[PHYSICS]

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

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

    $$\sqrt{M_{He}} = x\sqrt{M_{H_2}}$$
    $$\sqrt{4} = x\sqrt{2}$$

    or $\quad x = \dfrac{2}{\sqrt{2}}$ or $x = \sqrt{2}$

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

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

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

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

    or $\quad T = \dfrac{n_1 T_1 + n_2 T_2}{n_1 + n_2}$

7. **(a)** $C = \sqrt{\dfrac{1^2 + 2^2 + 3^2 + 4^2}{4}}$

    $= \sqrt{\dfrac{1 + 4 + 9 + 16}{4}} = \sqrt{\dfrac{15}{2}}\,\text{km s}^{-1}$

8. **(c)** $C = \sqrt{\dfrac{3PV}{M}}$

    $= \sqrt{\dfrac{3 \times 24 \times 10^5 \times 10 \times 1000}{20}}\,\text{cm s}^{-1}$

    $= 6 \times 10^4\,\text{cm s}^{-1} = 600\,\text{m s}^{-1}$

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

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

[v_{rms} = root mean square velocity]

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

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

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

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

$\therefore \quad v_{\text{rms}} = \sqrt{\dfrac{(9 \times 10^6)^2 + (1 \times 10^6)^2}{2}} = \sqrt{41} \times 10^6$ 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, $\dfrac{(v_{\text{rms}})_1}{(v_{\text{rms}})_2} = \sqrt{\dfrac{T_1}{T_2}}$

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

$\qquad T_1 = 27°C = 27 + 273 = 300$ K

$\qquad T_2 = 127°C = 127 + 273 = 400$ K

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

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

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

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

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

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

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

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

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

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.

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

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 ΔT = change in absolute temperature.

Then, kinetic energy of molecules due to velocity v_0,

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

Increase in translational kinetic energy

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

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

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

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

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

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

$$N = \dfrac{3}{2}\dfrac{PV}{\overline{K}}$$

$$n = \dfrac{N}{N_A} = \dfrac{3}{2}\dfrac{PV}{\overline{K}N_A}$$

15. **(a)** The gas temperature must be that implied by
We know kinetic energy of a molecule is given by

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

$$\Rightarrow \quad T = \frac{2}{3}\left(\frac{3.60\times10^{-22}}{1.38\times10^{-23}\,\text{J/K}}\right) = 17.4\,\text{K}$$

Now

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

$$n = \frac{(1.20\times10^5\,\text{N/m}^2)(4.00\times10^{-3}\,\text{m}^3)}{(8.314\,\text{j/mol.K})(17.4\,\text{K})} = 3.32\,\text{mol}$$

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

Density ρ 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)** $\rho \propto \dfrac{1}{V}$ in the process AB.

(V decreases, ρ increases).

In isothermal process ($T =$ constant), $\rho \propto \dfrac{1}{V}$.
As V decreases, P increases.

So, for ρ increase, V decreases (ρV curve) and P increases for V decreases ($P - T$ curve)

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

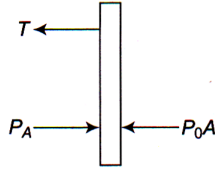But process $B - C$ is shown by a dot in (ρ-V) curve.

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)** $T + P_0 A = PA$

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

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

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 $V \propto T$

$$\Rightarrow \quad \frac{V}{T} = \text{constant} \quad \begin{bmatrix} V = \text{Volume of gas} \\ T = \text{Temperature of gas} \end{bmatrix}$$

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

Here we must convert the given temperature in kelvin.

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

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

But $V_1 = 100$ cc

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

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

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

Final volume of the gas $V_2 = 2$ litre

According the Gay Lussac law we have

$$\frac{P_1 V_1}{T} = \frac{p_2 V_2}{T_2} \text{ or } \frac{76 \times 1}{300} = \frac{152 \times 2}{T_2}$$

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

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

Final volume of gas $= V_2$

Initial temperature of gas $T_1 = 27°C = 300$ K

Final temperature of gas $T_2 = 54°C = 327$ 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)** $PV = \mu RT$, $PV = \dfrac{n}{N} \times h NT$ or $n = \dfrac{PV}{kT}$

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

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

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

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

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

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

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

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

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

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

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

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

    $V = \text{constant}$

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

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

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

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

    $$X = 3P_0$$

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

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

    $\therefore \quad \dfrac{PV}{R} = V + T_0$ \qquad (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, $n_1 = \dfrac{P_1 V}{RT_1}$; $n_2 = \dfrac{P_2 V}{RT_2}$

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

where V is the volume of each vessel.

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

[CHEMISTRY]

46. **(c)** : Rate $= -\dfrac{d[N_2O_5]}{dt} = +\dfrac{1}{2}\dfrac{d[NO_2]}{dt}$

$$= 2\dfrac{d[O_2]}{dt}$$

Given $\dfrac{-d[N_2O_5]}{dt} = 6{\cdot}25 \times 10^{-3}$ mol L^{-1} s^{-1}

Rate of formation of NO_2

$$= \dfrac{[NO_2]}{dt} = -2\dfrac{d[N_2O_5]}{dt}$$
$$= 2 \times 6{\cdot}25 \times 10^{-3} \text{ mol L}^{-1}\text{ s}^{-1}$$
$$= 12{\cdot}50 \times 10^{-3} \text{ mol L}^{-1}\text{ s}^{-1}$$
$$= 1{\cdot}25 \times 10^{-2} \text{ mol L}^{-1}\text{ s}^{-1}$$

Rate of formation of O_2

$$= \dfrac{d[O_2]}{dt} = -\dfrac{1}{2}\dfrac{d[N_2O_5]}{dt}$$

$$= \dfrac{1}{2} \times 6{\cdot}25 \times 10^{-3} \text{ mol L}^{-1}\text{ s}^{-1}$$
$$= 3{\cdot}125 \times 10^{-3} \text{ mol L}^{-1}\text{ s}^{-1}$$

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

48. **(b)** : $A_2 (g) \longrightarrow B (g) + \dfrac{1}{2}C (g)$
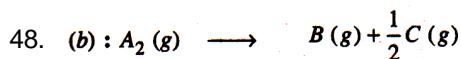
$\quad$ 100 $\qquad\qquad$ 0 $\qquad$ 0

$\quad$ $100 - p$ $\qquad\qquad$ p $\qquad$ $\dfrac{1}{2}p$

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

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

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

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

$$= \sqrt{2}\, k [Y][Z]^{1/2} = 1{\cdot}414\, k [Y][Z]^{1/2}.$$

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

51. (a) : Rate of reaction $= \dfrac{1}{4} \dfrac{\Delta[NO_2]}{\Delta t}$

$$= \dfrac{1}{4} \times \dfrac{[5 \cdot 2 \times 10^{-3}\,M]}{100\,s} = 1 \cdot 3 \times 10^{-5}\,M\,s^{-1}.$$

52. (b) : Rate = $k\,[NO]^2\,[O_2]$. Initially rate = $ka^2 b$. If volume is reduced to half, concentration are doubled so that new rate

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

53. (b) : For zero order reaction, $k = \dfrac{1}{t}\{[A]_0 - [A]\}$

or $[A] = -\,kt + [A]_0$. Thus, plot of $[A]$ vs t is linear with −ve slope ($= -\,k$).

54. (c) : From slow step, rate = $k\,[B_2]\,[A]$.

From 1st eqn, $K_{eq} = \dfrac{[A]^2}{[A_2]}$

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 $= 1\dfrac{1}{2}.$

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

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

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

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

58.

59. (d) : $k = \dfrac{2 \cdot 303}{32} \log \dfrac{a}{a - 0 \cdot 99\,a}$

$= \dfrac{2 \cdot 303}{32} \log 10^2 = \dfrac{2 \cdot 303}{16}\,min^{-1}$

$t_{99 \cdot 9\%} = \dfrac{2 \cdot 303}{k} \log \dfrac{a}{a - 0 \cdot 999\,a}$

$= \dfrac{2 \cdot 303}{k} \log 10^3 = \dfrac{3 \times 2 \cdot 303}{2 \cdot 303} \times 16$

$= 48\ min.$

60. (b) : 0.08 mol L^{-1} to 0.01 mol L^{-1} involves 3 half-lives.

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

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

$= 0.0516.$

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

or $\dfrac{a-x}{a} = 0.887$ or $1 - \dfrac{x}{a} = 0.887$

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

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

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

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

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

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

i.e., $t_{90\%} = 3.3$ 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., $= 2 \times 15$ min $= 30$ min.

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

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,

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

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

or $\log \dfrac{500}{P_t} = 0.0088$ or $\dfrac{500}{P_t} = 1.021$

or $P_t = 490$ 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 $\propto [A]^2$.
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^+ ion. As H_2SO_4 is stronger acid than HCl and moreover H^+ ions produced from 0·05 M H_2SO_4 are double than those produced from 0·05 M HCl, therefore $k_2 >> k_1$ or $k_1 < k_2$.

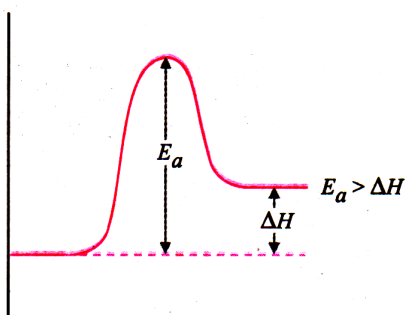70. (b) : $k = \dfrac{2\cdot303}{t}\log\dfrac{V_\infty}{V_\infty - V_t}$

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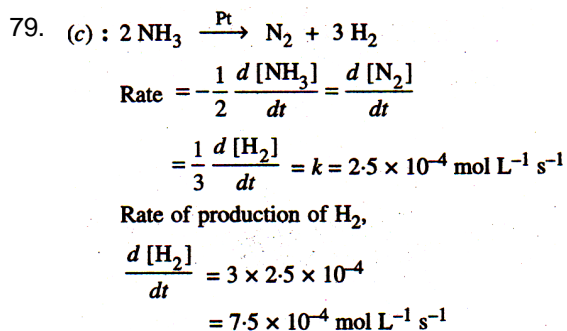
76. (d) : Activation energy of a particular reaction is constant temperature.

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 $= \dfrac{k_1}{k_1 + k_2}$

$= \dfrac{1\cdot26\times10^{-4}}{1\cdot26\times10^{-4} + 3\cdot8\times10^{-5}}\times100$

$= \dfrac{1\cdot26\times10^{-4}}{10^{-4}(1\cdot26 + 0\cdot38)}\times100 = 76\cdot83\%$.

AVIRAL CLASSES
CREATING SCHOLARS

79. (c) : $2\,NH_3 \xrightarrow{\text{Pt}} N_2 + 3\,H_2$

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

$$= \frac{1}{3}\frac{d\,[H_2]}{dt} = k = 2{\cdot}5 \times 10^{-4}\ \text{mol L}^{-1}\ \text{s}^{-1}$$

Rate of production of H_2,

$$\frac{d\,[H_2]}{dt} = 3 \times 2{\cdot}5 \times 10^{-4}$$

$$= 7{\cdot}5 \times 10^{-4}\ \text{mol L}^{-1}\ \text{s}^{-1}$$

80. (a) : In case I, fraction of A reacted $= \dfrac{0{\cdot}6}{0{\cdot}8} = \dfrac{3}{4}$

In case II, fraction of A reacted $= \dfrac{0{\cdot}675}{0{\cdot}9} = \dfrac{3}{4}$

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 $= k\,[NOBr_2]\,[NO]$

From step I, $K_{eq} = \dfrac{[NOBr_2]}{[NO]\,[Br_3]}$

or $[NOBr_2] = K_{eq}\,[NO]\,[Br_2]$

Substituting in eqn. (i), we get

Rate $= k\,K_{eq}\,[NO]^2\,[Br_2] = k'\,[NO]^2\,[Br_2]$

Hence, order with respect to NO is 2.

83. (c) : $10{\cdot}8$ g of $N_2O_5 = \dfrac{10{\cdot}8}{108}$ mole $= 0{\cdot}1$ mole

No. of half-lives in $9{\cdot}6$ h $= \dfrac{9{\cdot}6}{2{\cdot}4} = 4$

Amount left after 4 half-lives

$$= \frac{1}{2^4} \times 0{\cdot}1 = \frac{0{\cdot}1}{16}\ \text{mole}$$

Moles of N_2O_5 decomposed

$$= 0{\cdot}1 - \frac{0{\cdot}1}{16} = \frac{1{\cdot}5}{16}\ \text{mole}$$

Moles of O_2 formed $= \dfrac{1}{2} \times \dfrac{1{\cdot}5}{16} = \dfrac{1{\cdot}5}{32}$

Volume of O_2 at STP $= \dfrac{1{\cdot}5}{32} \times 22{\cdot}4\ \text{L} = 1{\cdot}05\ \text{L}$

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 frist order reaction does not depend upon initial concentration. It is equal to ln $2/k$.

86. (c) : Diagram (c) represents enothermic reaction with

$E_a = 200 - 150 = 50$ kJ and

$\Delta H = 50 - 150 = -100$ 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 $k = 6 - \dfrac{2000}{T}$

Comparing with log $k = \log\, A - \dfrac{E_a}{2.303\, RT}$

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

and $\dfrac{E_a}{2.303\, R} = 2000$

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

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

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

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

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

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

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