

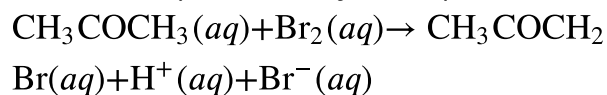
Q1 The rate law for a reaction between A and B is given by: $\text{Rate} = k[A]^n[B]^m$. On doubling the concentration of A and halving the concentration of B, the ratio of new rate to the earlier rate of the reaction will be as :

- (A) $n - m$
 (B) $2^{(n-m)}$
 (C) $2^{1/(n+m)}$
 (D) $2^{(m-n)}$

Q2 The rate of a gaseous reaction is given by the expression $k[A][B]$, if the volume of the reaction vessel is suddenly reduced to $1/4$ th of the initial volume, the reaction rate relating to original rate will be:

- (A) $1/10$ (B) $1/8$
 (C) 8 (D) 16

Q3 The bromination of acetone that occurs in acid solution is represented by this equation.



These kinetic data were obtained for given reaction concentrations.

Initial concentrations, M

$[\text{CH}_3\text{COCH}_3]$	$[\text{Br}_2]$	$[\text{H}^+]$	Initial rate, disappearance of Br_2 , Ms^{-1}
0.30	0.05	0.05	5.7×10^{-5}
0.30	0.10	0.05	5.7×10^{-5}
0.30	0.10	0.10	1.2×10^{-4}
0.40	0.05	0.20	3.1×10^{-4}

(A) $\text{rate} = k[\text{CH}_3\text{COCH}_3][\text{Br}_2]$

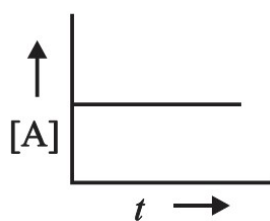
(B) $\text{rate} = k[\text{CH}_3\text{COCH}_3][\text{Br}_2][\text{H}^+]^2$

(C) $\text{rate} = k[\text{CH}_3\text{COCH}_3][\text{Br}_2][\text{H}^+]$

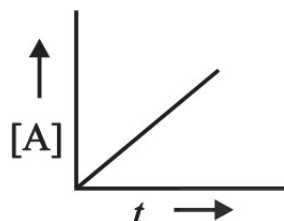
(D) $\text{rate} = k[\text{CH}_3\text{COCH}_3][\text{H}^+]$



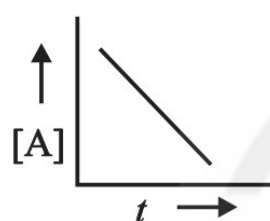
Q4 Which plot represents zero order reaction for



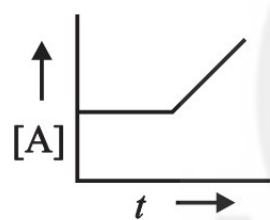
(A)



(B)



(C)



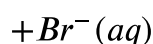
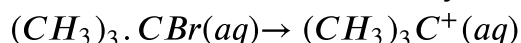
(D)

Q5 A chemical reaction is catalyzed by a catalyst X.

Hence X

- (A) Reduces enthalpy of the reaction
- (B) Does not affect equilibrium constant of reaction
- (C) Decreases rate constant of the reaction
- (D) Increases activation energy of the reaction

Q6 For the elementary step,



the molecularity is

- (A) Zero
- (B) 1
- (C) 2
- (D) cannot be ascertained

Q7 In a second order reaction when the concentrations of both the reactants are equal, the reaction is completed in 500 s. How long will it take for the reaction to go to 60% completion?

- (A) 1000 s
- (B) 300 s
- (C) 3000 s
- (D) 2000 s

Q8 The rate constant of a reaction is $1.2 \times 10^{-5} \text{ mol}^{-2} \text{ litre}^2 \text{ s}^{-1}$. The order of the reaction is:

- (A) Zero
- (B) 1
- (C) 2
- (D) 3

Q9 The decomposition of hydrocarbon follows the equation

$$k = (4.5 \times 10^{11} \text{ s}^{-1}) e^{-28000 \text{ K/T}}$$

Calculate the activation energy E_a .

- (A) 232.79 kJ mol⁻¹
- (B) 425.25 kJ mol⁻¹
- (C) 300 kJ mol⁻¹
- (D) 885.2 kJ mol⁻¹

Q10 If the rate constant for a first-order reaction is k , the time (t) required for the completion of 99% of the reaction is given by:

- (A) $t = \frac{2.303}{k}$
- (B) $t = \frac{0.693}{k}$
- (C) $t = \frac{4.606}{k}$
- (D) $t = \frac{6.909}{k}$

Q11 During the decomposition of H_2O_2 to give oxygen, 48 g O_2 is formed per minute at a certain point of time. The rate of formation of water at this point is

- (A) 0.75 mol min⁻¹
- (B) 1.5 mol min⁻¹
- (C) 2.25 mol min⁻¹
- (D) 3.0 mol min⁻¹

Q12 The temperature coefficient of most of the reactions lies between

- (A) 1 and 3
- (B) 2 and 3
- (C) 1 and 4
- (D) 2 and 4



- Q13** During the kinetic study of the reaction, $2A + B \rightarrow C + D$, following results were obtained

Run	[A]/mol L ⁻¹	[B]/mol L ⁻¹	Initial rate of formation of D/mol L ⁻¹ min ⁻¹
I.	0.1	0.1	6.0×10^{-3}
II.	0.3	0.2	7.2×10^{-2}
III.	0.3	0.4	2.88×10^{-1}
IV.	0.4	0.1	2.40×10^{-2}

Based on the above data which one of the following is correct?

- (A) Rate = $K[A]^2[B]$
 (B) Rate = $K[A][B]$
 (C) Rate = $K[A]^2[B]^2$
 (D) Rate = $K[A][B]^2$

- Q14** The following rate data were obtained at 303K for the following reaction: $2A + B \rightarrow C + D$

Experiment	[A]	[B]	Initial rate of formation of D
I	0.1	0.1	6.0×10^{-3}
II	0.3	0.2	7.2×10^{-2}
III	0.3	0.4	2.88×10^{-1}
IV	0.4	0.1	2.4×10^{-2}

The order of the reaction is:

- (A) 2 (B) 3
 (C) 1 (D) 4

- Q15** For a chemical reaction $2X + Y \rightarrow Z$, the rate of appearance of Z is $0.05 \text{ mol L}^{-1} \text{ min}^{-1}$. The rate of disappearance of X will be

- (A) $0.05 \text{ mol L}^{-1} \text{ h}^{-1}$
 (B) $0.05 \text{ mol L}^{-1} \text{ h}^{-1}$
 (C) $0.1 \text{ mol L}^{-1} \text{ h}^{-1}$
 (D) $0.25 \text{ mol L}^{-1} \text{ h}^{-1}$

- Q16** The reaction, $2X \rightarrow B$ is a zeroth order reaction. If the initial concentration of X is 0.2M, the half-life is 6 h. When the initial concentration of X is 0.5M, the time required to reach its final concentration of 0.2M will be:
 (A) 7.2 h
 (B) 18.0 h
 (C) 12.0 h
 (D) 9.0 h

- Q17** The rate constants k_1 and k_2 for two different reactions are $10^{16}e^{-2000/T}$ and $10^{15}e^{-1000/T}$, respectively. The temperature at which $k_1 = k_2$ is
 (A) $\frac{2000}{2.303} K$ (B) 2000K
 (C) $\frac{1000}{2.303} K$ (D) 1000K

- Q18** The value of rate constant for a first order reaction is $2.303 \times 10^{-2} \text{ sec}^{-1}$. What will be the time required to reduce the concentration to $\frac{1}{10}$ th of the initial concentration?
 (A) 100 sec (B) 10 sec
 (C) 2303 sec (D) 23.03 sec

- Q19** The half-life period of a first order reaction is 1 min 40 s. Calculate its rate constant
 (A) $6.93 \times 10^{-3} \text{ min}^{-1}$
 (B) $6.93 \times 10^{-3} \text{ s}^{-1}$
 (C) $6.93 \times 10^{-3} \text{ s}$
 (D) $6.93 \times 10^3 \text{ s}$

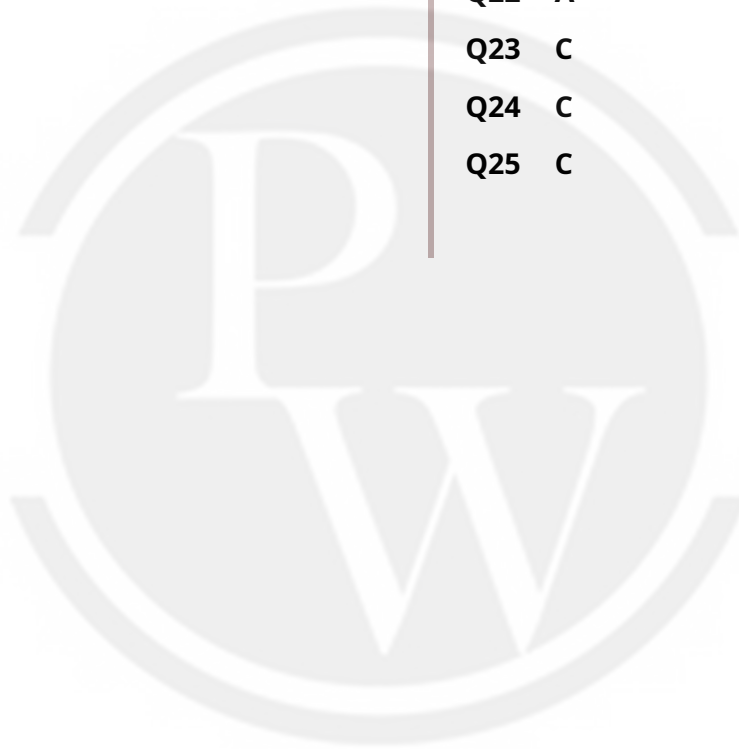
- Q20** For a reaction $aA \rightarrow bB$ when $[A] = 2.2 \text{ mM}$, the rate was found to be 2.4 mM s^{-1} . On reducing concentration of [A] to half, the rate changes to 0.6 mM s^{-1} . The order of reaction with respect to A is:
 (A) 1.5 (B) 2.0
 (C) 2.5 (D) 3.0



Answer Key

Q1 B
Q2 D
Q3 D
Q4 C
Q5 B
Q6 B
Q7 C
Q8 D
Q9 A
Q10 C
Q11 D
Q12 B
Q13 D

Q14 B
Q15 C
Q16 B
Q17 C
Q18 A
Q19 C
Q20 B
Q21 B
Q22 A
Q23 C
Q24 C
Q25 C



Q6 Text Solution:

Molecularity represents the number of molecules of reactants taking part in an elementary step of reaction.

Video Solution:**Q7 Text Solution:**

$$k = \frac{1}{t} \left[\frac{x}{a(a-x)} \right]$$

$$k = \frac{1}{500} \left[\frac{0.2a}{a(a-0.2a)} \right]$$

$$k = \frac{1}{2000a}$$

$$\frac{1}{2000a} = \frac{1}{t} \left[\frac{0.6a}{a(a-0.6a)} \right]$$

$$t = 3000s.$$

Video Solution:**Q8 Text Solution:**

unit of K is given as $[\text{Conc}]^{1-n} \text{ time}^{-1}$.
The units of k correspond to third order reaction.

Video Solution:**Q9 Text Solution:**

According to Arrhenius equation
 $k = Ae^{-E_a/RT}$... (i)

According to the available data
 $k = (4.5 \times 10^{11} s^{-1}) e^{-28000 K/T}$... (ii)

On comparing both equations :

$$-\frac{E_a}{RT} = \frac{-28000K}{T}$$

$$E_a = (28000K) \times R = (28000K)$$

$$\times (8.314 K^{-1} J \text{ mol}^{-1})$$

$$= 232792 \text{ kJ mol}^{-1}$$

$$E_a = 232.792 \text{ kJ mol}^{-1}$$

Video Solution:**Q10 Text Solution:**

First order rate constant is given as,

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

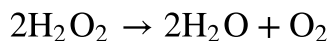
99% completed reaction

$$k = \frac{2.303}{t} \log \frac{100}{1} = \frac{2.303}{t} \log 10^2$$

$$k = \frac{2.303}{t} \times 2 \log 10$$

$$t = \frac{2.303}{k} \times 2 = \frac{4.606}{k} = \frac{4.606}{k}$$

Video Solution:**Q11 Text Solution:**



Rate of formation of

$$\text{O}_2 = \frac{48}{32} = 1.5 \text{ mol min}^{-1}$$

$$\text{Rate} = -\frac{1}{2} \frac{d[\text{H}_2\text{O}_2]}{dt} = \frac{1}{2} \frac{d[\text{H}_2\text{O}]}{dt} = \frac{d[\text{O}_2]}{dt}$$

Hence, Rate of formation of

$$\text{H}_2\text{O} = 2 \times \text{Rate of formation of O}_2$$

$$= 2 \times 1.5$$

$$= 3.0 \text{ mol min}^{-1}$$

Video Solution:



Q12 Text Solution:

Temperature coefficient is the ratio of two velocity constant having the difference of 10°C . For most of the reaction the value of temperature coefficient lies between 2 and 3

Video Solution:



Q13 Text Solution:

According to table:

when concentration of (1) = constant (II, III)
then (2) = doubles, r = four times.

i.e., order w.r.t. (2) = 2

when concentration of (2) = constant (I & IV)

(A) = four times, r = four times.

i.e., order w.r.t. (1) = 1

$$r = k[\text{A}]^1[\text{B}]^2$$

Video Solution:



Q14 Text Solution:

Inspection of table shows that $\text{Rate} = k[\text{A}] [\text{B}]^2$
 \therefore Order = 3.

Video Solution:



Q15 Text Solution:

For the reaction,



$$\text{Rate} = -\frac{1}{2} \frac{d[X]}{dt} = \frac{d[Z]}{dt}$$

$$= 0.05 \text{ mol L}^{-1} \text{ min}^{-1}$$

$$-\frac{1}{2} \frac{d[X]}{dt} = 0.05$$

$$-\frac{d[X]}{dt} = 2 \times 0.05$$

$$= 0.1 \text{ mol L}^{-1} \text{ min}^{-1}$$

Video Solution:**Q16 Text Solution:**

For a zero-order reaction,

$$t_{1/2} = \frac{[X]_0}{2k} \implies k = \frac{0.2}{2 \times 6} = \frac{0.1}{6} \text{ M/h.}$$

Time to drop from 0.5 M to 0.2 M:

$$t = \frac{[X]_0 - [X]_t}{k} = \frac{0.5 - 0.2}{\frac{0.1}{6}} = \frac{0.3 \times 6}{0.1} \approx 18 \text{ h.}$$

Video Solution:**Q17 Text Solution:**

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

$$k_1 = 10^{16} e^{-2000/T}, \quad k_2 = 10^{15} e^{-1000/T}$$

$$\log k_1 = 16 \log 10 - \frac{2000}{2.303T} \text{ and}$$

$$\log k_2 = 15 \log 10 - \frac{1000}{2.303T}$$

Since $k_1 = k_2$, and $\log 10 = 1$

$$16 - \frac{2000}{2.303T} = 15 - \frac{1000}{2.303T}$$

$$(16 - 15)T = -\frac{1000}{2.303T} + \frac{2000}{2.303T}$$

$$T = \frac{1000}{2.303} \text{ K}$$

Video Solution:**Q18 Text Solution:**

Given, rate constant for first order reaction:

$$k = 2.303 \times 10^{-2} \text{ s}^{-1}$$

Let, initial concentration be a.

Then, final concentration $(a - x) = \frac{1}{10}$ of an a

Integrated rate law expression for first order reaction is

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

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

On substituting values, we get:

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

$$= 10^2 \log_{10}$$

$$= 10^2 = 100 \text{ s}$$

Video Solution:**Q19 Text Solution:**

$$t_3 = 100 \text{ S}$$

$$K = \frac{0.693}{t_{1/2}} = \frac{0.693}{100}$$

$$k = 6.93 \times 10^{-3} \text{ s}^{-1}$$

Video Solution:

Q20 Text Solution:

$$r = K[A]^n$$

$$2.4 = K[2.2]^n$$

$$0.6 = K[1.1]^n$$

$$\therefore 4 = (2)^n \text{ or } n = 2$$

Video Solution:**Q21 Text Solution:**

Because half life is independent of initial concentration.

Video Solution:**Q22 Text Solution:**

$k = Ae^{-E_a/RT}$, As temperature increases increases exponentially.

Video Solution:**Q23 Text Solution:**

Molecular velocity increases with temperature as $V_1 > V_2$

Hence, T_1 must be greater than T_2 .

Video Solution:**Q24 Text Solution:**

$$t_3 = 100S$$

$$K = \frac{0.693}{t_2^1} = \frac{0.693}{100}$$

$$k = 6.93 \times 10^{-3} s^{-1}$$

Video Solution:**Q25 Text Solution:**

The definition of activation energy.

Video Solution:
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[iOS App](#)
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