# Speed Test-46

**1.** **(d)** $k=\frac{2.303}{t}\log\frac{a}{a-x}=\frac{2.303}{24}\log\frac{1}{\frac{1}{8}}=\frac{2.303}{24}\log 8$

**2.** **(b)** Since doubling the concentration of *B* does not change half life, the reaction is of 1st order w.r.t. *B*.
Order of reaction with respect to $A=1$ because rate of reaction doubles when concentration of A is doubled keeping concentration of *B* constant.
$\therefore$ Order of reaction $=1+1=2$ and units of second order reaction are L $mol^{-1}$ $sec^{-1}$.

**3.** **(d)** The molecularity of a reaction is the number of reactant molecules taking part in a single step of the reaction. Thus the reaction involving two different reactant can never be unimolecular.

**4.** **(a)** Let the rate law be $r=k[A]^x[B]^y$

Divide (3) by (1) $\frac{0.10}{0.10}=\frac{[0.024]^x[0.035]^y}{[0.012]^x[0.035]^y}$

$\therefore 1=[2]^x, x=0$

Divide (2) by (3) $\frac{0.80}{0.10}=\frac{[0.024]^x[0.070]^y}{[0.024]^x[0.035]^y}$

$\therefore 8=(2)^y, y=3$

Hence rate equation, $R=k[A]^0[B]^3=k[B]^3$

**5.** **(a)** If we write rate of reaction in terms of concentration of $NH_3$ and $H_2$, then

Rate of reaction $=\frac{1}{2}\frac{d[NH_3]}{dt}=-\frac{1}{3}\frac{d[H_2]}{dt}$

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

**6.** **(a)** As doubling the initial conc. doubles the rate of reaction, order =1

**7.** **(b)** Rate law has to be determined experimentally as $Cl_2$ is raised to power $\frac{1}{2}$ in rate law whereas its stichiometric coefficient in balanced chemical equation is 1.

**8.** **(a)**

**9.** **(c)** $\frac{(t_{1/2})_1}{(t_{1/2})_2}=\left(\frac{a_2}{a_1}\right)^{n-1}$ ; $\frac{120}{240}=\left(\frac{4\times10^{-2}}{8\times10^{-2}}\right)^{n-1}$ ; $n=2$

**10.** **(b)**

**11.** **(b)**

**12.** **(c)** Activation energy is lowered in presence of +ve catalyst.

**13.** **(d)** Overall order = sum of orders w.r.t each reactant.
Let the order be *x* and *y* for G and H respectively

| Exp.No. | [G]mole $litre^{-1}$ | [H]mole $litre^{-1}$ | rate(mole $litre^{-}time^{-1}$) |
|---|---|---|---|
| 1 | *a* | *b* | *r* |
| 2 | 2*a* | 2*b* | 8*r* |
| 3 | 2*a* | *b* | 2*r* |

$\because$ For (1) and (3), the rate is doubled when conc. of G is doubled keeping that of H constant i.e.,
rate $\propto [G]$ $\therefore x=1$
From (2) and (3), $y=2$
$\therefore$ Overall order is 3.

**14.** **(b)** For a first order reaction

$k=\frac{2.303}{t}\log\frac{a}{a-x}=\frac{2.303}{40}\log\frac{0.1}{0.025}$

$=\frac{2.303}{40}\log 4=\frac{2.303\times0.6020}{40}$

$=3.47\times10^{-2}\ min^{-1}$

Rate $=k[A]=3.47\times10^{-2}\times0.01$
$=3.47\times10^{-4}$ M/min

**15.** **(c)** Third order

**16.** **(c)** For a first order reaction

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

when $t=t_{1/2}$

$k=\frac{2.303}{t_{1/2}}\log_{10}\frac{a}{a-a/2}$

or $t_{1/2}=\frac{2.303}{k}\log_{10}2=\frac{\ln 2}{k}$

**17.** **(d)** The integrated rate equations are different for the reactions of different reaction orders. We shall determine these equations only for zero and first order chemical reactions.

**18.** **(b)** $T_2=T$(say), $T_1=25°C=298K$,
$E_a=104.4\ kJ\ mol^{-1}=104.4\times10^3\ J\ mol^{-1}$
$k_1=3\times10^{-4}, k_2=?$,

$\log\frac{k_2}{k_1}=\frac{E_a}{2.303R}\left[\frac{1}{T_1}-\frac{1}{T_2}\right]$

$\log\frac{k_2}{3\times10^{-4}}=\frac{104.4\times10^3 J\ mol^{-1}}{2.303\times(8.314\ J\ K^{-1}mol^{-1})}\left[\frac{1}{298}-\frac{1}{T}\right]$

As $T\to\infty, \frac{1}{T}\to0$

$$\therefore \log\frac{k_2}{3\times10^{-4}}=\frac{104.4\times10^3\ \text{J mol}^{-1}}{2.303\times8.314\times298}$$

$$\log\frac{k_2}{3\times10^{-4}}=18.297, \frac{k_2}{3\times10^{-4}}=1.98\times10^{18}$$

$$k_2=(1.98\times10^{18})\times(3\times10^{-4})=6\times10^{14}\text{s}^{-1}$$

**19.** **(c)** Adsorption lowers the activation energy.

**20.** **(a)** For the change $2A+3B\rightarrow$ products

$$-\frac{1}{2}\frac{d[A]}{dt}=-\frac{1}{3}\frac{d[B]}{dt}\quad ,\frac{1}{2}r_1=\frac{1}{3}r_2;\quad 3r_1=2r_2$$

**21.** **(c)** The $t_{1/2}$ is 15 minutes. To fall the concentration from 0.1 to 0.025 we need two half lives i.e., 30 minutes.

**22.** **(b)** According to Arrhenius equation

$$\ln\frac{k_2}{k_1}=\frac{E_a}{R}\left(\frac{1}{T_1}-\frac{1}{T_2}\right)$$

$$=-\frac{E_a}{R}\left(\frac{1}{T_2}-\frac{1}{T_1}\right)$$

$$\ln\frac{k_1}{k_2}=-\frac{E_a}{R}\left(\frac{1}{T_1}-\frac{1}{T_2}\right)$$

**23.** **(c)** $t_{1/4}=\frac{2.303}{k}\log\frac{1}{3/4}=\frac{2.303}{k}\log\frac{4}{3}$

$$=\frac{2.303}{k}(\log 4-\log 3)=\frac{2.303}{k}(2\log 2-\log 3)$$

$$=\frac{2.303}{k}(2\times0.301-0.4771)=\frac{0.29}{k}$$

**24.** **(d)** For a zero order reaction,

$t_{1/2}\propto a_0$ (initial concentration or initial pressure)

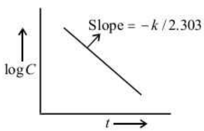
$(t_{1/2})_1\propto P_1$

$(t_{1/2})_2\propto P_2$

$$\frac{(t_{1/2})_2}{(t_{1/2})_1}=\frac{P_2}{P_1},\ \frac{(t_{1/2})_2}{45}=\frac{16}{4}$$

$$(t_{1/2})_2=\frac{16}{4}\times45=180\text{ min}$$

**25.** **(b)**




**26.** **(b)** Plot given is for zero order reaction.

**27.** **(c)** Since for every 10ºC rise in temperature rate doubles for 50ºC rise in temperature increase in reaction rate $=2^5=32$ times

**28.** **(a)** $k=\frac{2.303}{1386}\log\frac{100}{100-75}$ On solving we get . $k=10^{-3}$

**29.** **(a)**

**30.** **(a)** Presence of catalyst does not affect enthalpy change of reaction $\Delta H_R=E_f-E_b=180-200=-20$ kJ/mol

**31.** **(c)** Given $t_{1/2}=15$ minutes

Total time $(T)=1$ hr $=60$ min

From $T=n\times t_{1/2}$

$$n=\frac{60}{15}=4$$

Now from the formula $\frac{N}{N_0}=\left(\frac{1}{2}\right)^n$

$$=\left(\frac{1}{2}\right)^4=\frac{1}{16}$$

Where $N_0$ = initial amout

N = amount left after time $t$

hence the amount of substance left after 1 hour will be $\frac{1}{16}$

**32.** **(d)** $\text{Rate}_1=k[A]^n[B]^m$

$$\text{Rate}_2=k[2A]^n\left[\frac{1}{2}B\right]^m$$

$$\therefore\ \frac{\text{Rate}_2}{\text{Rate}_1}=\frac{k[2A]^n\left[\frac{1}{2}B\right]^m}{k[A]^n[B]^m}=(2)^n\left(\frac{1}{2}\right)^m$$

$$=2^n\ .\ (2)^{-m}=2^{n-m}$$

**33.** **(a)** On adding eq. (i) and eq. (ii) we get

$$O_3(g)+O(g)\longrightarrow 2O_2(g)$$

Hence overall rate constant $=K_i\times K_{ii}$

$=5.2\times10^9\times2.6\times10^{10}\approx1.4\times10^{20}\ \text{mol}^{-1}\ \text{L s}^{-1}$

$$O_3(g)+Cl^\bullet(g)\longrightarrow O_2(g)+ClO^\bullet(g),\ K_i$$

$$ClO^\bullet(g)+O^\bullet(g)\longrightarrow O_2(g)+Cl^\bullet(g),\ K_{ii}$$

$$O_3(g)+O^\bullet(g)\longrightarrow 2O_2(g),\ K_{Rate}=K_i\times K_{ii}$$

**34.** **(b)** $k=Ae^{-E_a/RT}$ $\log k=\log A-\frac{E_a}{2.303R}.\frac{1}{T}$

Plot of log k Vs. $\frac{1}{T}$

Straight line Slope $=\frac{-E_a}{2.303R}$

**35.** **(b)** $r=k[O_2][NO]^2$. When the volume is reduced to 1/2, the conc. will double

$\therefore$ New rate $=k[2O_2][2NO]^2=8k[O^2][NO]^2$

The new rate increases to eight times of its initial.

**36. (d)** From rate law

$$-\frac{1}{2}\frac{dSO_2}{dt}=-\frac{dO_2}{dt}=\frac{1}{2}\frac{dSO_3}{dt}$$

$$\therefore -\frac{dSO_2}{dt}=-2\times\frac{dO_2}{dt}$$
$$=-2\times 2.5\times 10^{-4}$$
$$=-5\times 10^{-4}\text{ mol L}^{-1}\text{ s}^{-1}$$

**37. (c)** $k_1=A_1e^{-E_{a_1}/RT}$ .........(i)

$k_2=A_2e^{-E_{a_2}/RT}$ ........(ii)

On dividing eqn (i) from eqn. (ii)

$\frac{k_1}{k_2}=\frac{A_1}{A_2}(E_{a_2}-E_{a_1})/RT$ ........(iii)

Given $E_{a_2}=2E_{a_1}$

On substituting this value in eqn. (iii)

$k_1=k_2A\times e^{E_{a_1}/RT}$

**38. (d)** Enthalpy of reaction $(\Delta H)=E_{a_{(f)}}-E_{a_{(b)}}$ for an endothermic reaction $\Delta H$ = +ve hence for $\Delta H$ to be positive

$E_{a_{(b)}}<E_{a_{(f)}}$

**39. (d)**

**40. (a)** $\ln\frac{K_1}{K_2}=\frac{E_a}{R}\left(\frac{1}{T_1},-\frac{1}{T_2}\right)$

$$\ln 4=\frac{E_a}{8.314}\left(\frac{310-300}{310\times 300}\right)$$

$$2\ln 2=\frac{E_a}{8.314}\left(\frac{310-300}{310\times 300}\right)$$

$$E_a=\frac{0.693\times 2\times 8.314\times 300\times 310}{10}=107.2\text{ kJ/mol}$$

**41. (d)** Since the slow step is the rate determining step hence if we consider option (A) we find

Rate $=k[Cl_2][H_2S]$

Now if we consider option (B) we find

Rate $=k[Cl_2][HS^-]$ ...(i)

For equation,

$H_2S \rightleftharpoons H^+ + HS^-$

$$K=\frac{[H^+][HS^-]}{H_2S}$$

or $[HS^-]=\frac{K[H_2S]}{H^+}$

Substituting this value in equation (i) we find

$$\text{Rate}=k[Cl_2]K\frac{[H_2S]}{H^+}=k'\frac{[Cl_2][H_2S]}{[H^+]}$$

Thus slow step should involve 1 molecule of $Cl_2$ and 1 molecule of $H_2S$.

hence only, mechanism (A) is consistent with the given rate equation.

**42. (a)** Arrhenius equation is given by

$k=Ae^{-E_a/(2.303RT)}$

Taking log on both sides, we get

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

Arrhenius plot a graph between log $k$ and $\frac{1}{T}$ whose slope is $\frac{-E_a}{2.303R}$.

**43. (d)** The graph show that reaction is exothermic.

$$\log k=\frac{-\Delta H}{RT}+1$$

For exothermic reaction $\Delta H<0$

$\therefore$ log k Vs $\frac{1}{T}$ would be negative straight line with positive slope.

**44. (a)** As per Arrhenius equation $(k=Ae^{-E_a/RT})$, the rate constant increases exponentially with temperature.

**45. (a)** Rate law for first order reaction $=k[N_2O_5]$

| | $2N_2O_5(g)\longrightarrow$ | $4NO_2(g)$ | $+O_2(g)$ |
|---|---|---|---|
| t = 0 min (Pressure in mm Hg) | 50 | 0 | 0 |
| t = 30 min (Pressure in mm Hg) | $50-2p$ | $4p$ | $p$ |

Total pressure $50-2p+4p+p=50+3p$
$=87.5$ mm Hg
$\therefore$ P = 12.5 mm Hg
$\therefore$ $P_0$ = 50 & P (t = 30 min)
= 25 for $N_2O_5$ reactant

$$\therefore\ k=\frac{2.303}{30\text{ min}}\times\log\left(\frac{50}{25}\right)=\frac{2.303}{60\text{ min}}\times\log\left(\frac{50}{x}\right)$$

On solving $x$ = 12.5 mm Hg = $50-2p$
$\therefore$ P = 18.75 mm Hg
$\therefore$ Total pressure = 50 + 3P = 106.25 mm Hg