

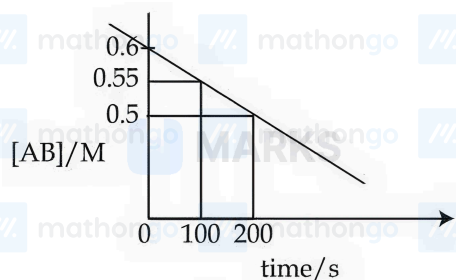
Q1. JEE Main 2026 (21 January Shift 2)

Decomposition of A is a first order reaction at T(K) and is given by $A(g) \longrightarrow B(g) + C(g)$. In a closed 1 L vessel, 1 bar A(g) is allowed to decompose at T(K). After 100 minutes, the total pressure was 1.5 bar. What is the rate constant (in min^{-1}) of the reaction ? ($\log 2 = 0.3$)

- (1) 6.9×10^{-3} (2) 6.9×10^{-4}
(3) 6.9×10^{-1} (4) 6.9×10^{-2}

Q2. JEE Main 2026 (23 January Shift 1)

For the thermal decomposition of reactant AB(g), the following plot is constructed.



The half life of the reaction is 'x' min. $x =$ _____ min. (Nearest integer)

Q3. JEE Main 2026 (28 January Shift 2)

$A \longrightarrow B$ (first reaction)

$C \longrightarrow D$ (second reaction)

Consider the above two first-order reactions. The rate constant for first reaction at 500 K is double of the same at 300 K. At 500 K, 50% of the reaction becomes complete in 2 hour. The activation energy of the second reaction is half of that of first reaction. If the rate constant at 500 K of the second reaction becomes double of the rate constant of first reaction at the same temperature; then rate constant for the second reaction at 300 K is _____ $\times 10^{-1} \text{ hour}^{-1}$ (nearest integer).

Q4. JEE Main 2026 (24 January Shift 2)

The half-life of ^{65}Zn is 245 days. After x days, 75% of original activity remained. The value of x in days is _____. (Nearest integer)

(Given: $\log 3 = 0.4771$ and $\log 2 = 0.3010$)

Q5. JEE Main 2026 (22 January Shift 2)

Correct statements regarding Arrhenius equation among the following are :

A. Factor $e^{-E_a/RT}$ corresponds to fraction of molecules having kinetic energy less than E_a .

B. At a given temperature, lower the E_a , faster is the reaction.

C. Increase in temperature by about 10°C doubles the rate of reaction.

D. Plot of $\log k$ vs $\frac{1}{T}$ gives a straight line with slope $= -\frac{E_a}{R}$.

Choose the correct answer from the options given below :

(1) B and D Only

(2) A and B Only

(3) B and C Only

(4) A and C Only

Q6. JEE Main 2026 (22 January Shift 2)

Consider $A \xrightarrow{k_1} B$ and $C \xrightarrow{k_2} D$ are two reactions. If the rate constant (k_1) of the $A \rightarrow B$ reaction can be expressed by the following equation $\log_{10} k = 14.34 - \frac{1.5 \times 10^4}{T/K}$ and activation energy of $C \rightarrow D$ reaction (E_{a2}) is $\frac{1}{5}$ th of the $A \rightarrow B$ reaction (E_{a1}), then the value of (E_{a2}) is _____ kJmol^{-1} . (Nearest Integer)

Q7. JEE Main 2026 (22 January Shift 1)

$A \rightarrow \text{product}$ (First order reaction). Three sets of experiment were performed for a reaction under similar experimental conditions:

Run 1 $\Rightarrow$ 100 mL of 10 M solution of reactant A

Run 2 $\Rightarrow$ 200 mL of 10 M solution of reactant A

Run 3 $\Rightarrow$ 100 mL of 10 M solution of reactant A + 100 mL of H_2O added.

The correct variation of rate of reaction is

(1) Run 1 < Run 2 < Run 3

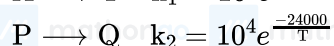
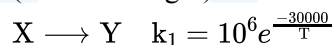
(2) Run 3 < Run 1 < Run 2

(3) Run 3 < Run 1 = Run 2

(4) Run 1 = Run 2 = Run 3

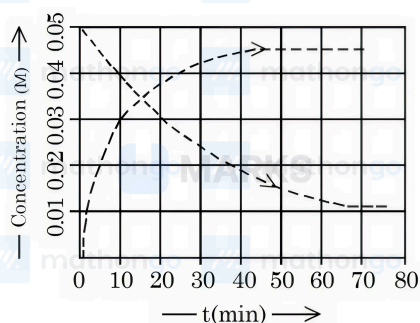
Q8. JEE Main 2026 (22 January Shift 1)

The temperature at which the rate constants of the given below two gaseous reactions become equal is _____ K. (Nearest integer)



Given: $\ln 10 = 2.303$

Q9. JEE Main 2026 (23 January Shift 2)



Given above is the concentration vs time plot for a dissociation reaction : $A \rightarrow nB$. Based on the data of the initial phase of the reaction (initial 10 min), the value of n is ____.

- (1) 3 (2) 2 (3) 4 (4) 5

Q10. JEE Main 2026 (23 January Shift 2)

Observe the following reactions at $T(K)$.

I. $A \rightarrow \text{products}$.

II. $5Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$

Both the reactions are started at 10.00 am. The rates of these reactions at 10.10 am are same. The value of $-\frac{\Delta[Br^-]}{\Delta t}$ at 10.10 am is $2 \times 10^{-4} \text{ mol L}^{-1} \text{ min}^{-1}$. The concentration of A at 10.10 am is $10^{-2} \text{ mol L}^{-1}$. What is the first order rate constant (in min^{-1}) of reaction I ?

- (1) 10^{-2} (2) 10^{-3} (3) 2×10^{-3} (4) 4×10^{-3}

Q11. JEE Main 2026 (21 January Shift 1)

Pre-exponential factors of two different reactions of same order are identical. Let activation energy of first reaction exceeds the activation energy of second reaction by 20 kJ mol^{-1} . If k_1 and k_2 are the rate constants of first and second reaction respectively at 300 K, then $\ln \frac{k_2}{k_1}$ will be _____. (nearest integer) [$R = 8.3 \text{ J K}^{-1} \text{ mol}^{-1}$]

Q12. JEE Main 2026 (28 January Shift 1)

An organic compound undergoes first order decomposition. The time taken for decomposition to $(\frac{1}{8})^{\text{th}}$ and $(\frac{1}{10})^{\text{th}}$ of its initial concentration are $t_{1/8}$ and $t_{1/10}$ respectively. What is the value of $\frac{t_{1/8}}{t_{1/10}} \times 10$? ($\log 2 = 0.3$)

- (1) 3 (2) 9 (3) 0.9 (4) 30

Q13. JEE Main 2026 (24 January Shift 1)

At 27°C in presence of a catalyst, activation energy of a reaction is lowered by 10 kJ mol^{-1} . The logarithm of ratio of $\frac{k(\text{catalysed})}{k(\text{uncatalysed})}$ is....

(Consider that the frequency factor for both the reactions is same)

- (1) 1.741 (2) 17.41 (3) 3.482 (4) 0.1741

ANSWERS AND SOLUTIONS

1. (1) 2. 10 3. 5 4. 102 5. (3) 6. 57 7. (3) 8. 1303
9. (1) 10. (4) 11. 8 12. (2) 13. (1)

1. (1) For the reaction $A(g) \rightarrow B(g) + C(g)$, initially pressure of A is 1 bar.

Initially: total pressure = 1 bar.

After 100 min: total pressure = 1.5 bar.

Using initial and final pressures, if x bar of A decomposes, then A becomes (1-x), B = x, C = x.

Total pressure = 1-x+x+x = 1+x = 1.5, so x = 0.5 bar.

For a first-order reaction: $k = \frac{2.303}{t} \log \frac{[A]_0}{[A]}$.

Using pressures instead: $k = \frac{2.303}{100} \log \frac{1}{0.5}$

$$= \frac{2.303}{100} \times \log 2 = \frac{2.303 \times 0.3}{100} = \frac{0.691}{100}$$

$$= 6.91 \times 10^{-3} \text{ min}^{-1} \approx 6.9 \times 10^{-3} \text{ min}^{-1}$$

2. (10) From the concentration vs. time graph, the thermal decomposition of AB(g) shows a linear decrease in concentration, indicating zeroth-order kinetics.

Initial concentration: $[AB]_0 = 0.6 \text{ M}$ at $t = 0 \text{ s}$.

Reading from the graph: $[AB] = 0.5 \text{ M}$ at $t = 200 \text{ s}$.

Rate of decomposition: $\Delta[AB]/\Delta t = (0.6 - 0.5) \text{ M} / 200 \text{ s} = 0.0005 \text{ M/s}$.

For zeroth-order kinetics: $[AB] = [AB]_0 - kt$.

Half-life occurs when $[AB] = [AB]_0/2 = 0.3 \text{ M}$.

$$0.3 = 0.6 - 0.0005 \times t_{1/2}$$

$$t_{1/2} = 0.3 / 0.0005 = 600 \text{ s} = 10 \text{ min.}$$

3. (5) For $A \xrightarrow{K_1} B$

$$\ln(2) = \frac{E_{a1}}{R} \left[\frac{1}{300} - \frac{1}{500} \right]$$

$$E_{a1} = \frac{\ln 2 \times R \times 1500}{2}$$

$$E_{a2} = \frac{E_{a1}}{2} = \frac{\ln 2 \times R \times 1500}{4}$$

$$(K_1)_{\text{at } 500 \text{ K}} = \frac{\ln 2}{2}$$

$$(K_2)_{\text{at } 500 \text{ K}} = \ln 2$$

Now for $C \xrightarrow{K_2} D$

$$\ln \left[\frac{(K_2)_{\text{at } 500 \text{ K}}}{(K_2)_{\text{at } 300 \text{ K}}} \right] = \left(\frac{\ln 2 \times R \times 1500}{4} \right) \times \frac{1}{R} \times \left[\frac{1}{300} - \frac{1}{500} \right]$$

$$(K_2)_{\text{at } 300 \text{ K}} = \frac{\ln 2}{\sqrt{2}} = 0.49$$

$$(K_2)_{\text{at } 300 \text{ K}} = 4.9 \times 10^{-1}.$$

So, answer is 5.

4. (102) Using radioactive decay: $A = A_0 \left(\frac{1}{2}\right)^{t/t_{1/2}}$.

With 75% activity remaining: $0.75 = \left(\frac{1}{2}\right)^{x/245}$.

Taking log base 10: $\log(0.75) = \frac{x}{245} \log(0.5)$.

Calculate: $\log(0.75) = \log 3 - \log 4 = 0.4771 - 2(0.3010) = -0.1249$ and $\log(0.5) = -0.3010$.

Thus $-0.1249 = \frac{x}{245} \times (-0.3010)$, giving $x = \frac{0.1249 \times 245}{0.3010} = 101.66 \approx 102$ days (nearest integer).

5. (3) The Arrhenius equation is given by $k = Ae^{-E_a/RT}$.

Statement A: The factor $e^{-E_a/RT}$ represents the fraction of molecules having kinetic energy equal to or greater than the activation energy E_a . Thus, statement A is incorrect.

Statement B: From the equation, as E_a decreases, the value of k increases, leading to a faster reaction rate. Thus, statement B is correct.

Statement C: For many reactions, the temperature coefficient is approximately 2, meaning the rate of reaction doubles for every 10°C rise in temperature. Thus, statement C is correct.

Statement D: Taking the natural logarithm of the Arrhenius equation, $\ln k = \ln A - \frac{E_a}{RT}$. Converting to base 10, $\log k = \log A - \frac{E_a}{2.303RT}$. A plot of $\log k$ vs $\frac{1}{T}$ gives a straight line with slope $= -\frac{E_a}{2.303R}$. The statement says slope $= -\frac{E_a}{R}$, which is only true for a $\ln k$ vs $\frac{1}{T}$ plot. Thus, statement D is incorrect.

Therefore, statements B and C are correct.

6. (57) From the Arrhenius equation:

$$\log_{10} k = \log_{10} A - \frac{E_a}{2.303RT}$$

Comparing with given equation $\log_{10} k_1 = 14.34 - \frac{1.5 \times 10^4}{T}$, we get: $\frac{E_{a_1}}{2.303R} = 1.5 \times 10^4 \text{ K}$

$$E_{a_1} = 1.5 \times 10^4 \times 2.303 \times 8.314$$

$$= 1.5 \times 10^4 \times 0.01913$$

$$= 287 \text{ kJ/mol}$$

Since $E_{a_2} = \frac{1}{5} E_{a_1}$:

$$E_{a_2} = \frac{287}{5} = 57.4 \approx 57 \text{ kJ/mol}$$

7. (3) For a first-order reaction, rate $= k[A]$, where rate is directly proportional to reactant concentration.

Run 1: 100 mL of 10 M solution

Concentration = 10 M

$$\text{Rate: } r_1 = k \times 10$$

Run 2: 200 mL of 10 M solution

Concentration remains 10 M (increasing volume doesn't change concentration)

$$\text{Rate: } r_2 = k \times 10$$

$$\text{Therefore } r_1 = r_2$$

Run 3: 100 mL of 10 M solution + 100 mL of H₂O

$$\text{Total moles of A} = 10 \text{ M} \times 0.1 \text{ L} = 1 \text{ mol}$$

$$\text{Total volume} = 0.2 \text{ L}$$

$$\text{New concentration} = 1 \text{ mol} / 0.2 \text{ L} = 5 \text{ M}$$

$$\text{Rate: } r_3 = k \times 5$$

$$\text{Therefore } r_3 < r_1 = r_2$$

8. (1303) To find the temperature at which $k_1 = k_2$,

$$\text{set the rate constant equations equal: } 10^6 e^{-30000/T} = 10^4 e^{-24000/T}$$

$$\text{Dividing both sides by } 10^4: 100 = e^{-24000/T+30000/T} = e^{6000/T}$$

$$\text{Taking natural logarithm: } \ln(100) = \frac{6000}{T}$$

$$\text{Since } \ln(100) = 2 \ln(10) = 2 \times 2.303 = 4.606: 4.606 = \frac{6000}{T}$$

$$\text{Solving for T: } T = \frac{6000}{4.606} = 1302.5 \text{ K}$$

Rounding to the nearest integer gives T = 1303 K.

9. (1) $A \rightarrow nB$

$$\text{At } t = 0: [A] = 0.05 \text{ M}, [B] = 0$$

$$\text{At } t = 10 \text{ min: } [A] = 0.04 \text{ M}, [B] = 0.01 \times n$$

$$\text{From the graph, } [B] \text{ at } t = 10 \text{ min} = 0.03 \text{ M}$$

$$0.01 \times n = 0.03$$

$$n = 3$$

10. (4) For reaction II:



The overall reaction rate is:

$$r = -\frac{1}{5} \frac{\Delta[\text{Br}^-]}{\Delta t} = -\frac{1}{3} \frac{\Delta[\text{Br}_2]}{\Delta t}$$

$$\text{Given } \frac{\Delta[\text{Br}^-]}{\Delta t} = 2 \times 10^{-4} \text{ mol L}^{-1} \text{ min}^{-1}$$

$$\text{The reaction rate is } r = \frac{2 \times 10^{-4}}{5} = 4 \times 10^{-5} \text{ mol L}^{-1} \text{ min}^{-1}$$

Since the rates of both reactions are equal at 10.10 am and reaction I is first-order:

$$r = k[A], \text{ so } 4 \times 10^{-5} = k \times 10^{-2}, \text{ giving } k = 4 \times 10^{-3} \text{ min}^{-1}$$

11. (8) Given: $A_1 = A_2$, $E_{a1} = E_{a2} + 20 \text{ kJ/mol}$, $T = 300 \text{ K}$, $R = 8.3 \text{ J K}^{-1} \text{ mol}^{-1}$

Using Arrhenius equation: $k = Ae^{-E_a/RT}$

$$\frac{k_2}{k_1} = \frac{e^{-E_{a2}/RT}}{e^{-E_{a1}/RT}} = e^{(E_{a1}-E_{a2})/RT}$$

$$E_{a1} - E_{a2} = 20 \text{ kJ/mol} = 20000 \text{ J/mol}$$

$$\frac{k_2}{k_1} = e^{\frac{20000}{8.3 \times 300}} = e^{\frac{20000}{2490}}$$

$$\ln \frac{k_2}{k_1} = \frac{20000}{2490} = 8.03 \approx 8$$

12. (2) For first-order decomposition with rate equation $[A] = [A]_0 e^{-kt}$, the time for concentration to reduce to fraction f is: $t = \frac{1}{k} \ln \frac{1}{f} = \frac{\ln(1/f)}{k}$

$$\text{For reduction to } 1/8: t_{1/8} = \frac{\ln 8}{k} = \frac{3 \ln 2}{k}$$

$$\text{For reduction to } 1/10: t_{1/10} = \frac{\ln 10}{k}$$

$$\text{Therefore: } \frac{t_{1/8}}{t_{1/10}} = \frac{\ln 8}{\ln 10} = \frac{3 \log 2}{\log 10} = \frac{3 \times 0.3}{1} = 0.9$$

$$\text{Thus: } \frac{t_{1/8}}{t_{1/10}} \times 10 = 0.9 \times 10 = 9$$

13. (1) Using the Arrhenius equation: $k = Ae^{-E_a/RT}$

$$\text{The ratio of rate constants is: } \frac{k_{\text{catalysed}}}{k_{\text{uncatalysed}}} = \frac{Ae^{-E_a^c/RT}}{Ae^{-E_a^u/RT}} = e^{(E_a^u - E_a^c)/RT} = e^{\Delta E_a/RT}$$

Since the catalyst lowers activation energy by 10 kJ/mol: $\Delta E_a = 10 \text{ kJ/mol} = 10000 \text{ J/mol}$

At $T = 27^\circ\text{C} = 300 \text{ K}$, $R = 8.314 \text{ J/(mol}\cdot\text{K)}$:

$$\frac{k_c}{k_u} = e^{10000/(8.314 \times 300)} = e^{10000/2494.2} = e^{4.009}$$

$$\text{Taking logarithm base 10: } \log_{10} \frac{k_c}{k_u} = 4.009 \times \log_{10}(e) = 4.009 \times 0.4343 = 1.741$$