Previous Year Questions

Subject	Chemistry
Class	12
Topic Name	Chemical Kinetics

1. The rate of a chemical reaction doubles for every 10^0 C rise of temperature. If the temperature is raised by 50^0 C, the rate of the reaction increases by about:-

(1) 32 times

(2) 64 times

(3) 10 times

(4) 24 times

Solution:

Given the rate of chemical reaction doubles for every 10^0 C rise of temperature. If the temperature is raised by 50^0 C, the rate of reaction increases by 2^n times. Here n = 5.

So $2^n = 2^5 = 32$.

Hence option (1) is the answer.

2. The time for half-life period of a certain reaction A → Products is 1 hour when the initial concentration of the reactant 'A' is 2.0 mol L^{-1}, How much time does it take for its concentration to come from 0.50 to 0.25 mol L^{-1} if it is a zero-order reaction?

(1) 1 h

(2) 4 h

(3) 0.5 h

(4) 0.25 h

Solution:

The half life period for a zero order reaction is given by $t_{1/2} = [A_0]/2k$

A_0 is the initial concentration of the reactant.

$K_0 = [A_0]/2t_{1/2}$

$= 2/2\times1$

$= 1$ mol $L^{-1}h^{-1}$

Rate constant for a zero order reaction is given by $k = (1/t) [A_0 - A]$

$t = (1/k) [A_0 - A]$

= 0.50-0.25/1

= 0.25 h

Hence option (4) is the answer.

3. For the reaction, 2A + B → products, when the concentrations of A and B both were doubled the rate of the reaction increased from 0.3 mol L^{-1} s^{-1} to 2.4 mol L^{-1} s^{-1}. When the concentration of A alone is doubled, the rate increases from 0.3 mol L^{-1} s^{-1} to 0.6 mol L^{-1} s^{-1}. Which one of the following statements is correct?

(1) Order of the reaction with respect to B is 2.

(2) Order of the reaction with respect to B is 1.

(3) Order of the reaction with respect to A is 2.

(4) Total order of the reaction is 4.

Solution:

If concentration of [A] is doubled, then the rate will be doubled, so the order of A is 1.

Then again if the concentration of A and B both were doubled, the rate will increase 8 times. Rate = $[2A] [2B]^2 = 8[A] [B]^2$

So the order of B is two.

So, the overall order is 3.

Hence option (1) is the answer.

4. For the equilibrium, A(g) ⇌ B(g), ΔH is −40 kJ/mol. If the ratio of the activation energies of the forward (E_f) and reverse (E_b) reactions is 2/3 then:

(1) E_f = 60kJ / mol; E_b = 100 kJ/mol

(2) E_f = 30kJ / mol; E_b = 70 kJ/mol

(3) E_f = 80kJ / mol; E_b = 120 kJ/mol

(4) E_f = 70kJ / mol; E_b = 30 kJ/mol

Solution:

A(g) ⇌ B(g)

Given ΔH = -40KJ/mol

$E_f / E_b = ⅔$

$\therefore E_f = ⅔ E_b$

$\Delta H = E_b - E_f$

$-40 = E_b - ⅔ E_b$

$-40 = \frac{1}{3} E_b$

$\therefore E_b = -40 \times 3 = -120$ KJ/mol

$E_f = E_b - \Delta H$

$= -120 - (-40)$

$= -120+40$

$= -80$ KJ/mol

Hence option (3) is the answer.

5. The rate law for the reaction below is given by the expression k[A] [B]

A + B → Product

If the concentration of B is increased from 0.1 to 0.3 mole, keeping the value of A at 0.1 mole, the rate constant will be

(1) 3 k

(2) 9 k

(3) k/3

(4) k

Solution:

The rate constant varies with the temperature only. It is independent of the concentration of reactants. So the rate constant will be k.

Hence option (4) is the answer.

6. Decomposition of X exhibits a rate constant of 0.05 mg/year. How many years are required for the decomposition of 5 mg of X into 2.5 mg?

(1) 25

(2) 50

(3) 20

(4) 40

Solution:

According to the unit of rate constant it is a zero-order reaction.

$t_{1/2} = a_0 / 2k$

$= 5/2 \times 0.05$

$= 50$ years

Hence option (2) is the answer.

7. The formation of gas at the surface of tungsten due to adsorption is the reaction of order

(1) 0

(2) 1

(3) 2

(4) insufficient data.

Solution:

Adsorption on the metal surface do not depend on the concentration of gas.

So it is a zero-order reaction.

Hence option (1) is the answer.

8. At 518^0 C, the rate of decomposition of a sample of gaseous acetaldehyde, initially at a pressure of 363 Torr, was 1.00 s^{-1} when 5% had reacted and 0.5 Torr s^{-1} when 33% had reacted. The order of the reaction is :

(1) 3

(2) 1

(3) 0

(4) 2

Solution:

r_1 = 1 torr/sec

When 5% is reacted, 95% is unreacted.

r_2 = 0.5 torr/sec

When 33% is reacted, (67% is unreacted)

m = order of reaction,

unreacted = a-x

$r_1/r_2 = [(a-x_1)/(a-x_2)]^m$

$1/0.5 = (0.95/0.67)^m$

$2 = (1.414)^m$

$\Rightarrow 2 = \sqrt{2}^m$

$\Rightarrow m = 2$

So the order of the reaction is 2.

Hence option (4) is the answer.

9.The rate of a reaction doubles when its temperature changes from 300 K to 310 K. Activation energy of such a reaction will be (R = 8.314KJmol-1 and log 2=0.301)

(1) 53.6 $KJmol^{-1}$

(2) 48.6 $KJmol^{-1}$

(3) 58.5 $KJmol^{-1}$

(4) 60.5 $KJmol^{-1}$

Solution:

According to Arrhenius equation,

$\ln k_2/k_1 = -(E_a/2.303R)(1/T_2 - 1/T_1)$

$r_2/r_1 = k_2/k_1 = 2$

Given $T_1 = 300K$

$T_2 = 310$ K

$\log(2) = -(E_a/2.303 \times 8.314)\ [(1/310) - (1/300)]$

$0.301 = -E_a/19.147\ [(300-310)/93000\ ]$

$0.301 = -E_a/19.147\ [-10/93000\ ]$

$E_a = 0.301 \times 19.147 \times 9300$

= 53598.19 J

= 53.6 kJ/mol

Hence option (1) is the answer.

10. If 50% of a reaction occurs in 100 seconds and 75% of the reaction occurs in 200 seconds, the order of this reaction is

(1) 1

(2) 2

(3) zero

(4) 3

Solution:

$t_{1/2}$ = 100 second (50% reaction)

After 200 seconds, 75% of reaction will be completed,

i.e., $t_{75\%}$ = 200 seconds.

So, it follows first-order kinetics as the half-life is independent of concentration and follows the relation $t_{3/4} = 2 \times t_{1/2}$

Hence option (1) is the answer.

11. Higher-order (>3) reactions are rare due to:

(1) shifting of equilibrium towards reactants due to elastic collision

(2) loss of active species on collision

(3) low probability of simultaneous collision of all the reacting species

(4) increase in entropy and activation energy as more molecules are involved.

Solution:

Higher-order (>3) reactions are rare due to the low probability of simultaneous collision of all the reacting species.

Hence option (3) is the answer.

12. For the reaction A + 2B → C, rate is given by R = [A] $[B]^2$ then the order of the reaction is

(1) 3

(2) 6

(3) 5

(4) 7

Solution:

The order of the reaction is the sum of the power of the concentration terms in rate law expression. R = [A] $[B]^2$

So, order of reaction = 1+2 = 3

Hence option (1) is the answer.

13. In a first-order reaction, the concentration of the reactant decreases from 0.8 M to 0.4 M in 15 minutes. The time taken for the concentration to change from 0.1 M to 0.025 M is

(a) 30 minutes

(b) 15 minutes

(c) 7.5 minutes

(d) 60 minutes.

Solution:

Gine that the concentration of the reactant decreases from 0.8 M to 0.4 M in 15 minutes.

$t_{1/2}$ = 15 minute.

So, the concentration of reactants will fall from 0.1 M to 0.025 M in two half lives.

i.e., $2t_{1/2} = 2 \times 15 = 30$ minutes.

Hence option (1) is the answer.

14. A reaction was found to be second order with respect to the concentration of carbon monoxide. If the concentration of carbon monoxide is doubled, with everything else kept the same, the rate of reaction will be

(1) remain unchanged

(2) tripled

(3) increased by a factor of 4

(4) doubled.

Solution:

Given $r_1 = dx/dt = k[CO]^2$

$r_2 = k[2CO]^2 = 4k[CO]^2$

According to the rate law expression doubling the concentration of CO increases the rate by a factor of 4.

Hence option (3) is the answer.

15. The half-life period of a first-order chemical reaction is 6.93 minutes. The time required for the completion of 99% of the chemical reaction will be (log 2 = 0.301):

(1) 46.06 minutes

(2) 460.6 minutes

(3) 230.3 minutes

(4) 23.03 minutes

Solution:

Given half-life period $t_{1/2} = 6.93$ min.

Decay constant, $\lambda = 0.693/ t_{1/2}$

$= 0.693/ 6.93$

$= 0.1$ /min

Let t be the time required for completion of chemical reaction.

We have $t = (2.303/ \lambda) \log (a/a-x)$

$\therefore t = (2.303/0.1) \log (100/100-99)$

$\therefore t = (2.303/0.1) \log 100$

$\therefore t = (2.303/0.1) \times 2$

$t = 46.06$ min

Hence option (1) is the answer.

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