DPP
DAILY PRACTICE PROBLEMS

Class : XIIth
Date :

Solutions

Subject : CHEMISTRY
DPP No. : 2

Topic :- Chemical Kinetics

1 **(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=3.2×10^{-3}

$\therefore$ Rate constant at 300 K=$2 \times 3.2 \times 10^{-3}$

$= 6.4 \times 10^{-3}$

2 **(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

$\frac{-dc}{dt}$.

3 **(c)**

$$k = \frac{2.303}{t} log \frac{A_0}{A_t}$$

$$= \frac{2.303}{2 \times 10^4} log \frac{800}{50}$$

$= 1.386 \times 10^4 s^{-1}$

4 **(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.

5 **(c)**

Given, $R_1 = k[A]^2[B]$

According to equation $R_2 = k[3A]^2[B]$

$= k \times 9[A]^2 2[B]$

$= 18 \times k[A]^2[B]$
$= 18R_1$

7 **(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

8 **(c)**

For second order reaction, $\frac{dx}{dt}$ (rate) $\propto [A]^2$

∵ Rate Of reaction increases four times when concentration of reaction is increased two times.

∴ It is second order reaction.

9 **(b)**

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

$\frac{d[H_2]}{dt} = -0.3 \times 10^{-4} ms^{-1}$

$rate = -\frac{1}{3}\frac{d[H_2]}{dt} = +\frac{1}{2}\frac{d[NH_3]}{dt}$

$= \frac{d[NH_3]}{dt} = -\frac{2}{3}\frac{d[H_2]}{dt}$

$= -\frac{2}{3} \times (-0.3 \times 10^{-4})$

$= 0.2 \times 10^{-4}$

10 **(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.

11 **(d)**

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

$\because$ The activation energy for the forward reaction $= 50$ kcal

$\therefore$ The activation energy for the backward reaction$=50+22=72$ kcal.

12 **(d)**

Only those collisions are effective collisions which are energetic enough and cross over the threshold energy level.

13 **(b)**

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{480}\ s^{-1}$$

$$k = 1.44 \times 10^{-3}\ s^{-1}$$

14 **(b)**

It is a characteristic of zero order reaction.

15 **(a)**

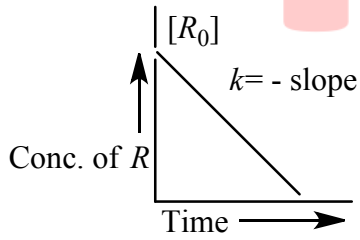
Follow review of order of reaction.

16 **(a)**

Average life is defined as, "reciprocal of decay constant." If decay constant for a reaction is λ then,

Average life$=\frac{1}{\lambda}$

17 **(d)**



[variation in the concentration V_s time plot for a zero order reaction]

18 **(a)**

Energy of activation does not depend on the stoichiometry of change. It is characteristic value for a chemical reaction.

19 **(d)**

Use ;$r = K[A]^m[B]^n$

20 **(a)**

Slow reaction rate indicates higher free energy of activation

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