CBSE Class 12 physics
Important Questions
Chapter 4
Chemical Kinetics

3 Marks Questions

1. For the reaction A+B $\rightarrow$ C+D, the rate of reaction doubles when the concentration of A doubles, provided the concentration of B is constant. To what order does A enter into the rate expression?

Ans. A+B $\rightarrow$ C+D

$Rate[A]^x$

Rate = 1 when [A] = 1 ---------- 1)

Rate = 2 when [A] = 2-----------2)

Dividing equation 2) by 1)

$$\frac{2}{1}\alpha\frac{(2)^x}{1^x}$$

$2^1 \alpha\ (2)^x$

$\therefore x = 1$

The reaction is first order reaction.

2. A chemical reaction 2A $\Leftrightarrow$ 4B+C in gas phase occurs in a closed vessel. The concentration of B is found to be increased by $5\times10^{-3} mole\ L^{-1}$ in 10 second. Calculate (i) the rate of appearance of B (ii) the rate of disappearance of A?

Ans. 2A $\rightarrow$ 4B +C

$$-\frac{1}{2}\frac{d[A]}{dt}=\frac{1}{4}\frac{d[B]}{dt}=\frac{d[C]}{dt}$$

i) Rate of disappearance of B

$$=\frac{5\times10^{-3}}{10\ 5}mol\ /L^{-1}=5\times\ 10^{4}\ mol\ L^{-1}\ s^{-1}$$

ii) $\frac{-d[A]}{dt}=\frac{2}{4}\frac{d[B]}{dt}=\frac{1}{2}\frac{d[B]}{dt}$

$$=\frac{1}{2}\times5\times10^{-4}\ mol\ L^{-1}\ s^{-1}=2.5\times10^{4}\ mol\ L^{-1}\ s^{-1}$$

3. For the following reactions, write the rate of reaction expression in terms of reactants and products?

i) $4NH_3(g)+5O_2(g)4\rightarrow NO(g)+6H_2O(g)$

ii) $2N_2O_5\ 2NO_2+O_2$

Ans.

In terms of reactant	In terms of products
i) $R_1 - \frac{-1}{4}\frac{\Delta[NH_3]}{\Delta t}$	$R_3=\frac{1}{4}\frac{\Delta[NO]}{\Delta t}$
$R_2=\frac{-1}{5}\frac{\Delta[O_2]}{\Delta t}$	$R_4=\frac{1}{6}\frac{\Delta[H_2O]}{\Delta t}$

$$\frac{1}{4}R_1=\frac{1}{5}R_2=\frac{1}{4}R_3=\frac{1}{6}R_4$$

$$-\frac{1}{4}\frac{\Delta[NH_3]}{\Delta t}=-\frac{1}{5}\frac{\Delta[O_2]}{\Delta t}=\frac{1}{4}\frac{\Delta[NO]}{\Delta t}=\frac{1}{6}\frac{\Delta[H_2O]}{\Delta t}$$

II) In terms of reactant	In terms of product

R1 = $-\dfrac{\Delta[N_2O_5]}{\Delta t}$	$R_2 = \dfrac{\Delta[NO_2]}{\Delta t}$
	$R_3 = \dfrac{\Delta[O_2]}{\Delta t}$

$$\frac{1}{2}R_1 = \frac{1}{2}R_2 = R_3 \text{ - } \frac{\Delta[N_2O_5]}{\Delta t} = \frac{1}{2}\ \frac{\Delta[NO_2]}{\Delta t} = \frac{\Delta[O_2]}{\Delta t}$$

4. The reaction $2N_2O_5(g) \rightarrow 2NO_2(g) + O_2(g)$ was studied and the following data were collected :

S.no (mol/L/min)	$[N_2O_5]\ mol\ L^{-1}$	Rate of disappearance of $[N_2O_5]$ (mol/L/min
1.	1.13×10^{-2}	34×10^{-5}
2.	$0.84\ \times10^{-2}$	25×10^{-5}
3.	0.62×10^{-2}	18×10^{-5}

Determine

i) The order

ii) The rate law.

iii) Rate constant for the reaction.

Ans. Let the order of reaction be x

$$Rate = K[N_2O_5]^x$$

i) From the data -

$34\times10^{-5} = (1.13\times\ 10^{-2})^x$ ----------1)

$25\times10^{-5} = (0.84\times\ 10^{-2})^x$ ----------2)

$18\times10^{-5} = \left(0.62\times10^{-2}\right)^{x}$ ----------3)

Dividing 1) by 2)

$$\frac{34\times10^{-5}}{25\times10^{-5}} = \left(\frac{1.13\times10^{-2}}{0.84\times10^{-2}}\right)^{x}$$

(1.36) = $(1.35)^{x}$

X=1

The order of reaction with respect with respect to N_2O_5 is 1

ii) Rate law $R = K\ [N_2O_5]$

iii) Rate constant , K = $\frac{Rate}{[N_2O_5]}$ = $\frac{18\times10^{-5}\ \text{mol/L/min}}{0.62\times10^{-2}\ \text{mol/L}}$ = $0.29\ min^{-1}$

5. The following experimental data was collected for the reaction:

$Cl_2(g)+2NO(g)\rightarrow 2NOCl(g)$

Trial	Intial conc. Of $Cl_2\ (mol/L)$	[NO] mol/L	Initial Rate,(mol/L/s)
1	0.10	0.010	1.2×10^{-4}
2	0.10	0.030	10.8×10^{-4}
3	0.20	0.030	21.6×10^{-4}

Construct the rate equation for the reaction.

Ans. Order of NO is 2

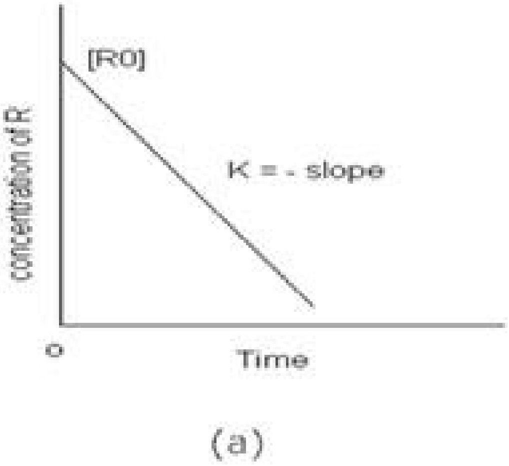
Rate law = $K\ [Cl_2]\ [NO]^2$

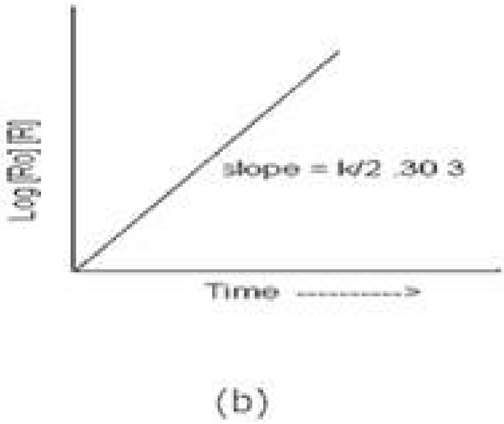
6. Draw a graph for

a) Concentration of reactant against time for a zero order reaction.

b) Log [Ro]/ [R] against time for a first order reaction.

Ans.





7. In general it is observed that the rate of a chemical reaction doubles with every 10^0 **rise in temperature. If this generalization holds for a reaction in the temperature range295K to 305K, what would be the activation energy for this reaction?** $\left(R = 8.314 Jk^{-1} mol^{-1}\right)$

Ans. $T_1 = 295K$ $T_2 = 305K$

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

$$K_2 = 2k_1$$

$$E_a = 2.303 \times 8.314 \times \left[\frac{305 \times 295}{305 - 295}\right] \log \frac{2k_1}{k_1}$$

= $2.303 \times 8.314 \times 8997.5K\ log2$

= 51855.2 J/mol (log2 = 0.3010)

8. The rate constant for a reaction is $1.5\times10^{7}s^{-1}$ **at** $50^{0}C$ **and** $4.5\times10^{7}s^{-1}$ **at** $100^{0}C$.

Calculate the value of activation energy for the reaction $R=8.314\ JK^{-1}mol^{-1}$ **?**

Ans. $\log\frac{k_2}{k_1}=\frac{Ea}{2.303_R}\left(\frac{T_2-T_1}{T_1T_2}\right)$

$$\log\frac{4.5\times10^{7}}{1.5\times10^{7}}=\frac{Ea}{2.303\times3.314}\left(\frac{373-323}{373\times323}\right)$$

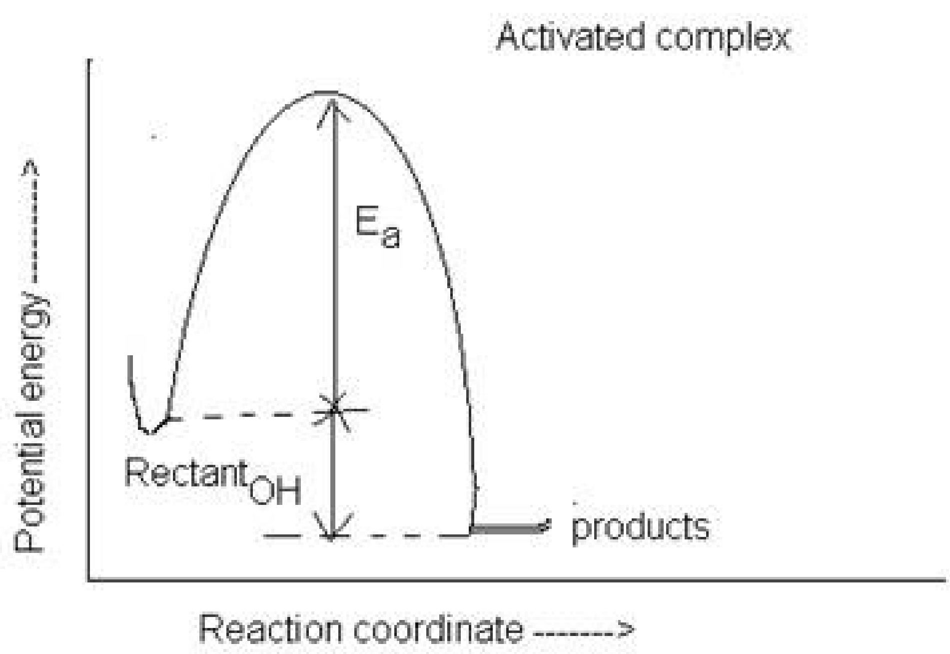
$$\log1.5=\frac{Ea}{2.303\times3.314}\left(\frac{50}{373\times323}\right)$$

$$Ea=\left(\frac{2.303\times3.314\times373\times323}{50}\right)\times\log1.5$$

$$=22\ KJ/mol$$

9. Plot a graph showing variation of potential energy with reaction. coordinate?

Ans.



10. The conversion of molecules X to Y follows second order kinetics. If concentration

of X is increased to three times how will it affect the rate of formation of Y?

Ans. The reaction $X \rightarrow Y$ follows second order kinetics.

Therefore, the rate equation for this reaction will be:

Rate = $k[X]^2$ (1)

Let [X] = *a* mol^{-1}, then equation (1) can be written as:

$\text{Rate}_1 = k.(a)_2$

= ka_2

If the concentration of X is increased to three times, then [X] = $3a\,mol\,L^{-1}$

Now, the rate equation will be:

Rate = $k(3a)_2$

$= 9(ka^2)$

Hence, the rate of formation will increase by 9 times.

11. A first order reaction has a rate constant $1.1510^{-3}\,S^{-1}$. How long will 5 g of this reactant take to reduce to 3 g?

Ans. From the question, we can write down the following information:

Initial amount = 5 g

Final concentration = 3 g

Rate constant = $1.15\,10^{-3}\,s^{-1}$

We know that for a 1storder reaction,

$$t = \frac{2.303}{k}\log\frac{[R]_0}{[R]}$$

$= \frac{2.303}{1.15 \times 10^{-3}} \log \frac{5}{3}$

$= \frac{2.303}{1.15 \times 10^{-3}} \times 0.2219$

= 444.38 s

= 444 s (approx)

12. For the reaction $R \rightarrow P$, the concentration of a reactant changes from 0.03 M to 0.02 M in 25 minutes. Calculate the average rate of reaction using units of time both in minutes and Pseconds.

Ans. Average rate of reaction $-\frac{\Delta[R]}{\Delta t}$

$= -\frac{[R]_2 - [R]_1}{t_2 - t_1}$

$= -\frac{0.02 - 0.03}{25} \text{M min}^{-1}$

$= -\frac{-0.01}{25} \text{M min}^{-1}$

$= 4 \times 10^{-4} \text{M min}^{-1}$

$= \frac{4 \times 10^{-4}}{60} \text{M s}^{-1}$

$= 6.67 \times 10^{-6} \text{M s}^{-1}$

13. For the reaction: $2A + B \rightarrow A_2B$ the rate = $k[A][B]_2$ with $k = 2.0 \times 10^{-6} \text{mol}^{-2}\text{L}^2\text{s}^{-1}$. Calculate the initial rate of the reaction when $[A] = 0.1 \text{mol L}^{-1}$, $[B] = 0.2 \text{mol L}^{-1}$. Calculate the rate of reaction after [A] is reduced to 0.06mol L^{-1}.

Ans. The initial rate of the reaction is

Rate = $k[A][B]_2$

$$=\left(2.0\times10^{-6}mol^{-2}L^2s^{-1}\right)\left(0.1mol\,L^{-1}\right)\left(0.2mol\,L^{-1}\right)^2$$

$$=8.0\times10^{-9}mol^{-2}L^2s^{-1}$$

When [A] is reduced from $0.1\ mol\ L^{-1}$ to $0.06\ mol^{-1}$, the concentration of A reacted = $\left(0.1-0.06\right)mol\,L^{-1}=0.004\,mol\,L^{-1}$

Therefore, concentration of B reacted $=\frac{1}{2}\times0.04\,mol\,L^{-1}=0.02mol\,L^{-1}$

Then, concentration of B available, $[B]=(0.2-0.02)mol\,L^{-1}$

$$=0.18mol\,L^{-1}$$

After [A] is reduced to $0.06\ mol^{-1}$, the rate of the reaction is given by,

Rate = $k[A][B]_2$

$$=\left(2.0\times10^{-6}mol^{-2}L^2s^{-1}\right)\left(0.06mol\,L^{-1}\right)\left(0.18mol\,L^{-1}\right)^2$$

$$=3.89\,mol\,L^{-1}s^{-1}$$

14. A reaction is second order with respect to a reactant. How is the rate of reaction affected if the concentration of the reactant is

(i) doubled (ii) reduced to half?

Ans. Letthe concentration of the reactant be [A] = *a*

Rate of reaction, $R=k\left[A\right]_2$

$$=ka^2$$

(i) If the concentration of the reactant is doubled, i.e. [A] = 2*a*, then the rate of the reaction would be

$R' = k(2a)^2$

$= 4ka^2$

= 4 R

Therefore, the rate of the reaction would increase by 4 times.

(ii) If the concentration of the reactant is reduced to half, i.e. $[A] = \frac{1}{2}a$, then the rate of the reaction would be $R' = k\left(\frac{1}{2}a\right)^2$

$= \frac{1}{4}ka^2$

$= \frac{1}{4}R$

Therefore, the rate of the reaction would be reduced to $\frac{1}{4}$th

15. A reaction is first order in A and second order in B.

(i) Write the differential rate equation.

(ii) How is the rate affected on increasing the concentration of B three times?

(iii) How is the rate affected when the concentrations of both A and B are doubled?

Ans. (i) The differential rate equation will be

$$-\frac{d[R]}{dt} = k[A][B]^2$$

(ii) If the concentration of B is increased three times, then

$$-\frac{d[R]}{dt}=k[A][3B]^2$$

$$=9.k[A][B]^2$$

Therefore, the rate of reaction will increase 9 times.

(iii) When the concentrations of both A and B are doubled,

$$-\frac{d[R]}{dt}=k[A][B]^2$$

$$=k[2A][2B]^2$$

$$=8.k[A][B]^2$$

Therefore, the rate of reaction will increase 8 times.

16. Calculate the half-life of a first order reaction from their rate constants given below:

(i) $200s^{-1}$ **(ii)** 2 min^{-1} **(iii) 4 years-1**

Ans. (i) Half life, $t_{1/2}=\frac{0.693}{k}$

$$=\frac{0.693}{200\text{min}^{-1}}$$

$=3.4\times10^{-3}s$ (approximately)

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

$$=\frac{0.693}{2\text{min}^{-1}}$$

= 0.35 min (approximately)

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

$$= \frac{0.693}{4 \text{years}^{-1}}$$

= 0.173 years (approximately)

17. The half-life for radioactive decay of 14C is 5730 years. An archaeological artifact wood had only 80% of the 14C found in a living tree. Estimate the age of the sample.

Ans. Here, $k = \frac{0.693}{t_{1/2}}$

$$= \frac{0.693}{5730} \text{years}^{-1}$$

It is known that,

$$t = \frac{2.303}{k} \log \frac{[R]_0}{[R]}$$

$$= \frac{2.303}{\frac{0.693}{5730}} \log \frac{100}{80}$$

= 1845 years (approximately)

Hence, the age of the sample is 1845 years.

18. The rate constant for a first order reaction is 60 s-1. How much time will it take to reduce the initial concentration of the reactant to its 1/16thvalue?

Ans. It is known that,

$$t = \frac{2.303}{k} \log \frac{[R]_0}{[R]}$$

$$= \frac{2.303}{60\,S^{-1}} \log \frac{1}{\frac{1}{16}}$$

$$= \frac{2.303}{60\,S^{-1}} \log 16$$

$= 4.6 \times 10^{-2}$ s (approximately)

Hence, the required time is 4.6×10^{-2}s.

19. A first order reaction takes 40 min for 30% decomposition. Calculate $t_{1/2}$.

Ans. For a first order reaction,

$$t_1 = \frac{2.303}{k} \log \frac{[R]_0}{[R]}$$

$$k = \frac{2.303}{40\,\text{min}} \log \frac{100}{100-30}$$

$$= \frac{2.303}{40\,\text{min}} \log \frac{10}{7}$$

$$= 8.918 \times 10^{-3}\,\text{min}^{-1}$$

Therefore, $t_{1/2}$ of the decomposition reaction is

$$t_{1/2} = \frac{0.693}{k}$$

$$= \frac{0.693}{8.918 \times 10^{-3}}\,\text{min}$$

= 77.7 min (approximately)

20. Consider a certain reaction A $\rightarrow$ Products with $k = 2.0 \times 10^{-2}\,s^{-1}$. Calculate the concentration of *A* remaining after 100 s if the initial concentration of *A* is $1.0\,\text{mol}\,L^{-1}$.

Ans. $k = 2.0 \times 10^{-2} s^{-1}$

T= 100 s

$[A]_0 = 1.0 \, mol \, L^{-1}$

Since the unit of *k* is s^{-1}, the given reaction is a first order reaction.

Therefore, $k = \frac{2.303}{t} \log \frac{[A]_0}{[A]}$

$$2.0 \times 10^{-2} s^{-1} = \frac{2.303}{\log s} \log \frac{1.0}{[A]}$$

$$2.0 \times 10^{-2} s^{-1} = \frac{2.303}{\log s} (-\log[A])$$

$$-\log[A] = \frac{2.0 \times 10^{-2} \times 100}{2.303}$$

$$[A] = \text{antilog}\left(-\frac{2.0 \times 10^{-2} \times 100}{2.303}\right)$$

= $0.135 \, mol \, L^{-1}$ (approximately)

Hence, the remaining concentration of A is $0.135 \, mol \, L^{-1}$.