}}=\frac{1}{2} \frac{\mathrm{d}[\mathrm{C}]}{\mathrm{dt}}$

Hence,

$\frac{-\mathrm{d}[\mathrm{~A}]}{\mathrm{dt}}=\frac{1}{2} \frac{\mathrm{d}[\mathrm{C}]}{\mathrm{dt}}$

$$\begin{aligned} \therefore \quad \frac{\mathrm{d}[\mathrm{C}]}{\mathrm{dt}} & =2 \times 1.4 \mathrm{~mol~dm}^{-3} \mathrm{~s}^{-1} \\ & =2.8 \mathrm{~mol~dm}^{-3} \mathrm{~s}^{-1} \end{aligned}$$

Question89

In a first order reaction if concentartion of reactant drops from 0.8 mol L^{-1} to 0.4 mol L^{-1} in 15 minute. What is the time required to drop concentration from 0.1 mol L^{-1} to 0.025 mol L^{-1}.

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Options:

A. 7.5 minute

B. 15 minute

C. 30 minute

D. 60 minute

Answer: C

Solution:

As the concentration of reactant is reduced to half i.e., from 0.8 mol L^{-1} to 0.4 mol L^{-1} in 15 minutes,

$\therefore \quad t_{1/2} = 15$ minutes

For a first order reaction,

$$\begin{array}{ccccc} 1 & \xrightarrow{t_{1/2}} & 1/2 & \xrightarrow{t_{1/2}} & 1/4 \\ [A]_0 & & [A]_t & & [A]_t \\ 1.0 & & 0.05 & & 0.025 \end{array}$$

$\therefore$ Time required to drop the concentration from
0.1 mol L^{-1} to 0.025 mol L^{-1} $= 2 \times t_{1/2}$
$= 2 \times 15$ minutes
$= 30$ minutes

Question90

Calculate the time required for reactant to decrease the concentration from 100% to 20%, if rate constant of first order reaction is 0.02303 hours $^{-1}$.

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Options:

A. 28 hour

B. 42 hour

C. 56 hour

D. 70 hour

Answer: D

Solution:

For a first-order reaction, the relationship between the rate constant (k) and the concentration of reactant ($[A]$) over time (t) is given by the integrated rate law:

$\ln \left(\frac{[A]_0}{[A]}\right)=k t$

where:

$[A]_0$ is the initial concentration,

$[A]$ is the concentration at time t,

k is the rate constant.

Given that the concentration decreases from 100% to 20%, we have:

$[A]_0=100\%$

$[A]=20\%$

$k=0.02303 \text{ hours}^{-1}$

Substitute these values into the equation:

$\ln \left(\frac{100}{20}\right)=0.02303 \times t$

Simplify the fraction:

$\frac{100}{20}=5$

Thus, the equation becomes:

$\ln(5) = 0.02303 \times t$

Calculate $\ln(5)$:

$\ln(5) \approx 1.60944$

Now, solve for t:

$t = \frac{1.60944}{0.02303}$

$t \approx 69.889$ hours

This is approximately 70 hours. Therefore, the correct answer is:

Option D: 70 hour

Question91

What is the rate of formation of O_2 for the reaction stated below?

$$2\,\mathrm{N}_2\mathrm{O}_{5(8)} \longrightarrow 4\mathrm{NO}_{2(g)} + \mathrm{O}_{2(\ g)}$$
$$\left[\frac{\mathrm{d}\left[\mathrm{N}_2\mathrm{O}_5\right]}{\mathrm{dt}} = 0.02\ \mathrm{moldm}^{-3}\ \mathrm{s}^{-1}\right]$$

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Options:

A. $0.01\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

B. $0.02\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

C. $0.03\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

D. $0.04\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$

Answer: A

Solution:

For the reaction:

$2\,N_2O_5 \longrightarrow 4NO_2 + O_2$

We use the concept of reaction rates to relate the rate of change of concentration for the reactants and products. The rate of disappearance or formation is related to stoichiometry in the balanced chemical equation.

Given:

$$\frac{d[N_2O_5]}{dt} = -0.02\,\mathrm{mol\,dm^{-3}\,s^{-1}}$$

The negative sign indicates that N_2O_5 is being consumed. According to the balanced equation, 2 moles of N_2O_5 produce 1 mole of O_2.

Using stoichiometry, the rate of formation of O_2 is half the rate of disappearance of N_2O_5:

$$\frac{d[O_2]}{dt} = -\frac{1}{2} \times \frac{d[\,N_2O_5]}{dt}$$

Substitute the given rate into the equation:

$$\frac{d[O_2]}{dt} = -\frac{1}{2} \times (-0.02) = 0.01\,\mathrm{mol\,dm^{-3}\,s^{-1}}$$

Thus, the rate of formation of O_2 is:

Option A

$0.01\,\mathrm{mol\,dm^{-3}\,s^{-1}}$

Question92

Rate law for the reaction $2NO + Cl_2 \rightarrow 2NOCl$ is rate $= k[NO]^2\,[Cl_2]$. When will the value of k increase?

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Options:

A. by increasing temperature

B. by increasing [NO]

C. by increasing $[Cl_2]$

D. by increasing both [NO] and $[Cl_2]$

Answer: A

Solution:

The rate constant k in the rate law rate $= k[\mathrm{NO}]^2[\mathrm{Cl}_2]$ is a parameter that is independent of the concentrations of the reactants. It is influenced by temperature and the presence of a catalyst but not by changes in reactant concentrations. Therefore, the value of k will increase by:

Increasing the temperature: According to the Arrhenius equation,

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

where A is the frequency factor, E_a is the activation energy, R is the universal gas constant, and T is the temperature in Kelvin. As temperature T increases, the exponential factor $e^{-\frac{E_a}{RT}}$ increases, which results in an increase in the rate constant k.

Therefore, the correct option is:

Option A: by increasing temperature

Question93

Calculate the rate constant of the first order reaction if 80% of the reactant decomposes in 60 minutes.

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Options:

A. 2.68×10^{-2} minute $^{-1}$

B. 5.36×10^{-2} minute $^{-1}$

C. 1.34×10^{-2} minute $^{-1}$

D. 8.1×10^{-2} minute $^{-1}$

Answer: A

Solution:

$$\begin{aligned} \mathrm{k} &= \frac{2.303}{\mathrm{t}} \log_{10} \frac{[\mathrm{A}]_0}{[\mathrm{A}]_{\mathrm{t}}} \\ &= \frac{2.303}{60} \log_{10} \frac{(100)}{(20)} = \frac{2.303}{60} \log_{10} \frac{10}{2} \\ &= \frac{2.303}{60} [\log_{10}(10) - \log_{10}(2)] \\ &= \frac{2.303}{60} \times 0.699 = 2.68 \times 10^{-2} \text{ minute }^{-1} \end{aligned}$$

Question94

What is the value of rate constant for first order reaction if slope for the graph of rate versus concentration is 2.5×10^{-3} ?

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Options:

A. 2.5×10^{-3} time^{-1}

B. 5.0×10^{-3} time^{-1}

C. 7.5×10^{-3} time^{-1}

D. 1.25×10^{-3} time^{-1}

Answer: A

Solution:

The rate of a first-order reaction is directly proportional to the concentration of one reactant. The mathematical expression for the rate of a first-order reaction is:

$k[A]$

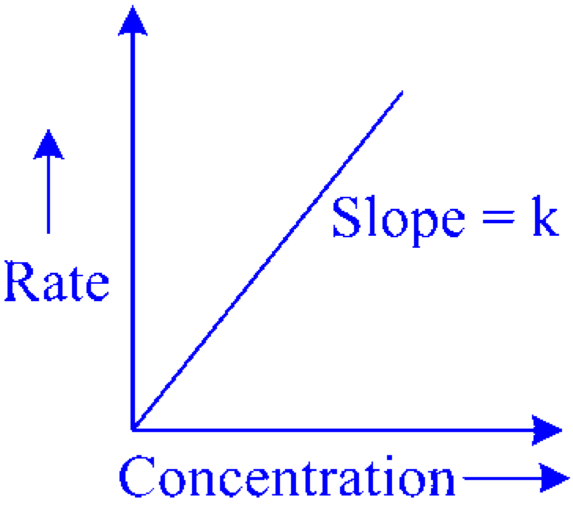
where:

- k is the rate constant
- $[A]$ is the concentration of the reactant.

The rate equation can be expressed as:

$\text{Rate} = k[A]$

Given that a graph of rate versus concentration, the slope of this graph for a first-order reaction would be equal to the rate constant, k, because the rate is directly proportional to the concentration $[A]$:



Slope $= k$

Therefore, if the slope for the graph of rate versus concentration is 2.5×10^{-3}, then the value of the rate constant for the first-order reaction is:

$k = 2.5 \times 10^{-3}$ time^{-1}

So, the correct answer is:

Option A2.5×10^{-3} time^{-1}

Question95

The rate law for the reaction $\mathrm{A} + \mathrm{B} \rightarrow$ product is rate $= \mathrm{k}[\mathrm{A}][\mathrm{B}]$. When will the rate of reaction increase by factor two?

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Options:

A. $[\mathrm{A}]$ and $[\mathrm{B}]$ both are doubled

B. $[\mathrm{A}]$ is doubled and $[\mathrm{B}]$ is kept constant

C. $[\mathrm{B}]$ is doubled and $[\mathrm{A}]$ is halved

D. [A] is kept constant [B] is halved

Answer: B

Solution:

$$\begin{aligned}
&\text{Rate } = \mathrm{k}[\mathrm{A}][\mathrm{B}] \\
&(\text{ Rate })_1 = \mathrm{k} \times 2[\mathrm{~A}][\mathrm{B}] \\
&\frac{(\text{ Rate })_1}{\text{Rate}} = \frac{\mathrm{k}2[\mathrm{~A}][\mathrm{B}]}{\mathrm{k}[\mathrm{A}][\mathrm{B}]} = 2 \\
&(\text{ Rate })_1 = 2 \times \text{ Rate}
\end{aligned}$$

Question96

Find the rate law for the reaction, $\mathrm{CHCl}_{3(\mathrm{~g})} + \mathrm{Cl}_{2(\mathrm{~g})} \rightarrow \mathrm{CCl}_{4(\mathrm{~g})} + \mathrm{HCl}_{(\mathrm{g})}$ if order of reaction with respect to $\mathrm{CHCl}_{\mathrm{a}(\mathrm{g})}$ is one and $\frac{1}{2}$ with $\mathrm{Cl}_{2(\mathrm{~g})}$.

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Options:

A. $\text{Rate} = \mathrm{k}\left[\mathrm{CHCl}_3\right]\left[\mathrm{Cl}_2\right]^{1/2}$

B. $\text{Rate} = \mathrm{k}\left[\mathrm{CHCl}_3\right]^2\left[\mathrm{Cl}_2\right]^{1/2}$

C. $\text{Rate} = \mathrm{k}\left[\mathrm{CHCl}_3\right]^{3/2}\left[\mathrm{Cl}_2\right]$

D. $\text{Rate} = \mathrm{k}\left[\mathrm{CHCl}_3\right]^{1/2}\left[\mathrm{Cl}_2\right]$

Answer: A

Solution:

The rate law for a reaction describes how the rate of the reaction depends on the concentration of the reactants. The rate law is determined experimentally and is generally given by the expression:

$\text{Rate} = k[\text{Reactant}_1]^m[\text{Reactant}_2]^n[\text{Reactant}_3]^p \ldots$

where k is the rate constant and m, n, p, etc., are the reaction orders with respect to each reactant.

For the given reaction,

$\mathrm{CHCl}_{3(\mathrm{g})} + \mathrm{Cl}_{2(\mathrm{g})} \rightarrow \mathrm{CCl}_{4(\mathrm{g})} + \mathrm{HCl}_{(\mathrm{g})}$, the order of reaction with respect to $CHCl_{3(g)}$ is one, and the order with respect to $Cl_{2(g)}$ is $\frac{1}{2}$. Therefore, the rate law can be expressed as:

$\text{Rate} = k[CHCl_3]^1[Cl_2]^{\frac{1}{2}}$.

Comparing this expression to the given options, the correct rate law corresponds to:

Option A:
$\text{Rate} = \text{k}\,[CHCl_3][Cl_2]^{1/2}$

Question97

The rate for reaction $2\text{ A} + \text{B} \rightarrow$ product is $6 \times 10^{-4}\text{ mol dm}^{-3}\text{ s}^{-1}$ Calculate the rate constant if the reaction is first order in A and zeroth order in B. [Given $[\text{A}] = [\text{B}] = 0.3\text{M}$]

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Options:

A. $1 \times 10^{-3}\text{ s}^{-1}$

B. $2 \times 10^{-3}\text{ s}^{-1}$

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

D. $4 \times 10^{-3}\text{ s}^{-1}$

Answer: B

Solution:

The rate law for a chemical reaction can be written based on the order of the reaction with respect to its reactants. For the given reaction:

$2\text{A} + \text{B} \rightarrow \text{product}$

and the information that it is first order in A (A) and zeroth order in B (B), the rate law is expressed as:

$\text{rate} = k[\text{A}]^1[\text{B}]^0$

Since the reaction is zeroth order in B, the concentration of B does not affect the rate and hence $[\text{B}]^0 = 1$. This simplifies the rate expression to:

$\text{rate} = k[\text{A}]^1$

The rate of the reaction is given as 6×10^{-4} mol dm^{-3} s^{-1} and the concentrations of A and B are both 0.3M (M = mol/dm³). Substituting the rate and the concentration of A into the simplified rate law gives us:

6×10^{-4} mol dm^{-3} $\mathrm{s}^{-1} = k(0.3\mathrm{M})^1$

Solving for the rate constant k:

$k = \frac{6 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}}{0.3\mathrm{M}}$

Performing the division:

$k = \frac{6 \times 10^{-4} \text{ s}^{-1}}{0.3} k = 2 \times 10^{-3} \text{ s}^{-1}$

Thus, the rate constant k for the reaction is $2 \times 10^{-3} \text{ s}^{-1}$, which corresponds to Option B.

Question98

Calculate half life of first order reaction if rate constant of reaction is $2.772 \times 10^{-3} \text{ s}^{-1}$

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Options:

A. 125 s

B. 250 s

C. 100 s

D. 150 s

Answer: B

Solution:

In the case of a first-order reaction, the half-life (which is the time taken for half of the reactant to be used up in the reaction) can be calculated using the following formula:

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

where $t_{1/2}$ is the half-life and k is the rate constant of the reaction.

Given that the rate constant $k = 2.772 \times 10^{-3} \text{ s}^{-1}$, we can plug this into our formula to find the half-life:

$t_{1/2} = \frac{0.693}{2.772 \times 10^{-3}\ \mathrm{s}^{-1}}$

To calculate this, first divide 0.693 by 2.772×10^{-3}:

$t_{1/2} = \frac{0.693}{2.772 \times 10^{-3}}$

$t_{1/2} = 250\ \mathrm{s}$

Therefore, the half-life of the reaction is 250 seconds. The correct option is:

Option B : 250 s

Question99

Which from following statements about rate constant is NOT true?

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Options:

A. It is independent of concentration.

B. It varies with temperature.

C. It is equal to rate of reaction at unit concentration of reactants.

D. It's unit is independent of order of reaction.

Answer: D

Solution:

Incorrect statement about rate constant is that it's unit is independent of order of reaction. It is because the unit of rate constant depends on the order of reaction. For nth order reaction, unit of rate constant is given by $\frac{\left[\mathrm{mol\ L}^{-1}\ \mathrm{s}^{-1}\right]}{\left[\mathrm{mol\ L}^{-1}\right]^{n}}$.

Question100

For the reaction,

$\mathrm{CH_3Br}(aq) + \mathrm{OH}^-(aq) \longrightarrow \mathrm{CH_3OH}(aq) + \mathrm{Br}^-(aq)$,

The rate law is rate $= k[\mathrm{CH_3Br}][\mathrm{OH}^-]$. What is change in rate of reaction if concentration of both reactants is doubled?

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Options:

A. Rate increases by factor 2

B. Rate increases by factor 4

C. Rate remains same

D. Rate decreases by factor 2

Answer: B

Solution:

Given, $\mathrm{CH_3Br}(aq) + \mathrm{OH}^-(aq) \longrightarrow \mathrm{CH_3OH}(aq) + \mathrm{Br}^-(aq)$

Rate $= k[2\mathrm{CH_3Br}][2\mathrm{OH}^-]$

When the concentration of both reactants are doubled, then new rate will be $= k'[2\mathrm{CH_3Br}][2\mathrm{OH}^-]$

Thus, the rate increases by factor 4.

Question101

For the reaction, $2A + 2B \longrightarrow 2C + D$, the rate law is expressed as rate $= k[A]^2[B]$.Calculate the rate constant if rate of reaction is $0.24\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$.

[$[A]$= $0.5\mathrm{M}$ and $[B] = 0.2\mathrm{M}$]

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Options:

A. $4.8 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

B. $9.6 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

C. $12.1 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

D. $14.4 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

Answer: A

Solution:

Given,

$$\begin{aligned}
& 2 A+2 B \longrightarrow 2 C+D \\
& \quad \text { Rate }=k[A]^2[B] \quad \text { ..... (i) } \\
& {[A]=0.5 \mathrm{M},[B]=0.2 \mathrm{M}}
\end{aligned}$$

Rate of reaction $=0.24 \mathrm{~mol} \mathrm{dm}^{-3} \mathrm{~s}^{-1}$

On substituting the given values in Eq. (i)

$0.24=k[0.5]^2[0.2] \Rightarrow k=4.8 \mathrm{~mol}^{-2} \mathrm{dm}^6 \mathrm{~s}^{-1}$

Question102

Find the rate of formation of $\mathrm{NO}_{2(\mathrm{~g})}$ in the following reaction.

$$2 \mathrm{~N}_2 \mathrm{O}_{5(\mathrm{~g})} \rightarrow 4 \mathrm{NO}_{2(\mathrm{~g})}+\mathrm{O}_{2(\mathrm{~g})}$$
$$\left[\frac{-\mathrm{d}\left[\mathrm{N}_2 \mathrm{O}_5\right]}{\mathrm{dt}}=0.02 \mathrm{~mol} \mathrm{dm}^{-3}\right]$$

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Options:

A. $0.01 \mathrm{~mol} \mathrm{dm}^{-3}$

B. $0.02 \mathrm{~mol} \mathrm{dm}^{-3}$

C. 0.03 mol dm^{-3}

D. 0.04 mol dm^{-3}

Answer: D

Solution:

$$\begin{aligned}
&\text{Rate of reaction } = -\frac{1}{2}\frac{d\left[N_2O_5\right]}{dt} \\
&= +\frac{1}{4}\frac{d\left[NO_2\right]}{dt} = \frac{d\left[O_2\right]}{dt} \\
&\text{Rate of formation of } NO_2 = \frac{d\left[NO_2\right]}{dt} \\
&= -\frac{4}{2}\frac{d\left[N_2O_5\right]}{dt} \\
&= -2\frac{d\left[N_2O_5\right]}{dt} \\
&= 2 \times 0.02 \\
&= 0.04 \text{ mol dm}^{-3}
\end{aligned}$$

Question103

Calculate the rate constant of the first order reaction if 20% of the reactant decomposes in 15 minutes.

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Options:

A. 1.488×10^{-2} minute $^{-1}$

B. 1.881×10^{-2} minute $^{-1}$

C. 1.984×10^{-2} minute $^{-1}$

D. 1.18×10^{-2} minute $^{-1}$

Answer: A

Solution:

20% of the reactant has decomposed.

So, if $[A]_0 = 100$, then $[A]_t = 100 - 20 = 80$

$$\begin{aligned} k &= \frac{2.303}{t} \log_{10} \frac{[A]_0}{[A]_t} \\ &= \frac{2.303}{15} \log_{10} \frac{100}{80} \\ &= \frac{2.303}{15} \log_{10} \frac{5}{4} \\ &= \frac{2.303}{15} \times (\log_{10} 5 - \log_{10} 4) \\ &= \frac{2.303}{15} \times (0.699 - 0.602) \\ &= 0.01488 = 1.488 \times 10^{-2} \text{ minute }^{-1} \end{aligned}$$

Question104

Which from following is the slope of the graph of $[A]_t$ versus time for zero order reaction?

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Options:

A. $-k$

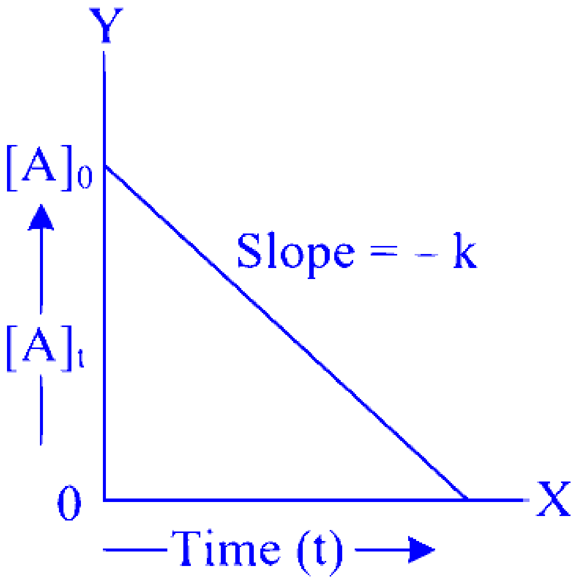
B. k

C. $\frac{k}{2.303}$

D. $\frac{-k}{2.303}$

Answer: A

Solution:



Variation of $[A]_t$ with time (t) for a zero order reaction

Question105

A first order reaction takes 23.03 minutes for 20% decomposition. Calculate its rate constant.

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Options:

A. 5.6×10^{-3} minute^{-1}

B. 4.5×10^{-3} minute^{-1}

C. 6.5×10^{-3} minute^{-1}

D. 9.69×10^{-3} minute^{-1}

Answer: D

Solution:

$[A]_0 = 100\%, [A]_t = 100 - 20 = 80\%$

Substitution of these in above

$$\begin{aligned}
\mathrm{k} &= \frac{2.303}{\mathrm{t}} \log_{10} \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}} \\
\mathrm{k} &= \frac{2.303}{\mathrm{t}} \log_{10} \frac{100}{80} \\
&= \frac{2.303}{23.03 \mathrm{~min}} \log_{10}(1.25) \\
&= \frac{2.303}{23.03 \mathrm{~min}} \times 0.0969 \\
&= 9.69 \times 10^{-3} \text{ minute }^{-1}
\end{aligned}$$

Question106

The rate law for the reaction $\mathrm{A} + \mathrm{B} \rightarrow \mathrm{C}$ at $25^{\circ}\mathrm{C}$ is given by rate $= k[A][B]^2$. Calculate the rate of reaction if rate constant at same temperature is $6.25 \mathrm{~mol}^{-2} \mathrm{~dm}^6 \mathrm{~s}^{-1}[[\mathrm{~A}] = 1\mathrm{M}, [\mathrm{B}] = 0.2\mathrm{M}]$

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Options:

A. $0.25 \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~s}^{-1}$

B. $0.5 \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~s}^{-1}$

C. $0.75 \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~s}^{-1}$

D. $1.25 \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~s}^{-1}$

Answer: A

Solution:

$$
\begin{aligned}
\text{Rate } &= k[A][B]^2 \\
&= 6.25 \text{ mol}^{-2} \text{ dm}^6 \text{ s}^{-1} \times 1 \text{ mol dm}^{-3} \\
&\quad \times 0.2 \text{ mol dm}^{-3} \times 0.2 \text{ mol dm}^{-3} \\
&= 0.25 \text{ mol dm}^{-3} \text{ s}^{-1}
\end{aligned}
$$

Question107

Calculate the time needed for reactant to decompose 99.9% if rate constant of first order reaction is 0.576 minute^{-1}.

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Options:

A. 8 minutes

B. 12 minutes

C. 16 minutes

D. 20 minutes

Answer: B

Solution:

99.9% of the reaction is complete.

So, if $[A]_0 = 100$, then $[A]_t = 100 - 99.9 = 0.1$

$$
\begin{aligned}
\mathrm{t} &= \frac{2.303}{\mathrm{k}} \log_{10} \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}} \\
&= \frac{2.303}{0.576} \log_{10} \frac{100}{0.1} = \frac{2.303}{0.576} \log_{10}(1000) \\
&= \frac{2.303}{0.576} \times 3 \\
&= 11.99 \approx 12 \text{ minutes}
\end{aligned}
$$

Question108

Calculate the rate constant of first order reaction if half life of reaction is 40 minutes.

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Options:

A. 1.733×10^{-2} minute $^{-1}$

B. 1.951×10^{-2} minute $^{-1}$

C. 1.423×10^{-2} minute $^{-1}$

D. 1.256×10^{-2} minute $^{-1}$

Answer: A

Solution:

For a first order reaction, $k=\frac{0.693}{t_{1/2}}$

$\mathrm{k}=\frac{0.693}{40}=1.733 \times 10^{-2}$ minute $^{-1}$

Question109

Identify rate law expression for $2\mathrm{NO}_{(\mathrm{g})}+\mathrm{Cl}_{2(\mathrm{~g})} \rightarrow 2\mathrm{NOCl}_{(\mathrm{g})}$ if the reaction is second order in NO and first order in Cl_2.

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Options:

A. Rate $=\mathrm{k}[\mathrm{NO}]^2\left[\mathrm{Cl}_2\right]$

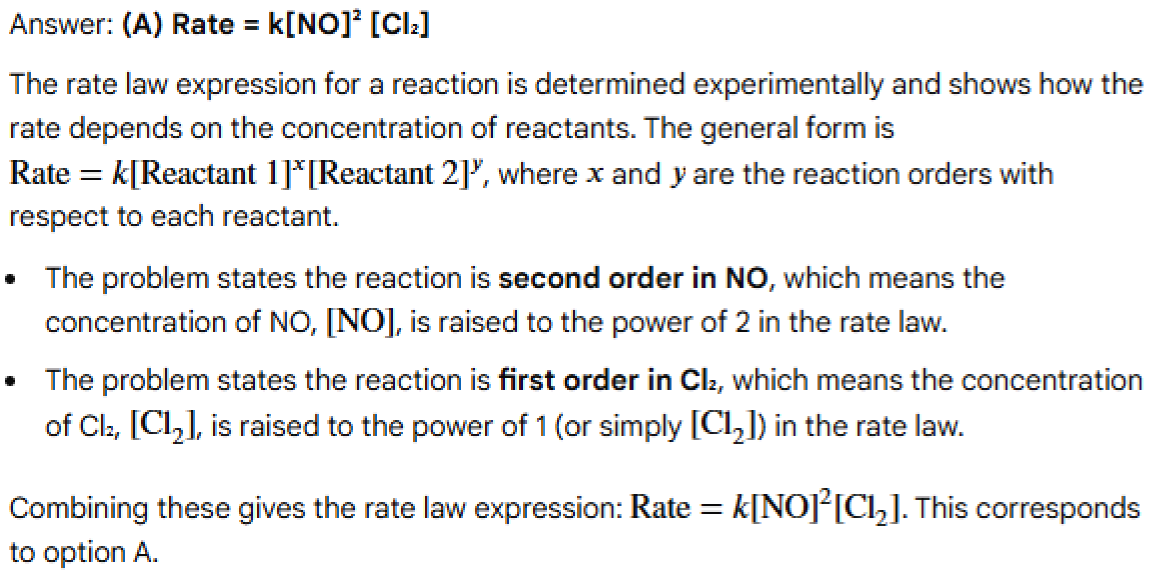
B. Rate $=\mathrm{k}[\mathrm{NO}]\left[\mathrm{Cl}_2\right]$

C. Rate $=\mathrm{k}[\mathrm{NO}]^2$

D. Rate $= \mathrm{k}\left[\mathrm{Cl}_2\right]$

Answer: A

Solution:



Answer: **(A) Rate = k[NO]² [Cl₂]**

The rate law expression for a reaction is determined experimentally and shows how the rate depends on the concentration of reactants. The general form is $\text{Rate} = k[\text{Reactant 1}]^x[\text{Reactant 2}]^y$, where x and y are the reaction orders with respect to each reactant.

- The problem states the reaction is **second order in NO**, which means the concentration of NO, $[\mathrm{NO}]$, is raised to the power of 2 in the rate law.
- The problem states the reaction is **first order in Cl₂**, which means the concentration of Cl₂, $[\mathrm{Cl}_2]$, is raised to the power of 1 (or simply $[\mathrm{Cl}_2]$) in the rate law.

Combining these gives the rate law expression: $\text{Rate} = k[\mathrm{NO}]^2[\mathrm{Cl}_2]$. This corresponds to option A.

Question110

For a reaction $\mathrm{A} + \mathrm{B} \rightarrow$ product, if $[\mathrm{A}]$ is doubled keeping $[\mathrm{B}]$ constant, the rate of reaction doubles. Calculate the order of reaction with respect to A.

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Options:

A. 0

B. 1/2

C. 1

D. 2

Answer: C

Solution:

$$\begin{aligned} & \mathrm{A}+\mathrm{B} \longrightarrow \text{ Product} \\ & \text{Rate } =\mathrm{k}[\mathrm{A}]^x[\mathrm{~B}]^{\mathrm{y}} \\ & \text{Rate }_1=\mathrm{k}[2 \mathrm{~A}]^x[\mathrm{~B}]^{\mathrm{y}} \\ \therefore \quad & \frac{\text{Rate }_1}{\text{Rate}}=\frac{\mathrm{k} 2^x[\mathrm{~A}]^x[\mathrm{~B}]^y}{\mathrm{k}[\mathrm{A}]^x[\mathrm{~B}]^{\mathrm{y}}} \end{aligned}$$

$$\therefore \quad \frac{2 \times \text{ Rate}}{\text{Rate}}=2^x$$

$$\therefore \quad 2^x=2$$

$$\therefore \quad x=1$$

$\therefore$ Order of reaction with respect to A = 1

Question111

Calculate the rate constant of the first order reaction if 80% of the reactant reacted in 15 minute.

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Options:

A. 0.11 minute^{-1}

B. 0.22 minute^{-1}

C. 0.34 minute^{-1}

D. 0.42 minute^{-1}

Answer: A

Solution:

80% of the reactant has reacted.

So, if $[A]_0=100$, then $[A]_{\mathrm{t}}=100-80=20$

$$\begin{aligned} k &= \frac{2.303}{t}\log_{10}\frac{[A]_0}{[A]_t} \\ &= \frac{2.303}{15}\log_{10}\frac{100}{20} \\ &= \frac{2.303}{15}\log_{10}(5) \\ &= \frac{2.303}{15} \times 0.699 \\ &= 0.1073 \\ &\approx 0.11 \text{ minute }^{-1} \end{aligned}$$

Question112

Identify the expression for average rate for following reaction.

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

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Options:

A. $\frac{-\Delta[N_2]}{\Delta t} = -\frac{1}{3}\frac{\Delta[H_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[NH_3]}{\Delta t}$

B. $-\frac{1}{3}\frac{\Delta[N_2]}{\Delta t} = \frac{\Delta[H_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[NH_3]}{\Delta t}$

C. $\frac{-\Delta[N_2]}{\Delta t} = \frac{-\Delta[H_2]}{\Delta t} = \frac{\Delta[NH_3]}{\Delta t}$

D. $-\frac{1}{2}\frac{\Delta[N_2]}{\Delta t} = \frac{-\Delta[H_2]}{\Delta t} = \frac{1}{3}\frac{\Delta[NH_3]}{\Delta t}$

Answer: A

Solution:

In the rate expression, always make sure to express the rates with the correct sign: Negative for the reactants and positive for the products.

Question113

For an elementary reaction

$2\,\mathrm{A}+\mathrm{B} \longrightarrow 3\mathrm{C}$

rate of appearance of C is $1.3 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}$, the rate of disappearance of A is:

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Options:

A. $1.3 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}$

B. $2.6 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}$

C. $5.2 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}$

D. $8.66 \times 10^{-5}\ \mathrm{mol\ L^{-1}\ s^{-1}}$

Answer: D

Solution:

$$\begin{aligned}
\text{Rate of reaction } &= -\frac{1}{2}\frac{d[A]}{dt}=-\frac{d[B]}{dt}=\frac{1}{3}\frac{d[C]}{dt}\\
\text{Rate of reaction } &= \frac{1}{3}\frac{d[C]}{dt}\\
&= \frac{1}{3} \times 1.3 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}\\
&= 0.433 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}
\end{aligned}$$

$\therefore$ Rate of disappearance of A

$$\begin{aligned}
&= 2 \times 0.433 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}\\
&= 0.866 \times 10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}\\
&= 8.66 \times 10^{-5}\ \mathrm{mol\ L^{-1}\ s^{-1}}
\end{aligned}$$

Question114

Slope of the graph between $\log \frac{[A]_0}{[A]_t}$ (y axis) and time (x axis) for first order reaction is equal to:

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Options:

A. $+\frac{\mathrm{k}}{2.303}$

B. k

C. $-\mathrm{k}$

D. $-\frac{2.303}{\mathrm{k}}$

Answer: A

Solution:

The integrated rate law for the first order reaction is

$k=\frac{2.303}{t} \log_{10} \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}$

$$\underset{\mathrm{y}}{\underset{\updownarrow}{\log_{10}\frac{[\mathrm{A}]_0}{[\mathrm{A}]_{\mathrm{t}}}}}=\underset{\mathrm{m}}{\underset{\updownarrow}{\frac{\mathrm{k}}{2.303}}}\,\underset{\mathrm{x}}{\underset{\updownarrow}{\mathrm{t}}}$$

The graph of $\log_{10} \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}$ versus time (t) is a straight line passing through origin with slope $(\mathrm{m})=+\frac{\mathrm{k}}{2.303}$.

Question115

What is half life time of a first order reaction if initial conc. of reactant is $0.01 \mathrm{~mol} \mathrm{~L}^{-1}$ and rate of reaction is $0.00352 \mathrm{~mol} \mathrm{~L}^{-1}$ minute $^{-1}$?

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Options:

A. 1.969 minute

B. 7.75 minute

C. 16.69 minute

D. 19.69 minute

Answer: A

Solution:

For a first order reaction, rate $= \mathrm{k}[\mathrm{A}]$

$\therefore \quad 0.00352 \mathrm{~mol} \mathrm{~L}^{-1} \text{ minute}^{-1} = \mathrm{k} \times 0.01 \mathrm{~mol} \mathrm{~L}^{-1}$

$\mathrm{k} = 0.352 \text{ minute }^{-1}$

Half-life period, $\left(\mathrm{t}_{1/2}\right) = \frac{0.693}{\mathrm{k}} = \frac{0.693}{0.352} = 1.969$ minute

---

Question116

Find the average rate of formation of $\mathrm{NO}_{2(\mathrm{~g})}$, in following reaction.

$$2 \mathrm{~N}_2\mathrm{O}_{5(\mathrm{~g})} \rightarrow 4\mathrm{NO}_{2(\mathrm{~g})} + \mathrm{O}_{2(\mathrm{~g})}$$
$$\left[-\frac{\Delta\left[\mathrm{N}_2\mathrm{O}_5\right]}{\Delta \mathrm{t}}\right] = x \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~s}^{-1}$$

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Options:

A. $x \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~s}^{-1}$

B. $\frac{x}{2} \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~s}^{-1}$

C. $2x$ mol $\mathrm{dm}^{-3}\ \mathrm{s}^{-1}$

D. $4x$ mol $\mathrm{dm}^{-3}\ \mathrm{s}^{-1}$

Answer: C

Solution:

$$\frac{-\Delta\left[\mathrm{N}_2\mathrm{O}_5\right]}{\Delta \mathrm{t}}=x \text{ moldm}^{-3}\ \mathrm{s}^{-1}$$
$$2\ \mathrm{N}_2\mathrm{O}_{5(\mathrm{~g})} \rightarrow 4\mathrm{NO}_{2(\mathrm{~g})}+\mathrm{O}_{2(\mathrm{~g})}$$

Average rate of reaction

$$\begin{aligned} & =-\frac{1}{2} \frac{\Delta\left[\mathrm{N}_2\mathrm{O}_5\right]}{\Delta \mathrm{t}}=\frac{1}{4} \frac{\Delta\left[\mathrm{NO}_2\right]}{\Delta \mathrm{t}}=\frac{\Delta\left[\mathrm{O}_2\right]}{\Delta \mathrm{t}} \\ \therefore \quad & \frac{\Delta\left[\mathrm{NO}_2\right]}{\Delta \mathrm{t}}=-\frac{4}{2} \frac{\Delta\left[\mathrm{N}_2\mathrm{O}_5\right]}{\Delta \mathrm{t}}=\frac{4}{2} \times x \\ & =2x \text{ moldm}^{-3}\ \mathrm{s}^{-1}\end{aligned}$$

$\therefore$ Average rate of formation of $\mathrm{NO}_{2(\mathrm{~g})}=2x$ mol $\mathrm{dm}^{-3}\ \mathrm{s}^{-1}$

[Note: In the question, $\frac{\Delta\left[\mathrm{N}_2\mathrm{O}_5\right]}{\Delta \mathrm{t}}$ is changed to $\frac{-\Delta\left[\mathrm{N}_2\mathrm{O}_5\right]}{\Delta \mathrm{t}}$ to apply appropriate textual concepts.]

Question117

Calculate the rate constant for the first order reaction, $\mathrm{A} \rightarrow \mathrm{B}$ if the rate of reaction is 5.4×10^{-6} mol $\mathrm{dm}^{-3}\ \mathrm{s}^{-1}$ and $[\mathrm{A}]=0.3\mathrm{M}$.

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Options:

A. $1.8 \times 10^{-5}\ \mathrm{s}^{-1}$

B. $1.5 \times 10^{-5}\ \mathrm{s}^{-1}$

C. $2.1 \times 10^{-5}\ \mathrm{s}^{-1}$

D. $2.4 \times 10^{-5}\ \mathrm{s}^{-1}$

Answer: A

Solution:

For the first order reaction, A $\longrightarrow$ B,

$$\begin{aligned}\text{Rate} &= \mathrm{k[A]}\\ \therefore \quad \mathrm{k} = \frac{\text{Rate}}{[\mathrm{A}]} &= \frac{5.4\times10^{-6}\ \mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}}{0.3\ \mathrm{mol\ dm}^{-3}}\\ &= 1.8\times10^{-5}\ \mathrm{s}^{-1}\end{aligned}$$

Question118

Time required for 90% completion of a first order reaction is 'x' minute. Calculate the time required to complete 99.9% of the reaction at same temperature.

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Options:

A. x minute

B. $2x$ minute

C. $3x$ minute

D. $\frac{x}{2}$ minute

Answer: C

Solution:

$$\begin{aligned}t_{90\%6} &= \frac{2.303}{k}\log_{10}\frac{[A]_0}{[A]_t} = \frac{2.303}{k}\log_{10}\frac{100}{10}\\ &= \frac{2.303}{k}\log_{10}10\end{aligned}$$

$$t_{99.9\%} = \frac{2.303}{k}\log_{10}\frac{[A]_0}{[A]_t} = \frac{2.303}{k}\log_{10}\frac{100}{0.1}$$

$$= \frac{2.303}{k}\log_{10} 1000$$

$$\frac{t_{999\%}}{t_{90\%}} = \frac{\frac{2.303}{k}\log_{10} 1000}{\frac{2.303}{k}\log_{10} 10} = \frac{\log_{10} 1000}{\log_{10} 10} = \frac{3}{1}$$

$$\therefore \quad t_{99.9\%} = 3 \times t_{90\%} = 3x \text{ minute} \qquad (\text{since, } t_{90\%} = x \text{ minute})$$

Question119

For the reaction, $3\,\mathrm{I} + \mathrm{S_2O_8^{2-}} \rightarrow \mathrm{I_3^-} + 2\mathrm{SO_4^{2-}}$, at a particular time t, $\frac{\mathrm{d}\left[\mathrm{SO_4^{2-}}\right]}{\mathrm{dt}}$ is $2.2 \times 10^{-2}\ \mathrm{mol\ dm^{-3}\ s^{-1}}$. What is the value of $-\frac{\mathrm{d}[\mathrm{I^-}]}{\mathrm{dt}}$?

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Options:

A. $1.1 \times 10^{-2}\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

B. $3.3 \times 10^{-2}\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

C. $4.4 \times 10^{-2}\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

D. $6.6 \times 10^{-2}\ \mathrm{mol\ dm^{-3}\ s^{-1}}$

Answer: B

Solution:

$3I + \mathrm{S_2O_8^{2-}} \rightarrow \mathrm{I_3^-} + 2\mathrm{SO_4^{2-}}$

$$\text{Rate of reaction } = -\frac{1}{3}\frac{\mathrm{d}\left[\mathrm{I^-}\right]}{\mathrm{dt}} = +\frac{1}{2}\frac{\mathrm{d}\left[\mathrm{SO_4^{2-}}\right]}{\mathrm{dt}}$$

$$-\frac{\mathrm{d}\left[\mathrm{I^-}\right]}{\mathrm{dt}} = \frac{3\,\mathrm{d}\left[\mathrm{SO_4^{2-}}\right]}{2} = \frac{3}{2} \times 2.2 \times 10^{-2}$$

$$= 3.3 \times 10^{-2}\ \mathrm{mol\ dm\ s^{-3}}$$

[Note: In the question, $\frac{\mathrm{d}[\mathrm{I^-}]}{\mathrm{dt}}$ is changed to $-\frac{\mathrm{d}[\mathrm{I^-}]}{\mathrm{dt}}$ to apply appropriate textual concepts.]

Question120

What is the half life of a first order reaction if rate constant is 4.2×10^{-2} per day?

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Options:

A. 5.0 day

B. 16.5 day

C. 28.0 day

D. 9.0 day

Answer: B

Solution:

For a first order reaction,

$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{4.2 \times 10^{-2}} = 16.5$ days

---

Question121

The reaction, $3ClO^- \rightarrow ClO_3^- + 2Cl^-$ occurs in two steps:

i. $2ClO^- \rightarrow ClO_2^- + Cl^-$

ii. $ClO_2^- + ClO^- \rightarrow ClO_3^- + Cl^-$,

the reaction intermediate is:

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Options:

A. Cl^{-}

B. ClO_2^{-}

C. ClO_3^{-}

D. ClO^{-}

Answer: B

Solution:

ClO_2^{-} is produced in step (i) and consumed in step (ii). Therefore, ClO_2^{-} is the reaction intermediate.

Question122

Calculate the rate constant of first order reaction if the concentration of the reactant decreases by 90% in 30 minutes.

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Options:

A. 7.7×10^{-2} minute $^{-1}$

B. 4.2×10^{-2} minute $^{-1}$

C. 2.1×10^{-2} minute $^{-1}$

D. 3.5×10^{-2} minute $^{-1}$

Answer: A

Solution:

Concentration decreases by 90%. Hence, 10% reactant is left after 30 minutes.

$$\begin{aligned} k &= \frac{2.303}{t}\log_{10}\frac{[A]_0}{[A]_t} \\ k &= \frac{2.303}{30}\log_{10}\frac{100}{10} \\ &= \frac{2.303}{30\ \text{min}} \\ &= 7.7\times10^{-2}\ \text{min}^{-1} \end{aligned}$$

Question123

The rate law for the reaction $\mathrm{A}+\mathrm{B}\rightarrow$ product is given by rate $=k[A][B]$ Calculate $[A]$ if rate of reaction and rate constant are $0.25\ \mathrm{moldm}^{-3}\ \mathrm{s}^{-1}$ and $6.25\ \mathrm{mol}^{-1}\mathrm{dm}^{3}\ \mathrm{s}^{-1}$ respectively $\left[[\mathrm{B}]=0.25\ \mathrm{moldm}^{-3}\right]$

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Options:

A. $0.22\ \mathrm{mol\ dm}^{3}$

B. $0.16\ \mathrm{mol\ dm}^{3}$

C. $0.30\ \mathrm{mol\ dm}^{3}$

D. $0.25\ \mathrm{mol\ dm}^{3}$

Answer: B

Solution:

$$\begin{aligned} &\text{rate } = \mathrm{k}[\mathrm{A}][\mathrm{B}] \\ \therefore\quad &[\mathrm{A}] = \frac{\text{rate}}{\mathrm{k}[\mathrm{B}]} = \frac{0.25\ \mathrm{moldm}^{-3}\ \mathrm{s}^{-1}}{6.25\ \mathrm{mol}^{-1}\mathrm{dm}^{3}\cdot \mathrm{s}^{-1}\times 0.25\ \mathrm{moldm}^{-3}} \\ &\qquad = 0.16\ \mathrm{mol\ dm}^{3} \end{aligned}$$

Question124

Find the average rate of formation $O_{2(g)}$ in the following reaction.

$$2NO_{2(g)} \rightarrow 2NO_{(g)} + O_{2(g)}$$
$$\left[-\frac{\Delta[NO]}{\Delta t}\right] = x \text{ moldm}^{-3}\text{ s}^{-1}$$

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Options:

A. $\frac{x}{2}$ mol dm^{-3} s^{-1}

B. x mol dm^{-3} s^{-1}

C. $2x$ mol dm^{-3} s^{-1}

D. $4x$ mol dm^{-3} s^{-1}

Answer: A

Solution:

Average rate of reaction

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

Rate of formation of $O_2 = \frac{\Delta[O_2]}{\Delta t}$

$$= \frac{1}{2}\frac{-\Delta[NO_2]}{\Delta t}$$
$$= \frac{1}{2} \times x = \frac{x}{2}\text{ mol dm}^{-3}\text{ s}^{-1}$$

Question125

Which from following is the slope of the graph of rate versus concentration of the reactant for first order reaction?

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Options:

A. $-\mathrm{k}$

B. k

C. $\frac{\mathrm{k}}{2.303}$

D. $\frac{-\mathrm{k}}{2.303}$

Answer: B

Solution:

For a first-order reaction, the rate of the reaction is directly proportional to the concentration of the reactant. The rate law for a first-order reaction can be written as :

$\text{Rate} = k[\text{Reactant}]$

where k is the rate constant.

The graph of rate versus concentration of the reactant for a first-order reaction would be a straight line with a slope equal to the rate constant k. This is because the rate is directly proportional to the concentration, making the slope of the line the rate constant itself.

Therefore, the correct answer is :

Option B : k

Question126

Calculate the amount of reactant in percent that remains after 60 minutes involved in first order reaction. $(\mathrm{k} = 0.02303 \text{ minute }^{-1})$

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Options:

A. 25%

B. 75%

C. 50%

D. 12.5%

Answer: A

Solution:

For a first order reaction,

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

$$t_{1/2}=\frac{0.693}{0.02303}=30\text{ min}$$

Percent of reactant that remains after $\mathrm{t}_{1/2}=50\%$

$2\times\mathrm{t}_{1/2}=60\text{ min}$

Therefore, percent of reactant that remains after $2t_{1/2}=25\%$

Question127

The rate for reaction $\mathrm{A}+\mathrm{B}\rightarrow$ product, is $1.8\times10^{-2}\text{ mol dm}^{-3}\text{ s}^{-1}$. Calculate the rate constant if the reaction is second order in A and first order in B. ($[\mathrm{A}]=0.2\mathrm{M}$; $[\mathrm{B}]=0.1\mathrm{M}$)

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Options:

A. $9.0\text{ mol}^{-2}\text{ dm}^{6}\text{ s}^{-1}$

B. $18.0\text{ mol}^{-2}\text{ dm}^{6}\text{ s}^{-1}$

C. $4.5\text{ mol}^{-2}\text{ dm}^{6}\text{ s}^{-1}$

D. $16.0\text{ mol}^{-2}\text{ dm}^{6}\text{ s}^{-1}$

Answer: C

Solution:

$$\text{Rate} = k[A]^2[B]$$

$$\therefore \quad k = \frac{\text{rate}}{[A]^2[B]}$$

$$k = \frac{1.8 \times 10^{-2} \text{mol dm}^{-3}\ \text{s}^{-1}}{\left(0.2 \text{ mol dm}^{-3}\right)^2 \times 0.1 \text{mol dm}^{-3}}$$

$$= 4.5 \text{ mol}^{-2} \text{dm}^6\ \text{s}^{-1}$$

Question128

For the reaction $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$, rate of disappearance of $N_{2(g)}$ is 2.22×10^{-3} mol dm^{-3}. What is the rate of appearance of $NH_{3(g)}$?

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Options:

A. 2.22×10^{-3} mol dm^{-3}

B. 1.11×10^{-3} mol dm^{-3}

C. 4.44×10^{-3} mol dm^{-3}

D. 3.33×10^{-3} mol dm^{-3}

Answer: C

Solution:

$$\frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt};$$

$$\therefore \frac{d[NH_3]}{dt} = 2\frac{d[N_2]}{dt}$$

$$\frac{d[NH_3]}{dt} = 2 \times 2.22 \times 10^{-3} = 4.44 \times 10^{-3}$$

Question129

Find the rate constant of first order reaction in second having half life of 2.5 hours.

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Options:

A. $4.3 \times 10^{-5} \mathrm{sec}^{-1}$

B. $7.7 \times 10^{-5} \mathrm{sec}^{-1}$

C. $6.9 \times 10^{-5} \mathrm{sec}^{-1}$

D. $8.4 \times 10^{-5} \mathrm{sec}^{-1}$

Answer: B

Solution:

$\lambda=\frac{0.693}{t_{1/2}}=\frac{0.693}{2.5 \times 60 \times 60}=7.7 \times 10^{-5} \mathrm{sec}^{-1}$

Question130

If rate of reaction is given as

$$\frac{1}{3} \frac{\mathrm{d}[\mathrm{x}]}{\mathrm{dt}}=-\frac{1}{2} \frac{\mathrm{d}[\mathrm{y}]}{\mathrm{dt}}=-\frac{\mathrm{d}[\mathrm{Z}]}{\mathrm{dt}},$$

the reaction can be represented as

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Options:

A. $2\mathrm{y}+Z \rightarrow 3\mathrm{x}$

B. $2\mathrm{y} \rightarrow 3\mathrm{x}+\mathrm{Z}$

C. $3x+2y \rightarrow Z$

D. $3x \rightarrow 2y+Z$

Answer: A

Solution:

The given rate of reaction is:

$$\frac{1}{3}\frac{\mathrm{d}[\mathrm{x}]}{\mathrm{dt}}=-\frac{1}{2}\frac{\mathrm{d}[\mathrm{y}]}{\mathrm{dt}}=-\frac{\mathrm{d}[\mathrm{Z}]}{\mathrm{dt}}$$

The negative sign in front of the derivatives of concentrations of Y and Z suggests these are reactants being consumed over time, while the positive sign for substance X indicates that it is a product being formed over time.

To relate the reaction rates of individual reactants and products to a balanced chemical equation, you have to equalize the rate of disappearance of the reactants with the rate of appearance of the products, taking into consideration their stoichiometric coefficients.

Using the stoichiometry of the balanced equation, we can understand that the coefficients of X, Y, and Z in the reaction equation would correspond to the ratios of the rates at which they appear or disappear.

For this, the coefficients of Y and Z can be taken as 2 and 1, respectively, which would mean that for every 1 mole of Z consumed, 2 moles of Y are consumed, as indicated by their rates being two times faster.

As per the given rate relationship, X is being formed thrice as fast as Z is being consumed and simultaneously 1.5 times (= 3/2) as fast as Y is being consumed. Adding these stoichiometric coefficients into the equation, you can balance them according to the reaction:

$2\mathrm{Y}+\mathrm{Z} \rightarrow 3\mathrm{X}$

Therefore, the correct representation of the reaction with stoichiometric coefficients is option A:

$2\mathrm{y}+\mathrm{Z} \rightarrow 3\mathrm{x}$

This is the only equation that correlates to the provided rate of reaction change for each compound, ensuring that the Law of Conservation of Mass is upheld, with reactants being converted into products at the rates described in the equation.

Question131

In a first reaction 60% of reactant decomposes in 4.606 min. What is half life of reaction? ($\mathbf{k = 0.1989\ min^{-1}}$)

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Options:

A. 3.48 min

B. 2.4 min

C. 3.0 min

D. 1.74 min

Answer: A

Solution:

To determine the half-life of the reaction, given the rate constant k and the percentage of reactant decomposed, we need to understand the relationship between these variables. The rate constant provided suggests that the reaction follows first-order kinetics.

For a first-order reaction, the relationship between the rate constant k and the half-life $t_{1/2}$ is given by:

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

Given the rate constant $k = 0.1989 \text{ min}^{-1}$, we can substitute this value into the equation to find the half-life:

$t_{1/2} = \frac{0.693}{0.1989}$

Let's perform the calculation:

$t_{1/2} = \frac{0.693}{0.1989} \approx 3.48 \text{ min}$

Therefore, the half-life of the reaction is approximately 3.48 minutes. Hence, the correct option is:

Option A: 3.48 min

Question132

For the reaction, $2 \mathrm{~A}+\mathrm{B} \rightarrow 2 \mathrm{C}$, rate of disappearance of A is $0.076 \mathrm{~mol~s}^{-1}$. What is the rate of disappearance of B ?

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Options:

A. $0.076 \mathrm{~mol~s}^{-1}$

B. $0.038 \mathrm{~mol~s}^{-1}$

C. $0.019 \mathrm{~mol~s}^{-1}$

D. $0.095 \mathrm{~mol~s}^{-1}$

Answer: B

Solution:

$$
\begin{aligned}
& 2 \mathrm{~A}+\mathrm{B} \longrightarrow 2 \mathrm{C} \\
& \quad \text { Rate }=-\frac{1}{2} \frac{\mathrm{d}[\mathrm{A}]}{\mathrm{dt}}=-\frac{\mathrm{d}[\mathrm{B}]}{\mathrm{dt}}=\frac{1}{2} \frac{\mathrm{d}[\mathrm{C}]}{\mathrm{dt}} \\
& \therefore \frac{\mathrm{d}[\mathrm{B}]}{\mathrm{dt}}=\frac{1}{2} \frac{\mathrm{d}[\mathrm{A}]}{\mathrm{dt}} \\
& \qquad=\frac{1}{2} \times 0.076 \mathrm{~mol~s}^{-1}=0.038 \mathrm{~mol~s}^{-1}
\end{aligned}
$$

Question133

Half-life and rate constant for first order reaction are related by equation,

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Options:

A. $t_{1/2}=\frac{k}{0.693}$

B. $k=\frac{[A]_0-[A]_1}{t_{1/2}}$

C. $t_{1/2}=\frac{0.693}{k}$

D. $t_{1/2}=\frac{[\mathrm{A}]_1-[\mathrm{A}]_0}{k}$

Answer: C

Solution:

For a first-order reaction, the relationship between the half-life ($t_{1/2}$) and the rate constant (k) is given by:

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

This relationship is derived from the integrated rate law for a first-order reaction and the definition of half-life (the time it takes for the concentration of the reactant to decrease to half of its initial value). None of the other options correctly express this relationship for a first-order reaction.

Therefore, the correct option is:

Option C

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

Question134

If decomposition of hydrogen peroxide is a first order reaction, it's rate law equation can be represented as

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Options:

A. $\mathrm{r} = \frac{\mathrm{k}}{[\mathrm{H}_2\mathrm{O}_2]}$

B. $\mathrm{r} = \mathrm{k}\left[\mathrm{H}_2\mathrm{O}_2\right]$

C. $\mathrm{r} = \mathrm{k}\frac{[\mathrm{H}_2\mathrm{O}][\mathrm{O}_2]^{1/2}}{[\mathrm{H}_2\mathrm{O}_2]}$

D. $\mathrm{r} = \mathrm{k}\frac{[\mathrm{H}_2\mathrm{O}_2]}{[\mathrm{H}_2\mathrm{O}_2][\mathrm{O}_2]^{1/2}}$

Answer: B

Solution:

$\mathrm{H}_2\mathrm{O}_{2(\mathrm{~g})} \longrightarrow \mathrm{H}_2\mathrm{O}_{(\mathrm{l})} + \frac{1}{2}\mathrm{O}_{2(\mathrm{~g})}$

It is first order reaction.

$\therefore \mathrm{r} = \mathrm{k}\left[\mathrm{H}_2\mathrm{O}_2\right]$

Question135

A reaction is first order with respective to A and second order with respective to B. What is the effect on reaction rate if concentration of B is increased 3 times?

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Options:

A. Rate increases 6 times

B. Rate increases 2 times

C. Rate increases 9 times

D. Rate increases 3 times

Answer: C

Solution:

To determine the effect on the reaction rate when the concentration of B is increased 3 times, we need to consider the rate law for the reaction. A reaction that is first order with respect to A and second order with respect to B has a rate law given by:

$\text{Rate} = k[\text{A}]^1[\text{B}]^2$

Here, k is the rate constant, $[\text{A}]$ is the concentration of reactant A, and $[\text{B}]$ is the concentration of reactant B.

According to the problem, the concentration of B is increased 3 times. Let's denote the initial concentration of B as $[\text{B}]_0$. After increasing the concentration, the new concentration of B becomes:

$[\text{B}]_{\text{new}} = 3[\text{B}]_0$

Now we substitute this new concentration into the rate law:

$\text{Rate}_{\text{new}} = k[\text{A}]_0(3[\text{B}]_0)^2$

Simplifying the expression, we get:

$\text{Rate}_{\text{new}} = k[\text{A}]_0 \cdot 9[\text{B}]_0^2$

Therefore:

$\text{Rate}_{\text{new}} = 9 \cdot k[\text{A}]_0[\text{B}]_0^2$

This shows that the new rate is 9 times the initial reaction rate. Thus, the reaction rate increases by a factor of 9 when the concentration of B is increased 3 times.

Therefore, the correct answer is:

Option C Rate increases 9 times

Question136

For the reaction $\mathrm{A}+\mathrm{B} \rightarrow$ product, rate of reaction is $3.6 \times 10^{-2} \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{sec}^{-1}$. When $[\mathrm{A}]=0.2 \mathrm{~moldm}^{-3}$ and $[\mathrm{B}]=0.1 \mathrm{~mol} \mathrm{~dm}^{-3}$, find rate constant of reaction if it is second order with respective to both reactants.

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Options:

A. $18 \mathrm{~mol}^{-3} \mathrm{dm}^9 \mathrm{sec}^{-1}$

B. $90 \mathrm{~mol}^{-3} \mathrm{dm}^9 \mathrm{sec}^{-1}$

C. $72 \mathrm{~mol}^{-3} \mathrm{dm}^9 \mathrm{sec}^{-1}$

D. $36 \mathrm{~mol}^{-3} \mathrm{~dm}^9 \mathrm{sec}^{-1}$

Answer: B

Solution:

$$
\begin{aligned}
& \mathrm{A}+\mathrm{B} \longrightarrow \text { Products } \\
& \mathrm{r}=\mathrm{k}[\mathrm{A}]^2[\mathrm{~B}]^2 \\
& \therefore \mathrm{k}=\frac{\mathrm{r}}{[\mathrm{A}]^2[\mathrm{~B}]^2} \\
& =\frac{3.6 \times 10^{-2} \mathrm{moldm}^{-3} \mathrm{sec}^{-1}}{[0.2]^2 \mathrm{~mol}^2 \mathrm{dm}^{-6} \times[0.1]^2 \mathrm{~mol}^2 \mathrm{dm}^{-6}} \\
& =\frac{3.6 \times 10^{-2}}{0.04 \times 0.01} \mathrm{~mol}^{-3} \mathrm{dm}^9 \mathrm{sec}^{-1}
\end{aligned}
$$

$= 90\ \mathrm{mol}^{-3}\mathrm{dm}^{9}\mathrm{sec}^{-1}$

Question137

Which of the following equations represents integrated rate law for zero order reaction?

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Options:

A. $k=\frac{[\mathrm{A}]_{\mathrm{t}}-[\mathrm{A}]_0}{\mathrm{t}}$

B. $\mathrm{k}=\frac{1}{\mathrm{t}} \log _{10} \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}$

C. $k=\frac{[\mathrm{A}]_0-[\mathrm{A}]_1}{\mathrm{t}}$

D. $k=\frac{t}{2.303} \times \log _{10} \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}$

Answer: C

Solution:

Explanation

For a zero-order reaction, the rate of the reaction is constant and does not depend on the concentration of the reactant:

$$\text{Rate} = k$$

The differential rate law is given by:

$$\text{Rate} = -\frac{d[\text{A}]}{dt}$$

Combining these two equations gives:

$$-\frac{d[\text{A}]}{dt} = k$$

Rearranging and integrating from time $t = 0$ (concentration $[\text{A}]_0$) to time t (concentration $[\text{A}]_t$):

$$-\int_{[\text{A}]_0}^{[\text{A}]_t} d[\text{A}] = \int_0^t k\,dt$$

$$-([\text{A}]_t - [\text{A}]_0) = k(t-0)$$

$$[\text{A}]_0 - [\text{A}]_t = kt$$

Solving for k gives the integrated rate law for a zero-order reaction:

$$k = \frac{[\text{A}]_0 - [\text{A}]_t}{t}$$

Question138

Ammonia and oxygen react at high temperature as

$$4\mathrm{NH}_{3(\,g)} + 5\mathrm{O}_{2(\,g)} \longrightarrow 4\mathrm{NO}_{(g)} + 6\mathrm{H}_2\mathrm{O}_{(g)}.$$

If rate of formation of $\mathrm{NO}_{(g)}$ is $3.6 \times 10^{-3}\ \mathrm{mol\ L^{-1}\ s^{-1}}$ then rate of disappearance of ammonia is

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Options:

A. $7.2 \times 10^{-3}\ \mathrm{mol\ L^{-1}\ s^{-1}}$

B. $1.2 \times 10^{-3}\ \mathrm{mol\ L^{-1}\ s^{-1}}$

C. $2.4 \times 10^{-3} \mathrm{~mol~L}^{-1} \mathrm{~s}^{-1}$

D. $3.6 \times 10^{-3} \mathrm{~mol~L}^{-1} \mathrm{~s}^{-1}$

Answer: D

Solution:

$\text{Rate of reaction} = -\frac{1}{4} \frac{d[NH_3]}{dt} = -\frac{1}{5} \frac{d[O_2]}{dt} = \frac{1}{4} \frac{d[NO]}{dt} = \frac{1}{6} \frac{d[H_2O]}{dt}$

$\therefore$ Rate of disappearance of NH_3 = Rate of formation of NO

$= 3.6 \times 10^{-3} \mathrm{~mol~L}^{-1} \mathrm{~s}^{-1}$

Question139

Which of the following represents integrated rate law equation for gas phase first order reaction, $A_{(g)} \rightarrow B_{(g)} + C_{(g)}$

if P_i = initial pressure of A

P = total pressure of reaction mixture at time ?

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Options:

A. $k = 2.303 \times \log_{10} \frac{P_i}{2P_i - P}$

B. $k = \frac{2.303}{t} \times \log_{10} \frac{P_i}{2P_i - P}$

C. $k = \frac{1}{t} \ln \frac{2P_i - P}{P_i}$

D. $k = \frac{2.303}{t} \times \log_{10} \frac{P_i - P}{P_i}$

Answer: B

Solution:

Answer: (B) $k = \frac{2.303}{t} \times \log_{10} \frac{P_i}{2P_i - P}$

The reaction is $A_{(g)} \rightarrow B_{(g)} + C_{(g)}$. The initial pressure of A is P_i. At time t, let the pressure of A that has reacted be x.

The pressures at time t are:

$$\begin{array}{l} A_{(g)} \rightarrow B_{(g)} + C_{(g)} \\ \text{Initial: } P_i \quad 0 \quad 0 \\ \text{At time } t\text{: } P_i - x \quad x \quad x \end{array}$$

The total pressure P at time t is the sum of the partial pressures:

$$P = (P_i - x) + x + x = P_i + x$$

So, $x = P - P_i$.

The pressure of A at time t is $P_A = P_i - x = P_i - (P - P_i) = 2P_i - P$.

For a first-order reaction, the integrated rate law is $k = \frac{2.303}{t} \log_{10} \frac{[A]_0}{[A]_t}$, which in terms of pressures is $k = \frac{2.303}{t} \log_{10} \frac{P_i}{P_A}$.

Substituting $P_A = 2P_i - P$ into the equation gives:

$$k = \frac{2.303}{t} \times \log_{10} \frac{P_i}{2P_i - P}$$

Question140

The rate law equation for a reaction between A, B and C is $\mathrm{r} = \mathrm{k}[\mathrm{A}][\mathrm{B}][\mathrm{C}]^2$, what will be ne rate of reaction if concentration of both A and B are doubled.

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Options:

A. 2r

B. 4r

C. 6r

D. 8r

Answer: B

Solution:

To determine the new rate of the reaction when the concentrations of both A and B are doubled, we start with the given rate law equation:

$r = k[A][B][C]^2$

Now, let's say initially, the concentrations of A, B, and C are $[A]$, $[B]$, and $[C]$ respectively. If the concentrations of A and B are doubled, the new concentrations will be:

$2[A]$ and $2[B]$.

We substitute these new concentrations back into the rate law equation to find the new rate r_{new}:

$r_{new} = k[(2[A])][(2[B])][C]^2$

Now we simplify this expression:

$r_{new} = k \cdot 2[A] \cdot 2[B] \cdot [C]^2$

$r_{new} = 4\left(k[A][B][C]^2\right)$

Since $r = k[A][B][C]^2$, we can substitute r for $k[A][B][C]^2$:

$r_{new} = 4r$

Therefore, the new rate of the reaction is 4r.

The correct answer is Option B: 4r.

Question141

Time required for 90% completion of a first order reaction is t. What is the time required for completion of 99% reaction?

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Options:

A. t

B. 2t

C. t1/2

D. 3t

Answer: B

Solution:

For first order reaction,

$$t = \frac{2.303}{k}\log_{10}\frac{[A]_0}{[A]_t}$$

Let t_1 be the time required for 90% completion and t_2 be the time required for 99% completion.

$$\therefore \frac{t_1}{t_2} = \frac{\log_{10}\frac{100}{10}}{\log_{10}\frac{100}{1}} = \frac{1}{2}$$

$$\therefore t_2 = 2t_1$$

Question142

The order of reaction for which the units of rate constant are $\mathrm{mol\ dm^{-3}\ s^{-1}}$ is

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Options:

A. 2

B. 1

C. 0

D. 3

Answer: C

Solution:

To determine the order of the reaction based on the units of the rate constant, we need to understand the relationship between the rate constant and the order of the reaction. The general form of the rate law for a reaction is given by:

$$\text{Rate} = k[A]^n$$

Here,

- k is the rate constant

- $[A]$ is the concentration of the reactant
- n is the order of the reaction

The units of the rate of reaction are typically $\mathrm{mol\,dm^{-3}\,s^{-1}}$ (molarity per second). The units of concentration are $\mathrm{mol\,dm^{-3}}$. When we look at the units of the rate constant, they can be derived based on the order of the reaction as follows:

Rearranging the rate law, we can express the units of the rate constant k as:

For a reaction of order n:

$$k = \frac{\text{Rate}}{[A]^n}$$

The units of the rate constant then would be:

$$\text{Units of } k = \frac{\left(\mathrm{mol\,dm^{-3}\,s^{-1}}\right)}{\left(\mathrm{mol\,dm^{-3}}\right)^n}$$

Simplifying further, we get:

$$\text{Units of } k = \left(\mathrm{mol\,dm^{-3}\,s^{-1}}\right)\left(\mathrm{mol^{-n}\,dm^{3n}}\right)$$

$$= \mathrm{mol^{1-n}\,dm^{3n-3}\,s^{-1}}$$

Given in the question, the units of the rate constant are $\mathrm{mol\,dm^{-3}\,s^{-1}}$. Therefore, by comparison:

$$\mathrm{mol^{1-n}\,dm^{3n-3}\,s^{-1}} = \mathrm{mol\,dm^{-3}\,s^{-1}}$$

Equating the exponents of mol and dm on both sides, we get:

$$1 - n = 1$$

$$3n - 3 = -3$$

Solving $1 - n = 1$:

$$n = 0$$

Therefore, the order of the reaction for which the units of the rate constant are $\mathrm{mol\,dm^{-3}\,s^{-1}}$ is zero. Hence, the correct answer is:

Option C: 0

Question143

What is rate constant of a first order reaction if 0.08 mole of reactant reduces to 0.02 mole in 23.03 minute?

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Options:

A. 0.2303 $\min^{-1}$

B. 1.6021 $\min^{-1}$

C. 0.4031 $\min^{-1}$

D. 0.06021 $\min^{-1}$

Answer: D

Solution:

For first order reaction,

$$\begin{aligned} k &= \frac{2.303}{t}\log_{10}\frac{[A]_0}{[A]_t} \\ k &= \frac{2.303}{23.03 \text{ min}}\log_{10}\frac{0.08}{0.02} \\ &= \frac{2.303}{23.03 \text{ min}}\log_{10} 4 = 0.1 \text{ min}^{-1} \times 0.6021 \\ \therefore k &= 0.06021 \text{ min}^{-1} \end{aligned}$$

Question144

Which of the following statements is NOT true for a reaction having rate law $r = k[H_2][I_2]$?

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Options:

A. The reaction is first order in H_2

B. Overall order of reaction is 1

C. The reaction is first order in I_2

D. Overall order of reaction is 2

Answer: B

Solution:

The given rate law for the reaction is: $\mathrm{r}=\mathrm{k}\left[\mathrm{H}_2\right]\left[\mathrm{I}_2\right]$.

To analyze each statement:

Option A: **The reaction is first order in** H_2. This statement is true. The rate law for H_2 is directly proportional to its concentration, meaning the reaction's rate increases linearly with an increase in the concentration of H_2. This is the definition of a first-order reaction in terms of a reactant.

Option B: **Overall order of reaction is 1**. This statement is not true. The overall order of a reaction is determined by adding the exponents of the concentration terms in the rate law. For the given reaction, the overall order is $1+1=2$ (since the reaction is first order with respect to both H_2 and I_2).

Option C: **The reaction is first order in** I_2. This statement is true. Similar to the explanation for H_2, the rate of the reaction is directly proportional to the concentration of I_2, making it also a first-order reaction with respect to I_2.

Option D: **Overall order of reaction is 2**. This statement is true. As explained under Option B, the overall order is the sum of the exponents of the concentration terms in the rate law, which is indeed 2 in this case.

Thus, the statement that is NOT true for the given reaction is Option B, **Overall order of reaction is 1**.

Question145

For the reaction $\mathrm{N}_{2(\mathrm{~g})}+3 \mathrm{H}_{2(\mathrm{~g})} \rightarrow 2 \mathrm{NH}_{3(\mathrm{~g})}$,what is the relation between $\frac{\mathrm{d}\left[\mathrm{N}_2\right]}{\mathrm{dt}}$ and $\frac{\mathrm{d}\left[\mathrm{H}_2\right]}{\mathrm{dt}}$?

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Options:

A. $3 \frac{\mathrm{d}\left[\mathrm{H}_2\right]}{\mathrm{dt}}=\frac{\mathrm{d}\left[\mathrm{N}_2\right]}{\mathrm{dt}}$

B. $\frac{\mathrm{d}\left[\mathrm{H}_2\right]}{\mathrm{dt}}=3 \frac{\mathrm{d}\left[\mathrm{N}_2\right]}{\mathrm{dt}}$

C. $3 \frac{\mathrm{d}\left[\mathrm{H}_2\right]}{\mathrm{dt}}=2 \frac{\mathrm{d}\left[\mathrm{N}_2\right]}{\mathrm{dt}}$

D. $\frac{\mathrm{d}\left[\mathrm{H}_2\right]}{\mathrm{dt}}=\frac{\mathrm{d}\left[\mathrm{N}_2\right]}{\mathrm{dt}}$

Answer: B

Solution:

$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$

$$\text{Rate of reaction} = -\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$$

$$\therefore \frac{d[H_2]}{dt} = 3\frac{d[N_2]}{dt}$$

Question146

For a reaction $A \rightarrow$ product, rate constant is $2 \times 10^{-2}\ s^{-1}$. The initial concentration of A is 1.0 mol dm^{-3}. What is the value of $\log \frac{1}{[A]_t}$ after 100 seconds?

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Options:

A. 0.430 mol dm^{-3}

B. 0.135 mol dm^{-3}

C. 0.270 mol dm^{-3}

D. 0.868 mol dm^{-3}

Answer: D

Solution:

$k = 2 \times 10^{-2}\ s^{-1}$, $[A]_0 = 1.0\ moldm^{-3}, t = 100\ s$, $\log \frac{1}{[A]_t} = ?$

For first order reaction,

$$k = \frac{2.303}{t}\log_{10}\frac{[A]_0}{[A]_t}$$

$$\therefore \log_{10}\frac{[A]_0}{[A]_t} = \frac{k \times t}{2.303}$$

$$\therefore \log_{10}\frac{1}{[A]_t} = \frac{2 \times 10^{-2} \times 100}{2.303} = 0.868\ \text{mol dm}^{-3}$$

Question147

Identify order of reaction if it's rate constant is x $\sec^{-1}$.

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Options:

A. 3

B. 2

C. 0

D. 1

Answer: D

Solution:

The order of a reaction can often be determined from the dimensions of the rate constant. For a general reaction of the form:

$aA + bB \rightarrow \text{Products}$

The rate law is given by:

$\text{Rate} = k[A]^m[B]^n$

where k is the rate constant, and m and n are the orders of the reaction with respect to reactants A and B, respectively. The overall order of the reaction is $m + n$.

For a zero-order reaction, the rate is independent of the concentration of reactants:

$\text{Rate} = k \quad \Rightarrow \quad k = \text{Rate}$

So, the units of k are the same as the units of rate, which are M $\sec^{-1}$ (where M represents molarity).

For a first-order reaction, the rate depends linearly on the concentration of one reactant:

$\text{Rate} = k[A] \quad \Rightarrow \quad k = \frac{\text{Rate}}{[A]}$

Thus, the units of k are $\sec^{-1}$.

For a second-order reaction, the rate depends on the concentration of two reactants, or on the square of the concentration of one reactant:

$\text{Rate} = k[A]^2 \quad \Rightarrow \quad k = \frac{\text{Rate}}{[A]^2}$

The units of k are $\mathrm{M}^{-1}\mathrm{sec}^{-1}$.

For a third-order reaction, the rate depends on the concentration of three reactants, or on the cube of the concentration of one reactant:

$$\text{Rate} = k[A]^3 \quad \Rightarrow \quad k = \frac{\text{Rate}}{[A]^3}$$

The units of k are $\mathrm{M}^{-2}\mathrm{sec}^{-1}$.

Therefore, since the given rate constant has units of sec^{-1}, it indicates that the reaction is first-order. Thus, the correct answer is:

Option D

1

Question148

For the reaction $2\mathrm{~A}+2\mathrm{~B} \rightarrow 2\mathrm{C}+\mathrm{D}$ if $\mathrm{r}=\mathrm{k}[\mathrm{A}]^2[\mathrm{~B}]^0$, then rate of reaction is

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Options:

A. inversely poportional to square of concentration of A

B. independent of concentration of A

C. independent of concentration of B

D. directly proportional to concentration of B

Answer: C

Solution:

Let's analyze the given rate law for the reaction $2\mathrm{~A}+2\mathrm{~B} \rightarrow 2\mathrm{C}+\mathrm{D}$. The rate of the reaction is given by:

$\mathrm{r}=\mathrm{k}[\mathrm{A}]^2[\mathrm{~B}]^0$

In this rate law, k is the rate constant, $[\mathrm{A}]$ is the concentration of reactant A, and $[\mathrm{B}]$ is the concentration of reactant B. Let’s examine each part of the rate law:

- $[\mathrm{~B}]^0$ means the rate is independent of the concentration of B, because any number to the power of zero is 1. Therefore, changing the concentration of B will not affect the rate of the reaction.

- $[\mathrm{A}]^2$ means the rate is directly proportional to the square of the concentration of A. Consequently, if the concentration of A is doubled, the rate of reaction will increase by a factor of 4 (since $2^2 = 4$).

With this information, we can understand how the rate of reaction is influenced by the concentrations of A and B:

1. *Option A: Inversely proportional to the square of concentration of* A - This option is incorrect because the rate is directly proportional to the square of the concentration of A, not inversely proportional.
2. *Option B: Independent of concentration of A* - This option is incorrect because the rate clearly depends on $[\mathrm{A}]^2$, so it is not independent of the concentration of A.
3. *Option C: Independent of concentration of* B - This option is correct because the rate is unaffected by the concentration of B, as shown by the term $[\mathrm{~B}]^0$ in the rate law equation.
4. *Option D: Directly proportional to concentration of B* - This option is incorrect because the rate of reaction does not depend on the concentration of B at all.

Therefore, the correct answer is:

Option C: Independent of concentration of B

Question149

For a first order reaction, intercept of the graph between $\log \frac{[\mathrm{A}]_{\mathrm{o}}}{[\mathrm{A}]_{\mathrm{t}}}$ (Y-axis) and conc. (X-axis) is equal to

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Options:

A. $-\frac{\mathrm{k}}{2.303 \mathrm{~K}}$

B. $-\log[\mathrm{A}]_{\mathrm{o}}$

C. zero

D. $\frac{2.303}{\mathrm{~K}}$

Answer: C

Solution:

Answer: **(C) zero**

The integrated rate law for a first-order reaction is given by:

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

Rearranging this equation to match the graph's axes (where $\log \frac{[A]_0}{[A]_t}$ is on the Y-axis and time (t) is on the X-axis):

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

This equation is in the form of a straight line, $y = mx + c$, where:

- $y = \log \frac{[A]_0}{[A]_t}$
- $m = \frac{k}{2.303}$ (slope)
- $x = t$ (time)
- $c = 0$ (intercept)

Therefore, the intercept of the graph between $\log \frac{[A]_0}{[A]_t}$ and time is equal to zero.

Question150

What is the half-life of a first order reaction if time required to decrease concentration of reactant from 1.0 M to 0.25 M is 10 hour?

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Options:

A. 12 hour

B. 4 hour

C. 5 hour

D. 10 hour

Answer: C

Solution:

$\mathrm{t} = 10$ hour, $[\mathrm{A}]_0 = 1.0\mathrm{M}$, $[\mathrm{A}]_\mathrm{t} = 0.25\mathrm{M}$

For first order reaction,

$$\begin{aligned} \mathrm{k} &= \frac{2.303}{\mathrm{t}}\log_{10}\frac{[\mathrm{A}]_0}{[\mathrm{A}]_\mathrm{t}} & &= \frac{2.303}{10\ \mathrm{h}}\log_{10}\frac{1.0}{0.25} \\ \text{(i)} \quad &= \frac{2.303}{10\ \mathrm{h}}\log_{10}4 & &= \frac{2.303 \times 0.6020}{10\ \mathrm{h}} \\ &= 0.139\ \mathrm{h}^{-1} \end{aligned}$$

$$\begin{aligned} \text{(ii)} \quad t_{1/2} &= \frac{0.693}{k} & &= \frac{0.693}{0.139\ \mathrm{h}^{-1}} \\ &= 4.98 \text{ hour} & &\approx 5 \text{ hours} \end{aligned}$$

Question151

What is the rate of disappearance of B in following reaction? $2\mathrm{A} + \mathrm{B} \rightarrow 3\mathrm{C}$, if rate of appearance of C is $1.3 \times 10^{-4}\ \mathrm{mol\ L}^{-1}\ \mathrm{s}^{-1}$.

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Options:

A. $4.33 \times 10^{-5}\ \mathrm{mol\ L}^{-1}\ \mathrm{s}^{-1}$

B. $8.6 \times 10^{-5}\ \mathrm{mol\ L}^{-1}\ \mathrm{s}^{-1}$

C. $2.6 \times 10^{-4}\ \mathrm{mol\ L}^{-1}\ \mathrm{s}^{-1}$

D. $5.2 \times 10^{-5}\ \mathrm{mol\ L}^{-1}\ \mathrm{s}^{-1}$

Answer: A

Solution:

$2A + B \rightarrow 3C$

$$\text{Rate} = -\frac{1}{2}\frac{d[A]}{dt} = -\frac{d[B]}{dt} = \frac{1}{3}\frac{d[C]}{dt}$$

$$\therefore \frac{d[B]}{dt} = \frac{1}{3} \times 1.3 \times 10^{-4} \text{ mol L}^{-1}\text{ s}^{-1}$$
$$= 4.33 \times 10^{-5} \text{ mol L}^{-1}\text{ s}^{-1}$$

Question152

What is the rate of appearance of Z in following reaction? $3x \rightarrow 2y + z$, if rate of disappearance of x is 0.072 mol s^{-1}

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Options:

A. 0.072 mol s^{-1}

B. 0.048 mol s^{-1}

C. 0.024 mol s^{-1}

D. 0.096 mol s^{-1}

Answer: C

Solution:

$3x \rightarrow 2y + z$

Rate of formation $= -\frac{1}{3}\frac{d[x]}{dt} = \frac{1}{2}\frac{d[y]}{dt} = \frac{d[z]}{dt}$

$\frac{d[z]}{dt} = \frac{1}{3}(0.072) = 0.024 \text{ mol s}^{-1}$

Question153

Slope of the graph between rate (Y-axis) and $[A]$(X-axis) for the first order reaction is equal to

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Options:

A. k

B. $\frac{2.303}{k}$

C. $\frac{k}{2.303}$

D. $-k$

Answer: A

Solution:

A plot of rate versus $[A]_t$ is a straight line passing through origin. This is shown in Fig. The slope of straight line = k.

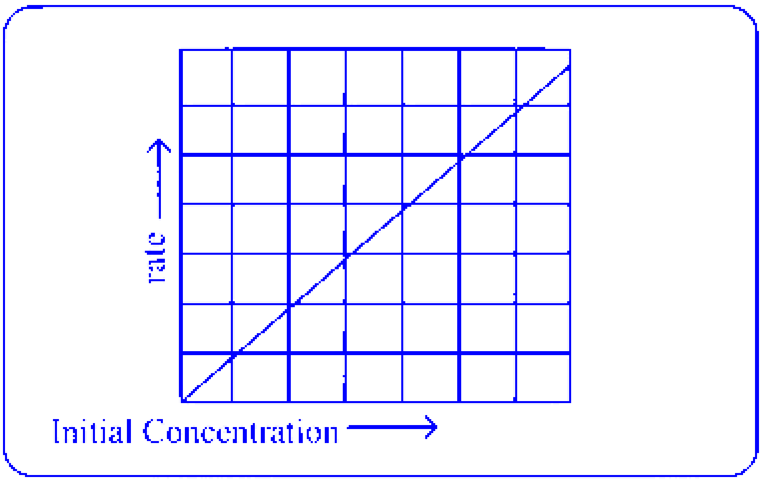

Fig. : Variation of rate with [A].

Question154

In a first order reaction, concentration of reactant is reduced to $(1/8)^{th}$ of concentration in 23.03 minutes. What is half-life period of reaction?

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Options:

A. 25 min

B. 7.7 min

C. 15 min

D. 30 min

Answer: B

Solution:

$[A]_0 = 1, [\,A]_t = \frac{1}{8}, t = 23.03 \text{ min}$

For first order reaction,

$$
\begin{aligned}
\text{(i)}\quad k &= \frac{2.303}{t} \log_{10} \frac{[A]_0}{[\,A]_t} \\
&= \frac{2.303}{23.03 \text{ min}} \log_{10} \frac{1}{(1/8)} \\
&= \frac{2.303}{23.03} \times \log_{10} 8 = \frac{2.303 \times 0.9031}{23.03} \\
&= 0.0903 \text{ min}^{-1}
\end{aligned}
$$

(ii) $t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.0903 \text{ min}^{-1}} = 7.7 \text{ min}$

Question155

For the reaction $2NO + Cl_2 \rightarrow 2NOCl$

What is the relation between $\frac{d[NO]}{dt}$ and $\frac{d[NOCl]}{dt}$?

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Options:

A. $\frac{d[NO]}{dt} = 2\frac{d[NOCl]}{dt}$

B. $\frac{d[NO]}{dt} = \frac{d[NOCl]}{dt}$

C. $\frac{1}{4}\frac{d[NO]}{dt} = \frac{d[NOCl]}{dt}$

D. $4\frac{d[NO]}{dt} = \frac{d[NOCl]}{dt}$

Answer: B

Solution:

$$2NO + Cl_2 \rightarrow 2NOCl$$

$$\text{Rate of reaction } = -\frac{1}{2}\frac{d[NO]}{dt} = -\frac{d\left[Cl_2\right]}{dt} = \frac{1}{2}\frac{d[NOCl]}{dt}$$

$$\therefore \frac{d[NO]}{dt} = \frac{d[NOCl]}{dt}$$

Question156

In a first order reaction concentration of reactant decreases from 20 m mol to 10 m mol in 1.151 min. What is rate constant?

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Options:

A. 1.15 min^{-1}

B. 3.0 min^{-1}

C. 5.50 min^{-1}

D. 0.60 min^{-1}

Answer: D

Solution:

$[A]_0 = 20 \text{ m mol}, [A]_t = 10 \text{ m mol}$

$t = 1.151 \text{ min}, k = ?$

For first order reaction,

$k=\frac{2.303}{t}\log_{10}\frac{[A]_0}{[A]_t}$

$\therefore k=\frac{2.303}{1.151\ \min^{-1}}\log_{10}\left(\frac{20}{10}\right)\quad=\frac{2.303\times\log 2}{1.151}\quad=\frac{2.303\times 0.3010}{1.151}=0.60\ \min^{-1}$

Question157

Which type of reaction order is followed by radioactive processes?

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Options:

A. 0

B. 1

C. 2

D. 1.5

Answer: B

Solution:

Radioactive processes follow the first order kinetics.

Question158

For the reaction, $3I^{-}_{(aq)}+S_2O^{2-}_{8(aq)}\longrightarrow I^{-}_{3(aq)}+2SO^{2-}_{4(aq)}$, rate of formation of SO_4^{2-} is $0.022\ \mathrm{moldm}^{-3}\mathrm{sec}^{-1}$. What is rate of formation of $I^{-}_{3(aq)}$?

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Options:

A. $0.022 \text{ mol dm}^{-3} \sec^{-1}$

B. $0.11 \text{ mol dm}^{-3} \sec^{-1}$

C. $0.011 \text{ mol dm}^{-3} \sec^{-1}$

D. $0.033 \text{ mol dm}^{-3} \sec^{-1}$

Answer: C

Solution:

$$3\mathrm{I}^{-}_{(\mathrm{aq})} + \mathrm{S}_2\mathrm{O}^{2-}_{8(\mathrm{aq})} \longrightarrow \mathrm{I}^{-}_{3(\mathrm{aq})} + 2\mathrm{SO}^{2-}_{4(\mathrm{aq})}$$
$$\therefore \frac{\mathrm{d}\left[\mathrm{I}_3^-\right]}{\mathrm{dt}} = \frac{1}{2}\frac{\mathrm{d}\left[\mathrm{SO}_4^{2-}\right]}{\mathrm{dt}} = \frac{1}{2} \times 0.022 = 0.011 \text{ mol dm}^{-3} \sec^{-1}$$

Question159

Instantaneous rate of a reaction is $-\frac{1}{2}\frac{d[x]}{dt} = -\frac{d[y]}{dt} = \frac{1}{2}\frac{d[z]}{dt}$, identify the reaction.

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Options:

A. $x - 2y \rightarrow 2z$

B. $2x + y \rightarrow 2z$

C. $2z + y \rightarrow 2x$

D. $2x - 2y \rightarrow z$

Answer: B

Solution:

Given the relationships :

$$-\frac{1}{2}\frac{d[x]}{dt} = -\frac{d[y]}{dt} = \frac{1}{2}\frac{d[z]}{dt}$$

This expression can be understood as follows :

1. The rate of disappearance of x is half that of y.
2. The rate of appearance of z is half the rate of disappearance of y.

To translate this into stoichiometric coefficients of a balanced chemical equation, let's analyze each component individually :

For compound x :

The negative sign indicates the disappearance (or consumption) of x. Given its relationship to y, we can infer that for every 2 moles of x that react, 1 mole of y will react.

$2x \rightarrow$ (the stoichiometric coefficient of x is 2)

For compound y :

The negative sign indicates the disappearance (or consumption) of y. Given its relationship to x and z, 1 mole of y is reacting for every 2 moles of x consumed and 2 moles of z produced.

$\rightarrow y \rightarrow$ (the stoichiometric coefficient of y is 1)

For compound z :

The positive sign indicates the production or appearance of z. Since the rate of appearance of z is half the rate of disappearance of y, 2 moles of z are produced for every mole of y that reacts.

$\rightarrow 2z$ (the stoichiometric coefficient of z is 2)

Combining all the components together, the balanced chemical equation is :

$2x + y \rightarrow 2z$

This represents a reaction where for every 2 moles of x and 1 mole of y that react, 2 moles of z are produced.

Therefore, Option B : $2x + y \rightarrow 2z$ is indeed the correct choice based on the given rate relationships.

Question160

Half life for a first order reaction is 6.93 hour. What is the time required for 80% completion of the reaction?

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Options:

A. 12 hours

B. 18 hours

C. 6 hours

D. 16 hours

Answer: D

Solution:

For first order reaction,

$k=\frac{0.693}{t_{\frac{1}{2}}}=\frac{0.693}{6.93}=0.1 \text{ hour }^{-1}$

Here, $[\mathrm{A}]_0=100,[\mathrm{~A}]_{\mathrm{t}}=100-80=20$

$$\text{Now, } \mathrm{k}=\frac{2.303}{\mathrm{t}} \log_{10} \frac{[\mathrm{A}]_0}{[\mathrm{~A}]_{\mathrm{t}}}$$

$$\therefore \mathrm{t}=\frac{2.303}{0.1} \log_{10} \frac{100}{20}=23.03 \times \log_{10} 5=23.03 \times 0.699=16.10 \text{ hours.}$$

Question161

For first order reaction the concentration of reactant decreases from 0.2 to 0.1 M in 100 minutes. What is the rate constant of the reaction?

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Options:

A. 144.3 min^{-1}

B. 69.3 min^{-1}

C. $6.93 \times 10^{-3} \text{ min}^{-1}$

D. 6.93 min^{-1}

Answer: C

Solution:

Given,

$$\begin{aligned}\text{Time}(t) &= 100 \text{ min} \\ [A_0] &= 0.2\text{M} \\ [A] &= 0.1\text{M}\end{aligned}$$

According first order reaction,

$$\text{rate constant } (k) = \frac{2.303}{t}\log\frac{[A_0]}{[A]}$$
$$= \frac{2.303}{100 \text{ min}}\log\frac{[0.2]}{[0.1]}$$
$$= \frac{2.303}{100\,\text{min}} \times 0.3010[\because \log 2 = 0.3010]$$
$$= 6.93 \times 10^{-3} \text{ min}^{-1}$$

Question162

A first order reaction is 25% completed in 40 minutes. What is the rate constant k for the reaction?

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Options:

A. $\frac{2.303 \times \log 1.33}{40}$

B. $2.303 \times \log \frac{4}{3}$

C. $\frac{2.303 \times \log 4}{40 \times 3}$

D. $\frac{2.303}{40} \times \log \frac{1}{4}$

Answer: A

Solution:

Given,

$$\begin{aligned}\text{Time}(t) &= 40 \text{ min} \\ a &= 100 \\ (a - x) &= 100 - 25 = 75\end{aligned}$$

First order reaction,

Rate $(k)=\frac{2.303}{t}\log\left(\frac{a}{a-x}\right)=\frac{2.303\times\log 1.33}{40}$

Question163

A first order reaction has a rate constant 0.00813 min^{-1}. How long will it take for 60% completion?

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Options:

A. 56.35 min

B. 98.7 min

C. 62.77 min

D. 112.7 min

Answer: D

Solution:

Given,

Rate constant $(k)=0.00813\ \mathrm{min}^{-1}$

According to first order reaction, the relation is

$$\begin{aligned} t &= \frac{2.303}{k}\log\frac{a}{a-x} \\ \text{As,}\quad a &= 100, (a-x)=100-60=40 \\ \therefore\quad t &= \frac{2.303}{0.00813}\log\frac{100}{40} \\ &= \frac{2.303}{0.00813}\times\log 2.5\quad[\log 2.5=0.3979] \\ &= 112.7\ \mathrm{min} \end{aligned}$$

Question164

What is the value of rate constant of first order reaction, if it takes 15 minutes for consumption of 20% of reactants?

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Options:

A. $1.38 \times 10^{-2} \mathrm{~min}^{-1}$

B. $1.48 \times 10^{-2} \mathrm{~min}^{-1}$

C. $1.07 \times 10^{-2} \mathrm{~min}^{-1}$

D. $1.84 \times 10^{-2} \mathrm{~min}^{-1}$

Answer: B

Solution:

Given, $t=15$ min for 20% of reactant to react.

For first order reaction,

Rate constant $(k)=\frac{.2.303}{t} \log \frac{[R_0]}{[R]}$

where,

$[R_0]=$ original amount of reactant

$[R]=$ reactant remaining unreacted

$$\begin{aligned} \text{So, } [R_0] &= 100 \\ [R] &= 80(20\% \text{ reacted }) \\ k &= \frac{2.303}{15 \text{ min}} \log \left[\frac{100}{80}\right] \\ &= 1.48 \times 10^{-2} \text{ min}^{-1} \end{aligned}$$

Question165

If concentration of reactant 'A' is increased by 10 times the rate of reaction becomes 100 times. What is the order of reaction, if rate law is, rate $=k[A]^x$?

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Options:

A. 3

B. 2

C. 4

D. 1

Answer: B

Solution:

Given,

$\text{Rate}_{\text{old}} = k[A]^x$ ..... (i)

If concentration of reactant 'A' is increased by 10 times that means,

$\text{Rate}_{\text{new}} = k[10A]^x$

Rate of reaction becomes 100 times,

$100 \times \text{Rate}_{\text{old}} = k[10A]^x$ ..... (ii)

From Eq. (i) and (ii),

$$\begin{aligned} 100 \times k[A]^x &= k[10A]^x \\ 100 &= 10^x \\ 10^2 &= 10^x \Rightarrow x = 2 \end{aligned}$$

Question166

A first order reaction has rate constant $1 \times 10^{-2}\ \mathrm{s}^{-1}$. What time will, it take for 20 g or reactant to reduce to 5 g ?

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Options:

A. 693.0 s

B. 138.6 s

C. 238.6 s

D. 346.5 s

Answer: B

Solution:

Given,

Rate constant $(k)=1 \times 10^{-2}\ \mathrm{s}^{-1}$

Initial concentration $(a)=20$ g

Final concentration $(a-x)=5 g$

For a first order reaction,

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

In terms of time,

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

Putting the values,

$t=\frac{2.303}{1 \times 10^{-2}} \log \left(\frac{20}{5}\right)=138.6\ \mathrm{s}$

Hence, the time required will be 138.6 s.

Question167

The integrated rate equation for first order reaction, $A \rightarrow$ product, is

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Options:

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

B. $k=\frac{2303}{t}+\log_{10}\frac{[A]_0}{[A]_t}$

C. $k=-\frac{1}{t}\ln\frac{[A]_t}{[A]_0}$

D. $k=2303t\log_{10}\frac{[A]_0}{[A]_t}$

Answer: C

Solution:

The integrated rate equation for first order reaction, $A \longrightarrow$ product is $k=-\frac{1}{t}\ln\frac{[A]_t}{[A]_0}$.

For 1st order reactions, $R \longrightarrow P$

Rate $=\frac{-d[R]}{dt}=k[R]$

On integrating this equation, we get

$\ln[R]=-kt+I$ .... (i)
At $t=0, R=[R]_0, \ln[R]_0=-k\times 0+I$
$\ln[R]_0=I$

Substituting the value of I in (i) and on rearrangement, we get

$$\ln[R]=-kt+\ln R_0$$
$$\text{and, } k=\frac{1}{t}\ln\frac{[R_0]}{[R]}$$

Question168

For the elementary reaction, $3H_2(g)+N_2(g)\longrightarrow 2NH_3(g)$ identify the correct relation among the following relations:

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Options:

A. $\frac{-3}{2}\frac{d[H_2(\mathrm{~g})]}{dt}=\frac{d[NH_3(\mathrm{~g})]}{dt}$

B. $\frac{-2}{3}\frac{d[H_2(\mathrm{~g})]}{dt}=\frac{d[NH_3(\mathrm{~g})]}{\mathrm{dt}}$

C. $\frac{d[\mathrm{NH}_3(g)]}{dt}=\frac{-1}{3}\frac{d[H_2(g)]}{dt}$

D. $\frac{-d[H_2(g)]}{dt}=\frac{d[\mathrm{NH}_3(g)]}{dt}$

Answer: B

Solution:

For elementary reaction,

$3\mathrm{H}_2(g)+\mathrm{N}_2(g)\longrightarrow 2\mathrm{NH}_3(g)$

$$\text{Rate of reaction }=\frac{-1}{3}\frac{d\left[\mathrm{H}_2\right]}{dt}=\frac{-d\,\mathrm{N}_2(g)}{dt}$$

$$=\frac{1}{2}\frac{\mathrm{NH}_3(g)}{dt}$$

From the given options, (b) is correct.

$-\frac{2}{3}\frac{d[\mathrm{H}_2(g)]}{dt}=\frac{d[\mathrm{NH}_3(g)]}{dt}$

Question169

The activation energy of a reaction is zero. Its rate constant at 280 K is $1.6\times10^{-6}\ \mathrm{s}^{-1}$, the rate constant at 300 K is

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Options:

A. $3.2\times10^{-6}\ \mathrm{s}^{-1}$

B. zero

C. $1.6\times10^{-6}\ \mathrm{s}^{-1}$

D. $1.6\times10^{-5}\ \mathrm{s}^{-1}$

Answer: C

Solution:

Given,

Activation energy of a reaction, $E_A = 0$

Rate constant, $k_1 = 1.6 \times 10^{-6}\ \mathrm{s}^{-1}$

Temperature, $T_1 = 280\ \mathrm{K}, T_2 = 300\ \mathrm{K}$

According to Arrhenius equation

$$\log \frac{k_2}{1.6 \times 10^{-6}} = \frac{0}{2.303R}\left[\frac{1}{280} - \frac{1}{300}\right]$$

$$\log \frac{k_2}{1.6 \times 10^{-6}} = 0$$

$$\frac{k_2}{1.6 \times 10^{-6}} = \text{antilog } 0$$

$$\frac{k_2}{1.6 \times 10^{-6}} = 1$$

$$\therefore \quad k_2 = 1.6 \times 10^{-6}\ \mathrm{s}^{-1}$$

Question170

For a chemical reaction rate law is, rate $= k[A]^2[B]$. If $[A]$ is doubled at constant $[B]$, the rate of reaction

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Options:

A. increases by a factor of 8

B. increases by a factor of 4

C. increases by a factor of 3

D. increases by a factor of 2

Answer: B

Solution:

Given,

Initial rate $= k[A]^2[B]$

If $[A]^1 = 2[A]$

Rate $k=[2A]^2[B]=4k[A]^2[B]$

Thus, if the concentration of A, $[A]$ is doubled and that of $[B]$ is constant, then initial rate increases by a factor of 4.

Question171

In the reaction, $H_2O_2(aq) \xrightarrow{I^-_{(aq)}} H_2O(I)+\frac{1}{2}O_2(g)$ **iodide ion acts as**

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Options:

A. homogenous catalyst

B. acid catalyst

C. heterogenous catalyst

D. enzyme catalyst

Answer: A

Solution:

For the reaction,

$H_2O_2(aq) \xrightarrow{I^-(aq)} H_2O(l)+\frac{1}{2}O_2(g)$

lodide ion acts a homogenous catalyst as the phase of $1^-$and reactant, H_2O_2 is aqueous (same).

Question172

For the elementary reaction $2SO_2(g)+O_2(g) \longrightarrow 2SO_3(g)$, **identify the correct among the following relations**

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Options:

A. $\frac{-d[\mathrm{SO}_2(g)]}{dt} = \frac{-d[\mathrm{O}_2(g)]}{dt}$

B. $\frac{+1}{2} \frac{d[\mathrm{SO}_3(g)]}{dt} = \frac{d[\mathrm{SO}_2(g)]}{dt}$

C. $\frac{+d[\mathrm{SO}_3(g)]}{dt} = \frac{-2d[\mathrm{O}_2(g)]}{dt}$

D. $\frac{+d[\mathrm{SO}_2(g)]}{dt} = \frac{-d[\mathrm{O}_2(g)]}{dt}$

Answer: C

Solution:

For an elementary reaction, $2\mathrm{SO}_2(\mathrm{g}) + \mathrm{O}_2(\mathrm{~g}) \longrightarrow 2\mathrm{SO}_3(\mathrm{~g})$, the instantaneous rate of reaction is expressed as,

$r_{\text{inst}} = \frac{-1}{2} d \frac{[\mathrm{SO}_2]}{dt} = -\frac{d[\mathrm{O}_2]}{dt} = \frac{1}{2} \frac{d[\mathrm{SO}_3]}{dt}$

Thus, $\frac{d[\mathrm{SO}_3]}{dt} = -\frac{2d[\mathrm{O}_2]}{dt}$.

Hence, option 'c' is correct.

Question173

Which among the following reaction is an example of a zero order reaction?

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Options:

A. $\mathrm{C}_{12}\mathrm{H}_{22}\mathrm{O}_{11}(aq) + \mathrm{H}_2\mathrm{O}(I) \longrightarrow \mathrm{C}_6\mathrm{H}_{12}\mathrm{O}_6(aq) + \mathrm{C}_6\mathrm{H}_{22}\mathrm{O}_6(aq)$

B. $2\mathrm{NH}_3(\mathrm{g}) \xrightarrow{\mathrm{Pt}} \mathrm{N}_2(g) + 3\mathrm{H}_2$

C. $2\mathrm{H}_2\mathrm{O}_2(I) \longrightarrow 2\mathrm{H}_2\mathrm{O}(I) + \mathrm{O}_2(g)$

D. $H_2(g) + I_2(g) \longrightarrow 2HI(g)$

Answer: B

Solution:

Decomposition of gaseous ammonia on platinum surface at high temperature and pressure is an example of zero order reaction. It is of first order when concentration of NH_3 is low.

$$2NH_3(g) \xrightarrow{\text{Pt catalyst}} N_2(g) + 3H_2(g)$$
$$\text{Rate } = k[NH_3]^{\circ} = k$$
