

**WEEKLY TEST MEDICAL PLUS - 01 TEST - 20 R & B**  
**SOLUTION Date 06-10-2019**

**[PHYSICS]**

1. (a) The mean kinetic energy of a gas depends only on temperature and is independent of molecular weight.

2. (c)  $C_{H_2} = xC_{He}$

$$\sqrt{M_{He}} = x\sqrt{M_{H_2}}$$

$$\sqrt{4} = x\sqrt{2}$$

or  $x = \frac{2}{\sqrt{2}}$  or  $x = \sqrt{2}$

3. (a)  $\frac{E_1}{E_2} = \frac{N \times \frac{3}{2}kT}{2N \times \frac{3}{2}kT} = \frac{1}{2}$

4. (a)  $\frac{1}{2}kT \times N = \frac{1}{2}RT$

5. (c)  $\frac{C_1}{C_2} = \sqrt{\frac{\rho_2}{\rho_1}} = \sqrt{\frac{9}{8}}$

6. (c)  $\frac{3}{2}kT_1 \times n_1 + \frac{3}{2}kT_2 \times n_2 = (n_1 + n_2) \frac{3}{2}kT$

or  $T = \frac{n_1T_1 + n_2T_2}{n_1 + n_2}$

7. (a)  $C = \sqrt{\frac{1^2 + 2^2 + 3^2 + 4^2}{4}}$   
 $= \sqrt{\frac{1 + 4 + 9 + 16}{4}} = \sqrt{\frac{15}{2}} \text{ kms}^{-1}$

8. (c)  $C = \sqrt{\frac{3PV}{M}}$   
 $= \sqrt{\frac{3 \times 24 \times 10^5 \times 10 \times 1000}{20}} \text{ cms}^{-1}$   
 $= 6 \times 10^4 \text{ cms}^{-1} = 600 \text{ ms}^{-1}$

9. (b) For  $n$ -molecules, we know that

$$v_{\text{rms}} = \sqrt{\frac{v_1^2 + v_2^2 + v_3^2 + \dots + v_n^2}{n}}$$

[ $v_{\text{rms}}$  = root mean square velocity]

where  $v_1, v_2, v_3 \dots v_n$  are individual velocities of  $n$ -molecules of the gas.

For two molecules,  $v_{\text{rms}} = \sqrt{\frac{v_1^2 + v_2^2}{2}}$

[ $v_1, v_2, v_3 \dots v_n$  are individual velocity]

Given,  $v_1 = 9 \times 10^6$  m/s and  $v_2 = 1 \times 10^6$  m/s

$$\therefore v_{\text{rms}} = \sqrt{\frac{(9 \times 10^6)^2 + (1 \times 10^6)^2}{2}} = \sqrt{41} \times 10^6 \text{ m/s}$$

10. (c) We know that for a given mass of a gas

$$v_{\text{rms}} = \sqrt{\frac{3RT}{M}}$$

where  $R$  is gas constant,  $T$  is temperature in kelvin and  $M$  is molar mass of the gas.

Clearly for a given gas,  $v_{\text{rms}} \propto \sqrt{T}$ , as  $R, M$  are constants.

Hence,  $\frac{(v_{\text{rms}})_1}{(v_{\text{rms}})_2} = \sqrt{\frac{T_1}{T_2}}$

Given,  $(v_{\text{rms}})_1 = 100$  m/s

$$T_1 = 27^\circ\text{C} = 27 + 273 = 300 \text{ K}$$

$$T_2 = 127^\circ\text{C} = 127 + 273 = 400 \text{ K}$$

$$\therefore \text{From Eq. (i)} \quad \frac{100}{(v_{\text{rms}})_2} = \sqrt{\frac{300}{400}} = \frac{\sqrt{3}}{2}$$

$$\Rightarrow (v_{\text{rms}})_2 = \frac{2 \times 100}{\sqrt{3}} = \frac{200}{\sqrt{3}} \text{ m/s}$$

11. (c)  $E_A = \frac{3}{2}kT_A; E_B = \frac{3}{2}k(2T_A)$

and  $E'_A = \frac{3}{2}kT; E'_B = \frac{3}{2}kT$

$$\frac{\Delta v_P}{v_P} = \left( \sqrt{\frac{T}{T_A}} - 1 \right) = 2 \frac{\Delta v_{\text{avg}}}{v_{\text{avg}}} = 2 \left( \sqrt{\frac{T}{2T_A}} - 1 \right)$$

Let  $T/T_A = x$ , the above equation reduces to

$$(\sqrt{x} - 1) = 2(\sqrt{x/2} - 1) \text{ or, } \sqrt{x}(\sqrt{2} - 1) = 1$$

or  $x = [1/(3 - 2\sqrt{2})] \text{ or, } x = 3 + 2\sqrt{2}$

12. (c) The kinetic energy of gas w.r.t. centre of mass of the system  $K.E. = \frac{5}{2}nRT$

Kinetic energy of gas w.r.t. ground = Kinetic energy of centre of mass w.r.t. ground + Kinetic energy of gas w.r.t. centre of mass.

$$K.E. = \frac{1}{2}MV^2 + \frac{5}{2}nRT$$

13. (b) When, the container stops, its total kinetic energy is transferred to gas molecules in the form of translational kinetic energy, thereby increasing the absolute temperature.

Assuming  $n$  = number of moles.

Given,  $m$  = molar mass of the gas.

If  $\Delta T$  = change in absolute temperature.

Then, kinetic energy of molecules due to velocity  $v_0$ ,

$$\Delta K_{\text{motion}} = \frac{1}{2}(mn)v_0^2 \quad (\text{i})$$

Increase in translational kinetic energy

$$\Delta K_{\text{translation}} = n \frac{3}{2}R(\Delta T) \quad (\text{ii})$$

According to kinetic theory Eqs. (i) and (ii) are equal

$$\Rightarrow \frac{1}{2}(mn)v_0^2 = n \frac{3}{2}R(\Delta T)$$

$$(mn)v_0^2 = n3R(\Delta T)$$

$$\Rightarrow \Delta T = \frac{(mn)v_0^2}{3nR} = \frac{mv_0^2}{3R}$$

14. (d)  $P = \frac{2N}{3V}\bar{K}$  from the kinetic-theory account for pressure.

$$N = \frac{3PV}{2\bar{K}}$$

$$n = \frac{N}{N_A} = \frac{3PV}{2\bar{K}N_A}$$



15. (a) The gas temperature must be that implied by

We know kinetic energy of a molecule is given by

$$\frac{1}{2}m_0\overline{v^2} = \frac{3}{2}k_B T \Rightarrow T = \frac{2}{3} \left( \frac{\frac{1}{2}m_0\overline{v^2}}{k_B} \right)$$

$$\Rightarrow T = \frac{2}{3} \left( \frac{3.60 \times 10^{-22}}{1.38 \times 10^{-23} \text{ J/K}} \right) = 17.4 \text{ K}$$

Now

$$PV = nRT \Rightarrow n = \frac{PV}{RT}$$

$$n = \frac{(1.20 \times 10^5 \text{ N/m}^2)(4.00 \times 10^{-3} \text{ m}^3)}{(8.314 \text{ J/mol.K})(17.4 \text{ K})} = 3.32 \text{ mol}$$

16. (c)  $\rho = \frac{PM}{RT}$

Density  $\rho$  remains constant when  $P/T$  or volume remains constant.

In graph (i) Pressure is increasing at constant temperature hence volume is decreasing so density is increasing. Graphs (ii) and (iii) volume is increasing hence, density is decreasing. Note that volume would had been constant in case the straight line in graph (iii) had passed through origin.

17. (c)  $\rho \propto \frac{1}{V}$  in the process  $AB$ .

( $V$  decreases,  $\rho$  increases).

In isothermal process ( $T = \text{constant}$ ),  $\rho \propto \frac{1}{V}$ .

As  $V$  decreases,  $P$  increases.

So, for  $\rho$  increase,  $V$  decreases ( $\rho V$  curve) and  $P$  increases for  $V$  decreases ( $P - T$  curve)

Options (b) and (c) may be correct.

But process  $B - C$  is shown by a dot in ( $\rho - V$ ) curve.

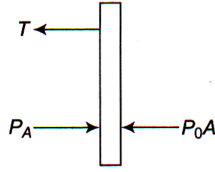
It means  $BC$  is a isochoric process.

Process  $BC$  in option (c) is a isochoric process because  $P \propto T$ .

So, option (c) is correct.

18. (c)  $T + P_0A = PA$

$$T = (P - P_0)A = \frac{3}{8}P_0A \text{ (given)}$$



$$P = \frac{3}{8}P_0 + P_0 = \frac{11}{8}P_0$$

Now volume is constant by string

So,  $P \propto T$

Initial temperature is  $T_0$ .

19. (a) We have to convert the given temperatures in kelvin.

If pressure of a given mass of the gas is kept constant, then  $V \propto T$

$$\Rightarrow \frac{V}{T} = \text{constant} \quad \left[ \begin{array}{l} V = \text{Volume of gas} \\ T = \text{Temperature of gas} \end{array} \right]$$

$$\Rightarrow \frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow V_2 = V_1 \left( \frac{T_2}{T_1} \right)$$

Here we must convert the given temperature in kelvin.

$$T_1 = 273 + 27 = 300 \text{ K}$$

$$T_2 = 273 + 327 = 600 \text{ K}$$

But  $V_1 = 100 \text{ cc}$

$$V_2 = V_1 \left( \frac{600}{300} \right) \Rightarrow V_2 = 2V_1$$

20. (b) Initial volume of the gas  $V_1 = 1 \text{ litre}$  Initial pressure of the gas  $P_1 = 70 \text{ cm of Hg}$

Final pressure of the gas  $P_2 = 2 \times 76 = 152 \text{ cm of Hg}$

Final volume of the gas  $V_2 = 2 \text{ litre}$

According the Gay Lussac law we have

$$\frac{P_1 V_1}{T} = \frac{P_2 V_2}{T_2} \text{ or } \frac{76 \times 1}{300} = \frac{152 \times 2}{T_2}$$

$$\frac{1}{300} = \frac{4}{T_2} \text{ or } T_2 = 4 \times 300 = 1200 \text{ K} = 927^\circ\text{C}$$

21. (a) Initial volume of gas =  $V_1$

Final volume of gas =  $V_2$

Initial temperature of gas  $T_1 = 27^\circ\text{C} = 300 \text{ K}$

Final temperature of gas  $T_2 = 54^\circ\text{C} = 327 \text{ K}$

Now from the Charles's law at constant pressure

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} = \frac{300}{327} = \frac{100}{109}$$

22. (d)  $PV = \mu RT, PV = \frac{n}{N} \times h NT$  or  $n = \frac{PV}{kT}$

23. (c) For a given pressure,  $V$  is small for  $T_1$ . Since  $V \propto T$ , therefore,  $T_1 < T_2$ .

24. (b)  $VP^3 = \text{constant} = k \Rightarrow P = \frac{k}{V^{1/3}}$

Also  $PV = \mu RT \Rightarrow \frac{k}{V^{1/3}} \cdot V = \mu RT \Rightarrow V^{2/3} = \frac{\mu RT}{k}$

Hence  $\left(\frac{V_1}{V_2}\right)^{2/3} = \frac{T_1}{T_2} \Rightarrow \left(\frac{V}{27V}\right)^{2/3} = \frac{T}{T_2} \Rightarrow T_2 = 9T$

25. (c)  $\mu = \mu_1 + \mu_2$

$$\frac{P(2V)}{RT_1} = \frac{P'v}{RT_1} + \frac{P'V}{RT_2} \Rightarrow \frac{2P}{RT_1} = \frac{P'}{R} \left[ \frac{T_2 + T_1}{T_1 T_2} \right]$$

$$P' = \frac{2PT_2}{(T_1 + T_2)} = \frac{2 \times 1 \times 600}{(300 + 600)} = \frac{4}{3} \text{ atm}$$

26. (d)  $\rho = \frac{PM}{RT}$

$\rho = \text{constant}$ , if  $(P/T)$  is constant

$(P/T)$  is constant in option (d).

27. (c)  $V = k \left( \frac{nRT}{VT} \right)^{0.33}$

$V^{1.33} = \text{constant}$

$V = \text{constant}$

28. (c)  $\rho = \frac{PM}{RT}$  gives  $\rho \propto (P/T)$

$$\rho_A = \rho_0 \text{ and } \rho_B = \frac{3}{2} \rho_0$$

$$(P/T)_B = \frac{3}{2} (P/T)_A$$

$$(X/2T_0) = \left(\frac{3}{2}\right) (P_0/T_0)$$

$$X = 3P_0$$

29. (a) Here,  $T = V \tan 45^\circ + T_0 \Rightarrow T = V + T_0$

And  $P = \frac{nRT}{V}$ ,

$\therefore \frac{PV}{R} = V + T_0$  (Since,  $n = 1$ )

$$(P - R)V = RT_0$$

Therefore, graph will be rectangular hyperbola.

30. (c) On combining the two vessels the total number of moles remains constant, i.e.,  $n = n_1 + n_2$

Using gas equation, we can write,  $n_1 = \frac{P_1 V}{RT_1}$ ;  $n_2 = \frac{P_2 V}{RT_2}$

and  $n = \frac{P(2V)}{RT}$

where  $V$  is the volume of each vessel.

Thus  $\frac{P(2V)}{RT} = \frac{P_1 V}{RT_1} + \frac{P_2 V}{RT_2}$  or  $\frac{P}{T} = \frac{1}{2} \left[ \frac{P_1}{T_1} + \frac{P_2}{T_2} \right]$

### [CHEMISTRY]

46. (c) : Rate =  $-\frac{d[\text{N}_2\text{O}_5]}{dt} = +\frac{1}{2} \frac{d[\text{NO}_2]}{dt}$   
 $= 2 \frac{d[\text{O}_2]}{dt}$

Given  $\frac{-d[\text{N}_2\text{O}_5]}{dt} = 6.25 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1}$

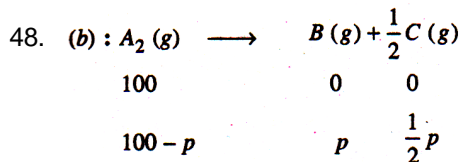
Rate of formation of  $\text{NO}_2$

$$\begin{aligned} &= \frac{[\text{NO}_2]}{dt} = -2 \frac{d[\text{N}_2\text{O}_5]}{dt} \\ &= 2 \times 6.25 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1} \\ &= 12.50 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1} \\ &= 1.25 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1} \end{aligned}$$

Rate of formation of  $\text{O}_2$

$$\begin{aligned} &= \frac{d[\text{O}_2]}{dt} = \frac{1}{2} \frac{d[\text{N}_2\text{O}_5]}{dt} \\ &= \frac{1}{2} \times 6.25 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1} \\ &= 3.125 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1} \end{aligned}$$

47. (b) : Minus signs are for reactants and positive signs for products. Dividing numbers are the coefficients.



$$100 - p + p + \frac{1}{2}p = 120 \text{ or } p = 40 \text{ mm}$$

$$\therefore -\frac{dp_{A_2}}{dt} = \frac{40}{5} = 8 \text{ mm min}^{-1}$$

49. (c) : Initially, Rate =  $k [Y] [Z]^{1/2}$

New rate =  $k [Y] [2Z]^{1/2}$

$$= \sqrt{2} k [Y] [Z]^{1/2} = 1.414 k [Y] [Z]^{1/2}$$



50. (d) : The rate of reaction is same as expressed in terms of any reactant or product.

$$51. (a) : \text{Rate of reaction} = \frac{1}{4} \frac{\Delta[\text{NO}_2]}{\Delta t}$$

$$= \frac{1}{4} \times \frac{[5.2 \times 10^{-3} \text{ M}]}{100 \text{ s}} = 1.3 \times 10^{-5} \text{ M s}^{-1}.$$

52. (b) : Rate =  $k [\text{NO}]^2 [\text{O}_2]$ . Initially rate =  $ka^2 b$ . If volume is reduced to half, concentration are doubled so that new rate

$$= k (2a)^2 (2b) = 8ka^2 b, \text{ i.e., 8 times.}$$

53. (b) : For zero order reaction,  $k = \frac{1}{t} \{[\text{A}]_0 - [\text{A}]\}$   
or  $[\text{A}] = -kt + [\text{A}]_0$ . Thus, plot of  $[\text{A}]$  vs  $t$  is linear with -ve slope ( $= -k$ ).

54. (c) : From slow step, rate =  $k [\text{B}_2] [\text{A}]$ .

$$\text{From 1st eqn, } K_{eq} = \frac{[\text{A}]^2}{[\text{A}_2]}$$

$$\text{or } [\text{A}] = \sqrt{K_{eq} [\text{A}_2]} = K_{eq}^{1/2} [\text{A}_2]^{1/2}$$

$$\text{Hence, rate} = k [\text{B}_2] K_{eq}^{1/2} [\text{A}_2]^{1/2}$$

$$= k' [\text{A}_2]^{1/2} [\text{B}_2].$$

$$\text{Hence, order} = 1\frac{1}{2}.$$

55. (c) : On the basis of given units of  $k$ , the reaction is of 3rd order.

56. (d) :  $r = k [\text{A}]^\alpha [\text{B}]^\beta = k a^\alpha b^\beta$ . If concentration of  $B$  is doubled,  $\frac{r}{4} = k a^\alpha (2b)^\beta$ . Dividing 2nd eqn. by 1st eqn.,

$$\frac{1}{4} = 2^\beta \text{ or } 2\beta = 2^{-2}. \text{ Hence, } \beta = -2.$$

57. (a) : As step I is the slowest, hence it is the rate determining step.

58.

$$59. (a) : k = \frac{2.303}{32} \log \frac{a}{a - 0.99a}$$

$$= \frac{2.303}{32} \log 10^2 = \frac{2.303}{16} \text{ min}^{-1}$$

$$t_{99.9\%} = \frac{2.303}{k} \log \frac{a}{a - 0.999a}$$

$$= \frac{2.303}{k} \log 10^3 = \frac{3 \times 2.303}{2.303} \times 16$$

$$= 48 \text{ min.}$$



60. (b) : 0.08 mol L<sup>-1</sup> to 0.01 mol L<sup>-1</sup> involves 3 half-lives.

$$61. (a) : k = \frac{2.303}{t} \log \frac{a}{a-x}$$

$$\text{or } \log \frac{a}{a-x} = \frac{kt}{2.303} = \frac{2.2 \times 10^{-5} \times 60 \times 90}{2.303}$$

$$= 0.0516.$$

$$\text{Hence, } \frac{a}{a-x} = \text{antilog } 0.0516 = 1.127.$$

$$\text{or } \frac{a-x}{a} = 0.887 \text{ or } 1 - \frac{x}{a} = 0.887$$

$$\text{or } \frac{x}{a} = 0.113 = 11.3\%.$$

$$62. (c) : t_{90\%} = \frac{2.303}{k} \log \frac{a}{a-0.9a}$$

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

$$t_{1/2} = \frac{2.303}{k} \log \frac{2}{a-a/2}$$

$$= \frac{2.303}{k} \log 2 = \frac{2.303}{k} \times 0.3010$$

$$\therefore t_{90\%}/t_{1/2} = \frac{1}{0.3010} = 3.3$$

$$\text{i.e., } t_{90\%} = 3.3 \text{ times } t_{1/2}.$$

63. (a) : Decrease in concentration from 0.8 M to 0.4 M in 15 minutes means  $t_{1/2} = 15$  minutes. Time taken for decrease in concentration from 0.1 M to 0.25 M means two half-lives, i.e.,  $= 2 \times 15 \text{ min} = 30 \text{ min}$ .

$$64. (c) : t_{1/2} \propto \frac{1}{a^{n-1}} \quad \text{For } n = 2, \quad t_{1/2} \propto \frac{1}{a}$$

65. (c) : At the point of intersection,  $[A] = [B]$ , i.e., half of the reactant has reacted. Hence, it represents  $t_{1/2}$ .

66. (a) : It  $P_t$  is the pressure after time  $t$ ,

$$k = \frac{2.303}{t} \log \frac{P_0}{P_t}$$

$$\therefore 3.38 \times 10^{-5} \text{ s}^{-1} = \frac{2.303}{600 \text{ s}} \log \frac{500 \text{ atm}}{P_t}$$

$$\text{or } \log \frac{500}{P_t} = 0.0088 \text{ or } \frac{500}{P_t} = 1.021$$

$$\text{or } P_t = 490 \text{ atm}$$



67. (a) : [A] is kept constant, [B] is doubled, rate is doubled. So rate  $\propto$  [B].

[B] is kept constant, [A] is tripled, rate becomes 9 times, so rate  $\propto$  [A]<sup>2</sup>.

Hence, rate law is

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

68. (a) :  $t_{1/2} \propto a^{1-n}$ . Hence,  $t_{1/2} \propto 1/a^3$  only when  $n = 4$ .

69. (c) : In presence of acid, hydrolysis of ethyl acetate is a pseudo-unimolecular reaction but the actual value of  $k$  depends upon the concentration of H<sup>+</sup> ion. As H<sub>2</sub>SO<sub>4</sub> is stronger acid than HCl and moreover H<sup>+</sup> ions produced from 0.05 M H<sub>2</sub>SO<sub>4</sub> are double than those produced from 0.05 M HCl, therefore  $k_2 >> k_1$  or  $k_1 < k_2$ .

70. (b) :  $k = \frac{2.303}{t} \log \frac{V_{\infty}}{V_{\infty} - V_t}$

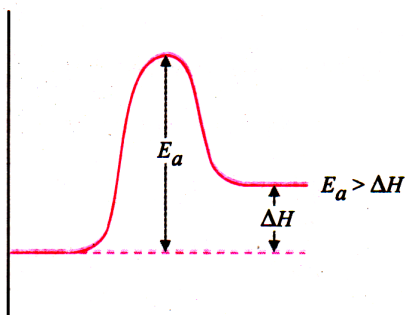
gives constant value of  $k$ .

Hence, it is 1st order reaction.

71. (b) : With increase in temperature, number of activated molecules increases.

72. (b) : As  $k' > k''$ ,  $E_a' < E_a''$  (Greater the rate constant, less in the activation energy).

73. (c) :



74. (d) : Lower the activation energy, faster is the reaction. Hence, relative ease of P, Q and R will be  $R > Q > P$ .

75. (d) : All the given statements are correct according to collision theory.

76. (d) : Activation energy of a particular reaction is constant temperature.

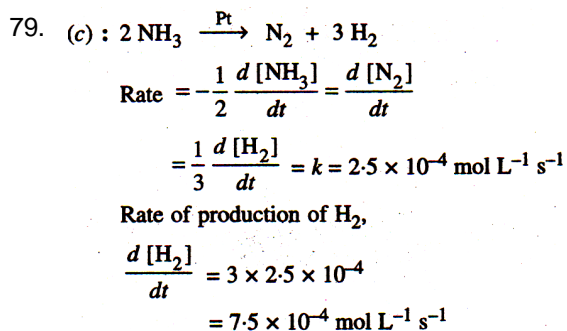
77. (c) : Small increase in  $e^{-E/RT}$  resulting in large increase in  $k$  is due to high value of  $A$ .

78. (d) : % of B =  $\frac{k_1}{k_1 + k_2}$

$$= \frac{1.26 \times 10^{-4}}{1.26 \times 10^{-4} + 3.8 \times 10^{-5}} \times 100$$

$$= \frac{1.26 \times 10^{-4}}{10^{-4}(1.26 + 0.38)} \times 100 = 76.83\%$$





80. (a) : In case I, fraction of A reacted =  $\frac{0.6}{0.8} = \frac{3}{4}$

In case II, fraction of A reacted =  $\frac{0.675}{0.9} = \frac{3}{4}$

For a first order reaction, time taken for the same fraction of reaction is independent of initial concentration.

81. (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.

82. (b) : Rate =  $k [\text{NOBr}_2] [\text{NO}]$

From step I,  $K_{eq} = \frac{[\text{NOBr}_2]}{[\text{NO}] [\text{Br}_2]}$

or  $[\text{NOBr}_2] = K_{eq} [\text{NO}] [\text{Br}_2]$

Substituting in eqn. (i), we get

$$\text{Rate} = k K_{eq} [\text{NO}]^2 [\text{Br}_2] = k' [\text{NO}]^2 [\text{Br}_2]$$

Hence, order with respect to NO is 2.

83. (c) :  $10.8 \text{ g of N}_2\text{O}_5 = \frac{10.8}{108} \text{ mole} = 0.1 \text{ mole}$

No. of half-lives in 9.6 h =  $\frac{9.6}{2.4} = 4$

Amount left after 4 half-lives

$$= \frac{1}{2^4} \times 0.1 = \frac{0.1}{16} \text{ mole}$$

Moles of  $\text{N}_2\text{O}_5$  decomposed

$$= 0.1 - \frac{0.1}{16} = \frac{1.5}{16} \text{ mole}$$

Moles of  $\text{O}_2$  formed =  $\frac{1}{2} \times \frac{1.5}{16} = \frac{1.5}{32}$

Volume of  $\text{O}_2$  at STP =  $\frac{1.5}{32} \times 22.4 \text{ L} = 1.05 \text{ L}$



84. (b) :  $r = K [A]^\alpha = k a^\alpha$   
 $1.837 r = k (1.5 a)^\alpha$   
 Dividing,  $1.837 = (1.5)^\alpha$   
 On solving, we get  $\alpha = 1.5$   
 Hence order = 1.5
85. (a) : Half-life of a first order reaction does not depend upon initial concentration. It is equal to  $\ln 2/k$ .
86. (c) : Diagram (c) represents endothermic reaction with

$$E_a = 200 - 150 = 50 \text{ kJ and}$$

$$\Delta H = 50 - 150 = -100 \text{ kJ.}$$

87. (d) : If  $E_a = 0$ ,  $k = A e^{-E_a/RT} = A e^0 = A$ . Hence,  $k$  becomes independent of  $T$ .

88. (d) : Given  $\log k = 6 - \frac{2000}{T}$

$$\text{Comparing with } \log k = \log A - \frac{E_a}{2.303 RT}$$

$$\log A = 6, \text{ i.e., } A = 10^6$$

$$\text{and } \frac{E_a}{2.303 R} = 2000$$

$$\text{or } E_a = 2000 \times 2.303 \times 8.314 \text{ J mol}^{-1}$$

$$= 38294 \text{ J mol}^{-1} \approx 38.3 \text{ kJ mol}^{-1}$$

89.

90. (c) :  $t_{\max}$  represents time corresponding to maximum concentration of the intermediate  $R_2$   
 The value of  $t_{\max}$  is given by the relation

$$t_{\max} = \frac{2.303}{(k_1 - k_2)} \log \frac{k_1}{k_2}$$

$$k_1 = 4.0 \times 10^{-2}$$

$$\therefore k_2 = 4.0 \times 10^{-2} \times 0.15 = 6 \times 10^{-3}$$

$$\therefore t_{\max} = \frac{2.303}{(4 - 0.6) 10^{-2}} \log \frac{4 \times 10^{-2}}{6 \times 10^{-3}}$$

$$= \frac{2.303}{3.4 \times 10^{-2}} \times 0.82 = 55.6 \text{ min}$$