DPP

DAILY PRACTICE PROBLEMS

Class : XIIth

Date :

Solutions

Subject : CHEMISTRY

DPP No. : 5

Topic :- Chemical Kinetics

1 **(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.

$-\int d[R] = k \int dt$

$-[R] = kt + \mathrm{I}$ ...(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 $[R] = -kt + [R]_0$

2 **(a)**

For first order reaction,

Half-life period $(t_{1/2}) = \frac{0.693}{k}$

Where, k=rate constant

$(t_{1/2}) = \frac{0.693}{69.3} s^{-1}$

$=0.01\ s^{-1}$

3 **(b)**

For nth order reaction :

$$t_{1/2} \propto \frac{1}{a^{n-1}}$$

For second order reaction

$$t_{1/2} = \frac{1}{ka} = \frac{1}{0.5 \times 0.2} = \frac{100}{10} = 10 \text{ min}$$

4 **(d)**

$r = K[\mathrm{CH_3COCH_3}]^a[\mathrm{Br_2}]^b[\mathrm{H^+}]^c$

$\therefore \quad 5.7 \times 10^{-5} = K[0.30]^a[0.05]^b[0.05]^c \quad ...(1)$

$5.7 \times 10^{-5} = K[0.30]^a[0.10]^b[0.5]^c \quad ...(2)$

$1.2 \times 10^{-4} = K[0.30]^a[0.10]^b[0.10]^c \quad ...(3)$

$3.1 \times 10^{-4} = K[0.40]^a[0.05]^b[0.20]^c \quad ...(4)$

By (1) and (2) $a = 1$

By (2) and (3) $b = 0$

By (3) and (4) $c = 1$

$\therefore \quad r = K[\mathrm{CH_3COCH_3}]^1[\mathrm{Br_2}]^0[\mathrm{H^+}]^1$

5 **(c)**

Unit of rate constant

$= \dfrac{time^{-1}}{conc^{(n-1)}}$

Where, n=order of reaction

Given, unit of rate constant $= L\,mol^{-1}s^{-1}$

$\therefore \; L\,mol^{-1}s^{-1} = \dfrac{(s)^{-1}}{(L\,mol^{-1})^{n-1}}$

$= \dfrac{(s)^{-1}}{(L\,mol^{-1})^{n-1}}$

$= s^{-1}(L\,mol^{-1})^{n-1}$

Or 1=n-1

Or n=2

$\therefore$ order of reaction =2

6 **(c)**

Activation energy of a chemical reaction can be determined by evaluating rate constants at two different temperatures

$\log\dfrac{k_2}{k_1} = \dfrac{E_a}{2.303\,R}\left(\dfrac{1}{T_1} - \dfrac{1}{T_2}\right)$

7 **(c)**

Molecularity can never be fractional.

9 **(d)**

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

For this reaction, rate

$(r_1) = k[SO_2]_1^2[O_2]_1 \quad ....(i)$

On doubling the volume of vessel, concentration would be half. Hence,

$\text{Rate}(r_2) = k\left(\frac{[SO_2]_1}{2}\right)^2\left(\frac{[O_2]_1}{2}\right) = \frac{r_1}{8}$

$\dfrac{r_1}{r_2} = 8{:}1$

10 **(c)**

$r = k[R\mathrm{Cl}]$

If $[RCl]=\frac{1}{2}$, then rate$=\frac{r}{2}$

11 **(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 $A \to B$ is the rate determining step of the reaction

12 **(b)**

$$k=\frac{2.303}{t}\log_{10}\frac{a}{a-x}$$

$$=\frac{2.303}{10}\log_{10}\frac{100}{80}$$

$$=\frac{2.303}{10}\ [\log 10-3\log 2]$$

$$=\frac{2.303}{10}\ [1-3\times 0.3010]$$

$k=0.0223$

13 **(d)**

$E_a(A\to B)=80\ \text{kJ mol}^{-1}$

Heat of reaction $(A\to B)=200\ \text{kJ mol}^{-1}$

For $(B\to A)$ backward reaction,

$E_a(B\to A)=E_a(A\to B)+$ heat of reaction

$=80+200=280\text{kJ mol}^{-1}$

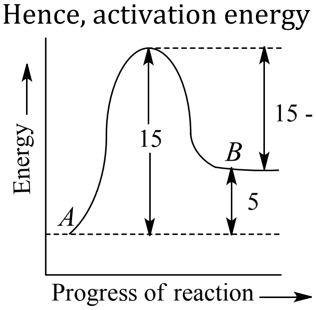
14 **(c)**

For endothermic reaction $A\to B$

Activation energy $=15$ kcal/mol

Energy of reaction $=5$ kcal/mol

Hence, activation energy for the reaction $B\to A$ is $15-5=10$ kcal/mol



15 **(d)**

For zero order $[A]_t=[A]_0-kt$

$0.5=[A]_0-2\times10^{-2}\times 25$

$\therefore \quad [A]_0=1.0$ M

16 **(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

17 **(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.

18 **(a)**

Temperature coefficient

$$= \frac{\text{rate of recation at } 35°C}{\text{rate of recation at } 25°C} = 2$$

Thus, increase in rate is two times, when temperature is increased 10℃. Hence, by the increase of 70℃(100-30=70℃), the increase in rate will be

$= (2)^7 \quad \because 70° = 7 \times 10°$

$= 128$ times

19 **(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%

20 **(b)**

$\frac{1}{2} A \rightarrow 2B$

Remember for $a\ A \rightarrow bB$

$-\frac{1}{a}\frac{d[A]}{dt}=\frac{1}{b}\frac{d[B]}{dt}$= Rate of reaction

For the given reaction

$-\frac{2d[A]}{dt}=\frac{1}{2}\frac{d[B]}{dt}$=rate of reaction

Rate of disappearance of A

$$=-\frac{d[A]}{dt}=\frac{1}{2\times 2}\frac{d[B]}{dt}$$

$$=\frac{1}{4}\frac{d[B]}{dt}$$

ANSWER-KEY										
Q.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
A.	**C**	**A**	**B**	**D**	**C**	**C**	**C**	**D**	**D**	**C**
Q.	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
A.	**A**	**B**	**D**	**C**	**D**	**B**	**D**	**A**	**D**	**B**