Collision Theory Chemistry Questions with Solutions

Q1: Why is the instantaneous rate of reaction preferred over the average rate? Or why is the rate of the reaction not uniform throughout?

Answer: This is because the rate of any reaction depends upon the reactant's concentration at the given time. As the concentration of each reactant decreases with time, the rate of reaction also keeps changing.

Q2. Can a reaction have zero activation energy (E_a)?

Answer: According to the Arrhenius equation: $k = A.e^{-Ea/RT}$
If $E_a = 0$, then, $k = A.e^0 = A$
This means that the rate constant is equal to the collision frequency and hence every collision leads to a chemical reaction which is not true. Hence, E_a cannot be zero.

Q3. A first order reaction takes 69.3 minutes to reach 50% of its completion. Determine the time required for 80% completion of the reaction.

Answer: From 1st order reaction we have: $k = 0.693 / t_{1/2}$
As $t_{1/2} = 69.3$ min, $k = 0.693 / 69.3$ min^{-1} = 0.01 min^{-1}
Hence, $k = 0.01$ min^{-1} for the reaction

Now,
$$t = \frac{2.303}{k} log\frac{[A_o]}{[A]} = \frac{2.303}{k} log\frac{a}{a - x}$$

Given: x = 80% of a = 0.8a
Hence,

$$t_{80\%} = \frac{2.303}{0.01\ min^{-1}} log\frac{a}{a - 0.8a} = 2.303 \times 10^{-2} \times log\frac{1}{0.2}$$

$$t_{80\%} = 1.6 \times 10^{-2}\ min^{-1}$$

Q4. From the given expressions, determine the order of the reactions and the dimension of rate constants.
   a. C_2H_5Cl (g) → C_2H_4 (g) + HCl (g); Rate = $k[C_2H_5Cl]$
   b. 3NO (g) → N_2O (g) + NO_2 (g); Rate = $k[NO]^2$

Answer: The order of the reaction and dimension of rate constant are hereunder.
   a. From the rate law expression, order of the reaction = 1
      Similarly, dimensions of k can be calculated from the rate law as:
      k = Rate / $[C_2H_5Cl]$ = mol L^{-1} s^{-1} / mol L^{-1} = s^{-1}

Hence, the dimension of k is s⁻¹.

b. From the rate law expression, order of the reaction = 2
   Similarly, dimensions of k can be calculated from the rate law as:
   k = Rate / [NO]² = mol L⁻¹ s⁻¹ / (mol L⁻¹)² = L mol⁻¹ s⁻¹
   Hence, the dimension of k is L mol⁻¹ s⁻¹.

Q5. Will the value of rate constant be feasible at extremely high temperature i.e. t = ∞?

Answer: At t = ∞, k = A.e⁻ᴱᵃ/ᴿᵀ = A.e⁻ᴱᵃ/ᴿ∞ = A.e⁰ = A
This implies that at t = ∞, the rate constant is equal to the frequency of collisions. However, this is not true as all the collisions may not lead the reactants into products.

Q6. Calculate the half-life period of a first order reaction given the rate constant of the reaction is 200 s⁻¹.

Answer: The half-life period of a first order reaction is given by:
$$t_{1/2} = 0.693/k$$
Hence, $t_{1/2}$ = 0.693 / 200 s⁻¹ = 0.346 x 10⁻² s = 3.46 x 10⁻³ s

Q7. An archaeological artefact found in a wooden piece was left with only 80% ¹⁴C. Estimate the age of the sample.

Answer: The half-life period for the radioactive decay of ¹⁴C is 5730 years. The Radioactive decay follows the first order kinetics.
Hence, the Decay constant (k) = 0.693 / $t_{1/2}$ = 0.693 / 5730 yr
Hence, the age of the sample can be calculated as:

$$t = \frac{2.303}{k} log\frac{[A]_o}{[A]}$$
$$t = \frac{2.303}{\frac{0.693}{5730\ yr}} log\frac{100}{80}$$
$$t = \frac{2.303 \times 5730}{0.693} \times 0.0969\ yr = 1845\ yr$$

Hence, the age of the sample is 1845 years.

Q8. In the rate expression, the term (-dx/dt) refers to
   a. Instantaneous rate of the reaction
   b. Concentration of reactants
   c. Average rate of reaction
   d. Increase in the concentration of reactants

Answer: (a.)

Explanation: The Instantaneous rate of the reaction is calculated for a very small value of concentration of reactants/ products in infinitesimally time period

Q9. The threshold energy is equal to
    a. Average kinetic energy of molecules of reactants
    b. Normal energy of reactants - activation energy
    c. Activation energy + normal energy of reactants
    d. Activation energy - normal energy of reactants

Answer: (c.)

Explanation: The activation energy is the excess energy that is acquired by the molecules in order to achieve the threshold energy.

Q10. The influence of temperature on the rate of the reaction is determined by
    a. Arrhenius equation
    b. Nernst's equation
    c. Gibbs-Helmholtz equation
    d. Van't Hoff equation

Answer: (a.)

Explanation: The Arrhenius equation has a direct relation between temperature and the rate constant.

$$k = A.e^{\frac{-E}{RT}}$$

Q11. Which plot can give the value of activation energy?
    a. C versus T
    b. 1/k versus T
    c. Log k versus 1/T
    d. k versus T

Answer: (c.)

**Explanation:** The slope of the log k versus 1/T curve gives the value of activation energy.
$$\text{Slope} = E_a / 2.303R$$

Q12. The number of molecules colliding per unit volume per second of the reaction mixture is known as _____.

Answer: Collision Frequency

Q13. Give one example when the molecularity and the order of the reaction are the same.

Answer: The example is:

$$CH_3COOC_2H_5 + NaOH \rightarrow CH_3COONa + C_2H_5OH$$
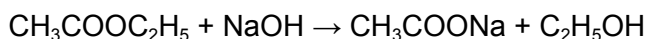
In this reaction, molecularity = order= 2

Q14. Verify the statement: Slow reactions require less energy of activation as compared to the fast reactions.

Answer: This statement is wrong as the energy of activation is inversely proportional to the rate of the reaction.

Q15. The experimentally determined rate expression for the reaction: $NO_2(g) + CO(g) \rightarrow CO_2(g) + NO(g)$ is Rate = $k[NO_2]^2$. Suggest the mechanism for the reaction.

Answer: As the rate depends upon the decomposition of NO_2,the possible mechanism can be:
Step 1: $NO_2 + NO_2 \rightarrow NO + NO_3$ (slow)
Step 2: $NO_3 + CO \rightarrow CO_2 + NO_2$ (fast)

Practise Questions on Carbonyl Compounds

Q1. Consider the following one-step reaction:

$$2(NO)\ (g) + O_2\ (g) \rightarrow 2NO_2\ (g)$$
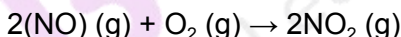
If the volume of the vessel is lowered to ⅓ of its original volume, what will be the effect on the rate of the reaction?

Answer: From the rate law, Rate = $k.[NO]^2[O_2]$ (i)
Let us assume the initial number of moles of NO = a, and that of O_2 = b.
Let us assume the initial volume of the vessel = V L
Hence, concentration, [NO] = a/V M, and [O_2] = b/V M
Now, from equation (i), Rate (r_1) = $k.(a/V)^2(b/V) = k.a^2b/V^3$

Now, new volume = V/3 L
Hence, new [NO] = a/(V/3) = 3a/V, and new [O_2] = b/(V/3) = 3b/V
New Rate (r_2) = $k.(3a/V)^2(3b/V) = k.27a^2b/V^3$
So, r_2/r_1 = 27
This implies that on decreasing the volume to ⅓, the rate of the reaction gets increased by 27 times.

Q2. Why are the reactions of higher order less in number?

Answer: The reactions depend upon the collision among the molecules of reactants. However, there are very less chances for a large number of molecules to collide simultaneously and hence undergo chemical reaction. Hence, the reactions of higher order are less.

Q3. Calculate the rate for the given reaction during the period when [A] decreases from 0.5 mol/L to 0.4 mol/L in 10 min.

$$2A \rightarrow Products$$

Answer: The average rate of reaction = $-\frac{1}{2}\{\Delta[A]/\Delta t\}$

Hence, $-\frac{1}{2}\{\Delta[A]/\Delta t\} = -\frac{1}{2}\{[A]_2 - [A]_1\}/(t_2 - t_1) = -\frac{1}{2}(0.4 M - 0.5 M)/10 min$

Average rate = $-\frac{1}{2}(-0.1 M/10 min) = 5 \times 10^{-3} M\ min^{-1}$

Q4. The rate expression for the reaction $CH_3OCH_3 \rightarrow CH_4 + H_2 + CO$ is given as:

$$Rate = k.[P_{CH_3OCH_3}]^{\frac{3}{2}}$$

Taking pressure in bar and time in minutes, calculate the units of the rate and the rate constant.

Answer: The unit of Rate in terms of "bar" = $-d[CH_3OCH_3]/dt = bar\ min^{-1}$

Hence, the unit of k:

$$k = \frac{Rate}{[P_{CH_3OCH_3}]^{\frac{3}{2}}} = \frac{bar\ min^{-1}}{bar^{\frac{3}{2}}} = bar^{\frac{-1}{2}}\ min^{-1}$$

The unit of Rate constant is $bar^{-\frac{1}{2}}\ min^{-1}$.

Q5. A reaction is first order in terms of A and zeroth order in terms of B. Fill in the blanks for the reaction in between A and B.

Experiment	[A]: mol L^{-1}	[B]: mol L^{-1}	Initial Rate: mol L^{-1} min^{-1}
I	0.1	0.1	2.0×10^{-2}
II	—	0.2	4.0×10^{-2}
III	0.4	0.4	—
IV	—	0.2	2.0×10^{-2}

Answer: Given: Rate = $k.[A]^1[B]^0 = k.[A]$

Hence, from expt. I, 2.0×10^{-2} mol L^{-1} min^{-1} = $k.(0.1$ mol $L^{-1})$

So, k = 0.2 min^{-1}

Now, from expt. II, 4.0×10^{-2} mol L^{-1} min^{-1} = 0.2 $min^{-1}.[A]$; [A] = 0.2 mol L^{-1}

From expt. III, Rate = 0.2 $min^{-1}.(0.4$ mol $L^{-1}) = 0.08$ mol L^{-1} min^{-1}

From expt. IV, 2.0×10^{-2} mol L^{-1} min^{-1} = 0.2 $min^{-1}.[A]$; [A] = 0.1 mol L^{-1}