CHEMICAL KINETICS

1. The activation energies of two reactions are E_1 and $E_2 (E_1 > E_2)$. If the temperature of the system is increased from T_1 to T_2, the rate constant of the reactions changes from k_1 to k_1 in the first reaction and k_2 and k_2 in the second reaction. Predict which of the following expression is correct?

   a) $\frac{k_1'}{k_1'} = \frac{k_2'}{k_2'}$ b) $\frac{k_1'}{k_1'} > \frac{k_2'}{k_2'}$ c) $\frac{k_1'}{k_1'} < \frac{k_2'}{k_2'}$ d) $\frac{k_1'}{k_1'} = \frac{k_2'}{k_2'} = 1$

2. Effective collisions are those in which molecules must:

   a) Have energy equal to or greater than the threshold energy
   b) Have proper orientation
   c) Acquire the energy of activation
   d) All of the above

3. Consider the following statements,
   The rate law for the acid catalysed hydrolysis of an ester being given as
   $Rate = k[H^+][ester] = k'[ester]$.
   If the acid concentration is doubled at constant ester concentration
   1. The second order rate constant, k is doubled.
   2. The pseudo first order rate constant, k is double.
   3. The rate of the reaction is doubled.
   Which of the above statements are correct?

   a) 1 and 2 b) 2 and 3 c) 1 and 3 d) 1,2 and 3

4. Half-life of two samples is 0.1 and 0.8 s. Their respective concentration is 400 and 50 respectively. The order of reaction is

   a) 0 b) 2 c) 1 d) 4

5. The units of rate of reaction are

   a) $\mathrm{L\ mol^{-1}\ s^{-1}}$ b) $\mathrm{mol\ L^{-1}\ s^{-1}}$ c) $\mathrm{mol\ s^{-1}}$ d) None of these

6. Units of rate constant of first and zero order reactions in terms of molarity M unit are respectively

   a) $\mathrm{s^{-1}, M\ s^{-1}}$ b) $\mathrm{s^{-1}, M}$ c) $\mathrm{M\ s^{-1}, s^{-1}}$ d) $\mathrm{M, s^{-1}}$

7. The half time of a second order reaction is:

   a) Inversely proportional to the square of the initial concentration of the reactants
   b) Inversely proportional to the initial concentration of the reactants
   c) Proportional to the initial concentration of reactants
   d) Independent of the initial concentration of reactants

8. $\frac{1}{[A]^2}$vs times are a straight line. Order of reaction is

   a) First b) Second c) Zero d) Third

9. For an endothermic reaction where, ΔH represents the enthalpy of the reaction in kJ/mol, the minimum value for the energy of activation will be

   a) Less than ΔH b) Zero c) More than ΔH d) Equal to ΔH

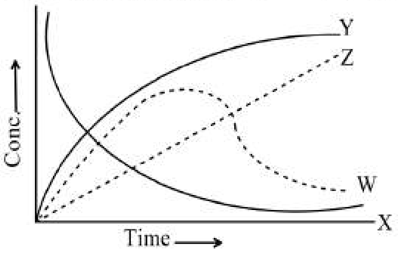
10. The unit of rate constant for a zero order reaction

    a) $\mathrm{L\ s^{-1}}$ b) $\mathrm{L\ mol^{-1}\ s^{-1}}$ c) $\mathrm{mol\ L^{-1}\ s^{-1}}$ d) $\mathrm{mol\ s^{-1}}$

11. What is the formula to find value of $t_{1/2}$ for a zero order reaction?

a) $\frac{k}{[R]_0}$ b) $\frac{2k}{[R]_0}$ c) $\frac{[R]_0}{2k}$ d) $\frac{0.693}{k}$

12. For the reaction, $A + B \longrightarrow C + D$. The variation of the concentration of the products is given by the curve:

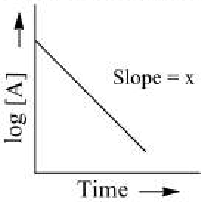


a) X b) Y c) Z d) W

13. Acid hydrolysis of sucrose is a

a) Pseudo first order reaction b) Zero order reaction
c) Second order reaction d) Unimolecular reaction

14. For a first order reaction the graph log $[A]$ vs t is given below



x is equal to

a) $\frac{0.693}{k}$ b) $\frac{k}{2.303}$ c) $-\frac{k}{2.303}$ d) log $[A]_0$

15. The rate constant of a first order reaction is $4 \times 10^{-3}\text{sec}^{-1}$. At a reactant concentration of 0.02 M, the rate of reaction would be:

a) 8×10^{-5} M sec^{-1} b) 4×10^{-3} M sec^{-1} c) 2×10^{-1} M sec^{-1} d) 4×10^{-1} M sec^{-1}

16. The rate constant for the reaction, $2N_2O_5 \rightarrow 4NO_2 + O_2$ is $3.0 \times 10^{-5}s^{-1}$. If the rate is $2.4 \times 10^{-5}molL^{-1}s^{-1}$ then the concentration of $N_2O_5(in\ mol\ L^{-1})$ is

a) 0.04 b) 0.8 c) 0.07 d) 1.4

17. Activation energy of a reaction is:

a) The energy released during the reaction
b) The energy evolved when activated complex is formed
c) Minimum amount of energy needed to overcome the potential barrier of reaction
d) The energy needed to form one mole of the product

18. The activation energy for a reaction is 9.0 Kcal/mol. The increase in the rate constant when its temperature is increased from 298 K to 308 K is:

a) 10% b) 100% c) 50% d) 63%

19. The rate of first order reaction, $A \longrightarrow$ Products, is 7.5×10^{-4}mol $\text{litre}^{-1}\text{sec}^{-1}$. If the concentration of A is 0.5 mol litre^{-1} the rate constant is:

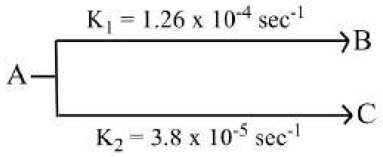
a) $3.75 \times 10^{-4}\text{sec}^{-1}$ b) $2.5 \times 10^{-5}\text{sec}^{-1}$ c) $1.5 \times 10^{-3}\text{sec}^{-1}$ d) $8.0 \times 10^{-4}\text{sec}^{-1}$

20. $2N_2O_5 \rightleftharpoons 4NO_2 + O_2$

For the above reaction which of the following is not correct above rates of reaction?

a) $\frac{-d[N_2O_5]}{dt} = 2\frac{d[O_2]}{dt}$ b) $\frac{-2d[N_2O_5]}{dt} = \frac{d[NO_2]}{dt}$
c) $\frac{d[NO_5]}{dt} = 4\frac{d[O_2]}{dt}$ d) $\frac{-2d[N_2O_5]}{dt} = 4\frac{d[NO_2]}{dt} = \frac{d[O_2]}{dt}$

21. A substance undergoes first order decomposition. The decomposition follows to parallel first order reactions as:



The percentage distribution of B and C are:
a) 80% B and 20% C
b) 76.83% B and 23.17% C
c) 90% B and 10% C
d) 60% B and 40% C

22. In Arrhenius plot intercept is equal to
a) $-E_a/R$ b) $\ln A$ c) $\ln k$ d) $\log_{10} a$

23. Half-life period of a first order reaction is 1386 seconds. The specific rate constant of the reaction is:
a) $5.0 \times 10^{-2} s^{-1}$ b) $5.0 \times 10^{-3} s^{-1}$ c) $0.5 \times 10^{-2} s^{-1}$ d) $0.5 \times 10^{-3} s^{-1}$

24. On addition of $AgNO_3$ to NaCl, white ppt. occurs:
a) Instantaneously
b) With a measurable speed
c) Slowly
d) None of these

25. Which is correct about zero order reaction?
a) Rate of reaction depends on decay constant.
b) Rate of reaction is independent of concentration.
c) Unit of rate constant is $conc^{-1}$
d) Unit of rate constant is $conc^{-1}\ time^{-1}$

26. The half-life period of a first order reaction is 1 min 40 s. Calculate its rate constant.
a) $6.93 \times 10^{-3} min^{-1}$ b) $6.93 \times 10^{-3} s^{-1}$ c) $6.93 \times 10^{-3} s$ d) $6.93 \times 10^{3} s$

27. The reaction $2A + B + C \rightarrow D + E$ is found to be first order in A, second in B and zero order in C. What is the effect on the rate of increasing concentration of A, B and C two times?
a) 72 times b) 8 times c) 24 times d) 36 times

28. In a reaction, the threshold energy is equal to:
a) Activation energy + normal energy of reactants
b) Activation energy – normal energy of reactants
c) Activation energy
d) Normal energy of reactants

29. Which one is not correct?
a) Rate of zero order reaction depends upon initial concentration of reactant
b) Rate of zero order reaction does not depend upon initial concentration of reactant
c) $t_{1/2}$ of first order reaction is independent of initial concentration of reaction
d) $t_{1/2}$ of zero order reaction is dependent of initial concentration of reaction

30. A reaction proceeds by first order, 75% of this reaction was completed in 32 min. the time required for 50% completion is
a) 8 min b) 16 min c) 20 min d) 24 min

31. The rate of the reaction
$CCl_3CHO + NO \rightarrow CHCl_3 + NO + CO$ is equal to rate $k[CCl_3CHO][NO]$. If concentration is expressed in mol/L. The unit of k is
a) $L\ mol^{-1}\ s^{-1}$ b) $mol\ L^{-1}\ s^{-1}$ c) $L^2\ mol^{-2}\ s^{-1}$ d) s^{-1}

32. Observe the following reaction,
$2A + B \rightarrow C$
The rate of formation of C is $2.2 \times 10^{-3} mol\ L^{-1}\ \overset{-1}{\text{min}}$.

What is the value of $-\frac{d[A]}{dt}$ $(mol\ L^{-1}\ \text{min}^{-1})$?

a) 2.2×10^{-3} b) 1.1×10^{-3} c) 4.4×10^{-3} d) 5.5×10^{-3}

33. The unit of rate constant of a third order chemical reaction is

a) $\text{mol}^{-1}\ \text{L}\ \text{s}^{-1}$ b) $\text{mol}^{-1}\ \text{s}^{-2}$

c) mol L d) $\text{s}^{-1}\ \text{mol}^{-2}\ \text{L}^2$

34. $CH_3COOC_2H_5 + H_2O \xrightarrow{H+} CH_3COOH + C_2H_5OH$ is an example of ...... order.

a) Zero b) Second c) Third d) Pseudo first order

35. Collision theory is applicable to

a) First order reactions b) Zero order reactions

c) Bimolecular reactions d) Intra-molecular reactions

36. The efficiency of an enzyme in catalyzing a reaction is due to its capacity

a) To form a strong enzyme substrate complex b) To decrease the bond energy of all substrate molecules

c) To change the shape of the substrate molecule d) To lower the activation energy of the reaction

37. The reaction

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

Which of the following does not express the reaction rate?

a) $\frac{d[D]}{dt}$ b) $-\frac{d[A]}{2dt}$ c) $-\frac{d[C]}{3dt}$ d) $-\frac{d[B]}{dt}$

38. If E_f and E_r are the activation energies of the forward and reverse reactions and the reaction is known to be exothermic then

a) $E_f < E_r$

b) $E_f > E_r$

c) $E_f = E_r$

d) No relation can be given between E_f and E_r as data are not sufficient

39. Milk turns sour at 40°C three times as faster as at 0°C. The energy of activation for souring of milk is:

a) 4.693 kcal b) 2.6 kcal c) 6.6 kcal d) None of these

40. Which plots will give the value of activation energy?

a) $K\ vs.\ T$ b) $1/K\ vs.\ T$ c) $\ln K\ vs.\ T$ d) $\ln K\ vs.\ \frac{1}{T}$

41. In a second order reaction when the concentration of both reactant 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

42. The rate constant (K) for the reaction $2A + B \longrightarrow$ Product was found to be 2.5×10^{-5} litre $\text{mol}^{-1}\text{sec}^{-1}$ after 15 sec, 2.60×10^{-5} litre $\text{mol}^{-1}\ \text{sec}^{-1}$ after 30 sec and 2.55×10^{-5}litre $\text{mol}^{-1}\ \text{sec}^{-1}$ after 50 sec. The order of reaction is:

a) 2 b) 3 c) Zero d) 1

43. The differential rate expression for the reaction $H_2 + I_2 \longrightarrow 2HI$ is:

a) $\frac{-d[H_2]}{dt} = \frac{-d[I_2]}{dt} = \frac{-d[HI]}{dt}$

b) $\frac{d[H_2]}{dt} = \frac{d[I_2]}{dt} = \frac{d[HI]}{dt}$

c) $\frac{1}{2}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[I_2]}{dt} = \frac{d[HI]}{dt}$

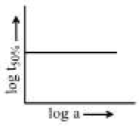
d) $-2\frac{d[H_2]}{dt} = -2\frac{d[I_2]}{dt} = \frac{d[HI]}{dt}$

44. For the elementary step,

$(CH_3)_3 \cdot CBr(aq) \longrightarrow (CH_3)_3C^+(aq) + Br^-(aq)$ the molecularity is:

a) Zero b) 1 c) 2 d) Cannot ascertained

45. A graph plotted between log $t_{50\%}$ $vs.$ $\log a$ concentration is a straight line. What conclusion can you draw from the given graph?



a) $n = 1, t_{1/2} = \frac{1}{K \cdot a}$ b) $n = 2, t_{1/2} = 1/a$ c) $n = 1, t_{1/2} = \frac{0.693}{K}$ d) None of the above

46. If a is the initial concentration then time required to decompose half of the substance for nth order is inversely proportional to:

a) a^n b) a^{n-1} c) a^{1-n} d) a^{n-2}

47. The hydrolysis of ethyl acetate,

$CH_3COOC_2H_5 + H_2O \xrightarrow{H^+} CH_3COOH + C_2H_5OH$ is:

a) First order b) Second order c) Third order d) Zero order

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

a) $\frac{1}{2^{m+n}}$ b) $(m + n)$ c) $(n - m)$ d) $2^{(n-m)}$

49. For the reaction

$H_2(g) + Br_2(g) \rightarrow 2HBr(g)$

The experimental data suggest

rate $= k[H_2][Br_2]^{1/2}$

the molecularity and order of the reaction are respectively

a) $1, \frac{1}{2}$ b) $1, 1$ c) $\frac{3}{2}, \frac{3}{2}$ d) $2, \frac{3}{2}$

50. The rate of reaction increases with temperature due to

a) Decrease in activation energy b) Increase in activation energy

c) Increase in collision frequency d) Increase in concentration

51. In a first order reaction, the concentration of the reactant is decreased from 1.0 M to 0.25 M in 20 minute. The rate constant of the reaction would be:

a) 10 min^{-1} b) 6.931 min^{-1} c) 0.6931 min^{-1} d) 0.06931 min^{-1}

52. The reaction obey I order with respect to H_2 and ICl both

$H_2(g) + 2ICl(g) \longrightarrow 2HCl(g) + I_2(g)$

Which of the following mechanism is in consistent with the given fact?

Mechanism A: $H_2(g) + 2ICl \longrightarrow 2HCl(g) + I_2(g)$

Mechanism B: (i) $H_2(g) + ICl(g) \xrightarrow{\text{slow}} HCl(g) + HI(g)$

(ii) $HI(g) + ICl(g) \longrightarrow HCl(g) + I_2$

a) A and B both b) Neither A nor B c) A only d) B only

53. Two reactions $A \longrightarrow$ products and $B \longrightarrow$ products have rate constants K_A and K_B at temperature T and activation energies E_A and E_B respectively. If $K_A > K_B$ and $E_A < E_B$ and assuming that A for both the reactions is same, then:

a) At higher temperatures K_A will be greater than K_B

b) At lower temperature K_A and K_B will differ more and $K_A > K_B$

c) As temperature rises K_A and K_B will be close to each other in magnitude

d) All of the above

54. The half life for a reaction ... of temperature.

a) Independent

b) Increased with increase

c) Decreased with increase

d) Dependent

55. The following mechanism has been proposed for the reaction of NO with Br_2 to form NOBr
$NO(g) + Br_2(g) \rightleftharpoons NOBr_2(g)$
$NOBr_2(g) + NO(g) \rightarrow 2NOBr(g)$
If the second step is the rate determining step, the order of the reaction with respect to NO(g) is
a) 1
b) 0
c) 3
d) 2

56. The unit and value of rate constant and that of rate of reaction are same for
a) Zero order
b) First order
c) Second order
d) Third order

57. According to collision theory of reaction rates:
a) Every collision between reactants leads to chemical reaction
b) Rate of reaction is proportional to velocity of molecules
c) All reactions which occur in gaseous phase are zero order reactions
d) Rate of reaction is directly proportional to collision frequency

58. Half-life of a reaction is found to be inversely proportional to the cube of initial concentration. The order of reaction is
a) 4
b) 3
c) 5
d) 2

59. A reaction involving two different reactants can never be
a) Bimolecular reaction
b) Second order reaction
c) First order reaction
d) Unimolecular reaction

60. For the non-equilibrium process, $A + B \longrightarrow$ Products, the rate is first order with respect to A and second order respect to B. If 1.0 mole each of A and B are introduced into a 1 litre vessel and the initial rate was 1.0×10^{-2} mol/litre-sec. The rate (in mol $\text{litre}^{-1}\ \text{sec}^{-1}$) when half of the reactants have been used:
a) 1.2×10^{-3}
b) 1.2×10^{-2}
c) 2.5×10^{-4}
d) None of these

61. The activation energy of a reaction is zero. The rate constant for the reaction
a) Decreases with decrease of temp
b) Increases with increase of temp
c) Decreases with increase of temp
d) Is nearly independent of temp

62. The burning of coal represented by the equation; $C(s) + O_2(g) \longrightarrow CO_2(g)$. The rate of this reaction is increased by:
a) Decrease in the concentration of oxygen
b) Powdering the lumps of coal
c) Decreasing the temperature
d) Providing inert atmosphere for burning

63. At room temperature, the reaction between NO and O_2 to give NO_2 is fast, while that between CO and O_2 is slow. It is due to:
a) CO is smaller in size than that of NO
b) CO is poisonous
The activation energy for the reaction,
c) $2NO + O_2 \longrightarrow 2NO_2$ is less than
$2CO + O_2 \longrightarrow 2CO_2$
d) None of the above

64. The rate of first order reaction is $1.5 \times 10^{-2} mol\ L^{-1}\ \min^{-1}$ at 0.5 M concentration of the reactant. The half-life of reaction is
a) 0.383 min
b) 23.1 min
c) 8.73 min
d) 7.53 min

65. The rate constant of a first order reaction at 27°C is $10^{-3}\ \text{min}^{-1}$. The temperature coefficient of this reaction is 2. What is the rate constant (in min^{-1}) at 17°C for this reaction?
a) 10^{-3}
b) 5×10^{-4}
c) 2×10^{-3}
d) 10^{-2}

66. The minimum energy required for the reacting molecules to undergo reaction is
a) Potential energy
b) Kinetic energy
c) Thermal energy
d) Activation energy

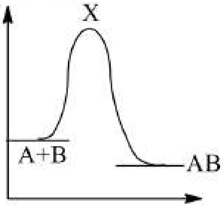
67. The decomposition of N_2O_5occur as $2N_2O_5 \rightarrow 4NO_2 + O_2$, and follows Ist order kinetics, hance
a) The reaction is unimolecular
b) The reaction is bimolecular
c) $t_{1/2} \propto a^0$
d) None of the above

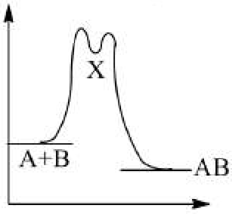
68. The rate of a chemical reaction doubles for every 10°C rise of temperature. If the temperature is raised by 50°C, the rate of the reaction increases by about
a) 10 times
b) 24 times
c) 32 times
d) 64 times

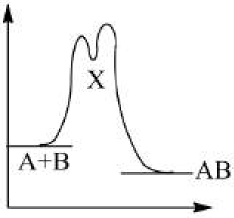
69. Which of the following statement is incorrect about the molecularity of a reaction?
a) Molecularity of a reaction is the number of molecules of the reactants presents in the balanced equation
b) Molecularity of a reaction is the number of molecules in the slowest step
c) Molecularity is always a whole number
d) There is no difference between order and molecularity of a reaction

70. For a reaction $A + B \longrightarrow$ Products, the rate of the reaction was doubled when the concentration of A was doubled. When the concentration of A and B were doubled, the rate was again doubled, the order of the reaction with respect to A and B are:
a) 1, 1
b) 2, 0
c) 1, 0
d) 0, 1

71. An exothermic chemical reaction occurs in two steps as follows
(I) $A + B \rightarrow X$ (fast)
(II) $X \rightarrow AB$ (slow)
The progress of the reaction can be best represented by

a)


b)


c)


d) All are correct

72. According to the Arrhenius equation a straight line is to be obtained by plotting the logarithm of the rate constant of a chemical reaction (log k) against
a) T
b) $\log T$
c) $\frac{1}{T}$
d) $\log \frac{1}{T}$

73. The rate constant is numerically the same for three reactions of first, second and third order respectively. Which one is true for rate of three reaction?
a) $r_1 = r_2 = r_3$
b) $r_1 > r_2 > r_3$
c) $r_1 < r_2 < r_3$
d) All of these

74. Mathematical expression for $t_{1/4}$ i.e., when (1/4)th reaction is over following first order kinetics can be given by
a) $t_{1/2} = \frac{2.303}{k} \log 4$
b) $t_{1/2} = \frac{2.303}{k} \log 2$
c) $t_{1/2} = \frac{2.303}{k} \log \frac{4}{3}$
d) $t_{1/2} = \frac{2.303}{k} \log \frac{3}{4}$

75. The rate of reaction:
$2NO + Cl_2 \longrightarrow 2NOCl$ is given by the rate, equation rate $= k[NO]^2[Cl_2]$. The value of the rate constant can be increased by:
a) Increasing the temperature
b) Increasing the concentration of NO
c) Increasing the concentration of the Cl_2
d) Doing all of these

76. A reaction was observed for 15 days and the percentage of the reactant remaining after the days indicated was recorded in the following table.

Time (days)	% Reactant remaining
0	100
2	50
4	39

6	25
8	21
10	18
12	15
14	12.5
15	10

Which one of following best describes the order and the half-life of the reaction?

Reaction order **Half-life (days)**

a) First 2
b) First 6
c) Second 2
d) Zero 6

77. In the reaction

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

The rate of appearance of bromine (Br_2) is related to rate of disappearance of bromide ions as following:

a) $\frac{d(Br_2)}{dt} = \frac{3}{5}\frac{d(Br^-)}{dt}$
b) $\frac{d(Br_2)}{dt} = -\frac{3}{5}\frac{d(Br^-)}{dt}$
c) $\frac{d(Br_2)}{dt} = -\frac{5}{3}\frac{d(Br^-)}{dt}$
d) $\frac{d(Br_2)}{dt} = \frac{5}{3}\frac{d(Br^-)}{dt}$

78. Which one of the following is a second order reaction?

a) $H_2 + Br_2 \rightarrow 2HBr$
b) $NH_4NO_3 \rightarrow N_2 + 3H_2O$
c) $H_2 + Cl_2 \xrightarrow{\text{Sunlight}} 2HCl$
d) $CH_3COOCH_3 + NaOH \rightarrow CH_3COONa + H_2O$

79. 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

80. In respect of the equation $k = Ae^{-E_a/RT}$ in chemical kinetics, which one of the statement is correct?

a) R is Rydberg constant
b) K is equilibrium constant
c) E_a is energy of activation
d) A is adsorption factor

81. The rate of chemical reaction (except zero order):

a) Decreases from moment to moment
b) Remains constant throughout
c) Is independent of the order of reaction
d) None of the above

82. For a zero order reaction

a) $t_{1/2} \propto R_0$
b) $t_{1/2} \propto 1/R_0$
c) $t_{1/2} \propto R_0^2$
d) $t_{1/2} \propto 1/R_0^2$

83. Effect of temperature on reaction rate is given by

a) Claisen-Clapeyron equation
b) Arrhenius equation
c) Gibbs Helmholtz equation
d) Kirchoff's equation

84. The Arrhenius equation expressing the effect of temperature on the rate constant of reaction is:

a) $K = \frac{E_a}{RT}$
b) $K = Ae^{-E_a/RT}$
c) $K = \log_e \frac{E_a}{RT}$
d) $K = e^{-Ea/RT}$

85. Find the two third life ($t_{1/2}$)of a first order reaction in which $k = 5.48 \times 10^{-14}$per second

a) $201 \times 10^{13} s$
b) $2.01 \times 10^{13} s$
c) $201 \times 10^{20} s$
d) $0.201 \times 10^{10} s$

86. $A + B \rightarrow$Product

If concentration of A is doubled, rate increases 4 times. If concentration of A and B are doubled, rate increases 8 times. The differential rate equation of the reaction will be

a) $\frac{dC}{dt} = kC_A \times C_B$
b) $\frac{dC}{dt} = kC_A^2 \times C_B^3$
c) $\frac{dC}{dt} = kC_A^2 \times C_B$
d) $\frac{dC}{dt} = kC_A^2 \times C_B^2$

87. For the reaction $A \rightarrow B$, the rate expression is $r = k[A]^n$. When the concentration of A is doubled, the rate of reaction is quadrupled. The value of n is

a) 1
b) Zero
c) 3
d) 2

88. The rate constant for the first order reaction is 60 s^{-1}. How much time will it take to reduce the concentration of the reaction to 1/16 M value ?

a) $4.6 \times 10^{-2} s$
b) $4.6 \times 10^{4} s$
c) $4.6 \times 10^{2} s$
d) $4.6 \times 10^{-4} s$

89. In the reaction,

$2N_2O_5 \rightarrow 4NO_2 + O_2$ initial pressure is 500 atm and rate constant k is $3.38 \times 10^{-5} s^{-1}$ after 10 min the final pressure of N_2O_5 is

a) 490 atm
b) 250 atm
c) 480 atm
d) 420 atm

90. For a chemical reaction, …... can never to a fraction

a) Order
b) Half life
c) Rate constant
d) Molecularity

91. The time taken for the completion of 3/4 of a first order reaction is

a) $(2.303/k) \log 3/4$
b) $(2.303/k) \log 4$
c) $(2.303/k) \log 1/4$
d) $(2.303/0.75) \log k$

92. $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$

What is the ratio of the rate of decomposition of N_2O_5 to rate of formation of O_2 ?

a) 1:2
b) 2:1
c) 1:4
d) 4:1

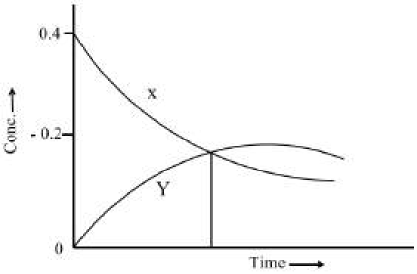
93. A first order reaction is 75% complete after 32 min. when was 50% of the reaction completed?

a) 16 min
b) 8 min
c) 4 min
d) 32 min

94. For a reaction, $A + 2B \rightarrow C$, rate is given by $+\frac{d[C]}{dt} = k[A][B]$,hence, the order of the reaction is

a) 3
b) 2
c) 1
d) 0

95. The accompanying figure depicts the change in concentration of species X and Y for the reaction $X \longrightarrow Y$, as a function of time. The point of intersection of the two curves represents:



a) $t_{1/2}$
b) $t_{3/4}$
c) $t_{2/3}$
d) Data is insufficient to predict

96. The rate constant of a reaction at temperature 200 K is 10 times less than the rate constant at 400 K. What is the activation energy (E_a) of the reaction?

a) 1842.4 R
b) 921.2 R
c) 460.6 R
d) 230.3 R

97. A zero order reaction is one:

a) In which reactants do not react
b) In which one of the reactants is in large excess
c) Whose rate does not change with time
d) Whose rate increases with time

98. In a first order reaction the $a/(a-x)$ was found to be 8 after 10 minute. The rate constant is:

a) $(2.303 \times 3 \log 2)/10$
b) $(2.303 \times 2 \log 3)/10$
c) $10 \times 2.303 \times 2 \log 3$
d) $10 \times 2.303 \times 3 \log 2$

99. If the rate of reaction A→ B doubles on increasing the concentration of A by 4 times, the order of the reaction is

a) 2
b) 1
c) $\frac{1}{2}$
d) 4

100. The rate of chemical reaction

a) Increase as the reaction proceeds
b) Decrease the reaction proceeds
c) May increase or decrease during reaction
d) Remains constant as the reaction proceeds

101. For zero order reaction, the integrated rate equation is

a) $kt = \frac{[A]}{[A]_0}$
b) $kt = [A] - [A]_0$
c) $[A] = -kt + [A]_0$
d) $[A] = kt - [A]_0$

102. The half-life period of a first order reaction is 69.3 s. what is the rate constant?

a) $0.01s^{-1}$ b) $0.1s^{-1}$ c) $1s^{-1}$ d) $10s^{-1}$

103. A reaction has a rate constant of 0.5 mol^{-1} dm^3 min^{-1}. If initial concentration of the reactant is 0.2 mol dm^{-3}, half-life of the reaction

a) 1.4 min b) 10 min c) 15 min d) 20 min

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

$\text{CH}_3\text{COCH}_3(aq) + \text{Br}_2(aq) \rightarrow \text{CH}_3\text{COCH}_2\text{Br}(aq) + \text{H}^+(aq) + \text{Br}^-(aq)$

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, $M\ \text{s}^{-1}$
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}

Based on these data, the rate equation is:

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

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

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

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

105. The rate constant for a chemical reaction has units $L\ mol^{-1}s^{-1}$,order of the reaction will be

a) 0 b) 1 c) 2 d) 3

106. Activation energy of a chemical reaction can be determined by

a) Evaluating rate constant at standard temperatures

b) Evaluating velocities of reaction at two different temperatures

c) Evaluating rate constants at two different temperatures

d) Changing concentration of reactants

107. Which statement about molecularity of a reaction is wrong?

a) It is the number of molecules of the reactants taking part in a single step of reaction

b) It is calculated from the reaction mechanism

c) It may be either whole number or fractional

d) None of the above

108. Arrhenius equation may not be represented as

a) $\ln\frac{A}{k} = \frac{E_a}{RT}$ b) $\frac{d\ln k}{dT} = \frac{E}{RT^2}$

c) $\log A = \log k + \frac{E_a}{2.303\ RT}$ d) $\log\left[-\frac{E_a}{RT}\right] = \frac{k}{A}$

109. The reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ is carried out in a $1dm^3$vessel and $2dm^3$vessel separately. The ratio of the reaction velocities will be

a) 1:8 b) 1:4 c) 4:1 d) 8:1

110. The rate for the reaction, $RCl + NaOH(aq) \rightarrow ROH + NaCl$ is given by rate =k[RCl],the freezing point of the reaction is

a) Unaffected by increasing the temperature of the reaction

b) Decreased on increasing the temperature of the reaction

c) Halved on reducing the concentration of RCl to half

d) Doubled on doubling the concentration of NaOh

111. In the sequence of reaction,

$\text{A} \xrightarrow{k_1} \text{B} \xrightarrow{k_2} \text{C} \xrightarrow{k_3} \text{D}$ $k_3 > k_2 > k_1$ then the rate determining step of reaction is

a) $A \rightarrow B$ b) $B \rightarrow C$ c) $C \rightarrow D$ d) $A \rightarrow D$

112. A first order reaction is 20% complete in 10 min. What is the rate constant of the reaction?

a) 0.223 b) 0.0223 c) 0.322 d) 0.0322

113. The activation energy of exothermic reaction $A \rightarrow B$ 80 kJ mol^{-1}. The heat of reaction is 200 kJmol^{-1} . The activation energy for the reaction $B \rightarrow A$(in kJ mol^{-1}) will be

a) 80 b) 120 c) 40 d) 280

114. An endothermic reaction $A \rightarrow B$ has an activation energy of 15 kcal/mol and the energy of reaction is 5 kcal/mol. The activation energy for the reaction $B \rightarrow A$ is

a) 20 kcal/mol b) 15 kcal/mol c) 10 kcal/mol d) Zero

115. K for a zero order reaction is 2×10^{-2} mol L^{-1}sec^{-1}. If the concentration of the reactant after 25 sec is 0.5 M, the initial concentration must have been:

a) 0.5 M b) 1.25 M c) 12.5 M d) 1.0 M

116. Rate constant for a reaction is $10^{-3}s^{-1}$. The to leave 25% reaction is

a) 693 s b) 1386 s c) 6930 s d) 2029 s

117. By increase in temperature by 10 K, the rate of reaction becomes double. How many times the rate of reaction will be if the temperature is increased from 303K to 353 K?

a) 4 b) 8 c) 16 d) 32

118. Temperature coefficient of a reaction is 2. When temperature is increased from 30°C to 100°C, rate of the reaction increases by

a) 128 times b) 100 times c) 500 times d) 250 times

119. The activation energy of a reaction is 9 kcal/mol. The increase in the rate constant when its temperature is raised from 295 to 300 K is approximately

a) 10% b) 50% c) 100% d) 28%

120. For a reaction $\frac{1}{2}A \rightarrow 2B$, rate of disappearance of 'A' is related to the rate of appearance of B by the expression

a) $-\frac{d[A]}{dt} = \frac{1}{2}\frac{d[B]}{dt}$ b) $-\frac{d[A]}{dt} = \frac{1}{4}\frac{d[B]}{dt}$ c) $-\frac{d[A]}{dt} = \frac{d[B]}{dt}$ d) $-\frac{d[A]}{dt} = 4\frac{d[B]}{dt}$

121. For a first order reaction, $(A) \rightarrow$ products, the concentration of A changes from 0.1 M to 0.025 M in 40 minutes. The rate of reaction when the concentration of A is 0.01 M, is:

a) 3.47×10^{-4} M/min b) 3.47×10^{-5} M/min c) 1.73×10^{-4} M/min d) 1.73×10^{-5} M/min

122. In a 1st order reaction, reactant concentration C varies with time t as

a) 1/C increases linearly with t b) Log C decreases linearly with t

c) C decreases with 1/t d) Log C decreases with 1/t

123. The rate constant of a zero order reaction is 0.2 mol dm^{-3} h^{-1}. If the concentration of the reactant after 30 min is 0.05 mol dm^{-3}. Then its initial concentration would be

a) 0.01 mol dm^{-3} b) 0.15 mol dm^{-3} c) 0.25 mol dm^{-3} d) 4.00 mol dm^{-3}

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

a) 0.05 mol L^{-1} h^{-1} b) 0.05 mol L^{-1} min^{-1} c) 0.1 mol L^{-1} min^{-1} d) 0.25 mol L^{-1} min^{-1}

125. For a reaction,$x(g) \rightarrow y(g) + z(g)$ the half-life period is 10 min. in what period of time would the concentration of X be reduce to 10% of original concentration?

a) 20 min b) 33 min c) 15 min d) 25 min

126. When is the activation energy for the decomposition of N_2O_5 as,

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

If the values of rate constant = 3.45×10^{-5} at 27°C and rate constant = 6.9×10^{-3} at 67°C?

a) 112.5 kJ b) 200 kJ c) 149.5 kJ d) 11.25 kJ

127. Rate equation for a second order reaction is:

a) $K = \frac{2.303}{t}\log\frac{a}{(a-x)}$ b) $K = \frac{1}{t}\log\frac{a}{a(a-x)}$ c) $K = \frac{1}{t}\cdot\frac{x}{a(a-x)}$ d) $K = \frac{1}{t^2}\cdot\frac{a}{(a-x)}$

128. Which of the following statements are correct?

1. Order of a reaction can be known from experimental result and not from the stoichiometry of reaction
2. Overall molecularity of a reaction may be determined in a manner similar to overall order of reaction
3. Overall order of reaction,
$A^m + B^n \rightarrow AB_x$ is $(m+n)$
4. Molecularity of a reaction refers to
(i) Molecularity of each of the elementary steps (slow steps) in a multistep reaction
(ii) Molecularity of that particular step in a single step reaction

Select the correct answer by using the codes given below

a) 1, 3 and 4 b) 1, 2 and 3 c) 2, 3 and 4 d) 1, 2 and 4

129. For the order reaction with rate constant 'K' and initial concentration 'a', the half-life period given by

a) $\frac{1n2}{k}$ b) $\frac{1}{ka}$ c) $\frac{3}{2k.a^2}$ d) None of these

130. For the reaction, $2A + B \rightarrow C + D$, the order of reaction is

a) One with respect[B] b) Two with respect to [A]
c) Three d) Cannot be predicted

131. Which expression is wrong for fist order reaction?

a) $k = \frac{2.303}{t}\log\left(\frac{A_0}{A_t}\right)$ b) $k = \frac{t}{2.303}\log\left(\frac{A_0}{A_t}\right)$
c) $-k = \frac{t}{2.303}\log\left(\frac{A_t}{A_0}\right)$ d) Rate=k[A]

132. For a first order reaction, the half-life period is

a) Dependent on the square of the initial concentration.
b) Dependent on first power of initial concentration.
c) Dependent on the square root of initial concentration.
d) Independent on initial concentration

133. Give relation between half reaction time $(t_{1/2})$and initial concentration of reactant for $(n-1)$order reaction.

a) $t_{1/2} \propto [R]_0$ b) $t_{1/2} \propto [R]_0^{2-n}$ c) $t_{1/2} \propto [R]_0^{n+1}$ d) $t_{1/2} \propto [R]_0^{n-2}$

134. For a first order reaction $A \longrightarrow$ Product, the initial concentration of A is 0.1 M and after 40 minute it becomes 0.025 M. Calculate the rate of reaction at reactant concentration of 0.01M:

a) $3.47 \times 10^{-4}\ M\ \text{min}^{-1}$
b) $3.47 \times 10^{-5}\ M\ \text{min}^{-1}$
c) $1.735 \times 10^{-6}\ M\ \text{min}^{-1}$
d) $1.735 \times 10^{-4}\ M\ \text{min}^{-1}$

135. Rate of reaction

a) Decreases with increase in temperature
b) Increases with increase in temperature
c) May increase or decrease with increase in temperature
d) Does not depends on temperature

136. For the first order reaction with the rate constant k, which expression gives the rate half-life period? (Initial conc. $= a$)

a) $\frac{3}{2}ka^2$ b) $\frac{1}{ka}$ c) $\frac{1}{k}$ d) $\frac{0.693}{k}$

137. The rate constant (k') of one of the reaction is found to be double that of the rate constant (k'') of another reaction. Then the relationship between the corresponding activation energies of the two reaction $(E'_a$ and $E''_a)$ can be represented as

a) $E'_a > E''_a$ b) $E'_a = 4E''$ c) $E'_1 = E''_a$ d) $E'_a < E''_a$

138. For a given reaction of first order, it takes 15 minute for the concentration to drop from 0.8 M litre^{-1} to 0.4 M litre^{-1}. The time required for the concentration to drop from 0.1 M litre^{-1} to 0.025 M litre^{-1} will be:

a) 60 minute b) 15 minute c) 7.5 minute d) 30 minute

139. The rate constant for a second order reaction is $8 \times 10^{-5} M^{-1} \text{min}^{-1}$. How long will it take a $1M$ solution to be reduced to $0.5M$?

a) 8.665×10^3 minute b) 8×10^{-5} minute c) 1.25×10^4 minute d) 4×10^{-5} minute

140. The slope in Arrhenius plot, is equal to

a) $-\frac{E_a}{2.303\,R}$ b) $\frac{E_a}{R}$ c) $-\frac{R}{2.303\,E_a}$ d) None of the above

141. 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) 2000 K c) $\frac{1000}{2.303}$ K d) 1000 K

142. If the volume of the vessel in which the reaction $2NO + O_2 \rightarrow 2NO_2$ is occurring is diminished to 1/3rd of its initial volume. The rate of the reaction will be increased by

a) 3 times b) 9 times c) 27 times d) 36 times

143. The time for half-life period of a creation reaction $A \rightarrow$products is 1 h. when the initial concentration of the reactant 'A', is 2.0 $mol\ L^{-1}$, how much time does it take for its concentration to come from 0.50 to 0.25 $mol\ L^{-1}$, if it is a zero order reaction?

a) 4 h b) 0.5 h c) 0.25 h d) 1 h

144. For a reaction $A + B \rightarrow C + D$, if the concentration of A is doubled without altering the concentration of B, the rate gets doubled. If the concentration of B is increased by nine times without altering the concentration of A, the rate gets tripled. The order of the reaction is

a) 2 b) 1 c) 3/2 d) 4/3

145. What fraction of a reactant showing first order remains after 40 minute if $t_{1/2}$ is 20 minute?

a) 1/4 b) 1/2 c) 1/8 d) 1/6

146. At 500 k, the half-life period of a gaseous reaction at an initial pressure of 80 kPa is 350 s. when the pressure is 40 kPa, the half-life period is 175 s. The order of the reaction is

a) Zero b) One c) Two d) Three

147. Which of the following statements are incorrect?

a) Rate of the reaction involving conversion of *ortho* hydrogen to *para* hydrogen $= -\frac{d[H_2]}{dt} = k[H_2]^{3/2}$

b) Rate of the reaction involving the thermal decomposition of acetaldehyde $= k[CH_3CHO]^{1/2}$

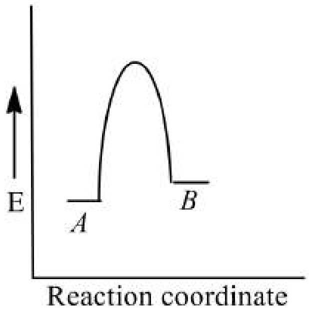
c) In the formation of phosgene from CO and Cl_2, the rate of the reaction $= k[CO][Cl_2]^{1/2}$

d) In the decomposition of H_2O_2, the rate of reaction $= k[H_2O_2]$

148. At 373 K, a gaseous reaction $A \rightarrow 2B + C$ is found to be of first order. Starting with pure A, the total pressure at the end of 10 min was 176 mm and after a long time when A was completely dissociated, it was 270 mm. The pressure of A at the end of 10 min was

a) 94 mm b) 47 mm c) 43 mm d) 90 mm

149. For a reversible reaction, $A \rightleftharpoons B$, which one of the following statements is wrong from the given energy profile diagram?



a) Activation energy of forward reaction is greater than backward reaction

b) The forward reaction is endothermic

c) The threshold energy is less than that of activation energy

d) The energy of activation of forward reaction is equal to the sum of heat of reaction and the energy of activation of backward reaction

150. Which one of the following is wrongly matched?

a) Saponification of $CH_3COOC_2H_5$ -second order reaction

b) Hydrolysis of CH_3COOCH_3 -pseudo unimolecular

c) Decomposition of H_2O_2 -first order reaction

d) Combination of H_2 and Br_2 to give HBr -first order reaction

151. For the reaction,

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

If the concentration of NO_2 increase by $5.2 \times 10^{-3} M$ in 100 s then the rate of the reactions

a) $1.3 \times 10^{-5} Ms^{-1}$ b) $0.5 \times 10^{-4} Ms^{-1}$ c) $7.6 \times 10^{-4} Ms^{-1}$ d) $2 \times 10^{-3} Ms^{-1}$

152. The rate of the reaction $A \rightarrow$ product, at the initial concentration of $3.24 \times 10^{-2} M$ is nine times its rate at another initial concentration of $1.2 \times 10^{-3} M$. The order of the reaction is

a) $\frac{1}{2}$ b) $\frac{3}{4}$ c) $\frac{3}{2}$ d) $\frac{2}{3}$

153. The half-life period for zero order reaction $A \rightarrow$ product, is 100 min. How long will it take in 80% completion?

a) 80 min b) 160 min c) 100 min d) 200 min

154. Consider the reaction $2A + B \rightarrow product$

When concentration of B alone was doubled, the half-life did not change. When the concentration of A alone was doubled, the rate increased by two times. The unit of rate constant for this reaction is

a) $L\ mol^{-1}s^{-1}$ b) No unit c) $mol\ L^{-1}s^{-1}$ d) s^{-1}

155. The expression for rate constant of a first order chemical reaction is

a) $k = \frac{1}{t} \cdot \frac{x}{a(a-x)}$ b) $k = \frac{2.303}{t} \log_{10} \frac{a}{(a-x)}$

c) $k = \frac{x}{t}$ d) $k = \frac{1}{2t}\left[\frac{1}{(a-x)^2} - \frac{1}{a^2}\right]$

156. In gaseous reactions important for the understanding of the upper atmosphere H_2O and O react bimolecularly to form two OH radicals. ΔH for this reaction is 72kJ at 500 K and E_a is 77 kJ mol^{-1}, then E_a for the bimolecular recombination of two OH radicals to form H_2O and O is:

a) 3 kJ mol^{-1} b) 4 kJ mol^{-1} c) 5 kJ mol^{-1} d) 7 kJ mol^{-1}

157. Activation energy of a reaction

a) Is independent of temperature

b) Increases with temperature

c) Gets doubled for every 10 degree rise in temperature

d) Decreases with temperature

158. For a I order reaction $A \longrightarrow B$ the reaction rate at reactant concentration $0.01M$ is found to be $2.0 \times 10^{-5}\ Ms^{-1}$. The half-life period of the reaction is:

a) 30s b) 300s c) 220s d) 347s

159. For a zero order reaction, the plot of concentration of reactant vs time is (intercept refers to concentration axis)

a) Liner with positive slope and zero intercept

b) Linear with negative slope and zero intercept

c) Linear with negative slope and non-zero intercept

d) Linear with positive slope and non-zero intercept

160. The rate of reaction between two reactants A and B decreases by a factor 4, if the concentration of reactant B is doubled. The order of this reaction with respect to B is:

a) -1 b) -2 c) 2 d) 1

161. The velocity constant of a reaction at 290 K was found to be 3.2×10^{-3} at 300 K, it will be

a) 6.4×10^{-3} b) 1.28×10^{-2} c) 9.6×10^{-3} d) 3.2×10^{-4}

162. The term $\frac{-dc}{dt}$ in a rate equation refers to

a) The decrease in concentration of the reactant with time
b) The concentration of the reactant
c) The change in concentration of the reactant
d) The velocity constant of the reaction

163. In a first order reaction the concentration of reactant decreases from 800 mol/dm^6 to 50 mol/dm^6 in 2×10^4 s. The rate constant of reaction in s^{-1} is

a) 2×10^4 b) 3.45×10^{-5} c) 1.386×10^{-4} d) 2×10^{-4}

164. Consider the chemical reaction,

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

The rate of this reaction can be expressed in terms of time derivative of concentration of $N_2(g)$, $H_2(g)$ and $NH_3(g)$. Identify the correct relationship amongst the rate expressions:

a) $\text{Rate} = -d[N_2]/dt = -\frac{1}{3}d[H_2]/dt = \frac{1}{2}d[NH_3]/dt$
b) $\text{Rate} = -d[N_2]/dt = -3d[H_2]/dt = 2d[NH_3]/dt$
c) $\text{Rate} = d[N_2]/dt = \frac{1}{3}d[H_2]/dt = \frac{1}{2}d[NH_3]/dt$
d) $\text{Rate} = -d[N_2]/dt = -d[H_2]/dt = d[NH_3]/dt$

165. Rate of reaction can be expressed by following rate expression, rate=$k[A]^2[B]$, if concentration of A is increased by 3 times and concentration of B is increased by 2 times, how many times rate of reaction increases?

a) 9 times b) 27 times c) 18 times d) 8 times

166. As the reaction progresses, the rate of reaction

a) Increases b) Decreases
c) Remains constant d) First increases, then decreases

167. The data for the reaction, A+B→C

Ex	$[A]_0$	$[B]_0$	Initial rate
1	0.012	0.035	0.10
2	0.024	0.070	0.80
3	0.024	0.035	0.10
4	0.012	0.070	0.80

The rate law corresponds to the above data is

a) $rate = k[B]^3$ b) $rate = k[B]^4$ c) $rate = k[A][B]^3$ d) $rate = k\,[A]^2[B]^2$

168. In a reaction, when the concentration of reactant is increased two times, the increase in rate of reaction was four times. Order of reaction is

a) Zero b) 1 c) 2 d) 3

169. For the reaction

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

The rate of change of concentration for hydrogen is $0.3 \times 10^{-4} Ms^{-1}$

The rates of change of concentration of ammonia is

a) -0.2×10^{-4} b) 0.2×10^{-4} c) 0.1×10^{-4} d) 0.3×10^{-4}

170. Which of the following statement is in accordance with collision theory?

Rate is directly proportional to collision frequency
Rate depend upon orientation of atoms
Temperature determines the rate

a) Only III b) Only I and II
c) Only II and III d) All of these

171. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + 22$ kcal.

The activation energy for the forward reaction is 50 kcal. What is the activation energy for the backward reaction?

a) -72 kcal b) -28 kcal c) +28 kcal d) +72 kcal

172. According to collision theory:
a) Collisions are sufficiently violent
b) All collision are responsible for reaction
c) All collisions are effective
d) Only highly energetic molecules have enough energy to react

173. The rate constant of a first order reaction whose half-life is 480 s is
a) $1.44\ s^{-1}$ b) $1.44 \times 10^{-3}\ s^{-1}$ c) $2.88 \times 10^{-3}\ s^{-1}$ d) $0.72 \times 10^{-3}\ s^{-1}$

174. $2A \rightarrow B + C$; It would be a zero order reaction when:
a) The rate of reaction is proportional to square of conc. of A
b) The rate of reaction remains same at any conc. of A
c) The rate remains unchanged at any conc. of B and C
d) The rate of reaction doubles if conc. of B is increased to double

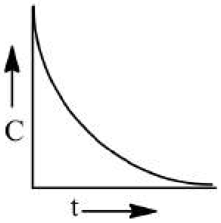
175. For a reaction $A + 2B \longrightarrow C$, rate is given by

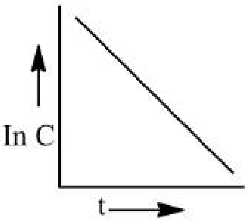
$$r = K[A][B]^2$$

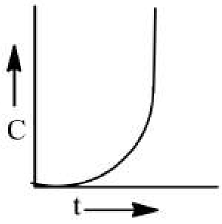
The order of reaction is:
a) 3 b) 6 c) 5 d) 7

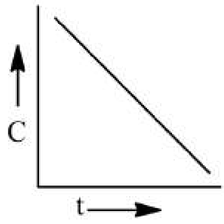
176. Rate constant for a reaction is λ. Average life is represent by
a) $\frac{1}{\lambda}$ b) $\frac{1n^2}{\lambda}$ c) $\frac{\lambda}{\sqrt{2}}$ d) $\frac{0.693}{\lambda}$

177. The plot between concentration versus time for a zero order reaction is represented by

a) 

b) 

c) 

d) 

178. For the decomposition of $N_2O_5(g)$, it is given that:
$2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$;
Activation energy E_a
$N_2O_5(g) \longrightarrow 2NO_2(g) + (1/2)O_2(g)$;
Activation energy E_a'
then;
a) $E_a = E_a'$ b) $E_a > E_a'$ c) $E_a < E_a'$ d) $E_a = 2E_a'$

179. During the kinetic study of the reaction $2A + B \longrightarrow C + D$ following results were obtained.

Run	$[A]$ in M	$[B]$ in M	Initial rate of formation of D in Ms^{-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}

On the basis of above data which one is correct:
a) $r = k[A]^2[B]$ b) $r = k[A][B]$ c) $r = k[A^2][B]^2$ d) $r = k[A][B]^2$

180. If the reaction rate at a given temperature becomes slower then
a) The free energy of activation is higher
b) The free energy of activation is lower
c) The entropy changes
d) The initial concentration of the reactants remains constant

181. The number of molecules of the reactants taking part in a single step of the reaction tells about:
a) Molecularity of the reaction
b) Mechanism of the reaction
c) Order of reaction
d) All of the above

182. For the reaction system,
$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
Volume is suddenly reduced to half its value by increasing the pressure on it. If the reaction is of first order with respect to O_2 and second order with respect to NO; the rate of reaction will
a) Diminish to one –fourth of its initial value
b) Diminish to one –eighth of its initial value
c) Increase to eight time of its initial value
d) Increase to four time of its initial value

183. The reaction,
$CH_3COOC_2H_5 + NaOH \longrightarrow CH_3COONa + C_2H_5OH$ is :
a) Biomolecular reaction
b) II order reaction
c) Both (a) and (b)
d) None of these

184. Which is correct relation in between $\frac{dc}{dt}, \frac{dn}{dt}$ and $\frac{dP}{dt}$, where c, n, and P, represent concentration, mole and pressure terms for gaseous phase reactant $A(g) \longrightarrow$ product?
a) $-\frac{dc}{dt} = -\frac{1}{V}\frac{dn}{dt} = -\frac{1}{RT}\frac{dP}{dt}$
b) $\frac{dc}{dt} = \frac{dn}{dt} = -\frac{dP}{dt}$
c) $\frac{dc}{dt} = \frac{RT}{V}\frac{dn}{dt} = -\frac{dP}{dt}$
d) All of the above

185. The rate constant of a reaction is found to be $3 \times 10^{-3} mol\ L^{-1}\ \text{min}^{-1}$. The order of reaction is
a) Zero
b) 1
c) 2
d) 1.5

186. A reactant (A) forms two products :
$A \overset{k_2}{\rightarrow} B$, Activation energy E_{a_1}
$A \overset{k_2}{\rightarrow} C$, Activation energy E_{a_2}
If $E_{a_2} = 2\ E_{a_1}$, than k_1 and k_2 are related as
a) $k_1 = 2k_2 e_{E_{a_2}}/RT$
b) $k_1 = k_2 e_{E_{a_1}}/RT$
c) $k_2 = k_1 e_{E_{a_2}}/RT$
d) $k_1 = Ak_2 e_{E_{a_1}}/RT$

187. For the reaction $2A + B \rightarrow A_2B$, the rate Law given is
a) $k[2A][B]$
b) $k[A]^3[B]$
c) $k[A][B]^3$
d) $k[A]^2[B]$

188. For producing the effective collisions the colliding molecules must have:
a) A certain minimum amount of energy
b) Energy lesser than threshold energy
c) Improper orientation
d) Proper orientation and energy equal or greater than threshold energy

189. The chemical reaction $2O_3 \rightarrow 3O_2$ proceeds as follows
$O_3 \rightleftharpoons O_2 + O$ (fast)
$O + O_3 \rightarrow 2O_2$ (slow)
The rate law expression should be
a) $r = k[O_3]^2$
b) $r = k[O_3]^2\ [O_2]^{-1}$
c) $r = k[O_3][O_2]$
d) Unpredictable

190. Two substances A and B are present such that $[A] = 4[B]$ and half-life of A is 5 minute and of B is 15 minute. If they start decaying at the same time following first order, how much time later will the concentration of both of them would be same?
a) 15 minute
b) 10 minute
c) 5 minute
d) 12 minute

191. A reaction involving A, B and C as reactants is found to obey the rate law, rate=$k[A]^x[B]^y[C]^z$. When the concentration of A, B and C are doubled separately, the rate is also found to increase two, zero and four times respectively. The overall order of the reaction is

a) 1
b) 2
c) 3
d) 4

192. The rate constant of nth order has units:

a) litre^{1-n}mol^{1-n} sec^{-1}
b) mol^{n-1} litre^{n-1} sec^{-1}
c) mol^{1-n} litre^{n-1} sec^{-1}
d) None of these

193. The reaction; N_2O_5 in $2NO_2 + 1/2\ O_2(g)$ is of first order for N_2O_5 with rate constant $6.2 \times 10^{-4} s^{-1}$. What is the value of rate of reaction when $[N_2O_5] = 1.25$ mole L^{-1}?

a) 5.15×10^{-5} mole $L^{-1}s^{-1}$
b) 6.35×10^{-3} mole $L^{-1}s^{-1}$
c) 7.75×10^{-4} mole $L^{-1}s^{-1}$
d) 3.85×10^{-4} mole $L^{-1}s^{-1}$

194. $t_{1/4}$ can be taken as the time taken for the concentration of reactant to drop to $\frac{3}{4}$ of its initial value. If the rate constant for a first order reaction is k the$t_{1/4}$ can be written as

a) 0.75 /k
b) 0.69 /k
c) 0.29 /k
d) 0.10 /k

195. In a chemical reaction two reactants take part. The rate of reaction is directly proportion to the concentration of one of them and inversely proportional to the concentration of the other. The order of reaction is

a) 0
b) 1
c) 2
d) 4

196. Which of the following is not the example of pseudo unimolecular reaction?

a) $CH_3COOC_2H_5 + H_2O \xrightarrow{H^+} CH_3COOH + C_2H_5OH$
b) $C_{12}H_{22}O_{11} + H_2O \xrightarrow{H^+} \underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{fructose}}{C_6H_{12}O_6}$
c) $CH_3COCl + H_2O \rightarrow CH_3COOH + HCl$
d) $CH_3COOC_2H_5 + H_2O \xrightarrow{OH^-} CH_3COOH + C_2H_5OH$

197. The differential rate law for the reaction,

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

a) $-\frac{d[NH_3]}{dt} = -\frac{d[O_2]}{dt} = -\frac{d[NO]}{dt} = -\frac{d[H_2O]}{dt}$
b) $\frac{d[NH_3]}{dt} = \frac{d[O_2]}{dt} = -\frac{1}{4}\frac{d[NO]}{dt} = -\frac{1}{6}\frac{d[H_2O]}{dt}$
c) $\frac{1}{4}\frac{d[NH_3]}{dt} = \frac{1}{5}\frac{d[O_2]}{dt} = \frac{1}{4}\frac{d[NO]}{dt} = \frac{1}{6}\frac{d[H_2O]}{dt}$
d) $-\frac{1}{4}\frac{d[NH_3]}{dt} = -\frac{1}{5}\frac{d[O_2]}{dt} = \frac{1}{4}\frac{d[NO]}{dt} = \frac{1}{6}\frac{d[H_2O]}{dt}$

198. 1 g of $_{79}Au^{198}$ $(t_{1/2} = 65\ h)$ give stable mercury by β- emission. What amount of mercury will left after 260 h?

a) 0.9375 g
b) 0.3758 g
c) 0.7586 g
d) 0.9000 g

199. The rate law for the chemical reaction

$2NO_2$CL→ $2NO_2 + CL_2$is rate =k[NO_2Cl].The rate determining step is

a) $2NO_2Cl \rightarrow 2NO_2 + 2Cl$
b) $NO_2 + Cl_2 \rightarrow NO_2Cl + Cl$
c) $NO_2Cl + Cl \rightarrow NO_2 + Cl_2$
d) $NO_2Cl \rightarrow NO_2 + Cl$

200. The rate law for the reaction

RCl + NaOH → ROH + NaCl is given by Rate = $k[R\text{Cl}]$. The rate of this reaction

a) Is doubled by doubling the concentration of NaOH
b) Is halved by reducing the concentration of RCl by one half
c) Is increased by increasing the temperature of the reaction
d) In unaffected by change in temperature

201. The rate constant of a reaction increases by 5% when its temperature is raised from 27°C to 28°C. The activation energy of the reaction is

a) 36.6 kJ/mol
b) 16.6 kJ/mol
c) 46.6 kJ/mol
d) 26.6 kJ/mol

202. .... of a reaction cannot be determined experimentally.

a) Order
b) Rate
c) Rate of constant
d) Molecularity

203. A first order reaction is carried out with an initial concentration of 10 mole per litre and 80% of the reactant changes into the product. Now if the same reaction is carried out with an initial concentration of 5 mol per litre the percentage of the reactant changing to the product is:

a) 40 b) 80 c) 160 d) Cannot be calculated

204. For a reaction of the type $A + B \rightarrow x$ products, it is observed that doubling the concentration of A causes the reaction rate (k_1) to be four times as great but doubling the amount of B does not affect the rate (k_2). The rate equation is

a) $k = k_1 + k_2$ b) $k = k_1 k_2$ c) $k = \frac{k_1}{k_2}$ d) $k^{1/2} = k_1 \times k_2$

205. Which increases on increase of temperature?

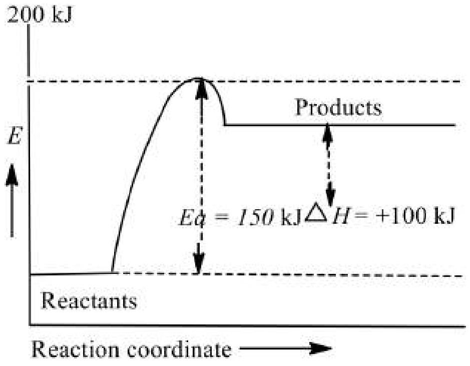
a) Energy of activation (E_a) b) Collision frequency (Z)

c) Rate constant (k) d) Both (a) and (c)

206. In the first order reaction, the concentration of the reactants is reduced to 25% in one hour. The half-life period of the reaction is

a) 2h b) 4h c) 1/2h d) 1/4h

207. In the given graph the activation energy, E_a for the reverse reaction will be



a) 150 kJ b) 50 kJ c) 200 kJ d) 100 kJ

208. What is the order of a reaction which has a rate expression rate $= k[A]^{3/2}[B]^{-1}$?

a) $\frac{3}{2}$ b) Zero c) $\frac{1}{2}$ d) None of these

209. For a fist order reaction, the concentration changes from 0.8 to 0.4 in 15 min. The time taken for the concentration to change from 0.1 M to 0.025 M is

a) 30 Min b) 15 Min c) 7.5 min d) 60 min

210. Give the hypothetical reaction mechanism

$A \xrightarrow{I} B \xrightarrow{II} C \xrightarrow{III} D \xrightarrow{IV} E$ and the date as

Species formed	Rate of its Formation
B	0.002 mol/h, per mole of A
C	0.030 mol/h, per mole of B
D	0.011 mol/h, per mole of C
E	0.420 mol/h, per mole of D

The rate determining step is

a) Step I b) Step II c) Step III d) Step IV

211. For $A + B \longrightarrow C + D, \Delta H = -20$ kJ mol^{-1} the activation energy of the forward reaction is 85kJ mol^{-1}. The activation energy for backward reaction is ...kJ mol^{-1}.

a) 65 b) 105 c) 85 d) 40

212. For the reaction $N_2 + 3H_2 \longrightarrow 2NH_3$, the rate $\frac{d[NH_3]}{dt} = 2 \times 10^{-4} M\ s^{-1}$. Therefore, the rate $-\frac{d[N_2]}{dt}$ is given as:

a) $10^{-4} M\ s^{-1}$ b) $10^{4} M\ s^{-1}$ c) $10^{-2} M\ s^{-1}$ d) $10^{-4} s\ M^{-1}$

213. The reaction $L \rightarrow M$ is started with 10.0 g of L. After 30 and 90 min, 5.0 g and 1.25 g of L respectively are left. The order of the reaction is

a) 0 b) 1 c) 2 d) 3

214. The rate of a reaction is expressed in different ways as follows

$$+\frac{1}{2}\frac{d[C]}{dt} = -\frac{1}{5}\frac{d[D]}{dt} = +\frac{1}{3}\frac{d[A]}{dt} = -\frac{d[B]}{dt}$$

The reaction is

a) $4A + B \rightarrow 2C + 3D$ b) $B + 5D \rightarrow 3A + 2C$

c) $4A + 2B \rightarrow 2C + 3D$ d) $B + \frac{1}{2}D \rightarrow 4A + 2C$

215. The rate of elementary reaction, $A \longrightarrow B$, increases by 100 times when the concentration of A is increased ten folds. The order of the reaction with respect to A is:

a) 1 b) 2 c) 10 d) 100

216. The differential rate law for the reaction $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$ is

a) $-\frac{d[NH_3]}{dt} = -\frac{d[O_2]}{dt} = -\frac{d[NO]}{dt} = -\frac{d[H_2O]}{dt}$ b) $\frac{d[NH_3]}{dt} = \frac{d[O_2]}{dt} = -\frac{1}{4}\frac{d[NO]}{dt} = -\frac{1}{6}\frac{d[H_2O]}{dt}$

c) $\frac{1}{4}\frac{d[NH_3]}{dt} = \frac{1}{5}\frac{d[O_2]}{dt} = \frac{1}{4}\frac{d[NO]}{dt} = \frac{1}{6}\frac{d[H_2O]}{dt}$ d) $-\frac{1}{4}\frac{d[NH_3]}{dt} = -\frac{1}{5}\frac{d[O_2]}{dt} = \frac{1}{4}\frac{d[NO]}{dt} = \frac{1}{6}\frac{d[H_2O]}{dt}$

217. For the reaction $A \rightarrow B$, when concentration of A is made 1.5 times, the rate of reaction becomes 1.837 times. The order of reaction is

a) 1 b) 1.5 c) 2 d) 2.5

218. For the reaction, $Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$, the initial concentration of I^- was 0.20 mol L^{-1} and the concentration after 20 min was 0.20 mol L^{-1}. Then the rate of formation of I_2 in mol L^{-1} would be

a) 1×10^{-4} b) 5×10^{-4} c) 1×10^{-3} d) 5×10^{-3}

219. The energies of activation for forward and reverse reactions for $A_2 + B_2 \rightleftharpoons 2AB$ are 180 kJ mol^{-1} and 200 kJ mol^{-1} respectively. The presence of a catalyst lowers the activation energy of both (forward and reverse) reactions by 100 kJ mol^{-1}. The enthalpy change of the reaction ($A_2 + B_2 \rightarrow 2AB$) in the presence of catalyst will be (in kJ mol^{-1})

a) 300 b) 120 c) 280 d) 20

220. Which statement is correct?

a) Reactions with low activation energy are usually exothermic

b) The rate law sometimes enable to deduce the mechanism of a reaction

c) The rate law for reaction is an algebraic expression relating the forward reaction rate to product concentration

d) Increase in the total pressure of a gas phase reaction increase the fraction of collisions effective in producing reactions

221. The temperature coefficient of a reaction is:

a) The rate constant at a fixed temperature

b) The ratio of rate constant at two temperature

c) The ratio of rate constant differing by 10° preferably 25° and 35°C

d) None of the above

222. The term $(-dc/dt)$ in rate equation refers to:

a) The concentration of a reactant

b) The decrease in concentration of the reactant with time

c) The velocity constant of reaction

d) None of the above

223. For a first order reaction, the initial concentration of a reactant is 0.05 M. After 45 min it is decreased by 0.015 M. calculation half reaction time $(t_{1/2})$

a) 87.42 min b) 25.90 min c) 78.72 min d) 77.20 min

224. The order of a gaseous phase reaction for which rate becomes half if volume of container having same amount of reactant is doubled is:

a) 1 b) 1/2 c) 2 d) 1/3

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

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

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

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

The relation in between K_1, K_2 and K_3 is:

a) $2K_1 = K_2 = 4K_3$ b) $K_1 = K_2 = K_3$ c) $2K_1 = 4K_2 = K_3$ d) None of these

226. Temperature dependent equation can be written as

a) $\text{In}\, k = \text{In}\, A - e^{E_a/RT}$ b) $\text{In}\, k = \text{In}\, A + e^{E_a/RT}$

c) $\text{In}\, k = \text{In}\, A - e^{RT/E_a}$ d) None of these

227. A first order reaction is 50% complete in 30 min at 27°C and in 10 min at 47°C. The energy of activation of the reaction is

a) 52.8 kJ b) 23.6 kJ c) 29.5 kJ d) 43.8 kJ

228. Increase in the concentration of the reactants leads to the change in

a) Activation energy b) Heat of reaction

c) Collision frequency d) Threshold energy

229. When a graph is plotted between ln k and $1/T$ for a first order reaction, a straight line is obtained. The slope of the line is equal to

a) $-\frac{E_a}{2.303}$ b) $-\frac{E_a}{2.303R}$ c) $-\frac{2.303}{E_aR}$ d) $-\frac{E_a}{R}$

230. For the two gaseous reactions, following data are given

$A \rightarrow B; k_1 = 10^{10}e^{-20,000/T}$

$C \rightarrow D; k_2 = 10^{12}\, e^{-24,606/T}$

The temperature at which k_1 becomes equal to k_2 is

a) 400 K b) 1000 K c) 800 K d) 1500 K

231. Hydrogenation of vegetable ghee at 25°C reduces pressure of H_2 from 2 atm to 1.2 atm in 50 minute. The rate of reaction in terms of molarity per second is:

a) 1.09×10^{-6} b) 1.09×10^{-5} c) 1.09×10^{-7} d) 1.09×10^{-8}

232. In the reaction $A + B \rightarrow$ products, if B is taken in excess, then it is an example of

a) Second order reaction b) Zero order reaction

c) Pseudounimolecular reaction d) First order reaction

233. The rate of a chemical reaction depends upon:

a) Time b) Pressure c) Concentration d) All of these

234. The rate constant for the reaction,

$2N_2O_5 \rightarrow 4NO_2 + O_2$ is $3.0 \times 10^{-5}\ s^{-1}$. If the rate is 2.40×10^{-5} then the concentration of N_2O_5 (in mol/L) is

a) 1.4 b) 1.2 c) 0.04 d) 0.8

235. The following data are for the decomposition of ammonium nitrite in aqueous solution.

Vol.of N_2 in cc	**Time (min)**
6.25	10
9.00	15
11.40	20
13.65	25
33.05	Infinity

The order of reaction is

a) Zero b) One c) Two d) Three

236. For a reaction $A \longrightarrow$ Product, rate law is $-\frac{d[A]}{dt} = K[A]_0$. The concentration of A left after time t when $t = \frac{1}{K}$ is:

a) $\frac{[A]_0}{e}$ b) $[A]_0 \times e$ c) $\frac{[A]_0}{e^2}$ d) $\frac{1}{[A]_0}$

237. For a first order reaction $A \rightarrow P$, the temperature (T) dependent rate constant (k) was found to follow the equation.

$\log k = -(2000)/T + 6.0$

The pre-exponential factor A and the activation energy E_a, respectively, are

a) $1.0 \times 10^6 s^{-1}$ and 9.2 kJ mol^{-1} b) 6.0 s^{-1} and 16.6 kJ mol^{-1}

c) $1.0 \times 10^6 s^{-1}$ and 16.6 kJ mol^{-1} d) $1.0 \times 10^6 s^{-1}$ and 38.3 kJ mol^{-1}

238. In Arrhenius equation $K = Ae^{-E_a/RT}$, the quantity $-E_a/RT$ is referred as:

a) Boltzmann factor b) Frequency factor c) Activation factor d) None of these

239. Among the following reaction, the fastest one is:

a) Burning of coal

b) Rusting of iron in moist air

c) Conversion of monoclinic sulphur to rhombic sulphur

d) Precipitation of silver chloride by mixing silver nitrate and sodium chloride solutions

240. The following homogeneous gaseous reactions were experimentally found to be second order overall.

1.$2NO \rightarrow N_2 + O_2$ 2.$3O_2 \rightarrow 2O_3$

3.$N_2O_3 \rightarrow NO + NO_2$ 4. $H_2 + I_2 \rightarrow 2HI$

Which of these are most likely to be elementary reaction that occur in one step?

a) 3 only b) 1 and 3 c) 1 and 4 d) 3 and 4

241. Consider a reaction; $aG + bH \rightarrow Products$

When concentration of both the reactants G and H is doubled, the rate increases by eight times. However, when concentration of G is doubled keeping the concentration of H fixed, the rate is doubled. The overall of the reaction,

a) 0 b) 1 c) 2 d) 3

242. The rate constant for a zero order reaction is

a) $k = \ln\frac{c_0 - c_t}{2t}$ b) $k = \frac{c_0 - c_t}{t}$ c) $k = \frac{c_0}{c_t}$ d) $k = \frac{c_0}{2t}$

243. The ratio of the times for 99.9% of the reaction to complete and half of the reaction to complete is

a) 2 b) 4 c) 8 d) 10

244. The activation energy for a simple chemical reaction $A \rightarrow B$ is E_a in forward direction. The activation energy for the reverse reaction

a) Is negative of E_a b) Is always less than E_a

c) Can be less than or more than E_a d) Is always double of E_a

245. After how many second will the concentration of the reactant in a first order reaction be halved if the rate constant is $1.155 \times 10^{-3} s^{-1}$?

a) 600 b) 100 c) 60 d) 10

246. For the reaction $A + 2B \rightarrow C$, the rate of reaction at a given instant can be given by

a) $+\frac{d[A]}{dt} = +\frac{1}{2}\frac{d[B]}{dt} = +\frac{d[C]}{dt}$ b) $\frac{d[A]}{dt} = -\frac{1}{2}\frac{d[B]}{dt} = -\frac{d[C]}{dt}$

c) $-\frac{d[A]}{dt} = -\frac{1}{2}\frac{d[B]}{dt} = +\frac{d[C]}{dt}$ d) $+\frac{d[A]}{dt} = +\frac{1}{2}\frac{d[B]}{dt} = +\frac{d[C]}{dt}$

247. Which of the following theory is not related to chemical kinetics?

a) Collision theory

b) Activated complex theory

c) Absolute reaction rate theory

d) VSPER theory

248. For the chemical change $A \rightarrow B$ it is found that the rate of reaction doubles when the concentration is increased by 4 times. The order of the reaction is

a) One b) Two c) Half d) None of these

249. If a certain reaction is first order with respect to A, second order with respect to B and zero order with respect to C then what is the order of reaction?

a) First b) Second c) Third d) Zero

250. If a plot of $\log_{10} C$ versus t give a straight line for a given reaction, then the reaction is

a) Zero order b) First order c) Second order d) Third order

251. The given reaction

$2FeCl_3 + SnCl_2 \rightarrow 2FeCl_2 + SnCl_4$ is an example of

a) Third order reaction b) Second order reaction

c) First order reaction d) None of these

252. The order for the reaction, $H_2 + Cl_2 \overset{hv}{\rightarrow} 2HCl$ over water is:

a) 0 b) 1 c) 2 d) 3

253. The inversion of cane sugar into glucose and fructose is:

a) I order b) II order c) III order d) Zero order

254. For the reaction, $N_2 + 3H_2 \longrightarrow 2NH_3$, if $\frac{d[NH_3]}{dt} = 2 \times 10^{-4}$ mol $L^{-1}s^{-1}$, the value of $\frac{-d[H_2]}{dt}$ would be:

a) 1×10^{-4} mol $L^{-1}s^{-1}$ b) 3×10^{-4} mol $L^{-1}s^{-1}$ c) 4×10^{-4} mol $L^{-1}s^{-1}$ d) 6×10^{-4} mol $L^{-1}s^{-1}$

255. For the reaction $\underset{(g)}{N_2O_5} \longrightarrow \underset{(g)}{2NO_2} + \underset{(g)}{\frac{1}{2}O_2}$, the rate of disappearance of N_2O_5 is 6.25×10^{-3}mol $L^{-1}s^{-1}$. The rate of formation of NO_2 and O_2 will be respectively.

a) 6.25×10^{-3}mol $L^{-1}s^{-1}$ and 6.25×10^{-3}mol $L^{-1}s^{-1}$

b) 1.25×10^{-2}mol $L^{-1}s^{-1}$ and 3.125×10^{-3}mol $L^{-1}s^{-1}$

c) 6.25×10^{-3}mol $L^{-1}s^{-1}$ and 3.125×10^{-3}mol $L^{-1}s^{-1}$

d) 1.25×10^{-2}mol $L^{-1}s^{-1}$ and 6.25×10^{-3}mol $L^{-1}s^{-1}$

256. If the concentration units are reduced by n times, then the value of rate constant of first order will:

a) Increase by n times

b) Decrease by factor of n

c) Not change

d) None of these

257. Unit of frequency factor (A) is

a) mol/L b) mol/L× s

c) Depends upon order of reaction d) It does not have any unit

258. The ionic reactions are usually very fast because:

a) It does not involve bond breaking

b) The energy of activation between charged ions is greater than that between neutral molecules

c) Collision frequency is very low

d) The reactions are highly exothermic

259. In the first order reaction, 75% of the reactant gets disappeared in 1.386h. The rate constant of the reaction is

a) $3.0 \times 10^{-3}s^{-1}$ b) $2.8 \times 10^{-4}s^{-1}$ c) $17.2 \times 10^{-3}s^{-1}$ d) $1.8 \times 10^{-3}s^{-1}$

260. Number of mole of a substance present in 1 litre volume is known as:

a) Activity b) Molar concentration c) Active mass d) None of these

261. The rate of a reaction is doubled when temperature increases by 10°C. If temperature is increased by 100°C, then rate of reaction will become

a) 64 times b) 256 times c) 512 times d) 1024 times

262. For a hypothetical reaction

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

$d[C]/dt$ is equal to

a) $\frac{3d[A]}{dt}$ b) $-\frac{3}{2}\frac{d[B]}{dt}$ c) $-\frac{d[B]}{dt}$ d) $-\frac{d[A]}{dt}$

263. The half-life of two samples is 0.1 and 0.4 s. Their reactive concentration is 200 and 50 respectively. What is the order of reaction?

a) 0 b) 2 c) 1 d) 4

264. Consider following two reaction,

$A \rightarrow product -\frac{d[A]}{dt} = k_1[A]^0$

$B \rightarrow product -\frac{d[B]}{dt} = k_2[B]^0$

$k_1 and\ k_2$ are expressed in term of molarity ($mol\ L^{-1}$) and time (s^{-1}) as

a) $s^{-1}, M\ s^{-1}L^{-1}$ b) $Ms^{-1}, M\ s^{-1}$ c) $s^{-1}, M^{-1}s^{-1}$ d) $Ms^{-1}, L\ s^{-1}$

265. The reaction, $NO + (1/2)O_2 \longrightarrow NO_2$ exhibits:

a) Small negative temperature coefficient

b) Decrease in value of K with temperature

c) Decrease in value of rate with temperature

d) All of the above

266. Consider the reaction,

$Cl_2(aq) + H_2S(aq) \rightarrow S(s) + 2H^+(aq) + 2Cl^-(aq)$ the rate equation for this reaction is , rate = $k[Cl_2][H_2S]$

Which of these mechanisms is/are consistent with this rate equation?

I. $Cl_2 + H_2S(aq) \rightarrow H^+Cl^- + Cl^+ + HS^-(slow)$

$Cl^+ + HS^- \rightarrow H^+Cl^- + S(fast)$

II. $H_2S \leftrightarrow H^+ + HS^-(fast\ equilibrium)$

$Cl_2 + HS^- \rightarrow 2Cl^- + H^+S(slow)$

a) (B) only b) Both (A) and (B) c) Neither (A) nor (B) d) (A)Only

267. In a reaction $2A \longrightarrow$ Products; the concentration of A decreases from 0.5 mol litre^{-1} to 0.4 mol litre^{-1} in 10 minute. The rate of reaction during this interval is:

a) 0.05 M min^{-1} b) 0.005 M min^{-1} c) 0.5 M min^{-1} d) 5 M min^{-1}

268. What is the two third life of a first order reaction having $k = 5.48 \times 10^{-14}\ s^{-1}$?

a) 2.01×10^{13} s b) 2.01×10^{12} s c) 4.02×10^{13} s d) 4.02×10^{26} s

269. In a Ist order reaction the concentration of reactant decreases from 800 mol/dm^3 to 50 mol/dm^3 in 2×10^4 s. The rate constant of reaction in S^{-1} is

a) 2×10^4 b) 3.45×10^{-5} c) 1.386×10^{-4} d) 2×10^{-4}

270. The rate constant of a first order reaction is $6.9 \times 10^{-3}s^{-1}$. How much time will it take to reduce the initial concentration to its 1/8th value ?

a) 100s b) 200s c) 300s d) 400s

271. In a reaction, the rate expression is, rate = $K[A][B]^{2/3}[C]^0$, the order of reaction is:

a) 1 b) 2 c) 5/3 d) Zero

272. In the Synthesis of ammonia by Haber process, if 60 moles of ammonia is obtained in one hour, then the rate of disappearance of nitrogen is

a) 30 mol/min b) 6 mol/min c) 0.5 mol/min d) 60 mol/min

273. Half-life period of second order reaction is

a) Proportional to initial concentration of reactants

b) Independent of initial concentration of reactants

c) Inversely proportional to initial concentration of reactants

d) None of the above

274. A reactant with initial concentration 1.386 mol litre^{-1} showing first order change takes 40 minute to become half. If it shows zero order change taking 20 minute to becomes half under the similar conditions, the ratio, $\frac{K_1}{K_0}$ for first order and zero order kinetics will be:

a) 0.5 mol^{-1} litre	b) 1.0 mol/litre	c) 1.5 mol/litre	d) 2.0 mol^{-1} litrte

275. The reason for almost doubling the rate of reaction on increasing the temperature of the reaction system by 10°C is

a) The value of threshold energy increases

b) Collision frequency increases

c) The fraction of the molecules having energy equal to threshold energy increases

d) Activation energy decreases

276. Plot of $\log(a - x)$ versus time t is straight line. This indicates that the reaction is of:

a) Second order	b) First order	c) Zero order	d) Third order

277. The $t_{1/2}$ of the first order reaction is

a) Dependent of initial concentration	b) Directly proportional to initial concentration

c) indirectly proportional to initial concentration	d) Independent of initial concentration

278. In a zero-order reaction for every 10° rise of temperature, the rate is doubled. If the temperature is increased from 10°C to 100°C, the rate of the reaction will become:

a) 64 times	b) 128 times	c) 256 times	d) 512 times

279. The time required for 100% completion of a zero order reaction is

a) ak	b) $\frac{a}{2k}$	c) $\frac{a}{k}$	d) $\frac{2k}{a}$

280. The reaction, $2N_2O_5 \rightleftharpoons 2N_2O_4 + O_2$ is

a) Bimolecular and second order	b) Unimolecular and first order

c) Bimolecular and first order	d) Bimolecular and zero order

281. The thermal decomposition of a compound is of first order. If a sample of the compound decompose 50% in 120 min. What time will it take to undergo 90% decomposition?

a) Nearly 400 min	b) Nearly 45 min	c) Nearly 480 min	d) Nearly 240 min

282. Which one of the following statements for the order of a reaction is incorrect?

a) Order of reaction is always a whole number

b) Order can be determined only experimentally

c) Order is not influenced by stoichiometric coefficient of the reactants

d) Order of reaction is sum of power to the concentration terms of reactants to express the rate of reaction

283. The rate of chemical reaction depends on the nature of chemical reactions, because:

a) The threshold energy level differs from one reaction to another

b) Some of the reactants are solid at room temperature

c) Some of the reactants are coloured

d) All of the above

284. If the rate of reaction between A and B is given by, rate $= K[A][B]^n$, then the reaction is:

a) First order in A

b) nth order in B

c) Overall order is $(1 + n)$

d) All are correct

285. In a reaction, $A + B \longrightarrow$ Product, rate is doubled when the concentration of B is doubled, and rate increases by a factor of 8 when the concentrations of both the reactants (A and B) are doubled, rate law for the reaction can be written as:

a) Rate $= k[A][B]$	b) Rate $= k[A]^2[B]$	c) Rate $= k[A][B]^2$	d) Rate $= k[A]^2[B]^2$

286. Combustion of carbon is exothermic, but coal stored in coal depots does not burn automatically because of:

a) High threshold energy barrier

b) Kinetic stability of coal

c) Higher energy of activation needed for burning

d) All of the above

287. A drop of a solution (volume $= 0.05$ mL) contains 6×10^{-7} mol of H^+. If the rate of disappearance of H^+ is 6.0×10^5 mol/L × s, how long will it take for H^+ to disappear from the drop

a) 8.0×10^{-8} s b) 2.0×10^{-8} s c) 6.0×10^{-6} s d) 2.0×10^{-2} s

288. The inversion of cane sugar into glucose and fructose according to the equation $C_{12}H_{22}O_{11} + H_2O \rightarrow C_6H_{12}O_6 + C_6H_{12}O_6$ is an example of

a) First order reaction b) Third order reaction
c) Second order reaction d) Zero order reaction

289. In the reversible reaction

$$2NO_2 \underset{k_2}{\overset{k_1}{\rightleftharpoons}} N_2O_4$$

The rate of disappearance of NO_2 is equal to

a) $\frac{2k_1}{k_2}[NO_2]^2$ b) $2k_1[NO_2]^2 - 2k_2[N_2O_4]$
c) $2k_2[NO_2]^2 - k_2[N_2O_4]$ d) $(2k_1 - k_2)[NO_2]$

290. For the reaction, $N_2 + 3H_2 \rightleftharpoons 2NH_3$ the rate of change of concentration for hydrogen is -0.3×10^{-4} Ms^{-1}. The rate of change of concentration of ammonia is

a) -0.2×10^{-4} b) 0.2×10^{-4} c) 0.1×10^{-4} d) 0.3×10^{-4}

291. In the reaction; $A + 2B \longrightarrow 3C + D$ which of the following expressions does not describe changes in the concentration of various species as a function of time?

a) $\frac{d[C]}{dt} = \frac{-3d[A]}{dt}$ b) $\frac{3d[D]}{dt} = \frac{d[C]}{dt}$ c) $\frac{3d[B]}{dt} = \frac{-2d[C]}{dt}$ d) $\frac{2d[B]}{dt} = \frac{d[A]}{dt}$

292. The order of reaction, with respect to one of the reacting component Y, is zero. In implies that

a) The reaction is going on at a constant rate.
b) The rate of reaction does not very with temperature.
c) The reaction rate is independent of the concentration of Y.
d) The rate of formation of the activated complex is zero.

293. The rate of disappearance of SO_2 in the reaction; $2SO_2 + O_2 \rightarrow 2SO_3$ is 1.28×10^{-3}g/sec. Then the rate of formation of SO_3 is:

a) 0.64×10^{-3}g/sec b) 0.80×10^{-3}g/sec c) 1.28×10^{-1}g/sec d) 1.60×10^{-3}g/sec

294. For the first order reaction half-life is 14 s, the time required for the initial concentration to reduce to 1/8 of its value is

a) $(14)^3 s$ b) 28 s c) 42 s d) $(14)^2 s$

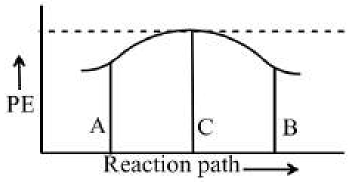
295. Given that K is the rate constant for some order of any reaction at temp. T then the value of $\lim_{T\to\infty} \log K =$ (where A is the Arrhenius constsnt):

a) $A/2.303$ b) A c) $2.303\ A$ d) $\log A$

296. The rate constant of a first order reaction is 3×10^{-6} per second and initial concentration is 0.10 M. Then the initial rate of reaction is

a) $3 \times 10^{-6} ms^{-1}$ b) $3 \times 10^{-8} ms^{-1}$ c) $3 \times 10^{-7} ms^{-1}$ d) $3 \times 10^{-9} ms^{-1}$

297. With respect to the figure given below which of the following statements is correct?



a) E_a for the forward reaction is $C - B$
b) E_a for the forward reaction is $B - A$
c) $E_{a(\text{forward})} < E_{a(\text{backward})}$
d) E_a (for reverse reaction) $= C - A$

298. For a first order reaction, $A \rightarrow$ products, the rate of reaction at $[A] = 0.2$ M is 1.0×10^{-2} mol $L^{-1}min^{-1}$. The half-life period for the reaction is

a) 476 s b) 496 s c) 832 s d) 242 s

299. From the following data, the activation energy for the reaction (cal/mol):

$H_2 + I_2 \longrightarrow 2HI$

T, K	$1T, K^{-1}$	$\log_{10} K$
769	1.3×10^{-3}	2.9
667	1.5×10^{-3}	1.1

a) 4×10^4 b) 2×10^4 c) 8×10^4 d) 3×10^4

300. An elementary reaction is given as $2P + Q \longrightarrow$ products. If concentration of Q is kept constant and concentration of P is doubled then rate of reaction is:

a) Doubled b) Halved c) Quadrupled d) Remains same

301. The hydrolysis of ethyl acetate was carried out separately with 0.05 M HCl and 0.05 M H_2SO_4. The rate constants were found to be k_1 and k_2 respectively then

a) $k_1 < k_2$ b) $k_1 > k_2$ c) $k_1 = k_2$ d) $k_2 = 2k_1$

302. Which one of the following statement for order of reaction is not correct ?

a) Order can be determined experimentally
b) Order of reaction is equal to sum of the power of concentration terms in differential rate law
c) It is not affected with stoichiometric coefficients of the reactants
d) Order cannot be fractional

303. For a single step, reaction, $A + 2B \rightarrow$ Products, the molecularity is

a) 0 b) 1 c) 2 d) 3

304. Which of the following statement is correct for a reaction $X + 2Y \longrightarrow$ Products?

a) The rate of disappearance of $X =$ twice the rate of disappearance of Y
b) The rate of disappearance of $X = 1/2$ rate of appearance of products
c) The rate of appearance of products $= 1/2$ the rate of disappearance of Y
d) The rate of appearance of products $= 1/2$ the rate of disappearance of X

305. For the reaction $A \longrightarrow B$, the rate law is; rate $= K[A]$. Which of the following statements is incorrect?

a) The reaction follows first order kinetics
b) The $t_{1/2}$ of reaction depends upon initial concentration of reactant
c) K is constant for the reaction at a constant temperature
d) The rate law provides a simple way of predicting the concentration of reactants and products at any times after the start of the reaction

306. For a first order reaction $k = 100\ s^{-1}$. The time for completion of 50% reaction is

a) 1 milli-second b) 4 milli-second c) 7 mili-second d) 10 milli-second

307. Pieces of wood burn faster than a log of wood of the same mass because

a) Surface area of log of wood is larger and needs more time to burn
b) Pieces of wood have large surface area
c) All pieces of wood catch fire at the same time
d) Block of wood has higher density than pieces of the same wood

308. Which statement is not correct?

a) For endothermic reactions, heat of reaction is lesser than energy of activation
b) For exothermic reactions, heat of reaction is more than energy of activation
c) For exothermic reactions energy of activation is less in forward reaction than in backward reaction
d) For endothermic reactions energy of activation is more in forward reaction than in backward reaction

309. Which statement is true?

a) Endothermic reactions have higher activation energies than exothermic reactions
b) The specific rate constant for a reaction is independent of the concentration of the reacting species
c) There is a single rate determining step in any reaction mechanism
d) None of the above

310. If $k_1 =$ rate contant at temperature T_1 and $k_2 =$rate constant at temperature T_2 for a first order reaction, then which of the following relations is correct?

(E_a : activation energy)

a) $\log\frac{k_1}{k_2} = \frac{2.303\,E_a}{R}\left(\frac{T_2 - T_1}{T_1T_2}\right)$ b) $\log\frac{k_2}{k_1} = \frac{E_a}{2.303R}\left(\frac{T_2 - T_1}{T_1T_2}\right)$
c) $\log\frac{k_2}{k_1} = \frac{E_a}{2.303R}\left(\frac{T_1T_2}{T_2 + T_1}\right)$ d) $\log\frac{k_1}{k_2} = \frac{E_a}{2.303\,R}\left(\frac{T_1T_2}{T_2 - T_1}\right)$

311. What is the half-life of Na-24 if 2×10^{-4} g sample of it disintegrate at the rate of 7.0×10^{12} atoms per s?
a) 4.97×10^5 s b) 4.97×10^4 s c) 4.97×10^6 s d) 4.97×10^2 s

312. The concentration of a reactant X decreases from 01 M to 0.005 m in 40 min. If the reaction follows first order kinetics, the rate of the reaction when the concentration of X is 0.01 M will be
a) $1.73 \times 10^{-4} M min^{-1}$ b) $3.47 \times 10^{-4} M min^{-1}$
c) $3.47 \times 10^{-5} M\ min^{-1}$ d) $7.5 \times 10^{-4} M\ min^{-1}$

313. For zero order reaction the integrated rate equation is
a) $kt = \frac{[A]}{[A]_0}$ b) $kt = [A] - [A]_0$ c) $[A] = -kt + [A]_0$ d) $[A] = kt - [A]_0$

314. The rate equation for the reaction $2A + B \rightarrow C$ is found to be
Rate=k[A][B]
The correct statement in relation to this reaction that the
a) Unit of k must be s^{-1}
b) $t_{1/2}$ is constant
c) Rate of formation of C is twice the rate of disappearance of A
d) Value of k is independent of the initial concentration of A and B

315. The unit of rate constant of second order reaction
a) $time^{-1}$ b) $conc^{-1} time^{-1}$ c) $conc\ time^{-1}$ d) $conc^{-2} time^{-1}$

316. The rate constant for the first order reaction is 6 s^{-1}. How much time will it take to reduce the concentration of the reactant to $\frac{1}{16}$ th value?
a) 4.6×10^{-2} s b) 4.6×10^4 s c) 4.6×10^2 s d) 4.6×10^{-4} s

317. Rate constant of a chemical reaction can be kept constant by:
a) Stirring the compounds
b) Keeping the temperature constant
c) Both (a) and (b)
d) None of the above

318. The rate of a chemical reaction doubled for every 10°C rise in temperature. If the temperature is increased by 60°C the rate of reaction increase by
a) 20 times b) 32 times c) 64 times d) 128 times

319. If 'a' is the initial concentration of the reactant, the half-life period of the reaction of nth order is proportional to
a) a^{n-1} b) a^n c) a^{1-n} d) a^{n+1}

320. Rate of reaction depends upon
a) temperature b) catalyst c) concentration d) All of these

321. For a reaction, the rate of reaction was found to increase about 1.8 times when the temperature was increased by 10°C. The increase in rate is due to:
a) Increase in number of active molecules
b) Increase in activation energy of reactants
c) Decrease in activation energy of reactants
d) Increase in the number of collisions between reacting molecules

322. A hypothetical reaction $A_2 + B_2 \rightarrow 2AB$ follows the mechanism as given below,
$A_2 \rightleftharpoons A + A\ (fast)$
$A + B_2 \longrightarrow AB + B$ (slow)
$A + B \rightarrow AB$ (fast)

The order of the overall reaction is

a) 2
b) 1
c) $1\frac{1}{2}$
d) 0

323. A chemical reaction proceeds following formula

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

Which of the following process will increase the rate of reaction?

a) Lowering of E_a
b) Lowering of P
c) Lowering of Z
d) Independent of all the above factors

324. In the respect of the equation $k = Ae^{-E_a/RT}$ in chemical kinetics, which one of the following statements is correct?

a) K is equilibrium constant
b) A is adsorption factor
c) E_a is energy of activation
d) R is Rydberg constant

325. For the reaction $2NO_2 + F_2 \longrightarrow 2NO_2F$, following mechanism has been provided,

$NO_2 + F_2 \xrightarrow{\text{Slow}} NO_2F + F$

$NO_2 + F \xrightarrow{\text{Fast}} NO_2F$

Thus, rate expression of the above reaction can be written as:

a) $r = K[NO_2]^2[F_2]$
b) $r = K[NO_2][F_2]$
c) $r = K[NO_2]$
d) $r = K[F_2]$

326. For the reaction:

$[Cu(NH_3)_4]^{2+} + H_2O \rightleftharpoons [Cu(NH_3)_3H_2O]^{2+} + NH_3$, the net rate of reaction at any time is given by, net rate $= 2.0 \times 10^{-4}[Cu(NH_3)_4]^{2+}[H_2O] - 3.0 \times 10^{5}[Cu(NH_3)_3H_2O]^{2+} \cdot [NH_3]$

Then correct statement is/are:

a) Rate constant for forward reaction $= 2 \times 10^{-4}$
b) Rate constant for backward reaction $= 3 \times 10^{5}$
c) Equilibrium constant for the reaction $= 6.6 \times 10^{-10}$
d) All of the above

327. For a reaction between gaseous compounds,

$2A + B \rightarrow C + D$

The reaction rate=k[A][B]. If the volume of the container is made $\frac{1}{4}$ of the initial, then what will be the rate of reaction as compared to the initial rate?

a) 16 times
b) 4 times
c) 1/8 times
d) 1/16 times

328. The rate constant for a first order reaction whose half-life, is 480 s is

a) $2.88 \times 10^{-3} s^{-1}$
b) $2.72 \times 10^{-3} s^{-1}$
c) $1.44 \times 10^{-3} s^{-1}$
d) $1.44 s^{-1}$

329. If X is the total number of collisions which a gas molecule register with others per unit time under particular conditions, then the collision frequency of the gas containing N molecules per unit volume is

a) X/N
b) NX
c) $2\,NX$
d) $NX/2$

330. For a reaction, the rate constant is $2.34 s^{-1}$. The half-life period for reaction is

a) 0.30 s
b) 0.60 s
c) 3.3 s
d) Data is insufficient

331. If "a" and "$t_{1/2}$" are initial concentration of reactant and half-life of a zero order reaction respectively, which of the following is correct ?

a) $t_{1/2} \propto \frac{1}{a}$
b) $t_{1/2} \propto a$
c) $t_{1/2} \propto \frac{1}{a^2}$
d) $t_{1/2} \propto a^2$

332. The temperature dependence of rate constant (k) of a chemical reaction is written in terms of Arrhenius equation, $k = Ae^{-E*RT}$. Activation energy (E^*) of the reaction can be calculated by plotting

a) $\log k \; vs \frac{1}{T}$
b) $\log k \; vs \frac{1}{\log T}$
c) $k \; vs \; T$
d) $k \; vs \frac{1}{\log T}$

333. The rate constant of a reaction is given by $k = 2.1 \times 10^{10} \exp(-2700\, RT)$.

It means that

a) $\log k \; vs \; 1/T$ will be a curved line with slope $= -\frac{2700}{2.303}$

b) $\log k$ vs $1/T$ will be a straight line with intercept on $\log k$ axis $= \log 2.1 \times 10^{10}$

c) The number of effective collisions are 2.1×10^{10} $\text{cm}^{-3}\ \text{s}^{-1}$

d) Half-life of the reaction increases with increase of temperature

334. The unit of the rate of a second order reaction

a) time^{-1} b) $\text{mol L}^{-1}\text{time}^{-1}$ c) $\text{L mol}^{-1}\text{time}^{-1}$ d) $\text{L}^2\ \text{mol}^{-2}\ \text{time}^{-1}$

335. Rate of a reaction can be expressed of by following rate expression
Rate=$k[A]^2[B]$, if concentration of A is reduced by half by what times concentration of B is to be increased to have same rate of reaction?

a) 4 times b) 2 times c) ¼ times d) 8 times

336. Select the intermediate in the following reaction mechanism:

$O_3(g) \rightleftharpoons O_2(g) + O(g)$

$O(g) + O_3(g) \longrightarrow 2O_2(g)$

a) $O_3(g)$ b) $O(g)$ c) $O_2(g)$ d) None of these

337. For a certain reaction of order $'n'$ the time for half change $t_{1/2}$ is given by; $t_{1/2} = \frac{2-\sqrt{2}}{K} \times c_0^{1/2}$, where K is rate constant c_0 is initial concentration. The value of n is:

a) 1 b) 2 c) 0 d) 0.5

338. For a reaction between A and B, the initial rate of reaction is measured for various initial concentrations A and B. the data provided are

	[A]	[B]	Initial reaction rate
1	0.20 M	0.30 M	5×10^{-5}
2	0.20 M	0.10 M	5×10^{-5}
3	0.40 M	0.05 M	1×10^{-5}

The overall order of the reaction is

a) One b) Two c) Two and half d) Three

339. Which order of reaction obeys the relation $t_{1/2} = 1/Ka$?

a) First b) Second c) Third d) Zero

340. How much faster would a reaction proceed at 25°C than at 0°C if the activation energy is 65 kJ?

a) 2 times b) 16 times c) 11 times d) 6 times

341. The activation energy of a reaction at a given temperature is found to be $2.303RT$ J mol^{-1} . The ratio of rate constant to the Arrhenius factor is

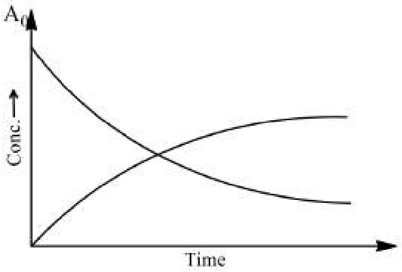
a) 0.01 b) 0.1 c) 0.02 d) 0.001

342. Consider an endothermic reaction $X \rightarrow Y$ with the activation energies E_b and E_f for the backward and forward reactions respectively. In general

a) There is no definite relation between E_b and E_f b) $E_b = E_f$

c) $E_b > E_f$ d) $E_b < E_f$

343. For the reaction $A \longrightarrow nB$, at the point of intersection of two curves show, the $[B]$ is can be given by:



a) $\frac{nA_0}{2}$ b) $\frac{A_0}{n-1}$ c) $\frac{nA_0}{n+1}$ d) $\left[\frac{n-1}{n+1}\right]A_0$

344. The elementary step of the reaction, $2Na + Cl_2 = 2NaCl$ is found to follow III order kinetics, its molecularity is:

a) 1 b) 2 c) 3 d) 4

345. Following mechanism has been proposed for a reaction,

$2A + B \longrightarrow D + E$

$A + B \longrightarrow C + D$ ...(Slow)

$A + C \longrightarrow E$ ...(fast)

The rate law expression for the reaction is:

a) $r = K[A]^2[B]$ b) $r = K[A][B]$ c) $r = K[A]^2$ d) $r = K[A][C]$

346. Two first order reaction have half-lives in the ratio 8 : 1. Calculate the ratio of time intervals $t_1 : t_2$. The time t_1 and t_2 are the time period for (1/4)th and (3/4)th completion

a) 1 : 0.602 b) 2 : 301 c) 0.256 : 0.603 d) 0.2 : 0.301

347. Order of a reaction can be

a) Fractional b) Zero c) Integer d) All of these

348. The half-life period for a zero order reaction is equal to

a) $\frac{2k}{[A]_0}$ b) $\frac{[A]_0}{2k}$ c) $\frac{0.693}{k}$ d) $\frac{0.693}{k[A]_0}$

349. In a reaction $A + B \rightarrow C$, the rate expression is $R = k[A][B]^2$. If the concentration of both the reaction is doubled at constant volume then the rate of the reaction will be

a) Eight time b) Double c) Quadruple d) Triple

350. For a gaseous reaction, the units of rate of rate of reaction are

a) L atm s^{-1} b) atm s^{-1} c) atm mol^{-1} s^{-1} d) mol s^{-1}

351. The rate constant is given by the equation $K = Ae^{-E_a/RT}$ which factor should register a decrease for the reaction to proceed more rapidly?

a) T b) Z c) A d) E_a

352. The activation energy for most of the reactions is approximately 50 kJ mol^{-1}. The value of temperature coefficient for such reactions is

a) > 2 b) >3 c) <1 d) >4

353. The half-life period for a first order reaction is 693 s. The rate constant of this reaction would be

a) 0.1 s^{-1} b) 0.01 s^{-1} c) 0.001 s^{-1} d) 0.0001 s^{-1}

354. For the reaction $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ under certain condition of temperature and partial pressure of the reactants, the rate of formation of NH_3 is 10^{-3}kg hr^{-1}. The rate of conversion of H_2 under same condition is:

a) 1.5×10^{-3}kg hr^{-1} b) 1.76×10^{-4}kg hr^{-1} c) 2×10^{-3}kg hr^{-1} d) 3×10^{-3}kg hr^{-1}

355. In a first order reaction $A \longrightarrow B$, if K is the rate constant initial concentration of the reactant is 0.5 M, then half-life is:

a) $\frac{\ln 2}{K}$ b) $\frac{\ln 2}{K\sqrt{0.5}}$ c) $\frac{\log_{10} 2}{K}$ d) $\frac{0.693}{0.5\,K}$

356. A reaction $A \rightarrow B$ follows a second order kinetics. Doubling the concentration of A will increase the rate of formation of B by a factor of

a) ¼ b) 4 c) ½ d) 2

357. With increase in temperature, rate of reaction

a) increases b) decreases

c) Remains same d) May increase or decrease

358. Which of the following statement is not correct?

a) In zero order reaction the rate of the reaction remains constant throughout.

b) A second order reaction would become a pseudo first order reaction when one of the reactant is taken in large excess.

c) The value of first order rate constant expends on the units of the concentration term used.

d) In a first order reaction the plot of log (a-x) vs time gives a straight line.

359. A drop of solution (volume 0.05 mL) contains 3.0×10^{-6} mole of H^+. If the rate constant of disappearance of H^+ is 1.0×10^{-7}mol $litre^{-1}sec^{-1}$. How long would it take for H^+ in drop to disappear?

a) 6×10^{-8}sec b) 6×10^{-7}sec c) 6×10^{-9}sec d) 6×10^{-10}sec

360. For the reaction, $2A + B \rightarrow$ products, the active mass of B is kept constant, and that of A is doubled. The rate of reaction will be then

a) Decrease 4 times b) Decrease 2 times c) Increase 4 times d) Increase 2 times

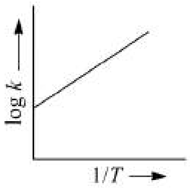
361. Which of the following expression is correct for second order reaction. (C_0 refers to initial concentration of reactant)?

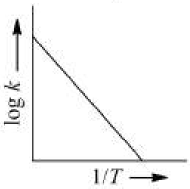
a) $t_{1/2} \propto C_0$ b) $t_{1/2} \propto C_0^0$ c) $t_{1/2} \propto C_0^{-1}$ d) $t_{1/2} \propto C_0^{-2}$

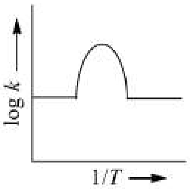
362. A first order reaction is 10% complete in 20 min. The time taken for 19% completion is

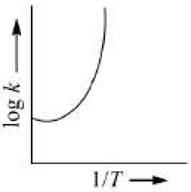
a) 30 min b) 40 Min c) 50 min d) 38 min

363. A graph plotted between $\log k$ *versus* $\frac{1}{T}$ for calculating activation energy is shown by

a)


b)


c)


d)


364. In the following reaction $A \rightarrow B + C$, rate constant is $0.001\ \text{Ms}^{-1}$. Half-life and completion time of the given reaction are

a) 500 s, 1000 s b) 500 s, 750 s c) 250 s, 500 s d) 300 s, 600 s

365. A reaction was found to order with respect to the concentration of carbon monoxide. If the concentration of carbon monoxide is doubled, with everything else kept the same, the rate of reaction will

a) Remain unchanged b) Triple
c) Increases by factor of four d) Double

366. The following data were obtained the first order decomposition of $2\,A(g) \rightarrow B(g) + C(S)$ at a constant volume and at a particular temperature

S N	Time	Total pressure in Pascal
1	At the end of 10 min	300
2	After completion	200

The rate constant in min^{-1} *is*

a) 0.0693 b) 69.3 c) 6.93 d) 6.93×10^{-4}

367. According to Arrhenius equation, the rate constant (k) is related to temperature (T) as

a) $\ln \frac{k_2}{k_1} = \frac{E_a}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$ b) $\ln \frac{k_2}{k_1} = -\frac{E_a}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$
c) $\ln \frac{k_2}{k_1} = \frac{E_a}{R}\left[\frac{1}{T_1} + \frac{1}{T_2}\right]$ d) $\ln \frac{k_2}{k_1} = \frac{E_a}{R}\left[\frac{1}{T_1} + \frac{1}{T_2}\right]$

368. Inversion of cane-sugar in dilute acid is a

a) Bimolecular reaction b) Pseudo-unimolecular reaction
c) Unimolecular reaction d) Trimolecular reaction

369. Consider the reaction

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

The equality relationship between $\frac{d[NH_3]}{dt}$ and $-\frac{d[H_2]}{dt}$ is:

a) $+\frac{d[NH_3]}{dt} = -\frac{3}{2}\frac{d[H_2]}{dt}$
b) $\frac{d[NH_3]}{dt} = -\frac{d[H_2]}{dt}$
c) $\frac{d[NH_3]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt}$
d) $+\frac{d[NH_3]}{dt} = -\frac{2}{3}\frac{d[H_2]}{dt}$

370. For a reaction $aA \longrightarrow bB$ when $[A] = 2.2$ mM, the rate was found to be $2.4\ \mathrm{mMs^{-1}}$. On reducing concentration of $[A]$ to half, the rate changes to $0.6\ \mathrm{mMs^{-1}}$. The order of reaction with respect to A is:

a) 1.5 b) 2.0 c) 2.5 d) 3.0

371. The units of the rate constant of a second order reaction are

a) $mol^{-1}L^{-1}s^{-1}$ b) $mol^{-1}Ls^{-1}$ c) $mol^{-1}Ls$ d) $mol\ L^{-1}s^{-1}$

372. A follows first order reaction , A→ product Concentration of A, change from 0.1 M to 0.025 M in 40 min. find the rate of reaction of A when concentration of A is 0.01 M.

a) $3.47\times 10^{-4}M\ \overset{-1}{\min}$

b) $3.47\times 10^{-5}M\ \overset{-1}{\min}$

c) $1.73\times 10^{-4}M\ \overset{-1}{\min}$

d) $1.73\times 10^{-5}M\ \overset{-1}{\min}$

373. In the reaction $2A + B \rightarrow A_2B$, if the concentration of A is doubled and of B is halved, then the rate of the reaction will

a) Increase by two times b) Decrease by two times

c) Increase by four times d) Remain the same

374. Energy of activation of an exothermic reaction is

a) Negative b) Positive c) Zero d) Can't be predict

375. For a reaction, the rate constant is $2.34s^{-1}$. The half-life period for the reaction is

a) 0.30 s b) 0.60 s c) 3.3 s d) Data is insufficient

376. The rate of a reaction get doubles when the temperature changes from 7°C to 17°C. By what factor will it change for the temperature change from 17°C to 27°C?

a) 1.81 b) 1.71 c) 1.91 d) 1.76

377. Arrhenius equation is

a) $\Delta H = \Delta E + \Delta n_g\ RT$ b) $\Delta G = \Delta H - T.\Delta S$ c) $k = Ae^{-E_a/RT}$ d) None of these

378. Which rate expression suggests an over all order of 0.5 for the reaction involving substances X, Y, Z?

a) Rate $= K[X][Y][Z]$

b) Rate $= K[X]^{0.5}[Y]^{0.5}[Z]^{0.5}$

c) Rate $= K[X]^{1.5}[Y]^{-1}[Z]^{0}$

d) Rate $= K[X][Y]^{0}/[Z]^{2}$

379. Rate of a reaction can be expressed by Arrhenius equation as

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

In this equation, E represents

a) The energy above which all the colliding molecules will react

b) The energy below which colliding molecules will not react

c) The total energy of the reacting molecules at a temperature, T

d) The fraction of molecules with energy greater than the activation energy of the reaction

380. The minimum energy required for a molecule to take part in a reaction is called

a) Threshold energy b) Nuclear energy c) Potential energy d) Kinetic energy

381. The rate of reaction becomes 2 times for every 10°C rise in temperature. How the rate of reaction will increase when temperature is increased from 30°C to 80°C?

a) 16 b) 32 c) 64 d) 128

382. In a gaseous phase reaction:

$A_2(g) \longrightarrow B(g) + (1/2)C(g)$, the increase in pressure from 100 mm to 120 mm is noticed in 5 minute. The rate of disappearance of A_2 in mm $\min^{-1}$ is:

a) 4 b) 8 c) 16 d) 2

383. The unit of rate constant of second order reaction is

a) Mol/Ls b) L / Mol s c) $L^2/\,mol^2$ d) Per second

384. The given reaction,

$2FeCl_3 + SnCl_2 \longrightarrow SnCl_4 + 2FeCl_2$ is an example of:
a) First order reaction
b) Third order reaction
c) Second order reaction
d) None of these

385. For the reaction $A + B \rightarrow C$, it is found that doubling the concentration of A increases the rate by four times and doubling the concentration of B doubles the reaction rate. What is the overall order of the reaction?
a) 4
b) 3/2
c) 3
d) 1

386. The rate constant K_1 of a reaction is found to be double that of rate constant K_2 of another reaction. The relationship between corresponding activation energies of the two reactions at same temperature (E_1 and E_2) can be represented as:
a) $E_1 > E_2$
b) $E_1 < E_2$
c) $E_1 = E_2$
d) None of these

387. For the reaction, $H_2 + I_2 \rightarrow 2HI$, the differential rate law is
a) $-\frac{d[H_2]}{dt} = -\frac{d[I_2]}{dt} = \frac{1}{2}\frac{d[HI]}{dt}$
b) $-\frac{d[H_2]}{dt} = -2\frac{d[I_2]}{dt} = \frac{1}{2}\frac{d[HI]}{dt}$
c) $-\frac{d[H_2]}{dt} = -\frac{d[I_2]}{dt} = \frac{d[HI]}{dt}$
d) $-\frac{d[H_2]}{dt} = -\frac{d[I_2]}{dt} = -\frac{d[HI]}{dt}$

388. The time for half-life of a first order reaction is 1 hr. What is the time taken for 87.5% completion of the reaction?
a) 1 hour
b) 2 hour
c) 3 hour
d) 4 hour

389. DDT on exposure to water decomposes. Half-life is 10 yr. How much time it will take for its decomposition to 99%?
a) 50 yr
b) 70 yr
c) 500 yr
d) 700 yr

390. In Arrhenius equation, $k = Ae^{-E_a/RT}$; A may be called the rate constant at
a) Very low temperature
b) Zero activation energy
c) The boiling temperature of reaction mixture
d) All of the above

391. The phenomenon of emission of visible light as a result of chemical change is known as
a) Chemiluminescence
b) Fluorescence
c) Phosphorescence
d) Photosensitization

392. Chemical reactions with very high E_a values are generally
a) Very fast
b) Very slow
c) Moderately fast
d) Spontaneous

393. In the reaction $3A \rightarrow 2B$, rate of reaction $+\frac{d(B)}{dt}$ is equal to
a) $-\frac{1}{3}\frac{d[A]}{dt}$
b) $-\frac{2}{3}\frac{d[A]}{dt}$
c) $+\frac{2d[A]}{dt}$
d) $-\frac{3}{2}\frac{d[A]}{dt}$

394. If the half-time for a particular reaction is found to be constant and independent of the initial concentration of the reactants, then the ratio is of
a) First order
b) Zero order
c) Second order
d) None of these

395. Under the same reaction condition, initial concentration of 1.386 $mol\ dm^{-3}$ of a substance become half in 40 s and 20 s through first order and zero order kinetics respectively. Ratio $\left(\frac{K_1}{K_0}\right)$ of the rate constants for first order (k_1) and zero order (k_0) of the reaction is
a) 0.5 $mol^{-1}dm^{-3}$
b) 1.0 $mol^{-1}dm^{-3}$
c) 1.5 $mol^{-1}dm^{-3}$
d) 2 $mol^{-1}dm^{-3}$

396. The order of a reaction with rate equal to $kC_A^{3/2}\ C_B^{-1/2}$ is
a) 1
b) $-\frac{1}{2}$
c) $-\frac{3}{2}$
d) 2

397. For reaction $A \rightarrow B$, the rate constant $k_1 = A_1e^{-Ea_1/RT}$ and for the reaction $X \rightarrow Y$, the rate constant $k_2 = A_2e^{-Ea_2/RT}$. If $A_1 = 10^8, A_2 = 10^{10}$ and $Ea_1 = 600$ cal/mol, $Ea_2 = 1800$ cal/mol, then the temperature at which $k_1 = k_2$ is ($R = 2$ cal/k mol)

a) $\frac{200}{3.506}$ K b) $\frac{300}{2.303}$ K c) $\frac{400}{6.506}$ K d) $\frac{200}{5.204}$ K

398. Which of these does not influence the rate of reaction?
a) Nature of the reactants b) Concentration of the reactants
c) Temperature of the reaction d) Molecularity of the reaction

399. The concentration of R in the reaction R→P was measured as a function of time and the following data is obtained

[R](molar)	1.00	0.75	0.40	0.10
t (min)	0.00	0.05	0.12	0.18

The order of the reaction is
a) Zero b) First c) Second d) Third

400. A graph plotted between concentration of reactant consumed at any time (x) and time t is found to be a straight line passing through the origin. Thus, reaction is of:
a) First order b) Zero order c) Third order d) Second order

401. Rate constant of a reaction depends upon
a) Speed of reaction b) Concentration of the reactants
c) Pressure of the surrounding d) Temperature

402. An endothermic reaction $A \rightarrow B$ has an activation energy as x kg mol^{-1} of A. If energy change of the reaction is y kJ, the activation energy of the reverse reactions is
a) $-x$ b) $y - x$ c) $x + y$ d) $x - y$

403. Consider the following statement in respect of zero order reaction.
III. The rate of the reaction is independent of reactant concentration.
IV. The rate of the reaction is independent of temperature.
V. The rate constant of the reaction is independent of temperature.
VI. The rate constant of reaction is independent of reactant cogeneration.
Choose the correct statements/s
a) I only b) I and II only c) III and IV only d) I and IV only

404. If concentration of reactants is increased by $'X'$, the rate constant K becomes:
a) $e^{K/X}$ b) K/X c) K d) X/K

405. The correct expression for the rate of reaction of elementary reaction, $A + B \longrightarrow C$ is:
a) $\frac{d[C]}{dt} = K[A]$ b) $\frac{d[C]}{dt} = K[B]$ c) $\frac{-d[A]}{dt} = K[A][B]$ d) $\frac{-d[A]}{dt} = K[A]$

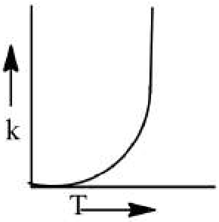
406. The threshold energy of a chemical reaction depends upon:
a) Nature of reacting species
b) Temperature
c) Concentration of species
d) Number of collisions per unit time or collision frequency

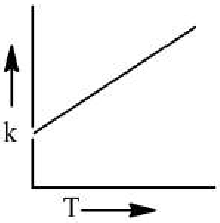
407. A first order reaction has a half-life period of 69.3 sec. At 0.10 mol litre^{-1} reactant concentration, rate will be:
a) 10^{-4} M sec^{-1} b) 10^{-3} M sec^{-1} c) 10^{-1} M sec^{-1} d) 6.93×10^{-1} M sec^{-1}

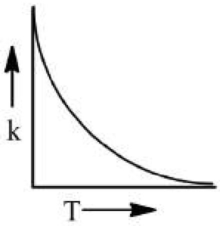
408. The rate constant for the reaction
$2N_2O_5 \rightarrow 4NO_2 + O_2$ is 3.0×10^{-5} s^{-1}. If the rate is 2.40×10^{-5} mol L^{-1} s^{-1}, then the concentration of N_2O_5 is mol L^{-1} is
a) 0.4 b) 0.8 c) 1.2 d) 3.2

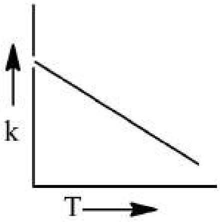
409. The enzyme catalysed reaction is faster than metal catalysed reaction because its activation energy is:
a) Greater b) Lower c) Same d) None of these

410. Plots showing the variation of the rate constant (k) with temperature (T) are given below. The plot that follows Arrhenius equation is

a)


b)


c)


d)


411. For the following homogeneous reaction, the unit of rate constant is$A + B \xrightarrow{k} C$

a) $mol^{-1}Ls^{-1}$ b) s^{-1} c) S d) $s^{-1}mol\ L^{-1}$

412. If $'a'$ is the initial concentration of a substance which reacts according to zero order kinetics and K is rate constant, the time for the reaction to go to completion is:

a) a/K b) $2/Ka$ c) K/a d) $2K/a$

413. What is the two third life of a first order reaction having $= 5.48 \times 10^{-14} s^{-1}$?

a) $2.01 \times 10^{11} s$ b) $2.01 \times 10^{13} s$ c) $8.08 \times 10^{13} s$ d) $16.04 \times 10^{11} s$

414. The velocity constant of a reaction at 290 K was found to be $3.2 \times 10^{-3} s^{-1}$. When the temperature is raised to 310 K, it will be about

a) 6.4×10^{-3} b) 3.2×10^{-4} c) 9.6×10^{-3} d) 1.28×10^{-2}

415. For fourth order reaction, what is the unit of K?

a) $\left(\frac{mol}{L}\right)^{-3} s^{-1}$ b) $\left(\frac{mol}{L}\right)^{+3} s^{-1}$ c) $\left(\frac{mol}{L}\right)^{-3} s$ d) $\left(\frac{mol}{L}\right)^{-3}$

416. The branch of chemistry which deals with the reaction rates and reaction mechanism is called:

a) Thermochemistry b) Photochemistry c) Analytical chemistry d) Chemical kinetics

417. For the reaction; $2N_2O_5 \longrightarrow 4NO_2 + O_2$, rate and rate constant are $1.02 \times 10^{-4} M\ sec^{-1}$ and $3.4 \times 10^{-5} sec^{-1} M$ respectively, then concentration of N_2O_5, at that time will be:

a) $1.732\ M$ b) $3\ M$ c) $1.02 \times 10^{-4} M$ d) $3.5 \times 10^{5} M$

418. In the following first order competing reactions:

$A + \text{Reagent} \longrightarrow \text{Product}$

$B + \text{Reagent} \longrightarrow \text{Product}$

Find the ratio of K_1/K_2 if only 50% of B will have been reacted when 94% of A has been reacted in same time is:

a) 4.06 b) 0.246 c) 2.06 d) 0.06

419. For which order half-life period is independent of initial concentration?

a) Zero b) First c) Second d) Third

420. For a given reaction, pressure of catalyst reduces the energy of activation by 2 kcal at 27°C. The rate of reaction will be increased by:

a) 20 times b) 14 times c) 28 times d) 2 times

421. For a reversible reaction $A \underset{K_2}{\overset{K_1}{\rightleftharpoons}} B$,

Ist order in both the directions, the rate of reaction is given by:

a) $K_1[A]$
b) $-K_2[B]$
c) $K_1[A] + K_2[B]$
d) $K_1[A] - K_2[B]$

422. A first order reaction is 20% complete in 10 min. Calculate the time for 75% completion of the reaction

a) 0.233 min b) 62.18 min c) 112.12 min d) 36.18 min

423. Order of radioactive disintegration reaction is

a) Zero b) First c) Second d) Third

424. Ethylene is produced by cyclobutane as:

$C_4H_8 \xrightarrow{\Delta} 2C_2H_4$

The rate constant is $2.48 \times 10^{-4} \text{sec}^{-1}$. In what time will the molar ratio of the ethylene to cyclobutane in reaction mixture attain the value 1?

a) 27.25 minute b) 28.25 minute c) 25 minute d) 20 minute

425. For a reaction for which the activation energies of forward and reverse reactions are equal?

a) $\Delta H = 0$ b) $\Delta S = 0$ c) The order is zero d) There is no catalyst

426. The half-life period of a first order chemical reaction is 6.93 min. the time required for the completion of 99% of the chemical reaction will be (log 2=0.302)

a) 230.3 min b) 23.03 min c) 46.06 min d) 460.6 min

427. The rate of the elementary reaction, $2NO + O_2 \longrightarrow 2NO_2$, when the volume of the reaction vessel is doubled:

a) Will grow eight times of its initial rate
b) Reduce to one-eight of its initial rate
c) Will grow four times of its initial rate
d) Reduce to one-fourth of its initial rate

428. For the reaction system

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ if the volume of the reaction vessel is reduced to one-third of its original volume, what will be the order of the reaction?

a) Diminished to one fourth of its initial value b) Diminished to one twenty seven of its initial value
c) Increase to twenty seven times of its initial value d) Increase to four times of its initial value

429. The rate constant of a second order reaction is $10^{-2} M\,s^{-1}$. The rate constant expressed in cc. $\text{molecule}^{-1}\text{min}^{-1}$ is:

a) 9.96×10^{-22} b) 9.96×10^{-23} c) 9.96×10^{-21} d) 1.004×10^{-24}

430. Radioactive decay is a

a) First order reaction b) Zero order reaction
c) Second order reaction d) Third order reaction

431. For the decomposition of N_2O_5 at a particular temperature, according to the equations

$2N_2O_5 \rightarrow 4NO_2 + O_2$

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

The activation energies are E_1 and E_2 respectively then

a) $E_1 > E_2$ b) $E_1 < E_2$ c) $E_1 = 2E_2$ d) $E_1 = E_2$

432. The rate of a gaseous reaction is equal to $k[A][B]$. The volume of the reaction vessel containing these gases is reduced by one-fourth of the initial volume. The rate of the reaction would be

a) $\frac{1}{16}$ b) $\frac{16}{1}$ c) $\frac{1}{8}$ d) $\frac{8}{1}$

433. The rate law of the reaction,

$A + 2B \longrightarrow$ Product.

Product is given by $\frac{d[P]}{dt} = K[A]^2 \cdot [B]$. If A is taken in large excess, the order of the reaction will be:

a) Zero b) 1 c) 2 d) 3

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

a) 444 s b) 402 s c) 442 s d) None of these

435. For a certain reaction a plot of $\frac{[c_0 - c]}{c}$ against time t, yields a straight line. c_0 and c are concentrations of reactant at $t = 0$ and $t = t$ respectively. The rate of reaction is:

a) 3 b) 0 c) 1 d) 2

436. The rate constant is doubled when temperature increases from 27°C to 37°C. Activation energy in kJ is

a) 34 b) 54 c) 100 d) 53

437. If the concentration of reactants is increased by $'x'$ then rate constant k becomes

a) $\ln \frac{k}{x}$ b) $\frac{k}{x}$ c) $k + x$ d) k

CHEMICAL KINETICS

: ANSWER KEY :

1)	b	2)	d	3)	b	4)	b	165)	c	166)	b	167)	c	168)	c
5)	b	6)	a	7)	b	8)	d	169)	b	170)	d	171)	d	172)	d
9)	c	10)	c	11)	c	12)	b	173)	b	174)	b	175)	a	176)	a
13)	a	14)	c	15)	a	16)	b	177)	d	178)	a	179)	d	180)	a
17)	c	18)	d	19)	c	20)	d	181)	a	182)	c	183)	c	184)	a
21)	b	22)	b	23)	d	24)	a	185)	a	186)	b	187)	d	188)	d
25)	b	26)	c	27)	b	28)	a	189)	b	190)	a	191)	c	192)	c
29)	a	30)	b	31)	a	32)	c	193)	c	194)	c	195)	a	196)	d
33)	d	34)	d	35)	c	36)	d	197)	d	198)	a	199)	d	200)	b
37)	c	38)	a	39)	a	40)	d	201)	a	202)	d	203)	b	204)	b
41)	c	42)	a	43)	d	44)	b	205)	d	206)	c	207)	b	208)	c
45)	c	46)	c	47)	a	48)	d	209)	a	210)	a	211)	b	212)	a
49)	d	50)	c	51)	d	52)	d	213)	b	214)	b	215)	b	216)	d
53)	d	54)	c	55)	d	56)	a	217)	b	218)	b	219)	d	220)	d
57)	d	58)	a	59)	d	60)	a	221)	c	222)	b	223)	a	224)	a
61)	d	62)	b	63)	c	64)	b	225)	a	226)	d	227)	d	228)	c
65)	b	66)	d	67)	c	68)	c	229)	d	230)	b	231)	b	232)	c
69)	d	70)	c	71)	c	72)	c	233)	d	234)	d	235)	b	236)	a
73)	d	74)	c	75)	a	76)	c	237)	d	238)	a	239)	d	240)	c
77)	b	78)	d	79)	b	80)	c	241)	d	242)	b	243)	d	244)	c
81)	a	82)	a	83)	b	84)	b	245)	a	246)	c	247)	d	248)	c
85)	b	86)	c	87)	d	88)	a	249)	c	250)	b	251)	a	252)	a
89)	a	90)	d	91)	b	92)	b	253)	a	254)	b	255)	b	256)	c
93)	a	94)	b	95)	a	96)	b	257)	c	258)	a	259)	b	260)	b
97)	c	98)	a	99)	c	100)	b	261)	d	262)	b	263)	b	264)	d
101)	c	102)	a	103)	b	104)	d	265)	d	266)	d	267)	b	268)	a
105)	c	106)	c	107)	c	108)	d	269)	c	270)	c	271)	c	272)	c
109)	d	110)	c	111)	a	112)	b	273)	c	274)	a	275)	b	276)	b
113)	d	114)	c	115)	d	116)	b	277)	d	278)	d	279)	c	280)	c
117)	d	118)	a	119)	d	120)	b	281)	a	282)	a	283)	a	284)	d
121)	a	122)	b	123)	b	124)	c	285)	b	286)	d	287)	b	288)	a
125)	b	126)	a	127)	c	128)	d	289)	b	290)	b	291)	d	292)	c
129)	a	130)	d	131)	b	132)	d	293)	d	294)	c	295)	d	296)	c
133)	b	134)	a	135)	c	136)	d	297)	c	298)	c	299)	a	300)	c
137)	d	138)	d	139)	c	140)	a	301)	a	302)	d	303)	d	304)	c
141)	c	142)	c	143)	c	144)	c	305)	b	306)	c	307)	b	308)	b
145)	a	146)	a	147)	b	148)	b	309)	b	310)	b	311)	a	312)	d
149)	c	150)	d	151)	a	152)	d	313)	c	314)	c	315)	b	316)	a
153)	b	154)	a	155)	b	156)	c	317)	b	318)	c	319)	c	320)	d
157)	d	158)	d	159)	c	160)	b	321)	a	322)	c	323)	a	324)	c
161)	a	162)	a	163)	c	164)	a	325)	b	326)	d	327)	a	328)	c

329)	d	330)	a	331)	b	332)	a
333)	b	334)	c	335)	a	336)	b
337)	d	338)	a	339)	b	340)	c
341)	b	342)	d	343)	c	344)	c
345)	b	346)	a	347)	d	348)	b
349)	a	350)	b	351)	d	352)	a
353)	c	354)	b	355)	a	356)	b
357)	a	358)	c	359)	c	360)	c
361)	c	362)	b	363)	b	364)	a
365)	c	366)	a	367)	a	368)	b
369)	d	370)	b	371)	b	372)	a
373)	a	374)	b	375)	a	376)	c
377)	c	378)	c	379)	b	380)	a
381)	b	382)	a	383)	b	384)	b
385)	c	386)	d	387)	a	388)	c

389)	b	390)	b	391)	a	392)	b
393)	b	394)	a	395)	a	396)	a
397)	b	398)	d	399)	a	400)	b
401)	d	402)	d	403)	d	404)	c
405)	c	406)	a	407)	b	408)	b
409)	b	410)	a	411)	a	412)	a
413)	b	414)	d	415)	a	416)	d
417)	b	418)	a	419)	b	420)	c
421)	d	422)	b	423)	b	424)	a
425)	a	426)	c	427)	b	428)	c
429)	a	430)	a	431)	d	432)	b
433)	b	434)	a	435)	d	436)	b
437)	d						

CHEMICAL KINETICS

: HINTS AND SOLUTIONS :

1 **(b)**

For first reaction,

$$E_1 = \frac{2.303RT_1T_2}{(T_1 - T_2)}\log\frac{k_1'}{k_1} \quad \text{... (i)}$$

For second reaction,

$$E_2 = \frac{2.303RT_1T_2}{(T_1 - T_2)}\log\frac{k_2'}{k_2} \quad \text{... (ii)}$$

Given, $E_1 > E_2$

$$\Rightarrow \frac{2.303RT_1T_2}{(T_1 - T_2)}\log\frac{k_1'}{k_1}$$

$$> \frac{2.303RT_1T_2}{(T_1 - T_2)}\log\frac{k_2'}{k_2}$$

$$\therefore \frac{k_1'}{k_1} > \frac{k_2'}{k_2}$$

2 **(d)**

These are the characteristics of effective collisions.

3 **(b)**

Pseudo first order rate constant is doubled as well as rate of reaction is doubled. It may be noted that in presence of acid, hydrolysis of ethyl acetate is a pseudo-unimolecular reaction but the actual value of k depends upon the concentration reaction but the actual value of k depends upon the concentration of H^+ions, otherwise rate constant of a reaction is constant at constant temperature.

4 **(b)**

We know that,$\frac{(t_{1/2})}{(t_{1/2})} = \left[\frac{a_2}{a_1}\right]^{n-1}$

Where, n=order of reaction

Given, $(t_{1/2}) = 0.1s, a_1 = 400$

$(t_{1/2}) = 0.8s, a_1 = 50$

On substituting the values

$$\frac{0.1}{08} = \left[\frac{50}{400}\right]^{n-1}$$

On taking log both sides

$$log\frac{0.1}{0.8} = (n-1)log\frac{50}{400}$$

$$log\frac{1}{8} = (n-1)log\frac{1}{8}$$

0.90=(n-1)0.90

n-1=1

n=2

6 **(a)**

For nth order reaction

$$k = (\text{mol L}^{-1})^{1-n}\ \text{s}^{-1}$$

For Ist order reaction

Unit of $k = \text{s}^{-1}$

For zero order reaction

Unit of $k = \text{mol L}^{-1}\ \text{s}^{-1} = \text{M s}^{-1}$

7 **(b)**

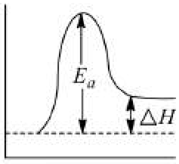
For II order reaction, $t_{1/2} = \frac{1}{Ka}$

8 **(d)**

If $\frac{1}{[A]^2}$vs times are a straight line then order of reaction is third.

9 **(c)**

For an endothermic reaction where ΔH represents the enthalpy of the reaction, the minimum value for the energy of activation is more than ΔH, ie, $E_a > \Delta H$



$E_a > \Delta H$

11 **(c)**

$$T_{\frac{1}{2}} = T_{50}, x = \frac{R}{2}$$

$$\therefore \quad T_{50} = \frac{R}{k_0}$$

So $T_{50} \propto R$

$$T_{50} \propto \frac{R}{k_0}$$

Therefore, the formula of $t_{1/2}$for a zero order reaction is $\frac{[R]_0}{2k}$

12 **(b)**

The curve Y shows a gradual increase in the concentration with time.

13 **(a)**

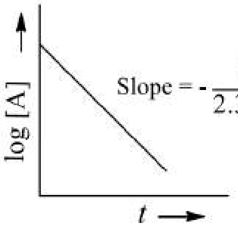
Acid hydrolysis of sucrose is a pseudo unimolecular or pseudo first order reaction. Hydrolysis of sucrose in presence of mineral acid is a biomolecular reaction. But as water is taken in large excess, so the rate of reaction only depends upon concentration of sucrose. Hence, order of the reaction is one.

Therefore, it is called a pseudo first order reaction.

14 **(c)**

For first order reaction

$$\log[A] = -\frac{kt}{2.303} + \log[A]_0$$



15 **(a)**

$r = K[A] = 4 \times 10^{-3} \times 0.02 = 8 \times 10^{-5} M\ \text{sec}^{-1}$

16 **(b)**

For the first order reaction,

$2N_2O_5 \rightarrow 2NO_2 + O_2$

$rate\ \frac{dx}{dt} = k(N_2O_5)\quad ...(i)$

Given,

$\frac{dx}{dt} = 2.400 \times 10^{-5} mol\ L^{-1} s^{-1}$

$k = 3.0 \times 10^{-5} s^{-1}$

$[N_2O_5] = ?$

Substituting these values in (i)

$2.4 \times 10^{-5} = 3.0 \times 10^{-5}[N_2O_5]$

$or\ [N_2O_5] = \frac{2.4 \times 10^{-5}}{3.0 \times 10^{-5}}$

$= 0.8 mol\ L^{-1}$

17 **(c)**

The definition of activation energy.

18 **(d)**

$$2.303 \log\frac{K_2}{K_1} = \frac{E_a}{R}\frac{[T_2 - T_1]}{T_1 T_2}$$

$$\therefore 2.303 \log\frac{K_2}{K_1} = \frac{9}{2 \times 10^{-3}}\left[\frac{10}{298 \times 308}\right];$$

$\therefore \frac{K_2}{K_1} = 1.63; i.e.,$ 63% increase

19 **(c)**

$r = K[A]^1;$

Thus, $7.5 \times 10^{-4} = K \times 0.5;$

$\therefore K = 15 \times 10^{-4} \text{sec}^{-1} = 1.5 \times 10^{-3} \text{sec}^{-1}$

20 **(d)**

$2N_2O_5 \rightleftharpoons 4NO_2 + O_2$

$Rate$

$$= \pm \frac{1}{stoichiometrics\ coefficient}\frac{d[product\ or\ reactant}{dt}$$

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

$$Or\quad -\frac{d[N_2O_5]}{dt} = 2\frac{d[O_2]}{dt}$$

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

$$Or\quad \frac{d[NO_2]}{dt} = \frac{4d[O_2]}{dt}$$

21 **(b)**

For parallel path reaction

$K_{\text{average}} = K_1 + K_2 = 1.26 \times 10^{-4} + 3.8 \times 10^{-5}$

$= 1.64 \times 10^{-4} \text{sec}^{-1}$

Also fractional yield of $B = \frac{K_B}{K_{\text{av}}} = \frac{1.26\times10^{-4}}{1.64\times10^{-4}} =$ 0.7683

Fractional yield of $A = \frac{K_A}{K_{\text{av}}} = \frac{3.8\times10^{-5}}{1.64\times10^{-4}} = 0.2317$

23 **(d)**

For first order : $K = \frac{0.693}{t_{1/2}} = \frac{0.693}{1386} = 5 \times 10^{-4} s^{-1}$

24 **(a)**

Ionic reactions are instantaneous one.

25 **(b)**

For zero order reaction, rate of reaction is independent of concentration i.e., rate of reaction does not depend upon the concentration of reactant.

$\frac{dx}{dt} = k[A]^0$

26 **(c)**

$t_{1/2} = 100s$

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

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

27 **(b)**

The rate law for the reaction is as

$r=\frac{dx}{dt}=k(A)(B)^2(C)^0=k(A)(B)^2$

on increasing the concentration of A,B and C two times.

$r'=\frac{dx}{dt}=k(2A)(2B)^2(2C)^0$

$=8k(A)(B)^2$

Thus, the rate increases eight times.

28 **(a)**

Activation energy is the needed by reactant molecules to gain threshold energy level.

29 **(a)**

The rate of zero order reaction is independent of the concentration of the reactants or the concentration of the reactant do not change with time. Thus, the rate of reaction remains constant.

$$\frac{dx}{dt} = k(a-x)^0 \Rightarrow \frac{dx}{dt} = k$$

Or Rate=k

30 **(b)**

For first order reaction,

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

Where, a= initial concentration

X= change in concentration during time'ť'.

If 75% of the reaction was completed in 32 min, then

$$k = \frac{2.303}{32}\log_{10}\frac{100}{100-75} = \frac{2.033}{32}\log_{10}4$$

$$k = 0.0433\text{min}^{-1}$$

Hence, time required for the completion of 50% reaction.

$$t = \frac{2.303}{0.0433}\log_{10}\frac{100}{100-75}$$

$$= \frac{2.033}{32}\log_{10}2 = 16\ min$$

31 **(a)**

For the reaction :

$CCl_3CHO + NO \rightarrow CHCl_3 + NO + CO$

$$\text{Rate} = \frac{dx}{dt} = k[CCl_3CHO][NO]$$

$$k = \frac{dx}{dt \times [CCl_3CHO][NO]}$$

$$= \frac{\text{mol/L}}{\text{s} \times \text{mol/L} \times \text{mol/L}}$$

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

32 **(c)**

$2A + B \rightarrow C$

Rate of reaction,

$$= -\frac{1}{2}\frac{d[A]}{dt} = -\frac{d[B]}{dt} = \frac{d[C]}{dt}$$

$$\therefore -\frac{d[A]}{dt} = 2\frac{d[C]}{dt}$$

$$= 2 \times 2.2 \times 10^{-3}$$

$$= 4.4 \times 10^{-3} mol\ L^{-1}\text{min}^{-1}$$

33 **(d)**

For third order reaction,

$$\text{rate} = k[A]^3$$

$$\frac{\text{mol L}^{-1}}{\text{s}} = k(\text{mol} - \text{L}^{-1})^3$$

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

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

34 **(d)**

$$CH_3COOC_2H_5 + H_2O \xrightarrow{H+} CH_3COOH + C_2H_5OH$$

Since, in this reaction, water is excess, it is an example of psedo first order reaction (as rate depends only on the concentration of $CH_3COOC_2H_5$).

36 **(d)**

The efficiency of an enzyme in catalyzing a reaction is due to its capacity to lower the activation energy of the reaction

37 **(c)**

The rate of reaction is:

$$= -\frac{1}{2}\frac{d[A]}{dt} = -\frac{d[B]}{dt} = \frac{1}{3}\frac{d[C]}{dt} = \frac{d[D]}{dt}$$

38 **(a)**

For exothermic reaction, activation energy of reverse reaction is greater than activation energy of forward reaction, *ie*, $E_f < E_r$

39 **(a)**

$$2.303\log\frac{K_2}{K_1} = \frac{E_a}{R}\frac{[T_2 - T_1]}{T_1 T_2}$$

$$\therefore \qquad 2.303\log 3 = \frac{E_a}{2}\frac{[313-273]}{313 \times 273}$$

$$\therefore \qquad E_a = 4693\ \text{cal}$$

40 **(d)**

$\ln K = \ln A - \frac{E_a}{RT}$ is Arrhenius equation. Thus plots of $\ln K$ *vs* $1/T$ will give slope $= E_a/R$.

41 **(c)**

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

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

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

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

$$t = 3000\ \text{s}$$

42 **(a)**

K does not change with time; also unit of K suggest it to be II order.

43 **(d)**

Follow review of rate of reaction.

44 **(b)**

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

45 **(c)**

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

or $t_{1/2} = Z(a)^{1-n}$

or $\log_{t_{1/2}} = \log Z + (1-n)\log a$

or $y = c + mx$

Thus, slope $= m = 1-n$ or $1-n=0 \therefore n=1$

and for I order reaction $t_{1/2} = \frac{0.693}{K}$.

46 **(c)**

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

47 **(a)**

A pseudounimolecular reaction.

48 **(d)**

Rate becomes x^y times if concentration is made x time of a reactant giving y^{th} order reaction.

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

Concentration of A is doubled hence x=2 ,y=n

and rate becomes= 2^n times

Concentration of B is halved ,hence x=$\frac{1}{2}$ and y=m

and rate becomes=$\left(\frac{1}{2}\right)^m$ times

Net rate becomes=$(2)^n \left(\frac{1}{2}\right)^m$times

=$(2)^{n-m}$times

49 **(d)**

For the reaction $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$

Rate of reaction $= k[H_2][Br_2]^{1/2}$

Molecularity of reaction $= 1+1=2$

Order of reaction $= 1 + \frac{1}{2} = \frac{3}{2}$

50 **(c)**

When heat energy is supplied, kinetic energy of reactant molecules increase. This will increase the number of collisions and ultimately rate of reaction will be enhanced.

51 **(d)**

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

$$\therefore \quad K = \frac{2.303}{20}\log\frac{1}{0.25}$$

$$= 0.06931 \text{ min}^{-1}$$

52 **(d)**

I step of mechanism B shows I order in both reactants.

53 **(d)**

$K_a = Ae^{-E_a/RT}$ and $K_b = Ae^{-E_b/RT}$

Also, $K_a > K_b$

$E_a < E_b$

Now notice that all the given facts are satisfied.

54 **(c)**

Half-life depends upon rate constant and rate constant (K) varies with temperature as $K = A \cdot e^{-E_n/RT}$; K increase with temperature. Also $t_{1/2} \propto \frac{1}{K}$

55 **(d)**

$Rate = k[NOBr_2][NO]$ ...(i)

But $NOBr_2$is in equilibrium

$$k_{eq} = \frac{[NOBr_2]}{[NO][Br_2]}$$

$[NOBr_2] = k_{eq}[NO][Br_2]$ ...(ii)

Putting the $[NOBr_2]$ in (i)

$rate = k.k_{eq}[NO][Br_2][NO]$

Hence, $rate = k.k_{eq}[NO]^2[Br_2]$

$rate = k'[NO]^2[Br_2]$

where, k'.Keq

the order, of reaction with respect to NO(g) is 2

56 **(a)**

For zero order reaction, for example,

$A \rightarrow B$

$$\frac{-d[A]}{dt} = k[A]^0$$

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

57 **(d)**

The increase in collision frequency brings in an increase in effective collisions and thus, rate of reaction increases.

58 **(a)**

$$t_{1/2} \propto \frac{1}{a^{n-1}}$$

When $n = 4$

$$t_{1/2} \propto \frac{1}{a^3}$$

Hence, order of reaction = 4

59 **(d)**

There are two different reactants (say A and B).

$A + B \rightarrow product$

Thus, it is a bimolecular reaction .

If $\frac{dx}{dt} = k[A][B]$

It is second order reaction

If $(\frac{dx}{dt}) = k[A]$

Or =k[B]

It is first order reaction .

Molecularity is independent of rate ,but is the sum of the reacting substance thus it cannot be unimolecular reaction .

60 **(a)**

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

$\therefore$ $10^{-2} = K[1][1]^2$

or $K = 10^{-2}$ litre^{2}mol^{-2}sec^{-1}

$\therefore$ rate II = $10^{-2}[0.5] \times [0.5]^2$

$= 1.2 \times 10^{-3}$mol/litre-sec

61 **(d)**

If $E_a = 0, k = Ae^{-E_a/RT} = Ae^0 = A$

Hence, k becomes independent of T

62 **(b)**

Larger is surface area, more is rate of reaction.

63 **(c)**

Reactions having lower energy of activation occurs more fast under similar experimental conditions.

64 **(b)**

For the first order reaction

$Rate\left(\frac{dx}{dt}\right) = k[A]$

[A]→concentration of reactant

K→rate constant

Given that,

$\frac{dx}{dt} = 1.5 \times 10^{-2} mol\ L^{-1}\text{min}^{-1}$

K=? and [A]=0.5 M

$1.5 \times 10^{-2} = k \times 0.5$

$k = \frac{1.5\times10^{-2}}{0.5}$ $=3 \times 10^{-3}\text{min}^{-1}$

For first order reaction,

Half-life period $t_{1/2} = \frac{0.693}{k} = \frac{0.693}{3\times10^{-2}}$

=23.1min

65 **(b)**

Temperature coefficient,

$= \frac{k_t + 10}{k_t}$

$2 = \frac{10^{-3}}{k_t}$

$k_t = \frac{10^{-3}}{2} = \frac{10 \times 10^{-4}}{2} = 5 \times 10^{-4}$

66 **(d)**

The minimum energy required by reaction molecules to undergo reaction is called activation energy.

67 **(c)**

For an nth order reaction

$t_{1/2} \propto \frac{1}{a^{n-1}}$

For 1st order reaction ,$t_{1/2} \propto \frac{1}{a^{1-1}} \propto a^0$

68 **(c)**

For every 10°C rise of temperature, rate is doubled. Thus, temperature coefficient of the reaction=2

When temperature is increased by 50°C, rate becomes

$=2^{50/10} = 2^5$ times=32 times

69 **(d)**

Order may or may not be equal to molecularity.

70 **(c)**

Use $r = K[A]^m[B]^n$

71 **(c)**

The reaction occurring in two steps has two activation energy peaks. The first step, being fast needs less activation energy. The second step being slow, needs more activation energy. Therefore, second peak will be higher than the first

73 **(d)**

$r_1 = K[A]^1$; $r_2 = K[A]^2, r_3 = K[A]^3$

if $[A] >$ t; $r_3 > r_2 > r_1$

74 **(c)**

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

If $t = t_{1/4}$; $x = a/4$

$\therefore t_{1/4} = \frac{2.303}{k} log \frac{a}{(a - 1/4)}$

$\frac{2.303}{k} \log \frac{4}{3}$

75 **(a)**

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

76 **(c)**

$100 \xrightarrow{2\ days} 50$

$50 \xrightarrow{4\ days} 25$

$25 \xrightarrow{8\ days} 12.5$

Hence, the order of reaction is second.

For second order reaction,

$$k = \frac{1}{2}\left[\frac{x}{a(a-x)}\right] = \frac{1}{2}\left[\frac{50}{100 \times 50}\right]$$
$$= \frac{1}{200}$$
$$t_{1/2} = \frac{1}{k.a}$$
$$\Rightarrow = \frac{1}{1/200.100}$$
$$= \frac{200}{100} = 2days$$

77 **(b)**

$$\frac{1}{3}\frac{d[\mathrm{Br}_2]}{dt} = -\frac{1}{5}\frac{d[\mathrm{Br}^-]}{dt}$$

78 **(d)**

The reaction is said to be of second order if its reaction rate is determined by the variation of two concentration terms of reactants.

$CH_3COOCH_3 + NaOH \rightarrow CH_3COONa + H_2O$

Is an example of second order reaction.

79 **(b)**

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

81 **(a)**

Rate is $\frac{dx}{dt}$ or $-\frac{d[A]}{dt}$;

Where x stands for product concentration and $[A]$ stands for reactant concentration. It continuously decreases with decrease in concentration of reactant with time.

82 **(a)**

For zero order reaction, $t_{1/2} \propto [R]_0$

83 **(b)**

Effect of temperature on reaction rate is given by Arrhenius equation

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

84 **(b)**

This is Arrhenius equation.

85 **(b)**

Let ,initial concentration=a

Final concentration=a-$\frac{2}{3}$a=$\frac{a}{3}$

$$t_{\frac{2}{3}} = \frac{2.303}{k} log \frac{a}{a/3}$$
$$= \frac{2.303}{5.48 \times 10^{-14}} \log 3$$

=2.01× 10^{13}s

86 **(c)**

Let the order with respect to A and B is x and y respectively.

Hence,

$Rate\ r = [A]^x[B]^y$ ...(i)

On doubling the concentration of A, rate increases 4 times,

$4r = [A]^x[B]^y$ ... (ii)

From Eqs. (i) and (ii)

$$\frac{1}{4} = \left(\frac{1}{2}\right)^x$$

∴ X=2

∴ order with respect to A is 2

If concentration of A and B both are doubled,

$8r = [2A]^x[2B]^y$ ... (iii)

From Eqs. (i) and (iii), we get

$$\frac{1}{8} = \frac{1}{(2)^x} \cdot \frac{1}{(2)^y} \qquad [\because x = 2]$$
$$\frac{1}{8} = \frac{1}{(2)^2} \cdot \frac{1}{(2)^y}$$
$$\frac{1}{8} = \frac{1}{4 \times 2^y}$$
$$2^y = 2$$

∴ Y=1

Hence , differential rate equation is

$r \propto [A]^2[B]^1$ or $\frac{dC}{dt} = kC_A^2 \times C_B$

[Where, C_A and C_B=concentrations of A and B]

87 **(d)**

$r = k[A]^n$ ...(i)

When concentration is doubled then

$4r = k(2A)^n$ ...(ii)

Divide Eq. (ii) by (i)

$4 = 2^n$

$n = 2$

88 **(a)**

$$t = \frac{0.693}{k} log \frac{[A]_0}{[A]}$$
$$= \frac{2.303}{60} log \frac{a}{\frac{a}{16}} = \frac{2.303}{60} \log 16$$
$$= \frac{2.303}{60} \times 1.204$$
$$= 0.0462s$$
$$= 4.6 \times 10^{-2}s$$

89 **(a)**

From the unit of rate constant ($i.e., s^{-1}$), it is clear that the reaction is of first order.

$2N_2O_5 \rightarrow 4NO_2 + O_2$

Hence, for first order reaction ,

$$k = \frac{2.303}{t} log \frac{p_0}{p_t}$$

$$\therefore 3.38 \times 10^{-5} = \frac{2.303}{10 \times 60} log \frac{500}{p_t}$$

Or $log \frac{500}{p_t}$=0.00880

$$\therefore \quad \frac{500}{p_t} = anti \log 0.00880$$

=1.02

$$p_t = \frac{500}{1.02} = 490\ atm$$

91 **(b)**

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

$$\because x = \frac{3}{4} a$$

$$\therefore \ t = \frac{2.303}{k} \log \frac{a}{a - \frac{3}{4} a}$$

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

92 **(b)**

$2N_2O_5 \rightarrow 4NO_2 + O_2$

Rate of decomposition of $N_2O_5 = -\frac{1}{2}\frac{k[N_2O_5]}{dt}$

Rate of formation of $NO_2 = \frac{1}{4}\frac{k[NO_2]}{dt}$

$$\therefore \frac{\text{rate of decompsition of } \mathrm{N_2O_5}}{\text{rate of formation of } \mathrm{NO_2}} = \frac{\frac{1}{2}\mathrm{k}\frac{[\mathrm{N_2O_5}]}{\mathrm{dt}}}{\frac{1}{4}\mathrm{k}\frac{[\mathrm{NO_2}]}{\mathrm{dt}}}$$

$$or \quad \frac{1}{2} k \frac{[N_2O_5]}{dt} \times \frac{4}{1} \frac{dt}{k[NO_2]}$$

$$= \frac{4}{2} = \frac{2}{1} = 2:1$$

93 **(a)**

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

Given, reaction is75% completed is 32 min

A=100,x=75

$\therefore \quad k = \frac{2.303}{32} log \frac{100}{100-75} \quad$ ...(1)

For 50% completion of reaction

A=100, x=50

$\therefore \quad k = \frac{2.303}{t} log \frac{100}{100-50} \quad$ ...(2)

$\because \quad LHSof\ Eq.(1) = Eq.(2)$

$\therefore \quad RHSof\ Eq.(1) = Eq.(2)$

$$\therefore \quad \frac{2.303}{32} log \frac{100}{100-75} = \frac{2.303}{t} log \frac{100}{100-50}$$

$$or \frac{2.303}{32} \log 4 = \frac{2.303}{t} \log 2$$

Or $\frac{t}{32} = \frac{\log 2}{\log 4}$ or $t = \frac{32 \times \log 2}{2 \log 2}$

$\therefore \quad t = 16\ min$

$\therefore \quad reaction\ will\ be\ 50\%\ completed\ in\ 16\ min$

94 **(b)**

Rate $\left(\frac{+d[C]}{dt}\right) = k[A][B]$

Thus, the order of reaction w.r.t. A=1

The order of reaction w.r.t.B=1

Total order of reaction=1+1=2

95 **(a)**

The intersection point indicates that half of the reactant X is converted into Y.

96 **(b)**

At $T_1 = 200\ K, T_2 = 400\ K, k_1 = k,\ k_2 = 10\ k$

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

$$\log \frac{10\ k}{k} = \frac{E_a}{2.303R}\left(\frac{400-200}{400 \times 200}\right)$$

$$E_a = 921.2\ R$$

97 **(c)**

Zero order reactions occur with constant rate.

98 **(a)**

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

Thus, $K = \frac{2.303}{10} \log 8 = (2.303 \times 3 \log 2)/10$

99 **(c)**

For the reaction $A \rightarrow B$

On increasing the concentration of reactant (i.e.,A) by 4 times , the rate of reaction becomes double ,hence order of reaction is$\frac{1}{2}$.

100 **(b)**

The rate of chemical reaction always decreases with time as reaction proceeds due to decrease in number of reactant molecules. Only for zero order reactions the rate of chemical reaction remains same.

101 **(c)**

For a zero order reaction,

$R \rightarrow$ product

Rate=$-\frac{d[R]}{dt} = k[R]^0 = k$

$-d[R] = k.dt$

Integrating the above equation.

$-\int d[R] = k \int dt$

$-[R] = kt + \mathrm{I} \quad$ ...(i)

Where, I is integration constant

At $t = 0, R = [R]_0$

$-[R]_0 = k \times 0 + \mathrm{I}$

$\mathrm{I} = -[R]_0$

Put this value in Eq. (i)

$-[R] = kt - [R]_0$

or $[R] = -kt + [R]_0$

102 **(a)**

For first order reaction,

Half-life period $(t_{1/2}) = \frac{0.693}{k}$

Where, k=rate constant

$(t_{1/2}) = \frac{0.693}{69.3} s^{-1}$

$=0.01\ s^{-1}$

103 **(b)**

For nth order reaction :

$$t_{1/2} \propto \frac{1}{a^{n-1}}$$

For second order reaction

$$t_{1/2} = \frac{1}{ka} = \frac{1}{0.5 \times 0.2} = \frac{100}{10} = 10 \text{ min}$$

104 **(d)**

$r = K[\mathrm{CH_3COCH_3}]^a[\mathrm{Br_2}]^b[\mathrm{H^+}]^c$

$\therefore \quad 5.7 \times 10^{-5}$

$= K[0.30]^a[0.05]^b[0.05]^c \qquad \dots(1)$

5.7×10^{-5}

$= K[0.30]^a[0.10]^b[0.5]^c \qquad \dots(2)$

1.2×10^{-4}

$= K[0.30]^a[0.10]^b[0.10]^c \qquad \dots(3)$

3.1×10^{-4}

$= K[0.40]^a[0.05]^b[0.20]^c \qquad \dots(4)$

By (1) and (2) $a = 1$

By (2) and (3) $b = 0$

By (3) and (4) $c = 1$

$\therefore \qquad r = K[\mathrm{CH_3COCH_3}]^1[\mathrm{Br_2}]^0[\mathrm{H^+}]^1$

105 **(c)**

Unit of rate constant

$= \frac{time^{-1}}{conc^{(n-1)}}$

Where, n=order of reaction

Given, unit of rate constant $= L\ mol^{-1}s^{-1}$

$\therefore\ L\ mol^{-1}s^{-1} = \frac{(s)^{-1}}{(L\ mol^{-1})^{n-1}}$

$= \frac{(s)^{-1}}{(L\ mol^{-1})^{n-1}}$

$= s^{-1}(L\ mol^{-1})^{n-1}$

Or 1=n-1

Or n=2

$\therefore$ order of reaction =2

106 **(c)**

Activation energy of a chemical reaction can be determined by evaluating rate constants at two different temperatures

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

107 **(c)**

Molecularity can never be fractional.

109 **(d)**

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

For this reaction, rate

$(r_1) = k[SO_2]_1^2[O_2]_1 \qquad \dots.(i)$

On doubling the volume of vessel, concentration would be half. Hence,

$\text{Rate}(r_2) = k\left(\frac{[SO_2]_1}{2}\right)^2\left(\frac{[O_2]_1}{2}\right) = \frac{r_1}{8}$

$\frac{r_1}{r_2} = 8:1$

110 **(c)**

$r = k[R\mathrm{Cl}]$

If $[R\mathrm{Cl}] = \frac{1}{2}$, then rate$=\frac{r}{2}$

111 **(a)**

$A \xrightarrow{k_1} B \xrightarrow{k_2} C \xrightarrow{k_3} D$

$\because k_3 > k_2 > k_1$

As k_1 is slowest hence $A \rightarrow B$ is the rate determining step of the reaction

112 **(b)**

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

$$= \frac{2.303}{10}\log_{10}\frac{100}{80}$$

$$= \frac{2.303}{10}[\log 10 - 3\log 2]$$

$$= \frac{2.303}{10}[1 - 3 \times 0.3010]$$

$k = 0.0223$

113 **(d)**

$E_a(A \rightarrow B) = 80 \text{ kJ mol}^{-1}$

Heat of reaction $(A \rightarrow B) = 200 \text{ kJ mol}^{-1}$

For $(B \rightarrow A)$ backward reaction,

$E_a(B \rightarrow A) = E_a(A \rightarrow B) + \text{heat of reaction}$

$= 80 + 200 = 280 \text{kJ mol}^{-1}$

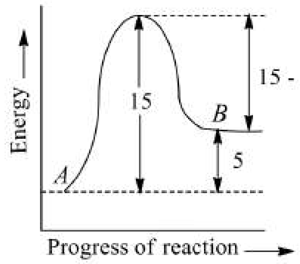
114 **(c)**

For endothermic reaction $A \rightarrow B$

Activation energy $= 15$ kcal/mol

Energy of reaction $= 5$ kcal/mol

Hence, activation energy for the reaction $B \rightarrow A$ is $15 - 5 = 10$ kcal/mol



115 **(d)**

For zero order $[A]_t = [A]_0 - kt$

$$0.5 = [A]_0 - 2 \times 10^{-2} \times 25$$

$$\therefore \quad [A]_0 = 1.0 \text{ M}$$

116 **(b)**

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

Where, k=rate constant=$10^{-3}s^{-1}$

a=initial amount=100

$a-x$=amount left after time t=25

t=time to leave 25% reaction

$$\therefore \ t = \frac{2.303}{10^{-3}} \log \frac{100}{25}$$

$$= \frac{2.303}{10^{-3}} \log 4$$

$$= \frac{2.303 \times 0.6020}{10^{-3}}$$

$= 1386$s

117 **(d)**

By increasing 10 K temperature the rate of reaction becomes double. When temperature is increased from 303 K to 353 K, the rate increases in steps of 10° and has been made 5 times. Hence, the rate of reaction should increases 2^5 times *i. e.*, 32 times.

118 **(a)**

Temperature coefficient

$$= \frac{\text{rate of recation at } 35°\text{C}}{\text{rate of recation at } 25°\text{C}} = 2$$

Thus, increase in rate is two times, when temperature is increased 10°C. Hence, by the increase of 70°C(100-30=70°C), the increase in rate will be

$= (2)^7 \quad \because 70° = 7 \times 10°$

$= 128$ times

119 **(d)**

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

$$\log \frac{k_2}{k_1} = \frac{9000}{2.303 \times 2}\left(\frac{5}{295 \times 300}\right)$$

$$\log \frac{k_2}{k_1} = 0.1103$$

$$\frac{k_2}{k_1} = 1.288, k_2 = 1.288\,k$$

ie, increase by 28.8%

120 **(b)**

$$\frac{1}{2}A \to 2B$$

Remember for $a\,A \to bB$

$$-\frac{1}{a}\frac{d[A]}{dt} = \frac{1}{b}\frac{d[B]}{dt} = \text{Rate of reaction}$$

For the given reaction

$$-\frac{2d[A]}{dt} = \frac{1}{2}\frac{d[B]}{dt} = \text{rate of reaction}$$

Rate of disappearance of A

$$= -\frac{d[A]}{dt} = \frac{1}{2 \times 2}\frac{d[B]}{dt}$$

$$= \frac{1}{4}\frac{d[B]}{dt}$$

121 **(a)**

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

$$= \frac{2.303}{40} \log \frac{0.1}{0.025} = 0.0347$$

$$\therefore r = K \times [A] = 0.0347 \times 0.01$$

$$= 3.47 \times 10^{-4} \ M/min$$

123 **(b)**

For zero order reaction

$$x = kt$$

$$= 0.2 \text{ mol dm}^{-3} \text{ h}^{-1} \times \frac{30}{60} \text{ h}$$

$$= 0.1 \text{ mol dm}^{-3}$$

Now, concentration $= 0.05 \text{ mol dm}^{-3}$

Hence, initial concentration $= 0.1 + 0.05 = 0.15 \text{ mol dm}^{-3}$

124 **(c)**

For the reaction,

$2X + Y \to Z$

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

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

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

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

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

125 **(b)**

$X(g) \to Y(g) + Z(g)$

The reaction is a first order reaction.

Hence ,

$$k = \frac{0.693}{t_{1/2}} = \frac{2.303}{t} log \frac{a}{(a-x)}$$

02234 $\frac{0.693}{10} = \frac{2.303}{t} log \frac{a}{a/10}$

Or $t = \frac{2.303 \times 10}{0.693} \times \log 10$

=33min

126 **(a)**

To be solved with the help of formula,

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

$T_1 = 273 + 27 = 300\ \mathrm{K}$

$T_2 = 273 + 67 = 340\ \mathrm{K}$

$$\log \frac{6.9 \times 10^{-3}}{3.45 \times 10^{-5}} = \frac{E_a}{2.303 \times 8.31} \left[\frac{340 - 300}{340 \times 300}\right]$$

$$\log 200 = \frac{E_a}{19.1379} \times \frac{40}{102000}$$

$$2.3010 = \frac{E_a}{19.14} \times \frac{4}{10200},$$

$$E_a = \frac{19.14 \times 10200 \times 2.3010}{4}$$

$= 112304.907\ \mathrm{J} = 112.3\ \mathrm{kJ}$

127 **(c)**

$\frac{dx}{dt} = K(a-x)^2$ is differential form of II order.

Integrate it to get (c).

128 **(d)**

$A^m + B^n \rightarrow ABx$

In this case,

Overall order of reaction $= m + n$

Hence, code 3 is wrong

129 **(a)**

For the first order reaction,

$$t_{1/2} = \frac{1n2}{k}$$

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

130 **(d)**

Order of reaction is an experimentally determined quantity and thus, cannot be predicted from the given equation.

131 **(b)**

The rate for first order reaction is expressed as

$A \rightarrow products$

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

Rate=k[A]

And the rate constant (k) is given as

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

$$or \ -k = \frac{t}{2.303} log \frac{[A_t]}{[A_0]}$$

132 **(d)**

$$t_{1/2} = \frac{1}{(a)^{n-1}}$$

Where, n=order of reaction

a= initial concentration

For first order reaction,

n=1

$$t_{1/2} = \frac{1}{a^{n-1}}$$

$$a = \frac{1}{a^0} = 0$$

Thus for a first order reaction,$t_{1/2}$ is independent of initial concentration.

133 **(b)**

Relation between ($t_{1/2}$)and initial concentration of reactant for $(n-1)$order reaction

$t_{1/2} \propto [R]_0^{2-n}$

134 **(a)**

$$K = \frac{2.303}{40} \log \frac{0.1}{0.025}$$

$\therefore K = 0.03466\ \mathrm{min}^{-1}$

rate $= K \times 0.01 = 0.03466 \times 0.01$

$= 3.47 \times 10^{-4} M\ \mathrm{min}^{-1}$

135 **(c)**

Rate of endothermic reactions increase with increase in temperature while that of exothermic reactions decrease with increasing temperature.

136 **(d)**

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

For half-life period, $x = \frac{a}{2}$

$$t = \frac{2.303}{k} \log_{10} \frac{a}{a - \frac{a}{2}}$$

$$t = \frac{2.303}{k} \log_{10} 2$$

$\because t = t_{1/2}$

$$t_{1/2} = \frac{2.303 \times 0.3010}{k}$$

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

137 **(d)**

$\because$ Rate constant $(k') >$ rate constant (k'')

Greater the rate constant lesser will be the activation energy

$\therefore E_a' < E_a''$

138 **(d)**

The reactant concentration drop from 0.8 to 0.4 M, $i.e.$, 50% takes place in 15 minute.

$$K = \frac{2.303}{15}\log\frac{0.8}{0.4} = \frac{0.693}{15} = 0.0462 \text{ min}^{-1}$$

$$\text{Also, } t = \frac{2.303}{K}\log\frac{0.1}{0.025} = \frac{2.303}{0.0462}\log\frac{0.1}{0.025}$$
$$= 30 \text{ min}$$

139 **(c)**

$$\text{For II order, } t = \frac{1}{Ka}\frac{x}{(a-x)}$$

$$\therefore \ t = \frac{1}{8\times10^{-5}\times1}\left(\frac{0.5}{0.5}\right)$$
$$= 1.25\times10^{-4} \text{ minute}$$

140 **(a)**

The Arrhenius equation can be written as

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

On comparing this equation with standard equation of straight line

$y = mx + c$, we get,

$y = \log k$

$x = \frac{1}{T}$

$m = -\frac{E_a}{2.303R}$

$c = \log A$

Hence, on plotting graph between $\log k$ (y-axis) and $\frac{l}{T}$ (x-axis), we get a line with slope equal to

$$m = -\frac{E_a}{2.303\,R}$$

141 **(c)**

$K_1 = 10^{16}e^{-2000/T}; K_2 = 10^{15}e^{-1000/T}$

if $\quad K_1 = K_2$ then $10^{16}e^{-2000/T}$

$= 10^{15}e^{-1000/T}$

or $\quad \log 10 - \frac{2000}{T} = -\frac{1000}{T}$ or $T = \frac{1000}{2.303}$K

142 **(c)**

As we know that, rate of reaction is directly proportional to concentration of reactant and inversely proportional to the volume of vessel.

$i.e.$, concentration $\alpha \propto \frac{n}{v}$

For a given reaction,

$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$

Rate of reaction $= k[NO]^2[O_2]$

If volume of vessel is reduced by $\frac{1}{3}rd$ of its initial value, then concentration of compound is increase by 3 times. Hence, the rate of reaction will be increased by 27 times.

143 **(c)**

For a zero order reaction

$$k_0 = \frac{[A_0]}{2t_{1/2}}$$

Since, $[A]_0 = 2\,M, t_{1/2} = 1\,h$

So, $k_0 = 1$ and

$$k_0 = \frac{\Delta x}{t} \ or \ t = \frac{0.50 - 0.25}{1} = 0.25\,h$$

144 **(c)**

$A + B \rightarrow C + D$

Rate $(r) = k[A]^a[B]^b$ ...(i)

$2r = k[2A]^a[B]^b$ ... (ii)

$3r = k[A]^a[9B]^b$ ...(iii)

Dividing eq.(ii) by eq.(i)

$2 = 2^a$ or a=1

Dividing eq (iii) by eq.(i)

$3 = 9^b \ or \ 3 = 3^{2b} \ or \ 2b = 1 \ or \ b = 1/2$

So order of reaction $= 1 + \frac{1}{2} = \frac{3}{2}$

145 **(a)**

$$N = \frac{N_0}{2^n}$$

$$n = \frac{T}{t_{1/2}} = \frac{40}{20} = 2$$

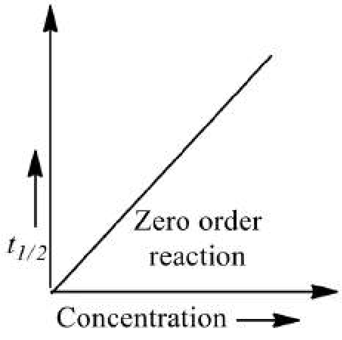
$$\therefore \quad N = \frac{N_0}{2^2} = \frac{N_0}{4}$$

146 **(a)**

$p_1 = 80 \text{ kPa}, (t_{1/2})_1 = 350s$

$p_1 = 40 \text{ kPa}, (t_{1/2})_1 = 175s$

$$\frac{80}{40} = \frac{350}{175} = 2$$



$$\because \frac{p_1}{p_2} = \frac{(t_{1/2})_1}{(t_{1/2})_2} = \frac{a_1}{a_2}$$

$(t_{1/2}) \propto a(zero\ order\ reaction)$

147 **(b)**

Thermal decomposition,

$CH_3CHO \xrightarrow{\Delta} CH_4 + CO$

$$\frac{dx}{dt} = k[CH_3CHO]^{3/2}$$

148 **(b)**

For the reaction : $A \rightarrow 2B + C$

$p_1 \quad 0 \quad 0$

After 10 min $\quad p_1 - p \quad 2p \quad p$

After long time $\quad 0 \quad 2p_1 \quad p_1$

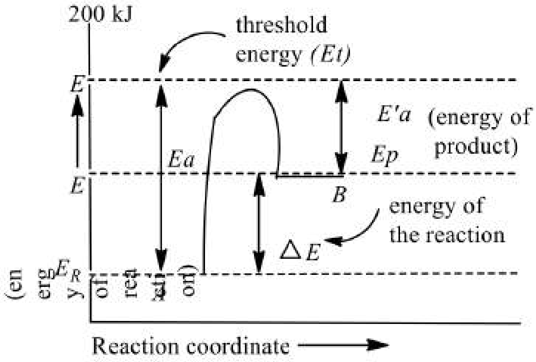
Total pressure $= (p_1 - p + 2p + p) = 176$ mm

Total pressure after long time

$= 2p_1 + p_1 = 270$ mm

Calculate the value of p from above two equation and then the difference of p_1 and p will be the pressure of A

149 **(c)**



Where,

E_a =activation energy of forward reaction

E_a''=activation energy of backward reaction

The above energy profile diagram shows that

$E_a > E_a''$

The potential energy of the product is greater than that of the reactant, so the reaction is endothermic.

$E_a = E_a'' + \Delta E$

$E_t = E_a$ or $E_t > E_a''$

150 **(d)**

Combination of $H_2$and B_{r2} to give HBr is zero order reaction as the rate of reaction is not affected by the concentration of reactants.

$H_2 + B_{r2} \overset{hv}{\rightarrow} 2HBr$

151 **(a)**

$$Rate\ of\ reaction = \frac{1}{4}\frac{d(NO_2)}{dt} = \frac{5.2 \times 10^{-3}}{4 \times 100}$$

$= 1.3 \times 10^{-5}\ Ms^{-1}$

152 **(d)**

$$9 = \left(\frac{3.24 \times 10^{-2}}{1.2 \times 10^{-3}}\right)^n$$

$9 = (3^3)^{2/3}$

order of the reaction is=2/3

153 **(b)**

For zero order reaction

$$k = \frac{a}{2t_{1/2}} = \frac{a}{2 \times 100} = \frac{a}{200}$$

When 80% completion take place

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

$$\frac{a}{200} = \frac{0.80a}{t}$$

$t = 200 \times 0.8 = 160$ min

154 **(a)**

$2A + B \rightarrow product$

[B] is doubled, half-life didn't change

Half-life is independent of change in concentration of reactant i.e., first order

First order w.r.t. to B

When [A] is doubled, rate increased by two times

⇒ First order w.r.t.A

Hence, net order of reaction =1+1=2

Unit for the rate constant=$conc.^{(1-n)}\ t^{-1}$

$= (mol^{-1})^{-1}.s^{-1}$

$L.mol^{-1}s^{-1}$

156 **(c)**

$H_2O + O \longrightarrow 2OH$; $\Delta H = 72$kJ at 500 K;

Given $E_a = 77$ kJmol^{-1}

$2OH \longrightarrow H_2O + O; E_{a_{B.R.}}$

For a reaction $E_{a_{F.R}} = \Delta H + E_{a_{B.R}}$

$\therefore 77 = 22 + E_{a_{B.R}}$

$\therefore E_{a_{B.R}} = 5$kJ mol^{-1}

157 **(d)**

According to Arrhenius equation, the relationship between the activation energy and temperature is

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

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

∴ Activation energy decreases with rise in temperature, thereby increasing the rate of the reaction.

158 **(d)**

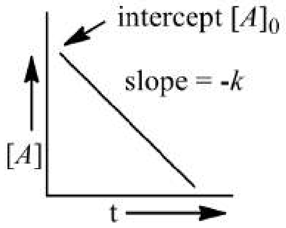
rate $= K[A]^1$

$$K = \frac{2.0 \times 10^{-5}}{0.01} = 2.0 \times 10^{-3}\text{s}^{-1}$$

$$\therefore \quad t_{1/2} = \frac{0.693}{2 \times 10^{-3}} = 347\text{s}$$

159 **(c)**

For a zero order reaction, the plot of concentration of reactant vs time is a straight line (linear) with a negative slope and non-zero intercept.



160 **(b)**

$r = K[A]^m[B]^n$;

Also, $\frac{r}{4} = K[A]^m[2B]^n$

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

∴ $n = -2$

161 **(a)**

The ratio of rate constant when temperature is raised 10°C, is called temperature coefficient. For most of the reaction, it has a value of 2.

Hence, for the given reaction,

Rate constant at 290 K$=3.2 \times 10^{-3}$

∴ Rate constant at 300 K$=2 \times 3.2 \times 10^{-3}$

$= 6.4 \times 10^{-3}$

162 **(a)**

$\frac{dc}{dt}$ represent the change in concentration of reactant with time. As, in a reaction, concentration of reactant always decrease with time hence, rate of reaction is represented as $\frac{-dc}{dt}$.

163 **(c)**

$k = \frac{2.303}{t} log \frac{A_0}{A_t}$

$= \frac{2.303}{2 \times 10^4} log \frac{800}{50}$

$= 1.386 \times 10^4 s^{-1}$

164 **(a)**

For, $N_2 + 3H_2 = 2NH_3$

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

Where, $-\frac{d[N_2]}{dt}$ is rate consumption of N_2(−ve sign)

$-\frac{d[H_2]}{dt}$ is rate of consumption of H_2(−ve sign)

$\frac{+d[NH_3]}{dt}$ is rate of formation of NH_3(+ve sign)

Individual rates become equal when each of these is divided by their respective stoichiometric coefficient.

165 **(c)**

Given, $R_1 = k[A]^2[B]$

According to equation $R_2 = k[3A]^2[B]$

$= k \times 9[A]^2 2[B]$

$= 18 \times k[A]^2[B]$

$= 18R_1$

167 **(c)**

For the reaction,

A+B →C

$Rate = k[A]_0^x.[B]_0^y$

$\frac{Rate_2}{Rate_3} = \frac{k(0.024)^x(0.070)^y}{k(0.024)^x(0.035)^y} = \frac{0.80}{0.10}$

$(2)^y = 8$

$y = 3$

$\frac{Rate_3}{Rate_1} = \frac{k(0.024)^x(0.035)^y}{k(0.012)^x(0.035)^y} = \frac{0.10}{0.10}$

$(2)^x = 1$

$x = 0$

$Rate = k[B]^3$, where, k=rate constant

168 **(c)**

For second order reaction, $\frac{dx}{dt}$ (rate)∝ $[A]^2$

∵ Rate Of reaction increases four times when concentration of reaction is increased two times.

∴ It is second order reaction.

169 **(b)**

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

$\frac{d[H_2]}{dt} = -0.3 \times 10^{-4} ms^{-1}$

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

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

$= -\frac{2}{3} \times (-0.3 \times 10^{-4})$

$= 0.2 \times 10^{-4}$

170 **(d)**

According to collision theory,

1. The reaction rate depends on collision frequency and effective collisions. For a molecule to have effective collision it should fulfill two conditions; proper orientation and sufficient energy.

2. The collision rate *i.e.*, the number of collisions taking place in unit volume is also termed as collision frequency (Z) and is given by

$$z = \frac{\pi n^2 \sigma^2 u_{av}}{\sqrt{2}}$$

3. Greater the temperature, greater will be the collision rate.

171 **(d)**

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + 22$ kcal.

∵The activation energy for the forward reaction = 50 kcal

∴ The activation energy for the backward reaction=50+22=72 kcal.

172 **(d)**

Only those collisions are effective collisions which are energetic enough and cross over the threshold energy level.

173 **(b)**

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{480}\ \mathrm{s^{-1}}$$

$$k = 1.44 \times 10^{-3}\ \mathrm{s^{-1}}$$

174 **(b)**

It is a characteristic of zero order reaction.

175 **(a)**

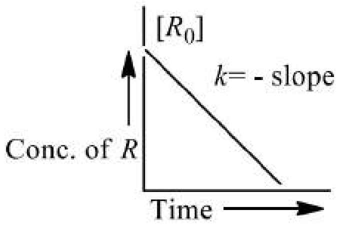
Follow review of order of reaction.

176 **(a)**

Average life is defined as, "reciprocal of decay constant." If decay constant for a reaction is λ then,

Average life=$\frac{1}{\lambda}$

177 **(d)**



[variation in the concentration V_s time plot for a zero order reaction]

178 **(a)**

Energy of activation does not depend on the stoichiometry of change. It is characteristic value for a chemical reaction.

179 **(d)**

Use ; $r = K[A]^m[B]^n$

180 **(a)**

Slow reaction rate indicates higher free energy of activation

181 **(a)**

It is the definition of molecularity.

182 **(c)**

$$\left(\frac{dx}{dt}\right) = k[NO]^2[o_2]$$

$$= k\left(\frac{n_{NO}}{V}\right)^2\left(\frac{n_{O_2}}{V}\right)$$

$$\left(\frac{dx}{dt}\right) = \frac{k}{V^3}(n_{No})^2(n_{O_2})$$

$$\left(\frac{dx}{dt}\right) = \frac{k(n_{No})^2(n_{O_2})}{\left(\frac{V}{2}\right)^3}$$

$$= 8\left(\frac{dx}{dt}\right)$$

183 **(c)**

No doubt order cannot be predicted by merely looking chemical reaction but this can be treated as standard example of II order reaction.

184 **(a)**

$$\text{Rate} = -\left[\frac{dc}{dt}\right] = \left[-\frac{dn}{dt}\right]\frac{1}{V} \qquad \left[\because c = \frac{n}{V}\right]$$

$$\therefore \qquad -\left[\frac{dc}{dt}\right] = -\frac{1}{RT}\left[\frac{dP}{dt}\right]. \qquad \left[c = \frac{P}{RT}\right]$$

185 **(a)**

For zero order reaction

Rate $=[A]^0 = k$

$$\frac{mol\ L^{-1}}{s} = k$$

K=$mol\ L^{-1}\ s^{-1}$

186 **(b)**

$$A \xrightarrow{k_1} B,$$

$$A \xrightarrow{k_2} C,$$

By Arrhenius equation,

$R_1 = A'e^{-E_a 1/RT}$and $k_2 = A'e^{-E_a 2/RT}$

(A' is Arrhenius constant) (Since, $E_{a2} = 2E_{a1}$)

$\therefore k_2 = A'e^{-2E_{a'}|RT}$

$$\frac{k_1}{k_2} = \frac{A'e^{-E_a 1|RT}}{A'e^{-2E_a|RT}} = e^{E_a 1|RT}$$

$\therefore k_1 = k_2 e^{E_a 1/RT}$

187 **(d)**

For the reaction, $2A + B \rightarrow A_2B$

According to rate laws,

Rate ∝ concentration of reactants

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

Where, k=rate constant

188 **(d)**

This is activation state and orientation concept for mechanism of reactions.

189 **(b)**

Rate depends upon the slowest step. Hence, from equation

$O + O_3 \rightarrow 2O_2$

$r = k[O_3][O]$

And from equation $O_3 \rightleftharpoons O_2 + O$

$$K_{eq} = \frac{[O_2][O]}{[O_3]}$$

$$[O] = \frac{K_{eq}[O_3]}{[O_2]}$$

$$\therefore r = k[O_3]\frac{K_{eq}[O_3]}{[O_2]}$$

$= k'[O_3]^2[O_2]^{-1}$

190 **(a)**

Amount of A left in n_1 halves $= \frac{[A_0]}{2^{n_1}}$

Amount of B left in n_2 halves $= \frac{[B_0]}{2^{n_2}}$

Also if $\frac{[A_0]}{2^{n_1}} = \frac{[B_0]}{2^{n_2}}$ when A decays to n_1 halves and B decays to n_2 halves.

$\because \quad [A_0] = 4[B_0]$

$$\therefore \quad 4 = \frac{2^{n_1}}{2^{n_2}} = (2)^{n_1 - n_2}$$

or $\quad (n_1 - n_2) = 2$

$\therefore \quad n_2$

$= n_1 - 2$ ... (i)

Now, $\quad T = n_1 \times t_{1/2A}$ and $T = n_2 \times t_{1/2B}$

$$\therefore \quad \frac{n_1 \times t_{1/2A}}{n_2 \times t_{1/2B}} = 1$$

$$\text{or} \quad \frac{n_1 \times 5}{n_2 \times 15} = 1$$

$$\text{or} \quad \frac{n_1}{n_2}$$

$= 3$ ... (ii)

$\therefore$ By Eqs. (i) and (ii) $n_1 = 3, n_2 = 1$

Thus, $\quad T = 3 \times 5 = 15$ minute

191 **(c)**

$\because$ On doubling the concentration of A, the rate of reaction becomes two times.

$\therefore$ The order of reaction w.r.t. A is 1

$\because$ On doubling the concentration of B, the rate of reaction does not change.

$\therefore$ the order of reaction respect to B is 0

$\because$ on doubling the concentration of C, the rate of reaction becomes four times

$\therefore$ the order of reaction w.r.t. C is 2

$\therefore$ the overall order of reaction=1+0+2=3

192 **(c)**

For nth order; unit of rate constant may be derived by

$$K = \frac{\text{rate}}{[\text{reactant}]^n}$$

193 **(c)**

$r = K[N_2O_5] = 6.2 \times 10^{-4} \times 1.25$

$= 7.75 \times 10^{-4}$M/s

194 **(c)**

A	$\rightarrow$ *product*	
Initially	a	0
After time t	(a-x)	x
After $t_{1/4}$	$\left(a - \frac{a}{4}\right)$	$\frac{a}{4}$

For the first order kinetics ,

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

$$k = \frac{2.303}{t_{1/4}} log\frac{a}{\frac{3a}{4}}$$

$$t_{1/4} = \frac{2.303 log\frac{4}{3}}{k}$$

$$= \frac{0.29}{k}$$

195 **(a)**

The order of reaction is zero. Suppose the following reaction take place .

$A + B \rightarrow product$

$\therefore rate = [A][B]^{-1}$

$\therefore$ order =1+(-1)=0

196 **(d)**

Pseudo first order reactions are those reactions which are not truly first order but show first order kinetics under specific conditions. For examples, acidic hydrolysis of an ester and hydrolysis of cane sugar.

197 **(d)**

The differential rate law for the reaction,

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$ is

$$Rate = -\frac{1}{4}\frac{d[NH_3]}{dt} = -\frac{1}{5}\frac{d[O_2]}{dt}$$

$$= +\frac{1}{4}\frac{d[NO]}{dt} = +\frac{1}{6}\frac{d[H_2O]}{dt}$$

198 **(a)**

$${}_{79}\text{Au}^{198} \xrightarrow{-B} {}_{80}\text{Hg}^{198}$$

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{65}$$

After 260 hr,

$$k = \frac{2.303}{260}\log\frac{a}{a-x}$$

$$\frac{0.693}{65}=\frac{2.303}{260}\log\frac{a}{a-x}$$

$$\frac{a}{a-x}=16$$

$$\frac{1}{1-x}=16$$

$$x=\frac{15}{16}\text{ g}=0.9375\text{ g}$$

199 **(d)**

Rate $=k[NO_2Cl]$

Hence ,rate determining step is

$NO_2CL \rightarrow NO_2+CL$

200 **(b)**

$\mathrm{RCl}+\mathrm{NaOH}\rightarrow \mathrm{ROH}+\mathrm{NaCl}$

Rate $=k[\mathrm{RCl}]$

For this reaction rate of reaction is depends upon the concentration of $R\mathrm{Cl}$

It means, the rate of reaction is halved by reducing the concentration of $R\mathrm{Cl}$ by one half

201 **(a)**

$$\log\frac{k_2}{k_1}=\frac{E_a}{2.303\,R}\times\frac{T_2-T_1}{T_1\cdot T_2}$$

$$\log\frac{105}{100}=\frac{E_a}{2.303\times 8.314}\times\frac{1}{300\times 301}$$

$E_a=36.65$ kJ

202 **(d)**

Molecularity of reaction is simply the number of molecules reacting in balanced chemical equation. It can be simply determined by examining balanced equation.

203 **(b)**

Time required to complete a definite fraction is independent of initial concentration.

204 **(b)**

For this reaction

$2A+B_2 \rightleftharpoons 2AB$

$$k=\frac{[AB]^2}{[A]^2[B_2]}$$

For this reaction,

$A+B_2 \rightleftharpoons AB+B$

$$k_1=\frac{[AB][B]}{[A][B_2]}$$

For this reaction,

$A+B \rightleftharpoons AB$

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

$$k_1\times k_2=\frac{[AB][B]}{[A][B_2]}\times\frac{[AB]}{[A][B]}$$

$$=\frac{[AB]^2}{[A]^2[B_2]}$$

$\therefore k_1\times k_2=k$

205 **(d)**

The collision frequency $(Z)$and rate contant(k) increase on increase of temperature. With the increase in temperature, the average kinetic energy of the molecules increases and this leads to an increase in number of collisions per unit time (Z). The rate constant (k) of a reaction is also increases with increase of temperature.

206 **(c)**

For first order reaction ,

$$\mathrm{k}=\frac{2.303}{t_{1/2}}log\frac{a}{\left(a-\frac{a}{2}\right)}=\frac{2.303}{t}log\frac{a}{(a-x)}$$

$$\frac{2.303}{t_{1/2}}\log 2=\frac{2.303}{1}log\frac{a}{a/4}$$

$$\frac{2.303\times 0.3010}{t_{1/2}}=2.303\times\log 4$$

$$t_{1/2}=\frac{0.3010}{\log 4}=\frac{0.3010}{0.6020}=\frac{1}{2}h$$

207 **(b)**

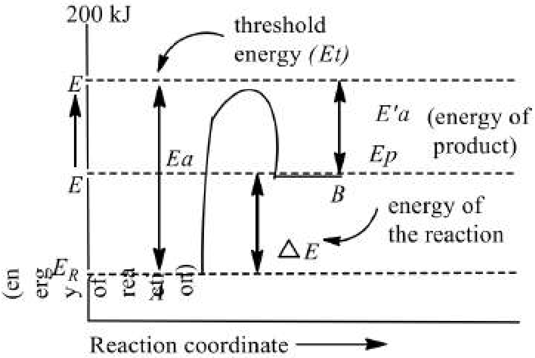
In the given graph,

$E_a=E_a''+\Delta H$

where, E_a'' =activation energy of reverse reaction

$150=E_a''+100$

$E_a''=50$ kJ



209 **(a)**

$T_{50}=15\,min$

$$k=\frac{2.303log2}{T_{50}}=\frac{2.303log2}{15}$$

a=0.1M

$(a-x)=0.025\,M$

For first order reaction,

$$k = \frac{2.303log2}{T_{50}}\log\left(\frac{a}{a-x}\right)$$

$$\frac{2.303log2}{15} = \frac{2\times 2.303log2}{t}log\frac{0.1}{0.025}$$

$$= \frac{2.303}{t}log4$$

$$\therefore \frac{2.303log2}{15} = \frac{2\times 2.303log2}{t}$$

∴ t=30 min

210 **(a)**

The slowest step is the rate determining step. Formation of B(i.e., step I) is the slowest step, therefore step I is the rate determining step.

211 **(b)**

For a reaction E_a for forward reaction $= E_a$ for backward reaction $+\Delta H$,

∴ $85 = A - 20$

or $A = 105$ kJ mol^{-1}.

212 **(a)**

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

213 **(b)**

10 g → 5 g → 2.5 g → 1.25 g takes 30 min in each step, $ie, t^{1/2}$ is independent of 'a' hence, it is a first order reaction

214 **(b)**

$B + 5D \rightarrow 3A + 2C$

$$-\frac{d[B]}{dt} = -\frac{1}{5}\frac{d[D]}{dt} = +\frac{1}{3}\frac{d[A]}{dt} = +\frac{1}{2}\frac{d[C]}{dt}$$

215 **(b)**

$r = K[A]^m$;

Also, $100r = K[10A]^m$;

$$\frac{1}{100} = \left(\frac{1}{10}\right)^m;$$

$\therefore m = 2$

216 **(d)**

$4\mathrm{NH_3(g)} + 5\mathrm{O_2(g)} \rightarrow 4\mathrm{NO(g)} + 6\mathrm{H_2O(g)}$

$$\text{Rate} = -\frac{1}{4}\frac{d[\mathrm{NH_3}]}{dt} = -\frac{1}{5}\frac{d[\mathrm{O_2}]}{dt} = +\frac{1}{4}\frac{d[\mathrm{NO}]}{dt}$$

$$= +\frac{1}{6}\frac{d[\mathrm{H_2O}]}{dt}$$

217 **(b)**

$r = k[A]^\alpha = ka^\alpha$

$1.837\,r = k(1.5a)^\alpha$

Dividing, $1.837 = (1.5)^\alpha$

$\alpha = 1.5$

Hence, order of reaction = 1.5

218 **(b)**

$\mathrm{Cl_2} + 2\mathrm{I^-} \rightarrow \mathrm{I_2} + 2\mathrm{Cl^-}$

Rate of formation of

$$I_2 = \frac{dI_2}{dt} = -\frac{1}{2}\frac{dI^-}{dt}$$

$$= \frac{1}{2}\times\frac{0.20-0.18}{20}$$

$$= \frac{1}{2}\times\frac{0.02}{20}$$

$$= 5\times 10^{-4} mol\ L^{-1}\mathrm{min}^{-1}$$

219 **(d)**

$A_2 + B_2 \rightleftharpoons 2AB$

E_a(forward)= 180 kJ mol^{-1}

E_a(backward)= 200 kJ mol^{-1}

In the presence of catalyst

E_a(forward)=180-100=80 kJ mol^{-1}

E_a(backward)=200-100=100 kJ mol^{-1}

$\Delta H = E_a(\text{forward}) - E_a(\text{backward})$

$= 80 - 100$

$= -20$ kJ mol^{-}

220 **(d)**

Increase in pressure or concentration brings in an increase in collision frequency as well as increase in effective collision. Recall that energy of activation is not at all related with exothermic or endothermic nature.

221 **(c)**

$\frac{K_{t+10}}{K_t} = 2,3$ $i.e.$, temperature coefficient is the ratio of rate constant at two temperatures differing by 10°C, preferably at 35°C and 25°C.

222 **(b)**

Negative sign indicates for the decrease in concentration with time.

223 **(a)**

For first order reaction

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

$$k = \frac{2.303}{45}log\frac{0.05}{0.035}$$

$$k = \frac{2.303}{45}\log 1.42$$

$$k = \frac{2.303\times 0.1523}{45} = 0.0078$$

$$t_{1/2} = \frac{0.693}{0.0078} = 88\ min$$

224 **(a)**

$r = K[A]^n$ where n is order of reaction when $V_F = 2V_{\text{initial}}$,

then $\frac{1}{2}r = K\left[\frac{A}{V}\right]^n$

∴ $2 = 2^n$ or $n = 1$

225 **(a)**

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

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

226 **(d)**

Arrhenius equation gives relation of rate constant with temperature.

$k=A.e^{-E_a/RT}$

On taking logarithm, we get

$$\text{In}\, k=\text{In}\, A-\frac{E_a}{RT}\text{In}\, e$$

or $\text{In}\, k=\text{In}\, A-\frac{E_a}{RT}$

227 **(d)**

$$k_{27°\text{C}}=\frac{0.693}{30}\ \text{min}^{-1},$$

$$k_{47°\text{C}}=\frac{0.693}{10}\ \text{min}^{-1}$$

$$\therefore k_{47°\text{C}}/\ k_{27°\text{C}}=3$$

Or $\log\frac{k_{47°\text{C}}}{k_{27°\text{C}}}=\log 3$

We know that

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

So, $E_a=\frac{\log 3\times 2.303\ R\times T_1T_2}{T_2-T_1}$

$$=\frac{0.4771\times 2.303\times 8.314\times 300\times 320}{20}$$

$=43.84$ kJ

228 **(c)**

Increase in the concentration of the reactants leads to the change in collision frequency because greater the concentration, greater is the collision frequency

229 **(d)**

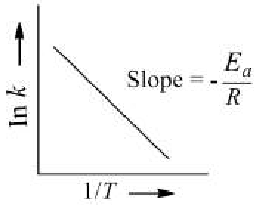
Arrhenius equation is:

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

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

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

$$\text{Slope}=-\frac{E_a}{R}$$



230 **(b)**

Given, $k_1 10^{10}e^{-20,000/T}$

$k_2=10^{12}e^{-24,606/T}$

$k_1=k_2$

$10^{10}e^{-20,000/T}=10^{12}e^{-24,606/T}$

$e^{\frac{-20,000}{T}+\frac{24,606}{T}}=10^2$

$e^{\frac{4,606}{T}}=10^2$

On taking log both sides,

$$\frac{4606}{2.303T}=\log 10^2$$

$$2\log 10\times T=\frac{4606}{2.303}$$

$$T=\frac{4606}{2.303\times 2}$$

$$=\frac{4606}{4.606}=1000\ \text{K}$$

231 **(b)**

The change in molarity $=\frac{n}{V}=\frac{\Delta P}{RT}=\frac{0.8}{0.0821\times 273}=$ 0.0327

$\therefore$ rate of reaction = change in molarity per sec

$$=\frac{0.0327}{50\times 60}=1.09\times 10^{-5}\ \text{mol litre}^{-1}\text{sec}^{-1}$$

233 **(d)**

The rate of reaction varies with time as well as with concentration and pressure.

234 **(d)**

For a first order reaction,

rate = k[reactant]

$$[\text{reactant}]=\frac{\text{rate}}{k}$$

$$=\frac{2.40\times 10^{-5}}{3.0\times 10^{-5}}=0.8$$

235 **(b)**

For first order reaction,

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

$$k_1=\frac{2.303}{10}log\frac{33.05}{33.05-6.25}$$

$=2\times 10^{-3}\text{min}^{-1}$

Constant values of k calculated for different times, shows first order reaction.

236 **(a)**

For I order reaction :

$[A]_t = [A]_0 \times e^{-Kt}$

$= [A]_0 e^{-K\times\frac{1}{K}} = \frac{[A]_0}{e}$

237 **(d)**

Comparing the slope and intercept of the given equation with the following Arrhenius equation :

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

Hence, $\log A = 6$ i.e., $A = 10^6 \mathrm{s}^{-1}$

Comparing slope gives E_a =38.3 kJ/mol

238 **(a)**

The factor $-E_a/RT$ is Boltzmann factor and A is frequency factor.

239 **(d)**

Ionic reactions are instantaneous.

240 **(c)**

$2NO \rightarrow N_2 + O_2$

Rate $=k[NO]^2$

Hence, order of reaction is 2

$H_2 + I_2 \rightarrow 2HI$

Rate $=k[H_2][I_2]$

Hence, order of reaction is (1+1)=2

Therefore these reactions are most likely to be elementary reaction that occurs in one step.

241 **(d)**

$aG + bH \rightarrow Products$

Suppose order of reaction =n

When concentration of both G and H doubled then rate increases by eight times.

$rate = k(reactants)^n$

$(8) = k(2)^n$

$(2)^3 = k(2)^n$

n=3

When concentration of G is doubled keeping the concentration of H fixed, the rate is doubled.

$Rate = [G]^1$

then,

$Rate = [G]^1[H]^2$

242 **(b)**

For a zero order reaction

$A \rightarrow B$

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

Or $-\int_{C_0}^{C_t} d[A] = k\int_{t=0}^{t=0} dt$

$(C_0 - C_t) = kt$

Or $k = (C_0 - C_t)/t$

243 **(d)**

$t_{99.9\%} = \frac{2.303}{k}\log\frac{a}{a-99.9a}$

$t_{50\%} = \frac{0.693}{k}$

(b) $k = \frac{1}{t}\ln\frac{100}{100-99.9} = \frac{1}{t}\ln\frac{100}{0.1}$

Or $\frac{\ln 2}{t_{1/2}} = \frac{1}{t}\ln 10^3$

Or $\frac{\log 2}{t_{1/2}} = \frac{1}{t}\log 10^3 = 3/t$

$t_{1/2} = \frac{\log 2}{3} \times t$

$\therefore t = 10t_{1/2}$

245 **(a)**

Rate constant $k = 1.155 \times 10^{-3} s^{-1}$

$k = \frac{2.303}{t} log\frac{a}{(a-x)} \quad \because a = a, (a-x) = \frac{a}{2}$

$t_{1/2} = \frac{2.303}{t} log\frac{a}{a/2}$

$= \frac{2.303}{1.155 \times 10^{-3}} log2$

$= \frac{2.303}{1.155 \times 10^{-3}} \times 0.3010$

$= \frac{0.693 \times 10^3}{1.155} = 600\ s$

Or

$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{1.155 \times 10^{-3}}$

= 600 s

246 **(c)**

For the reaction

$A + 2B \rightarrow C$

Rate of reaction at a given instant can be represented by

rate $= -\frac{d[A]}{dt} = -\frac{1}{2}\frac{d[B]}{dt} = +\frac{d[C]}{dt}$

247 **(d)**

VSEPR theory is for bonding concept.

248 **(c)**

$A \rightarrow B$

$$\frac{dx}{dt} = k[A]^n$$

$$\frac{dx'}{dt} = k[4A]^n$$

$$\frac{dx'/dt}{dx/dt} = \left(\frac{4}{1}\right)^n$$

$$2 = \left(\frac{4}{1}\right)^n$$

$$n = \frac{1}{2}$$

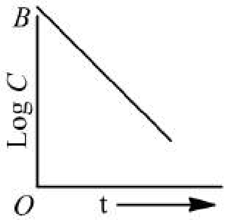
249 **(c)**

Rate law $\frac{dx}{dt} = k[A]^1[B]^2[C]^0 = k[A]^1[B]^2$

Hence, order of reaction $= 1 + 2 = 3$

250 **(b)**

A graph between the log concentration (log c) of reactant and time t for the first order reaction gives a straight line, whose slope is equal to $-\frac{k}{2.303}$



$$\log_{10} C_A = -\frac{kt}{2.303} + \log_{10}(C_A)_0$$

Hence, the order of the above reaction is one.

251 **(a)**

$2FeCl_3 + SnCl_2 \rightarrow 2FeCl_2 + SnCl_4$

It is a third order reaction. As the concentration of both $FeCl_3$ and $SnCl_2$ affect the rate of reaction

$\therefore \text{Rate} = k[FeCl_3]^2[SnCl_2]$

Hence, order of reaction $= 2 + 1 = 3$

252 **(a)**

The order of this reaction over water is zero and in general case it is two. This is an experimental fact.

253 **(a)**

$\text{Rate} = K[\text{sugar}][H_2O]^\circ$.

254 **(b)**

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

$$\therefore -\frac{d[H_2]}{dt} = \frac{3}{2} \times \frac{d[NH_3]}{dt} = \frac{3}{2} \times 2 \times 10^{-4}$$
$$= 3 \times 10^{-4}$$

255 **(b)**

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

256 **(c)**

Rate constant is characteristic constant of a reaction and depends only on temperature and catalyst.

257 **(c)**

Unit of A depends on unit of k.

258 **(a)**

Oppositely charged ions are attracted instantaneously to show reaction.

259 **(b)**

for first order reaction,

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

$$k = \frac{2.303}{1.386 \times 60 \times 60} log \frac{100}{25}$$

$$k = \frac{2.303 \times 0.6021}{1.386 \times 60 \times 60} = 2.8 \times 10^{-4} s^{-1}$$

260 **(b)**

$$\text{Molarity} = \frac{\text{Mole of solute}}{\text{Volume of solution in litre}}; \text{Also } a$$
$$= b \times c_m$$

261 **(d)**

We know that if temperature is increased 10°C then velocity increases 2 times.

$$\frac{r_1 + 10}{r_1} = 2$$

Because temperature increases 100°C hence, rate of reaction will increase 2^{10} times or 1024 times.

262 **(b)**

$$\frac{-d[A]}{dt} = -\frac{1}{2}\frac{d[B]}{dt} = \frac{1}{3}\frac{d[C]}{dt} = \frac{d[D]}{dt}$$

$$\frac{d[C]}{dt} = -\frac{3}{2}\frac{d[B]}{dt} = -\frac{3d[A]}{dt} = \frac{3d[D]}{dt}$$

263 **(b)**

We know, $\frac{t_1}{t_2} = \left[\frac{a_2}{a_1}\right]^{n-1}$

Where, t_1 and t_2 are the half-life periods when the initial concentration are a_1 and a_2

According to question

$t_1 = 0.1s, a_1 = 200$

$t_2 = 0.4s, a_2 = 50$

On substitution the values

$$\frac{0.1}{0.4} = \left[\frac{50}{200}\right]^{(n-1)}$$

On taking log on both sides

$$log \frac{0.1}{0.4} = (n-1) log \frac{50}{200}$$

$$0.60 = (n-1) \times 0.60$$

$$(n-1) = \frac{0.60}{0.60} = 1$$

$$n - 1 = 1$$

$$n = 1 + 1 = 2$$

264 **(d)**

A→ $product$

$$-\frac{d[A]}{dt} = k_1[A]^0$$

Units of $k_1 = Ms^{-1}$

B→ $product$

$$-\frac{d[B]}{dt} = k_2[B]$$

Units of $k_2 = s^{-1}$

265 **(d)**

Choice (c) is an exceptional case. The rate of reaction always increases with increase in temperature.

266 **(d)**

Slowest step is rate determining step, thus, in case (A), rate law is given as rate = $k[Cl_2][H_2S]$

While for the reaction given in case (B), rate law is given as rate= $k[H_2S][Cl_2][H^+]^{-1}$

Hence, only mechanism (A) is consistent with give rate law.

267 **(b)**

Rate of reaction

$$= \frac{1}{2}\left[\frac{-d[A]}{dt}\right] = \frac{1}{2}\left[\frac{0.5 - 0.4}{10}\right]$$

= 0.005 mol litre^{-1}min^{-1}.

268 **(a)**

For 2/3 of a reaction

$$[A]_0 = a, [A] = a - \frac{2}{3}a = \frac{a}{3}$$

$$\because t_{2/3} = \frac{2.303}{k}\log\frac{[A]_0}{[A]}$$

$$= \frac{2.303}{5.48 \times 10^{-14}}\log\frac{a}{a/3}$$

$$= \frac{2.303}{5.48 \times 10^{-14}}\log 3$$

$$t_{2/3} = 2.01 \times 10^{13}\ \text{s}$$

269 **(c)**

First calculate number of half-lives with

$$C_n = \frac{C_0}{2^n}$$

Then calculate $t_{1/2}$ as $t_{1/2} = \frac{\text{total time taken}}{\text{number of half-lives}}$ afterwards

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

$$50 = \frac{800}{2^n}$$

$$2^n = \frac{800}{50} = 16 = 2^4$$

$$n = 4$$

$$T = n \times t_{1/2}$$

$$t_{1/2} = \frac{2 \times 10^4}{4} = 0.5 \times 10^4\ \text{s}$$

$$k = \frac{0.693}{0.5 \times 10^4} = 1.386 \times 10^{-4}$$

270 **(c)**

Given ,N_0=initial concentration =1

N=concentration after time $t = \frac{1}{8}$

K=$6.9 \times 10^{-3}s^{-1}$

$$N = N_0\left(\frac{1}{2}\right)^n$$

Or

$$\frac{1}{8} = 1\left(\frac{1}{2}\right)^n$$

$$\therefore n = 3$$

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{6.9 \times 10^{-3}} = 100s$$

T=$n \times t_{1/2}$

=3× 100

=300s

∴ after 300s it will be reduced to 1/8 of original concentration.

271 **(c)**

Order of reaction is sum of powers raised on concentration terms in order to write rate expression.

272 **(c)**

In Haber's process, ammonia is synthesized as

$$N_2 + 3H_2 \xrightarrow[\text{High Pressure}]{\text{Fe,Mo,300°C}} 2NH_3$$

Rate of synthesis of ammonia=$\frac{60}{60}$

=1 mol/min

∴ Rate of disappearance of nitrogen, i.e.,

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

=$\frac{1}{2}$×= 0.5 mol/min

273 **(c)**

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

For second order reaction, $n = 2$

$$\therefore t_{1/2} \propto \frac{1}{a}$$

274 **(a)**

For I order : $(t_{1/2})_1 = \frac{0.693}{K_1}$

For zero order : $(t_{1/2})_0 = \frac{a}{2K_0}$

$$\therefore \frac{K_1}{K_0} \times \frac{1.386}{2 \times 0.693} = \frac{(t_{1/2})_0}{(t_{1/2})_1}$$

$$\frac{K_1}{K_0} = \frac{20 \times 2 \times 0.693}{40 \times 1.386} = 0.5\ \text{mol}^{-1}\text{litre}$$

275 **(b)**

On increasing the temperature the kinetic energy of the reacting molecules increases and hence, number of collisions increases. So, the rate of reaction will also be increased.

276 **(b)**

For I order reaction,

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

$$y = c + mx$$

277 **(d)**

The half-life period of a first order reaction is independent of the initial concentration of the reactant

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

278 **(d)**

$\frac{r_{(t+10)}}{r_t} = 2$ for each 10° rise in temperature

$$\therefore \frac{r_{100}}{r_{10}} = (2)^9 = 512 \text{ times}$$

279 **(c)**

For zero order reaction,

$[A] = -kt + [A]0$

Where, [A]0= initial concentration=a

[A]=remaining concentration= $a - a = 0$

On putting value of $[A]_0$ and [A],we get

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

280 **(c)**

$2N_2O_5 \rightleftharpoons 2N_2O_4 + O_2$ Bimolecular but of first order.

281 **(a)**

For first order reaction :

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

For Ist case :

$a = 100, x = 50, t = 120$ min

$$k = \frac{2.303}{120} \log \frac{100}{100 - 50}$$

$$= \frac{2.303}{120} \log 2$$

$$k = \frac{2.303 \times 0.3010}{120}$$

$k = 0.0057\ \text{min}^{-1}$

For the IInd case :

$$t = \frac{2.303}{0.0057} \log \frac{100}{100 - 90}$$

$$= \frac{2.303}{0.0057} \log 10$$

$t = 404$ min

≈ 400 min

282 **(a)**

Order of reaction may be fractional.

283 **(a)**

Difference in threshold energy barrier gives different values of energy of activation.

Also $K = Ae^{-E_a/RT}$ and rate $= K[\text{reactant}]^n$

284 **(d)**

Order of reaction is sum of powers raised on concentration terms in order to write rate expression.

285 **(b)**

Let, $r = K[A]^m[B]^n$

$r_1 = K[A]^m[2B]^n$

$r_2 = K[2A]^m[2B]^n$

Also, $\frac{r_1}{r} = 2$ and $\frac{r_2}{r} = 8$ (Given)

Therefore $m = 2$ and $n = 1$

$\therefore \quad r = K[A]^2[B]^1$

286 **(d)**

The higher threshold energy barrier prevents coal to burn spontaneously and provides kinetic stability to fuel.

287 **(b)**

$$[\text{H}^+] = \frac{6 \times 10^{-7}\ \text{mol}}{5 \times 10^{-5}\ \text{L}} = 1.2 \times 10^{-2}\text{M}$$

$$\text{rate} = \frac{dx}{dt} \text{ or } dt = \frac{dx}{\text{rate}} = \frac{1.2 \times 10^{-2}\text{M}}{6 \times 10^5\ \text{M/s}}$$

$$= 2 \times 10^{-8}\ \text{s}$$

289 **(b)**

$$2\text{NO}_2 \underset{k_2}{\overset{k_1}{\rightleftharpoons}} \text{N}_2\text{O}_4$$

$$\text{Rate} = -\frac{1}{2}\frac{d[\text{NO}_2]}{dt}$$

$$= k_1[\text{NO}_2]^2 - k_2[\text{N}_2\text{O}_4]$$

$\therefore$ Rate of disappearance of NO_2

$$ie, -\frac{d[\text{NO}_2]}{dt} = 2k_1[\text{NO}_2]^2 - 2k_2[\text{N}_2\text{O}_4]$$

290 **(b)**

$\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$

$$\frac{d[\text{H}_2]}{dt} = -0.3 \times 10^{-4}\ \text{Ms}^{-1}$$

$$\text{Rate} = -\frac{1}{3}\frac{d[\text{H}_2]}{dt} = +\frac{1}{2}\frac{d[\text{NH}_3]}{dt}$$

$$= \frac{d[\text{NH}_3]}{dt} = -\frac{2}{3}\frac{d[\text{H}_2]}{dt}$$

$= -\frac{2}{3} \times (-0.3 \times 10^{-4})$

$= 0.2 \times 10^{-4}\ \text{Ms}^{-1}$

291 **(d)**

For the given reaction,

$$\frac{-d[A]}{dt} = -\frac{1}{2}\frac{d[B]}{dt} = \frac{1}{3}\frac{d[C]}{dt} = \frac{d[D]}{dt}$$

292 **(c)**

The order of reaction with respect to a reacting compound Y, is zero. It means rate of reaction is independent of its concentration.

293 **(d)**

Rate of formation of SO_3 = rate of disappearance of SO_2

$$= 1.28 \times 10^{-3}\text{g/sec} = \frac{1.28 \times 10^{-3}}{64}\text{M/sec}$$

$$= \frac{1.28 \times 10^{-3}}{64} \times 80\ \text{g/sec}$$

$$= 1.60 \times 10^{-3}\text{g/sec}$$

294 **(c)**

$N = N_0 \times \left(\frac{1}{2}\right)^n$

$\frac{1}{8}N_0 = N_0 \times \left(\frac{1}{2}\right)^n$

N=3

$T = n \times t_{1/2}$

=3 × 14=42s

295 **(d)**

$\log_e K = \log_4 A - \frac{E_a}{RT}$; (Arrhenius eq.)

If $T \to \infty$, then $\log_e K = \log_e A$

296 **(c)**

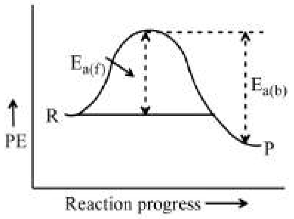
For first order reaction, rate=k[concentration]

$k = 3 \times 10^{-6}$per second, concentration=0.1M

Rate=$3 \times 10^{-6} \times 0.1$

$= 3 \times 10^{-7} ms^{-1}$

297 **(c)**



$i.e., E_{a_{(b)}} > E_{a_{(f)}}$

298 **(c)**

Rate $= k[A]$

$\therefore k = \frac{10^{-2}}{0.2} = 5 \times 10^{-2}$

Now, $t_{1/2} = \frac{0.693}{k} = \frac{0.693}{5 \times 10^{-2}} = 13.86$ min

299 **(a)**

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

$2.9 = \log A - \frac{E_a}{2.303R \times 769}$ ... (i)

$L1 = \log A = \frac{E_a}{2.303R \times 667}$ ... (ii)

Eqs. (i) − (ii)

$$1.8 = \frac{E_a}{R}\left[\frac{1}{667} - \frac{1}{769}\right]$$

$$1.8 = \frac{E_a}{R}\frac{[769 - 667]}{667 \times 769 \times 2.303}$$

$\therefore \quad E_a = 4.17 \times 10^4$ cal/mol.

300 **(c)**

$r = K[P]^2[Q]$

$r_1 = K[2P]^2[Q]$

$\therefore \frac{r_1}{r} = 4$

301 **(a)**

In the presence of acid, hydrolysis of ethyl acetate is a pseudo-unimolecular reaction but the actual value of k depends upon the concentration of H^+ ion. As H_2SO_4 is stronger acid than HCl and moreover H^+ ions produced from 0.05 M H_2SO_4 is double than 0.05 M HCl therefore, $k_1 < k_2$

304 **(c)**

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

305 **(b)**

For first order reaction $t_{1/2} \propto [a]^\circ$.

306 **(c)**

$$t_{50\%} = \frac{0.693}{k} = \frac{0.693}{100}\ \text{s}$$

$= 0.693 \times 10^{-3}$ s

$\approx$ 7 milli second

308 **(b)**

For exothermic reaction $E_a \gtrless \Delta E$. However, rest all are true.

We have

$E_{a_{(f)}} = E_{\text{Threshold}} - E_{\text{Reactant}}$

$E_{a_{(b)}} = E_{\text{Threshold}} - E_{\text{Product}}$

Thus, $E_{a_{(f)}} - E_{a_{(b)}} = E_{\text{Product}} - E_{\text{Reaction}} = \Delta E$

If, $\Delta E = -\text{ve}$ then $E_{a_{(f)}} < E_{(b)}$

If, $\Delta E = +\text{ve}$ then $E_{a_{(b)}} > E_{(f)}$

309 **(b)**

Rate constant is characteristic of a reaction.

310 **(b)**

For a first order reaction, Arrhenius equation is given as

$k = A.e^{-E_a/RT}$

where, E_a =energy of activation

$$\therefore \frac{k_2}{k_1} = \frac{e^{-E_a/RT_1}}{e^{-E_a/RT_2}}$$

$$= e^{E_a/RT_2 - E_a/RT_1}$$

$$= e^{\frac{E_a}{R}}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$= e^{\frac{E_a}{R}}\left(\frac{T_2 - T_1}{T_1 T_2}\right)$$

Taking log on both the sides, we get

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

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

311 **(a)**

24 g of Na has 6.02×10^{23} atoms

2×10^{-4} g of Na will have $= \frac{6.02 \times 10^{23}}{24} \times 2 \times 10^{-4}$ atoms

$= 0.502 \times 10^{19}$ atoms

$$k = \frac{\text{disintegration (in atom s}^{-1})}{\text{no. of atoms present}}$$

$$= \frac{7.0 \times 10^{12}}{0.502 \times 10^{19}}$$

$= 13.94 \times 10^{-7}\ \text{s}^{-1}$

Therefore, $t_{1/2} = \frac{0.693}{13.94 \times 10^{-7}} = 4.97 \times 10^5$ s

312 **(d)**

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

$$= \frac{2.303}{t} log\frac{0.1}{0.005}$$

$$= \frac{2.303}{t} log 20 = 0.075$$

Rate of reaction when concentration of X is 0.01 M will be

$0.075 \times 0.01 = 7.5 \times 10^{-4}\text{min}^{-1}$

313 **(c)**

For a zero order reaction

$A \rightarrow$ product

$$\text{Rate} = \frac{-d[A]}{dt} = k[A]^0 = k$$

$-d[A] = k \cdot dt$

Integrating the above equation

$$-\int d[A] = k\int dt$$

$-[A] = kt + \theta$ ...(i)

At $t = 0, A[A]_0$

$-[A]_0 = k \times 0 + \theta$

$\theta = -[A]_0$

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

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

314 **(c)**

$2A + B \rightarrow C$

Rate=k[A][B]

It represents second order reaction.

Thus , unit of k is $M^{-1}S^{-1}$

∴ (a) is false

T_{50} is dependent of concentration but not constant

∴ (b) is false

$$-\frac{1}{2}\frac{d[A]}{dt} = \frac{d[C]}{dt},$$

thus (C)is correct

315 **(b)**

For *nth* order reaction the unit of k is expressed as

$conc.^{(1-n)}\ time^{-1}$

Hence ,for second order ,the unit of rate constant is as

$conc.^{(1-2)}\ time^{-1}$

$conc.^{-1}\ time^{-1}$

316 **(a)**

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

$$= \frac{2.303}{60}\log\frac{a}{\frac{a}{16}} = \frac{2.303}{60}\log 16$$

$$= \frac{2.303}{60} \times 1.204$$

$= 0.0462$ s

$= 4.6 \times 10^{-2}$s

317 **(b)**

The rate constant of reaction depends upon temperature.

318 **(c)**

If the temperature is increased by 60°C then 10° increase has been made 6 times and therefore, rate will increase by $2^6 = 64$ times

320 **(d)**

Rate of reaction depends upon nature of reactants, concentration of the reactants, temperature and presence of catalyst.

321 **(a)**

No doubt an increase in temperature may increase no. of collisions also but this is the increase in number of effective collisions (with more energy) which decide the rate.

322 **(c)**

From slow step, rate $= k[B_2][A]$

From 1st equation $k_{eq} = \frac{[A]^2}{[A_2]}$

$Or\ [A] = \sqrt{k_{eq}[A_2]} = k_{eq}^{1/2} A_2^{1/2}$

Hence, rate $= k[B_2]\, k_{eq}^{1/2} A_2^{1/2}$

$= k'[A_2]^{1/2}[B_2]$

Hence, order

$= 1\frac{1}{2}$

323 **(a)**

According to formula $k = PZe^{-\frac{E_a}{RT}}$

Lowering of E_a (activation energy), raises the value of k.

324 **(c)**

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

k =rate constant

A =pre-exponential, frequency factor

E_a =activation energy

R =gas constant

T =temperature

325 **(b)**

Slowest step of mechanism decides the rate expression,

Thus, rate $= K[NO_2][F_2]$

326 **(d)**

Net rate of reaction = rate of forward reaction – rate of backward reaction

$$= K_f[\text{reactants}] - K_b[\text{Product}]$$

Also $K_c = \frac{K_f}{K_b}$ at equilibrium.

327 **(a)**

When volume is reduced to one fourth, concentration become four times. Hence, the rate of reaction becomes 16 times as compared to the initial rate.

328 **(c)**

We know that,

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

$or\ k = \frac{0.693}{t_{1/2}} = \frac{0.693}{480}$

$= 1.44 \times 10^{-3} s^{-1}$

329 **(d)**

Number of collisions of one molecules/unit time=X

Number of collision of N molecules/unit time=NX.

As in one collision, two molecules are involved, collision frequency=$NX/2$.

330 **(a)**

K=$2.34 s^{-1}$

Unit of k suggest that the reaction is of 1st order, hence for 1st order reaction,

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

$= \frac{0.693}{2.34} = 0.30s$

331 **(b)**

For zero order reaction integrated rate equation is

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

If $[A]_0 = a, [A] = \frac{a}{2}, t = t_{1/2}$

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

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

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

$\therefore\ t_{1/2} \propto a$

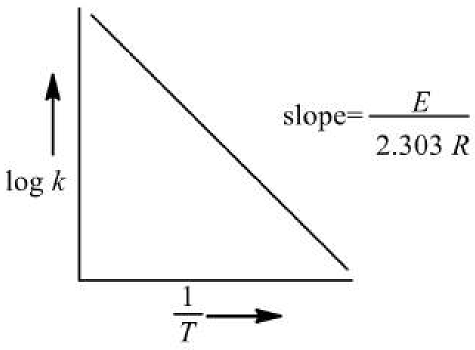
332 **(a)**

Arrhenius equation $k = Ae^{-\frac{E}{RT}}$

In $k = \ln A - \frac{E}{RT}$

or $\log k = \log A - \frac{E}{2.303\,RT}$

Hence, E is calculated with the help of slope of following.



333 **(b)**

$k = 2.1 \times 10^{10} \exp(-2700/RT)$

ie, $\log k$ *vs* $1/T$ will be straight line

Intercept of $\log k$ axis $= \log 2.1 \times 10^{10}$

334 **(c)**

$\frac{dA}{dt} = k[A]^2$

K= reaction rate constant

Then unit of second order reaction rate constant

$k = -\frac{d[A]}{dt} \times \frac{1}{[A]^2}$

$= \frac{1}{s.mol\ L^{-1}}$

$or\ L\ mol^{-1}s^{-1} or\ L\ mol^{-1}t^{-1}$

335 **(a)**

Given, $R_1 = k[A]^2[B]$

Let concentration of B is changed by x times.

Then,

$R_2 = k[A/2]^2[xB]$

$= k \times x/4[A]^2[B]$

$= x/4 \times k[A]^2[B]$

As $R_2 = R_1, thus\ x = 4$

336 **(b)**

The intermediates species is one which is formed and used up during the course of reaction.

337 **(d)**

Integrated rate expression for 1/2 order is

$-2c^{1/2} = Kt - 2c_0^{1/2}$.

338 **(a)**

If the order of reaction w.r.t. A is n and the order of reaction w.r.t. B is m, rate law become

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

From (1)

$5 \times 10^{-5} = [0.20]^n[0.30]^m$ ... (i)

From (2)

$5 \times 10^{-5} = [0.20]^n[0.10]^m$ ... (ii)

From (3)

$1 \times 10^{-4} = [0.40]^n[0.05]^m$ ... (iii)

$or\quad 10 \times 10^{-4} = [0.40]^n[0.05]^m$

from eqs.(i) and (ii)

$\frac{5 \times 10^{-5}}{5 \times 10^{-5}} = \left[\frac{0.20}{0.20}\right]^n \left[\frac{0.30}{0.10}\right]^m$

$1 = (3)^m \Rightarrow (3)^0 = (3)^m \Rightarrow m = 0$

From eqs. (iii) and (iv)

$\frac{5 \times 10^{-5}}{5 \times 10^{-5}} = \left[\frac{0.20}{0.40}\right]^n \left[\frac{0.10}{0.05}\right]^m$

$\frac{1}{2} = \left(\frac{1}{2}\right)^n \times \left(\frac{0.10}{0.05}\right)^0$

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

Overall order of the reaction$= n + m = 1 + 0 = 1$

339 **(b)**

For II order reaction, $t = \frac{1}{K}\frac{x}{a(a-x)}$;

if $x = a/2, t_{1/2} = \frac{1}{K \cdot a}$.

340 **(c)**

$$2.303 \log \frac{K_2}{K_1} = \frac{E_a}{R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$$

$$\therefore 2.303 \log \frac{K_2}{K_1} = \frac{65 \times 10^3}{8.314}\left[\frac{25}{298 \times 273}\right]$$

$$\therefore \qquad \frac{K_2}{K_1} = 11.05$$

341 **(b)**

Arrhenius equation is,

Rate constant, $k = Ae^{-E_a|RT}$

$k = Ae^{-2.303\ RT/RT}$

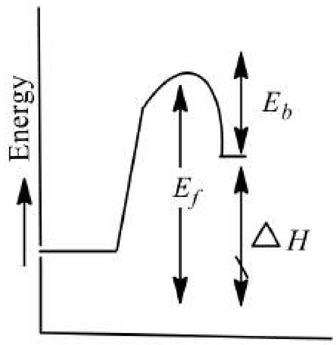
$\frac{k}{A} = e^{-2.303}$

On solving, we get

$\frac{k}{A} = 10^{-1}$

342 **(d)**

$X \rightarrow Y$ is an endothermic reaction ΔH=+ ve



E_b =energy of activation of backward reaction

E_f =energy of activation of forward reaction

ΔH = heat of reaction

Thus, $E_f = E_b + \Delta H$

Thus, $E_f > E_b$

343 **(c)**

$[A]_{\text{left}} = [B]_{\text{formed}} = n \times [A]_{\text{decayed}}$

$$A_0e^{-\lambda t} = n \times A_0[1 - e^{\lambda t}]$$

$$\therefore \quad e^{-\lambda t} = \frac{n}{n+1}$$

$$\therefore [B]_{\text{formed}} = n \times A_0 \times \left[1 - \frac{n}{n+1}\right] = \frac{nA_0}{n+1}$$

344 **(c)**

Three molecules are taking part in elementary step.

345 **(b)**

The rate expression is derived for slowest step of mechanism.

346 **(a)**

$$t_1 = \frac{(t_{1/2})_1}{0.693} \ln\left(\frac{1}{1-(1/4)}\right)$$

And $t_2 = \frac{(t_{1/2})_2}{0.693} \ln\left(\frac{1}{1-(3/4)}\right)$

$$\frac{t_1}{t_2} = \frac{8}{1} \times \frac{\ln(4/3)}{\ln 4}$$

$$= \frac{8 \times 0.125}{0.602} = \frac{0.001}{0.602}$$

$= 1 : 0.602$

348 **(b)**

For zero order reaction

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

For $t_{1/2}$, $[A] = [A]_0/2$

$\therefore kt_{1/2} = [A]_0 - [A]_0/2$

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

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

349 **(a)**

Given , $R = k[A][B]^2$

Where [A]=[2A]and B=[2B]

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

$= k[2A]4[B]^2$

$= 8k[A][B]^2$

R’=8R

Thus, the rate will become eight times

351 **(d)**

A decrease in E_a will increase rate constant K and thus rate of reaction increases.

352 **(a)**

The temperature coefficient is the ratio of two velocity constants having the difference of 10°C.

Temperature coefficient

$$= \frac{k_t + 10}{k_t}$$

For most of the reactions its value lies between 2 and 3.

353 **(c)**

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{693} = 0.001 \text{ s}^{-1}$$

354 **(b)**

$$-\frac{1}{3}\frac{d[\text{H}_2]}{dt} = \frac{1}{2}\frac{d[\text{NH}_3]}{dt}$$

$$\therefore -\frac{d[\text{H}_2]}{dt} = \frac{3}{2} \times \frac{d[\text{NH}_3]}{dt}$$

$$= \frac{3}{2} \times \frac{10^{-3} \times 10^3}{17} \text{mol hr}^{-1}$$

$$= \frac{3}{2} \times \frac{10^{-3} \times 10^3}{17} \times 2 \times 10^{-3} \text{ kg hr}^{-1}$$

$$= 1.76 \times 10^{-4} \text{kg hr}^{-1}$$

355 **(a)**

For I order reaction, $K = \frac{2.303 \log_{10} 2}{t_{1/2}}$

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

356 **(b)**

For second order reaction

$r = k[A]^2$ ... (1)

When concentration of A(reactant)is doubled

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

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

Comparing Eqs.(1)and(2)

r’ =4r

357 **(a)**

With increase in temperature reaction rate increases due to increase in number of molecules having threshold energy.

358 **(c)**

The value of first order rate constant expends on the units of the concentration term used is not the correct statement.

359 **(c)**

Since rate constant $= 1.0 \times 10^7$ mol litre^{-1} sec^{-1}

$\therefore$ Zero order reaction.

For zero order $t = \frac{x}{K} = \frac{\text{concentration used}}{\text{rate constant}}$ ... (i)

$\because$ 0.05 mL has 3×10^{-6}mole of H^+

$\therefore$ 1000 mL has $\frac{3 \times 10^{-6} \times 10^3}{0.05}$

$= 0.6 \times 10^{-1}$mol/litre of H^+

$\therefore$ By Eq. (i), $t = \frac{0.6 \times 10^{-1}}{1 \times 10^7} = 6 \times 10^{-9}$ second

360 **(c)**

$2A + B \rightarrow$ products

Rate of reaction, $r_1 = k[A]^2[B]$

If the concentration of A become double then the rate will be

$r_2 = k[2A]^2[B]$

$r_2 = 4k[A]^2[B]$

$r_2 = 4r_1$

361 **(c)**

From $t_{1/2} \propto C_0^{1-n}$

Where, n = order of reaction

For second order $n = 2$

$\therefore t_{1/2} \propto C_0^{1-2}$

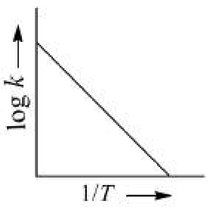
$t_{1/2} \propto C_0^{-1}$

362 **(b)**

10% of the reaction is completed in 20min .Next 20% of the reaction will be completed in next 20min. Hence, 10%+9%=19% of the reaction will be completed in 20+20=40min.

363 **(b)**

A graph plotted between $\log k$ *versus* $\frac{1}{T}$ for calculating activation energy is shown as:



From Arrhenius equation

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

364 **(a)**

$A \rightarrow B + C$

$$t_{1/2} = \frac{a}{2k} = \frac{1}{2 \times 0.001} = 500 \text{ s}$$

And $T = \frac{a}{k} = \frac{1}{0.001} = 1000$ s

365 **(c)**

$r \propto [CO]^2$

$r^1 \propto [2CO]^2$

$r^1 \propto 4[CO]^2$

Hence,

$$\frac{r^1 \propto 4[CO]^2}{r^1 \propto [CO]^2}$$

$r^1 \propto 4r$

366 **(a)**

$2A \rightarrow B + C$

2p 0 0

2p-x x x after 10 min

0 p p after completion

Given, P+P=200

After 10 min,

$2p - x + x + x = 300$ (given)

$2p + x = 300$

$x = 300 - 200 = 100$

$$k = \frac{2.303}{10} log \frac{200}{100} = 0.0693 \text{ min}^{-1}$$

367 **(a)**

Arrhenius equation is written as

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

Taking logarithm, above equation may be written as

$\text{In } k = \text{In } A - \frac{E_a}{R} \times \frac{1}{T}$

$$\therefore \text{In } k_1 = \text{In } A - \frac{E_a}{R} \times \frac{1}{T_1} \quad \text{... (i)}$$

$$\text{In } k_2 = \text{In } A - \frac{E_a}{R} \times \frac{1}{T_2} \quad \text{... (ii)}$$

Subtracting the Eq. (i) from Eq. (ii)

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

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

368 **(b)**

$$c_{12}H_{22}O_{11} + H_2O \xrightarrow{Dil.acid} C_6H_{12}O_6 + C_6 + H_{12}O_6$$

Excess glucose fructose

when one of the reactant is present in large excess, the second order reaction confirms to the first order and is knows as Pseudo-unimolecular reaction

369 **(d)**

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

370 **(b)**

$r = K[A]^n$

$2.4 = K[2.2]^n$

$0.6 = K[1.1]^n$

$\therefore 4 = (2)^n$ or $n = 2$

371 **(b)**

$$\frac{dx}{dt} = k[A]^2$$

$$k = \frac{dx}{dt[A]^2}$$

$$= \frac{mol(L)^2}{L \times s \times (mol)^2}$$

$mol^{-1}Ls^{-1}$

372 **(a)**

A→ *product* (first order reaction)

For first order reaction ,

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

At t= 40min,

$= \frac{2.303}{40} log_{10} \frac{0.1}{0.025} = \frac{2.303}{40} log_{10} 4$

$\frac{2.303}{40} \times 2 log_{10} 2 = \frac{2.303}{40} \times 2 \times 0.3010$

=0.0347 $\overset{-1}{\text{min}}$

At concentration of $A = 0.01\ M = [A]$

Rate $\frac{dx}{dt} k[A]$

$\frac{dx}{dt} = 0.0347 \times 0.01$

$= 3.47 \times 10^{-4} mol\ L^{-1}\ min^{-1}$

373 **(a)**

$2A + B \rightarrow A_2B$

$r_1 = k[A]^2[B]$

$r_2 = k[2A]^2[B/2]$

Or $r_2 = 2k[A]^2[B]$

$\therefore r_2 = 2r_1$

375 **(a)**

Given that, $k = 2.34 s^{-1}$ unit of 'k' suggest that the reaction is of Ist order , hence for Ist order reaction,

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

$= \frac{0.693}{2.34}$

=0.30 s

376 **(c)**

$2.303 \log \frac{K_2}{K_1} = \frac{E_a}{R} \left[\frac{T_2 - T_1}{T_1 T_2}\right]; T_1 = 280$ K,

$T_2 = 290$ K; $\frac{K_2}{K_1} = 2$

$2.303 \log \frac{K_2'}{K_1'} = \frac{E_a}{R} \left[\frac{T_2' - T_1'}{T_1' T_2'}\right]; T_1' = 290$ K; T_2

$= 310\ K$

377 **(c)**

Arrheniu equation gives relation between rate constant and temperature of a reaction. It can be written in many forms, as

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

Where, k = rate constant

E_a = activation energy

T =absorbption temperature

A = frequency factor

378 **(c)**

Order of reaction is sum of the power raised on concentration terms to express rate expression.

379 **(b)**

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

E is activation energy, it is that energy, which molecule must have to give the product.

381 **(b)**

$\frac{K_{t+10}}{K_t} = \frac{r_t + 10}{r_t} = 2;$

For an increase of temperature to 50°C, *i. e.*, 5 times, the rate increases by 2^5 times, *i. e.*, 32 times.

382 **(a)**

$A_2(g) \longrightarrow B(g) + \frac{1}{2}C(g)$

$\therefore \quad -\frac{d[A_2]}{dt} = 2.\frac{d[C]}{dt} = \frac{d[B]}{dt}$

Also the increase in pressure is due to the formation of C because one mole of A_2 gives one mole of B and half mole of C.

Thus,

$\frac{d[C]}{dt} = \frac{120 - 100}{5} = 4$ mm min^{-1}

383 **(b)**

The units of the rate constant for nth order reaction is $\left(\frac{mol}{L}\right)^{1-n} s^{-1}$

$\therefore$ for second order reaction,

Unit of rate constant$\left(\frac{mol}{L}\right)^{1-n} s^{-1}$

$mol^{-1} L s^{-1} = L\ /mol\ -s$

384 **(b)**

However, order should not be suggested from chemical equation. This question is not correct.

385 **(c)**

$A + B \rightarrow C$

On doubling the concentration of A, rate of reaction increases by 4 times.

$\therefore \quad rate \propto [A]^2$ (w.r.t. A)

However, on doubling the concentration of B rate of reaction increases two times.

$\therefore \quad rate \propto [B]^2$ (w. r. t. B)

Thus, overall order of reaction =2+1=3

386 **(d)**

$K_1 = A_1 e^{-E_1/RT}$ and $K_2 = A_2 e^{-E_2/RT}$

$\frac{K_1}{K_2} = \frac{A_1}{A_2} e^{(-E_1+E_2)/RT}$; A_1 and A_2 are not given.

387 **(a)**

$H_2 + I_2 \rightarrow 2HI$

Rate of reaction,

$= \frac{-d[H_2]}{dt} = \frac{-d[I_2]}{dt} = \frac{1}{2}\frac{d[HI]}{dt}$

Or $= \frac{-2d[H_2]}{dt} = \frac{-2d[I_2]}{dt} = \frac{d[HI]}{dt}$

388 **(c)**

87.5% completion means 3 half lives.

389 **(b)**

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

$$k=\frac{0.693}{10}\ \text{yr}^{-1}$$

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

$$=\frac{2.303\times 10}{0.693}\log\frac{a}{a-0.99a}$$

$$=\frac{23.03}{0.693}\log 100$$

$$t=\frac{23.03}{0.693}\times 2=66.46\ \text{yr}\approx 70\ \text{yr}$$

391 **(a)**

Luminescence is the emission of light by a substance for any reason other than rise in its temperature.

1. **Chemiluminescence** it is luminescence resulting from a chemical reaction, this is emission of visible light

2. **Phosphorescence** if the luminescence persists significantly after the existing cause is removed, it is called phosphorescence if it does not, and it is called **fluorescence.**

3. **Bioluminescence** It is luminescence produced by living organism *e. g.*, firefly.

392 **(b)**

Reactions having low E_a are fast reactions and reactions having high $E_a$are slow reactions.
If the E_a is high, then the number of effective collisions will be small and the reaction will be slow.

393 **(b)**

For the reaction,

$3A\rightarrow 2B$

$$Rate=-\frac{1}{3}\frac{d[A]}{dt}=+\frac{1}{2}\frac{d[B]}{dt}$$

$$\therefore\quad +\frac{d[B]}{dt}=-\frac{2}{3}\frac{d[A]}{dt}$$

394 **(a)**

For first order reaction the half-life period is independent of the initial concentration of the reactants.

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

395 **(a)**

First order kinetics,$k_1=\frac{0.693}{t_{1/2}}=\frac{0.693}{40}s^{-1}$

Zero order kinetics, $k_0=\frac{C_0}{2t_{1/2}}=\frac{1.386}{2\times20}$

Hence,$\frac{k_1}{k_0}=\frac{0.693}{1.386}=0.5$

396 **(a)**

Given, $r=KC_A^{3/2}c^{-1/2}$

$\therefore$ order of reaction $=\frac{3}{2}+\left(-\frac{1}{2}\right)$

$$=\frac{3-1}{2}=1$$

397 **(b)**

$$A_1\,.\,e^{-E_{a1}/RT}=A_2\ \cdot e^{-E_{a2}/RT}$$

$$\frac{A_2}{A_1}=e^{(E_{a_2}-E_{a_1})/RT}$$

$$10^2=\exp\left(\frac{1200}{RT}\right)$$

$$2\ln 10=\frac{1200}{2T}$$

$$T=\left(\frac{600}{4.606}\right)K$$

398 **(d)**

Nature and concentration of the reactants and temperature of the reaction influence the rate of reaction. But molecularity does not affect the rate of reaction as it includes the number of atoms, ions or molecules that must collide with one another to result into a chemical reaction.

399 **(a)**

$$k_I=\frac{\Delta[R]}{\Delta t}=\frac{0.25}{0.05}=5$$

$$k_{II}=\frac{\Delta[R]}{\Delta t}=\frac{0.60}{0.12}=5$$

So reaction must be zero order

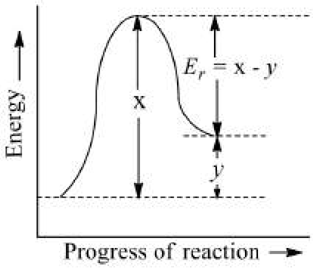
400 **(b)**

For zero order, $x=Kt, i.e., y=mx+c;$

$\therefore c=0$ or line passes through origin.

402 **(d)**

For an endothermic reaction



403 **(d)**

For a zero order reaction rate and rate constant are independent of reactant concentration

404 **(c)**

Rate constant is a characteristic constant for a given reaction.

405 **(c)**

The rate law for an elementary step can be given by simply observing the rate expression.

406 **(a)**

Threshold energy level is a characteristic of a reaction which may be however lowered, if catalyst is used.

407 **(b)**

$t_{1/2} = 69.3 \text{ sec} \quad \therefore K = \frac{0.693}{69.3} = 10^{-2}\text{sec}^{-1}$

Now, $r = K[A] = 10^{-2} \times 0.1 = 10^{-3} M \text{ sec}^{-1}$

408 **(b)**

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

$2.4 \times 10^{-5} = 3.0 \times 10^{-5}[\mathrm{N_2O_5}]$

$[\mathrm{N_2O_5}] = \frac{2.4 \times 10^{-5}}{3.0 \times 10^{-5}} = 0.8 \text{ mol L}^{-1}$

409 **(b)**

Enzyme catalysed reactions decreases energy of activation to greater extent.

410 **(a)**

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

411 **(a)**

For the homogenous gaseous reaction,

$A + B \xrightarrow{k} c$, *the rate law is*

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

Or

$k = \frac{dx}{dt \times [A][B]}$

$\frac{mol\ L^{-1}}{time \times mol\ L^{-1} \times mol\ L^{-1}}$

Or unit of rate constant 'k' is

$mol^{-1}\ L\ time^{-1} or\ mol^{-1} Ls^{-1}$.

412 **(a)**

For zero order reaction, $K = \frac{x}{t}$

If $x = a$ (complete reactant to react);

$t = \frac{a}{K}$

413 **(b)**

for two-third of a reaction ,

$[A]_0 = a, [A] = a - \frac{2}{3}a = \frac{a}{3}$

$t_{2/3} = \frac{2.303}{k} log \frac{[A]_0}{[A]}$

$= \frac{2.303}{k} log \frac{a}{\frac{a}{3}} = \frac{2.303}{k} \log 3$

$t_{2/3} = \frac{2.303 \times 0.4771}{5.48 \times 10^{-14}} = 2.01 \times 10^{13} s$

414 **(d)**

Rate constant almost gets doubled by the increase of 10°C in temperature. Hence, the rate constant at 310 K will be

$= 3.2 \times 10^{-3} \times (2)^2$

(∵increase in temperature=20 K)

$= 1.28 \times 10^{-2} \mathrm{J}^{-1}$

415 **(a)**

Unit of $k = conc^{1-n} t^{-1}$

$\left(\frac{mol}{L}\right)^{1-4} t^{-1}$

$\left(\frac{mol}{L}\right)^{-3} t^{-1}$

Therefore, the unit of k is

$\left(\frac{mol}{L}\right)^{-3} t^{-1}$ for fourth order reaction.

416 **(d)**

It is the definition of chemical kinetics.

417 **(b)**

$r = K[\mathrm{N_2O_5}]$

$\therefore [\mathrm{N_2O_5}] = \frac{r}{K} = \frac{1.02 \times 10^{-4}}{3.4 \times 10^{-5}} = 3\mathrm{M}$

418 **(a)**

For 50% B reacted, $K_2 = \frac{2.303}{t_2} \log \frac{100}{50}$

For 94% A reacted, $K_1 = \frac{2.303}{t_1} \log \frac{100}{6}$

$\therefore \quad \frac{K_2}{K_1} = \frac{t_2}{t_1} \times \frac{0.3010}{1.2218}$

Since $t_2 = t_1$, because 50% B has reacted when 94% A has reacted.

$\therefore \quad \frac{K_2}{K_1} = \frac{0.3010}{1.2218} = 0.246$ and $\frac{K_1}{K_2} = 4.06$

419 **(b)**

For first order half-life period is independent of initial concentration

420 **(c)**

$K_1 = Ae^{-Ea/RT}$

$K_2 = Ae^{-[E_a - 2]/RT}$

$$\frac{K_1}{K_2} = e^{2/RT} = e^{2/2\times10^{-3}\times300} = 28$$

421 **(d)**

Rate $= K_1[A] - K_2[B]$ for a reversible reaction of I order opposed by I order.

422 **(b)**

For Ist order reaction

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

If $[A]_0 = a, [A] = a - \frac{a\times20}{100} = 0.80a$

$t = 10$ min

$$k = \frac{2.303}{10}\log\frac{a}{0.80\,a} = \frac{2.303}{10}\times0.0969$$

$= 0.0223$

If $[A]_0 = a, [A] = a - \frac{a\times75}{100} = 0.25\,a$

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

$$= \frac{2.303}{0.0223}\log\frac{a}{0.25\,a}$$

$$= \frac{2.303}{0.0223}\times0.6021$$

$= 62.18$ min

423 **(b)**

Order of radioactive disintegration reaction is first. For example

${}_{88}Ra^{226} \rightarrow {}_{86}Rn^{222} + {}_2He^4$

Radium radon a-particle

424 **(a)**

$C_4H_8 \xrightarrow{\Delta} 2C_2H_4$

Mole at $t = 0$ a 0

Mole at $t = t$ $(a - x)$ $2x$

When $\frac{2x}{a-x} = 1$, then $x = \frac{a}{3}$

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

$$= \frac{2.303}{2.48\times10^{-4}}\log\frac{a}{a-\frac{a}{3}}$$

$= 1635.2$ second $= 27.25$ minute

425 **(a)**

If energy of activation for forward and backward reactions are same, reaction is neither exothermic not endothermic.

426 **(c)**

Half-life=6.93 min

$$k = \frac{0.693}{6.93} = 0.1$$

We know ,$k_1$for per cent completion

$$k_1 = \frac{2.303}{t}\log\left(\frac{100}{1}\right)$$

$$0.1 = \frac{2.303}{t}\log\frac{100}{1}$$

$$0.1 = \frac{2.303}{t}\log10^2$$

$$t = \frac{2.303\times2}{0.1} = 46.06$$

427 **(b)**

Rate $= K[\mathrm{NO}]^2[\mathrm{O}_2]^1$;

Concentration of each species are reduced by 1/2 on increasing volume to two times and thus, rate becomes 1/8 times of initial rate.

428 **(c)**

The rate of reaction is

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

When the volume is reduced to $\frac{1}{3}$, the concentration of each reactant is increased by 3 times

$$\text{rate}' = k[3\mathrm{NO}]^2[3\mathrm{O}_2]$$

$$= 27k[\mathrm{NO}]^2[\mathrm{O}_2]$$

$$\frac{\text{rate}'}{\text{rate}} = \frac{27k[\mathrm{NO}]^2[\mathrm{O}_2]}{k[\mathrm{NO}]^2[\mathrm{O}_2]}$$

$$\text{rate}' = 27\ \text{rate}$$

429 **(a)**

$K = 10^{-2}\text{mol}^{-1}\text{litre sec}^{-1}$

$$= \frac{10^{-2}\times1000\times60}{6.02\times10^{23}}\text{cc molecule}^{-1}\text{min}^{-1}$$

$$= 9.9618\times10^{-22}\ \text{cc molecule}^{-1}\ \text{min}^{-1}$$

431 **(d)**

Activation energy of a reaction is constant at constant temperature hence, $E_1 = E_2$

432 **(b)**

Rate $(r) = k[A][B] = kab$

When volume is reduced by one fourth then concentration becomes 4 times

Hence, $r' = k(4a)(4b)$

$= 16kab$

$\therefore r' = 16r$

433 **(b)**

The reactant taken in excess obeys zero order reaction;

$\therefore r = K[B]$.

434 **(a)**

For first order reaction,

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

$[A]_0 = 5g, [A] = 3g\ k = 1.15 \times 10^{-3} s^{-1}$

$$t = \frac{2.303}{1.15 \times 10^{-3}} \log \frac{5}{3}$$

$$= \frac{2.303}{1.15 \times 10^{-3}} [log 5 - log 3]$$

$$= \frac{2.303}{1.15 \times 10^{-3}} \times 0.2218$$

$= 444$s

435 **(d)**

For II order reaction $\frac{c_0 - c}{c} = K \cdot t \cdot c_0$.

436 **(b)**

We know,

Given, $T_1 = 27 + 273 = 300\ K K_1 = k$

$T_2 = 37 + 273 = 310K\ K_2 = 2k$

$R = 8.314 \times 10^{-3}$ kJ

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

$$2.303 \log \frac{2k}{k} = \frac{E_a}{8.314 \times 10^{-3}} \times \left(\frac{310 - 300}{310 \times 300} \right)$$

$$2.303 \log 2 = \frac{E_a}{8.314 \times 10^{-3}} \times \frac{10}{93000}$$

$$2.303 \times 0.3010 = \frac{E_a}{8.314 \times 10^{-3}} \times 0.000108$$

$$E_a = \frac{2.303 \times 0.3010 \times 8.314 \times 10^{-3}}{0.000108}$$

$$= \frac{5.7633 \times 10^{-3}}{0.000108}$$

$E_a = 53.363$ kJ

437 **(d)**

If the concentration of reactants is increased by x, then rate constant k remains same, *ie*, k because change of concentration has no effect on the rate constant