WEEKLY TEST RANKER'S BATCH-01 TEST - 04 Balliwala SOLUTION Date 06-10-2019

[PHYSICS]

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[CHEMISTRY]

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

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

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

Rate of formation of NO_2

$$= \frac{[NO_2]}{dt} = -2\frac{d[N_2O_5]}{dt}$$

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

$$= 12{\cdot}50 \times 10^{-3}\ \text{mol L}^{-1}\ \text{s}^{-1}$$

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

Rate of formation of O_2

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

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

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

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

48. (*b*) :
$$\begin{array}{lcll} A_2\,(g) & \longrightarrow & B\,(g) + \frac{1}{2}C\,(g) & \\ 100 & & 0 \qquad 0 & \\ 100 - p & & p \qquad \frac{1}{2}p & \end{array}$$

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

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

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

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

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

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

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

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

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

53. (*b*) : For zero order reaction, $k = \frac{1}{t}\{[A]_0 - [A]\}$
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, $K_{eq} = \frac{[A]^2}{[A_2]}$

or $\qquad [A] = \sqrt{K_{eq}\,[A_2]} = K_{eq}^{1/2}\,[A_2]^{1/2}$

Hence, rate $= k\,[B_2]\,K_{eq}^{1/2}\,[A_2]^{1/2}$

$$= k'\,[A_2]^{1/2}\,[B_2].$$

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

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, $\frac{r}{4} = k\,a^\alpha\,(2\,b)^\beta$. 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. (*d*) : $k = \frac{2.303}{32}\log\frac{a}{a - 0.99a}$

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

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

$$= \frac{2.303}{k}\log 10^3 = \frac{3\times2.303}{2.303}\times16$$

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

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

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

or $\log\frac{a}{a-x} = \frac{kt}{2.303} = \frac{2.2\times10^{-5}\times60\times90}{2.303}$

$$= 0.0516.$$

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

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

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

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

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

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

$$= \frac{2.303}{k}\log 2 = \frac{2.303}{k}\times0.3010$$

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

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 × 15 min = 30 min.

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

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,

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

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

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

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

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 $\propto [A]^2$.
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 $k_1 < k_2$.

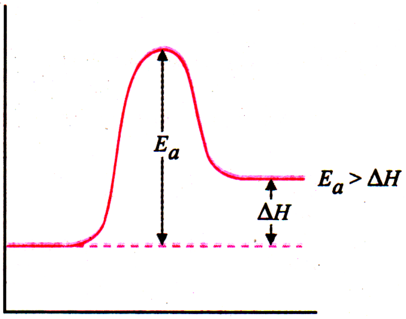
70. **(*b*) :** $k = \frac{2.303}{t}\log\frac{V_\infty}{V_\infty - V_t}$

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

$$= \frac{1{\cdot}26 \times 10^{-4}}{1{\cdot}26 \times 10^{-4} + 3{\cdot}8 \times 10^{-5}} \times 100$$

$$= \frac{1{\cdot}26 \times 10^{-4}}{10^{-4}(1{\cdot}26 + 0{\cdot}38)} \times 100 = 76{\cdot}83\%.$$

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

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

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

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 $= \frac{0{\cdot}6}{0{\cdot}8} = \frac{3}{4}$

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

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

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

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*) : 10·8 g of $N_2O_5 = \frac{10\cdot8}{108}$ mole = 0·1 mole

No. of half-lives in 9·6 h $= \frac{9\cdot6}{2\cdot4} = 4$

Amount left after 4 half-lives

$$= \frac{1}{2^4} \times 0\cdot1 = \frac{0\cdot1}{16} \text{ mole}$$

Moles of N_2O_5 decomposed

$$= 0\cdot1 - \frac{0\cdot1}{16} = \frac{1\cdot5}{16} \text{ mole}$$

Moles of O_2 formed $= \frac{1}{2} \times \frac{1\cdot5}{16} = \frac{1\cdot5}{32}$

Volume of O_2 at STP $= \frac{1\cdot5}{32} \times 22\cdot4 \text{ L} = 1\cdot05 \text{ L}$

84. (*b*) : $r = K[A]^\alpha = k\,a^\alpha$

$1\cdot837\, r = k\,(1\cdot5\,a)^\alpha$

Dividing, $1\cdot837 = (1\cdot5)^\alpha$

On solving, we get $\alpha = 1\cdot5$

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

$$E_a = 200 - 150 = 50 \text{ kJ and}$$
$$\Delta H = 50 - 150 = -100 \text{ kJ.}$$

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 $\log k = 6 - \frac{2000}{T}$

Comparing with $\log k = \log A - \frac{E_a}{2\cdot303\,RT}$

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

and $\frac{E_a}{2\cdot303\,R} = 2000$

or $E_a = 2000 \times 2\cdot303 \times 8\cdot314 \text{ J mol}^{-1}$
$= 38294 \text{ J mol}^{-1} \approx 38\cdot3 \text{ kJ mol}^{-1}$

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

$$t_{max} = \frac{2{\cdot}303}{(k_1 - k_2)} \log \frac{k_1}{k_2}$$

$$k_1 = 4{\cdot}0 \times 10^{-2}$$

$$\therefore \quad k_2 = 4{\cdot}0 \times 10^{-2} \times 0{\cdot}15 = 6 \times 10^{-3}$$

$$\therefore \quad t_{max} = \frac{2{\cdot}303}{(4 - 0{\cdot}6)\,10^{-2}} \log \frac{4\times10^{-2}}{6\times10^{-3}}$$

$$= \frac{2{\cdot}303}{3{\cdot}4\times10^{-2}} \times 0{\cdot}82 = 55{\cdot}6 \text{ min}$$