CBSE Test Paper-04

Class - 12 Chemistry (Chemical Kinetics)

1. As temperature increases, the reaction rate:
    a. First decreases then increases
    b. Increases
    c. Decreases
    d. Stays the same
2. Consider the reaction
    $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
    The equality relation between $\frac{d[NH_3]}{dt}$ and $\frac{-d[H_2]}{dt}$ is:
    a. $\frac{d[NH_3]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt}$
    b. $\frac{d[NH_3]}{dt} = -\frac{2}{3}\frac{d[H_2]}{dt}$
    c. $\frac{d[NH_3]}{dt} = -\frac{3}{2}\frac{d[H_2]}{dt}$
    d. $\frac{d[NH_3]}{dt} = -\frac{d[H_2]}{dt}$
3. Collision theory is applicable to
    a. First order reaction
    b. Pseudo order reaction
    c. Bimolecular reaction
    d. Zero order reaction
4. The expression which relates the rate of reaction to the concentration of the reactants are called
    a. Differential rate equation
    b. Rate law
    c. Half life
    d. Rate constant
5. The effect of temperature on reaction rate is given by
    a. Clausius Claperon equation
    b. Gibb’s Helmholtz equation
    c. Kirchoff’s equation
    d. Arrhenius equation
6. If half life period of a first order reaction is X and $\frac{3}{4}th$ life period of the same reaction

is Y, how are x and y related each other?

7. For the homogeneous decomposition of N_2O_5 into NO_2 and O_2;

   $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$

   Rate = k $[N_2O_5]$

   Find out the order of reaction with respect to N_2O_5.

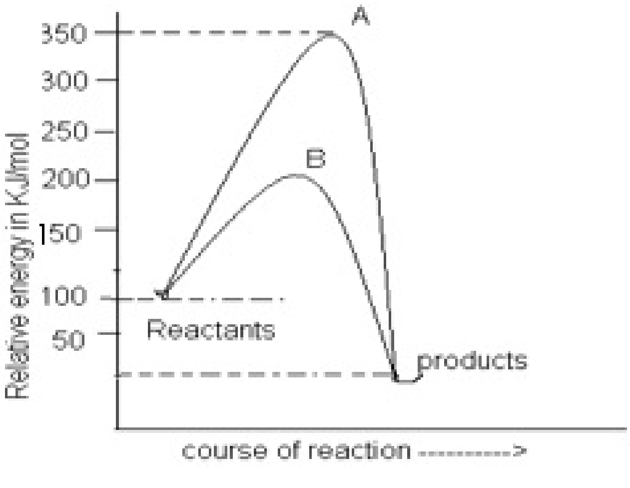
8. For a chemical reaction, what is the effect of catalyst on the rate of the reaction.

9. For the reaction : $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ express the rate of reaction in terms of formation of ammonia.

10. Differentiate between order and molecularity of a reaction?

11. From the figure,

    a. Calculate ΔE for the reaction, activation energy for forward reaction.
    b. Identify the curve for catalysed reaction.
    c. What is the energy of activation in the presence of catalyst?



12. Show that in case of first order reaction, the time required for 99.9% of the reaction to complete is 10 times that required for half of the reaction to take place. [log 2 = 0.301]

13. What is meant by order and molecularity of a reaction? Distinguish between order and molecularity.

14. Hydrogen peroxide, $H_2O_2(aq)$ decomposes to H_2O and O_2 in a reaction which is first order in H_2O_2 and has a rate constant, $k = 1.06 \times 10^{-3} min^{-1}$. Then:

i. How long will it take 15% of a sample of H_2O_2 to decompose?

ii. How long will it take 85% of a sample of H_2O_2 to decompose?

15. a. At 380° C the half life for the first order decomposition of H_2O_2 is 360 min. The energy of activation is 200 **kJ** mol^{-1}. Calculate the time required for 75% decomposition at 450°C.

b. Consider the following data for the reaction.

A + B $\rightarrow$products

Conc. of A (mol L^{-1})	**Conc. of B (mol L^{-1})**	**Initial rate (mol L^{-1}S^{-1})**
0.1	0.1	4.0×10^{-4}
0.2	0.2	1.6×10^{-3}
0.5	0.1	1.0×10^{-2}
0.5	0.5	1.0×10^{-2}

Calculate:

i. Order w.r.t A and B for the reaction

ii. Rate constant

iii. The reaction rate when conc. of A is 0.2 M and B is 0.35 M

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Solutions

1. b. Increases

   Explanation: with increase in temperature the effective molecular collisions increases, hence rate of reaction also increases.

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

   Explanation: For the given reaction

   $rate = -\frac{1}{2}\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$

3. c. Bimolecular reaction

   Explanation: According to Collision theory two molecules must collide with each other with proper orientation in order to bring chemical reaction.

4. b. Rate law

   Explanation: Rate law is defined as the mathematical relationship between rate of a reaction and the molar concentration of the reactants as determined experimentally.

5. d. Arrhenius equation

   Explanation: Arrhenius gives relation between rate contain t (k) and temperature (T) is as follows:

   $k = A\,e^{\frac{-E_a}{T}}$

   taking log on both sides, we get.

   $\ln k = \ln A - \frac{E_a}{T}$

   $Ln\,[R] = ln\,[R_o] - kt$

6. Y = 2X.

7. It is first order with respect to N_2O_5.

8. Catalyst increases the rate of reaction by providing alternative path with lower energy of activation.

9. Rate = $\frac{1}{2}\frac{d[NH_3]}{dt}$

10.

Order	**Molecularity**
i. It is an experimental quantity.	i. It is a theoretical concept.
ii. It is the sum of the powers to which the concentration terms are raised in the rate law equation.	ii. It is the number of species undergoing simultaneous collision in the reaction.
iii. It may have fractional values.	iii. It has only whole number values.
iv. It can be zero.	iv. It can not be zero.

11. a. From the figure-

Energy of reactants E_r = 100 kJ/mol

Energy of products, E_p = 25 kJ/mol

Energy of reaction=Change in energy=$\Delta E = E_P - E_R$

= 25 - 100 = -75 kJ/mol

Threshold energy, E_t = 350 kJ/mol

Activation energy= E_a = E_t - E_r

= 350 - 100

= 250 kJ/mol.

b. The curve B is for catalysed reaction because it represents the lower energy path for the reaction.

c. In the presence of catalyst, Threshold energy $E_t = 200\ kJ/mol$

Activation energy= E_a = 200 - 100

= 100 kJ/mol

12. $t99.9\% = \frac{2.303}{K}\log\frac{[R]_0}{\left([R]_0 - \frac{99.9}{100}[R]_0\right)}$

$\frac{2.303}{K}\log 1000$

$= \frac{6.909}{K}$

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

$\frac{t99.9\%}{t_{1/2}} = \frac{6.909}{K} \times \frac{K}{0.693} = 10$

13. **Order**

   i. It is sum of powers to which concentration terms are raised in rate law or rate equation.
   ii. It is determined experimentally.
   iii. It can be zero or even in fraction
   iv. The order of complex reaction can be determined and slowest step is rate determining step.

   Molecularity

   i. It is the number of molecules taking part in the reaction.
   ii. It is determined theoretically
   iii. It is always in whole number
   iv. Molecularity of each step is determined separately.

14. i. For 15% decomposition of $H_2 0_2$

   $[R]_0$ = 100 M, [R] = 100 - 15 = 85 M

   k = 1.06 × 10^{-3} min^{-1}

   For first order reaction

   $t = \frac{2.303}{k} \log \frac{[R_0]}{[R]}$

   $t = \frac{2.303}{1.06 \times 10^{-3}} \times \log \left[\frac{100}{85}\right]$

   $t = \frac{2.303}{1.06 \times 10^{-3}} \times [\log 100 - \log 85]$

   $t = \frac{2.303}{1.06 \times 10^{-3}} \times \log(2 - 1.9292)$

   $t = \frac{2.303}{1.06 \times 10^{-3}} \times (0.0706)$

   t = 153.38 min.

   ii. Similarly, for 85% decomposition of reaction

   $t = \left(\frac{2.303}{1.06 \times 10^{-3}}\right) \log\left(\frac{100}{15}\right)$

   t = 1790.325 min.

15. a. Let $t_{\frac{1}{2}}$ = half life = 360 min, k_1 = rate constant

   $k_1 = \frac{0.693}{t_{\frac{1}{2}}} = \frac{0.693}{360 \text{ min}}$

   $= 1.92 \times 10^{-3} \text{min}^{-1}$

$\log \frac{k_2}{k_1} = \frac{E_a}{2.303R}\left[\frac{T_2-T_1}{T_1T_2}\right]$

where k_2 = rate constant at temperature T_2(450^oC = 450+273 K = 723 K), k_1 = rate constant at temperature T_1(380^oC = 380+273 K = 653 K), E_a = Activation energy = 200 kJ/mol, R = 8.314 J/molK

$\log \frac{k_2}{k_1} = \frac{200000}{19.147} \times \frac{723-653}{723\times653}$

= 1.5487

$\frac{k_2}{k_1}$ = antilog 1.5487 = 35.38

$$k_2 = k_1 \times 35.38$$
$$= 1.92 \times 10^{-3} \times 35.38$$
$$= 6.792 \times 10^{-2}\text{min}^{-1}$$
$$t = \frac{2.303}{k_2}\log\frac{[R_0]}{[R]} = \frac{2.303}{k_2}\log\frac{[R_0]}{\frac{25}{100}[R_0]}$$

where $[R_0]$=Initial reactant concentration, [R]=reactant concentration at time t after 75% decomposition

t = 20.415 min

b. Let us consider the following rate equation

Rate = $[A]^x[B]^y$ where [A] and [B] are reactant concentrations, x and y are orders of reaction w.r.t. A and B respectively

From experiment 3

$$1 \times 10^{-2} = k(0.5)^x(0.1)^y \ldots (i)$$

From experiment 4

$$1.0 \times 10^{-2} = k(0.5)^x(0.5)^y \ldots (ii)$$

Dividing (i) by (ii) we get

$\frac{1\times10^{-2}}{1\times10^{-2}} - \frac{1}{5^y} \Rightarrow 5^y = 5^0$

y = 0

From experiment 1

$$4.0 \times 10^{-4} = k(0.1)^K(0.1)^y \ldots (iii)$$

From experiment 3

$$1.0 \times 10^{-2} = k(0.5)^K(0.1)^y \ldots (iv)$$

Dividing (iii) by (iv)

$\frac{1}{25} = \frac{1}{5^x}$

$\Rightarrow 5^x = 5^2$

x = 2

Order w.r.t A is 2 and B is 0

rate $= k[A]^2[B]^0$ where k = rate constant

$4.0 \times 10^{-4} = k(0.1)^2$ (from experiment 1)

$k = 4.0 \times 10^{-2} L\,mol^{-1} s^{-1}$

iii. When [A]=0.2 M, [B]=0.35 M

Rate $= k(0.2)^2(0.35)^0$

$= 4.0 \times 10^{-2} \times (0.2)^2$

$= 1.6 \times 10^{-3} mol\,L^{-1} s^{-1}$