
	INDIAN SCHOOL AL WADI AL KABIR		
Class: XII	Department: SCIENCE 2022 – 23 SUBJECT: CHEMISTRY	Date of submission: 10.11.2022	
Worksheet No: 08 WITH ANSWERS	Chapter: CHEMICAL KINETICS	Note: A4 FILE FORMAT	
NAME OF THE STUDENT	CLASS & SEC:	ROLL NO.	

MULTIPLE CHOICE QUESTIONS (1 M)

- If 60% of a first order reaction was completed in 60 minutes, 50% of the same reaction would be completed in approximately
 - 45 minutes
 - 60 minutes
 - 40 minutes
 - 50 minutes
- For a reaction $X + Y \rightarrow Z$, $\text{rate} \propto [X]$. What is (i) molecularity and (ii) order of reaction?
 - (i) 2, (ii) 1
 - (i) 2, (ii) 2
 - (i) 1, (ii) 1
 - (i) 1, (ii) 2
- Consider the reaction
 $2A + B \rightarrow \text{Products}$
 When the concentration of B was doubled, the half life did not change. When the concentration of A alone was doubled, the rate increased by two times. The unit of rate constant of this reaction is
 - no unit
 - $\text{molL}^{-1}\text{s}^{-1}$
 - s^{-1}
 - $\text{Lmol}^{-1}\text{s}^{-1}$
- Which of these does not influence the rate of the reaction?
 - Nature of reactants
 - Concentration of reactants
 - Temperature of reaction
 - Molecularity of the reaction
- Consider a reaction, $aG + bH \rightarrow \text{Products}$.
 When the concentration of both the reactants G and H is doubled the rate increases by eight times. However,

when the concentration of G is doubled keeping the concentration of H fixed, the rate is doubled. The overall order of the reaction is

- a. Zero
- b. One
- c. Two
- d. Three

6. For a reaction, $\frac{1}{2} A \rightarrow 2B$, the rate of disappearance of 'A' is related to the rate of appearance of 'B' by the expression:

a.

$$-\frac{d[A]}{dt} = \frac{d[B]}{dt}$$

b.

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

c.

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

d.

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

7. The rate constant of a reaction is $k=3.28 \times 10^{-4} \text{ s}^{-1}$. Find the order of the reaction.

- a. Zero order
- b. First order
- c. Second order
- d. Third order

The questions below consist of Assertion and Reason. Use the following key to choose the appropriate answer.

- a. If both **Assertion** and **Reason** are correct and **Reason** is the correct explanation of the **Assertion**.
- b. If both **Assertion** and **Reason** are correct and **Reason** is not the correct explanation of the **Assertion**.
- c. If **Assertion** is correct and **Reason** is incorrect.
- d. If **Assertion** is wrong and **Reason** is correct.

8. **Assertion** : Rate of reaction increases with increase in temperature.

Reason : Number of effective collisions increases with increase in temperature.

9. **Assertion**: The order and molecularity of a reaction are always the same.

Reason: Order is determined experimentally whereas molecularity by a balanced elementary reaction.

10. **Assertion**: In a first-order reaction, the concentration of the reactant is doubled, its half-life is also doubled.

Reason: The half-life of a reaction does not depend upon the initial concentration of the reactant in a first-order reaction.

1 MARK

11. Define rate of a reaction.

12. Write two factors that affect rate of a reaction.
13. What do you understand by the rate law and rate constant of a reaction?
14. Identify the order of a reaction if the units of rate constants are
- $L^{-1}mol s^{-1}$
 - $Lmol^{-1}s^{-1}$

2 MARKS

15. Show that in a first order reaction, time required for completion of 99.9 % is 10 times that of half-life ($t_{1/2}$) of the reaction.
16. A first order reaction takes 40 minutes for 30% decomposition. Calculate $t_{1/2}$ for this reaction.
17. For the reaction
- $$2N_2O_5 (g) \rightarrow 4NO_2 (g) + O_2 (g)$$
- The rate of formation of $NO_2 (g)$ is $2.8 \times 10^{-3} Ms^{-1}$. Calculate the rate of disappearance of $N_2O_5 (g)$.
18. Derive integrated rate equation for rate constant for first order reaction.
19. A first order reaction has a rate constant of 0.0051 min^{-1} . If we begin with 0,1 M concentration of the reactant, what concentration the reactant will be left after 3 hours?

3 MARKS

20. A reaction is second order in A and first order in B.
- Write the differential rate equation.
 - How is the rate affected on increasing the concentration of A three times?
 - How is the rate affected on increasing the concentration of both A and B two times?
21. The thermal decomposition of $HCOOH$ is a first order reaction with a rate constant of $2.4 \times 10^{-3} s^{-1}$ at a certain temperature. Calculate how long will it take for three- fourths of initial quantity of $HCOOH$ to decompose.
22. For a reaction
- $$2H_2O_2 \xrightarrow[\text{alkaline medium}]{I^-} 2H_2O + O_2$$
- the proposed mechanism is as given below:
- $H_2O_2 + I^- \rightarrow H_2O + IO^-$ (slow)
 - $H_2O_2 + IO^- \rightarrow H_2O + I^- + O_2$ (fast)

- a. Write the rate law for the reaction
- b. Write the overall order of the reaction.
- c. Out of steps (i) and (ii), which one is rate determining determining step?

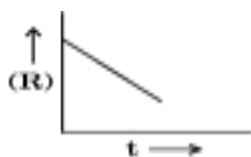
23. a. Differentiate between molecularity and order of a reaction.

- b. The decomposition of NH_3 on platinum surface is zero order reaction. If rate constant (k) is $4 \times 10^{-3} \text{Ms}^{-1}$, how long will it take to reduce the initial concentration of NH_3 from 0.1 M to 0.064 M.

5 MARKS

24. a. A first order reaction is 50% completed in 40 minutes at 300 K and in 20 minutes at 320 K. calculate the activation energy of the reaction. ($\log 4 = 0.6021$, $\log 2 = 0.3010$, $R = 8.314 \text{ JK}^{-1}\text{mol}^{-1}$).

- b. For a chemical reaction $\text{R} \rightarrow \text{P}$, the variation in the concentration of reactant (R) vs Time (t) plot is given as



- i. Predict the order of the reaction
- ii. What is the slope of the curve?

25. a. For the first order thermal decomposition, the following data were obtained



Time /sec	Total pressure/atm
0	0.3
300	0.5

Calculate the rate constant.

- b. The rate of a reaction quadruples when the temperature changes from 300 K to 320 K. Calculate the energy of activation of the reaction , assuming that it does not change with temperature. ($R = 8.314 \text{ JK}^{-1}\text{mol}^{-1}$)

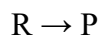
ANSWERS

1. a

2. a

- 3. d
- 4. d
- 5. d
- 6. d
- 7. b
- 8. a
- 9. d
- 10. d

11. Rate of a reaction is equal to change in concentration of reactant/product in unit time.



Rate of disappearance of R

$$= \frac{\text{Decrease in concentration of R}}{\text{Time taken}} = - \frac{\Delta[R]}{\Delta t}$$

Rate of appearance of P

$$= \frac{\text{Increase in concentration of P}}{\text{Time taken}} = + \frac{\Delta[P]}{\Delta t}$$

12. Factors

T, p (gaseous reactant/products), Concentration, Catalyst

13. A rate law is an expression showing the relationship of the reaction rate to the concentration of each reactant. The rate constant k is the proportionality constant relating the rate of reaction to concentration of reactants.

14. i. Zero order

ii. Second order

15.

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

$$t_{99.9\%} = \frac{2.303}{k} \log \left(\frac{100}{0.1} \right)$$

$$\Rightarrow t_{99.9\%} = \frac{2.303}{k} \log 1000$$

$$t_{99.9\%} = \frac{2.303}{k} (3)$$

$$\Rightarrow t_{99.9\%} = \frac{6.909}{k}$$

For half-life of the reaction

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

calculating $\rightarrow \frac{t}{t_{1/2}}$

$$t_{99.9\%} \approx 10 t_{1/2}$$

16

For a first order reaction,

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

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

$$t = 40 \text{ min}$$

$$[R]_0 = 100$$

$$[R] = 100 - 30 = 70$$

$$\therefore k = \frac{2.303}{40} \log \frac{100}{70}$$

$$\Rightarrow k = \frac{2.303}{40} (\log 10 - \log 7)$$

$$\Rightarrow k = \frac{2.303}{40} \times (1 - 0.845) = 8.918 \times 10^{-3} \text{ min}^{-1}$$

Therefore,

$$t = \frac{0.693}{k}$$

$$\Rightarrow t = \frac{0.693}{8.918 \times 10^{-3}} \approx 77.7 \text{ min}$$

17.

A rate of formation of $\text{NO}_2 = 2.8 \times 10^{-3} \text{ M s}^{-1}$

A rate of Disappearance of $\text{N}_2\text{O}_5 =$

$$-\frac{1}{2} \frac{d[\text{N}_2\text{O}_5]}{dt} = \frac{1}{4} \frac{d[\text{NO}_2]}{dt}$$

$$\frac{d[\text{N}_2\text{O}_5]}{dt} = \frac{1}{4} \times 2.8 \times 10^{-3} \times 2 = 1.4 \times 10^{-3}$$

$$= 1.4 \times 10^{-3} \text{ M s}^{-1}$$

18.

a) Consider a general first order reaction



The differential rate equation for given reaction can be written as

$$\text{Rate} = -\frac{d[\text{R}]}{dt} = k[\text{R}]^1$$

Rearrange above equation.

$$\frac{d[\text{R}]}{[\text{R}]} = -k \times dt$$

Integrating on both sides of the given equation

$$\int \frac{d[\text{R}]}{[\text{R}]} = -k \int dt$$

$$\ln[\text{R}] = -kt + I \quad \dots(1)$$

Where I is Integration constant

At $t = 0$ the concentration of reactant $[\text{R}] = [\text{R}]_0$ where $[\text{R}]_0$ is initial concentration of reactant

Substituting in equation (1) we get

$$\ln [R]_0 = (-K \times 0) + I$$

$$\ln [R]_0 = I \quad (2)$$

Substitute I value in equation (1)

$$\ln [R] = -Kt + \ln[R]_0$$

$$Kt = \ln[R]_0 - \ln R$$

$$Kt = \ln \frac{[R]_0}{[R]}$$

$$\text{or } K = \frac{1}{t} \ln \frac{[R]_0}{[R]}$$

$$\text{or } K = \frac{2.303}{t} \log \frac{[R]_0}{[R]}$$

19.

For a first order reaction

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

Here

$$t = 3 \text{ h} = 3 \times 60 \text{ min} = 180 \text{ min}$$

$$k = 0.0051 \text{ min}^{-1}, [R]_0 = 0.10 \text{ M}, [R] = ?$$

$$180 \text{ min} = \frac{2.303}{0.0051 \text{ min}^{-1}} \log \frac{0.10}{[R]}$$

$$\text{Log} \frac{0.1}{[R]} = \frac{180 \text{ min} \times 0.0051 \text{ min}^{-1}}{2.303} = \frac{918}{2303}$$

$$\log \frac{0.1}{[R]} = 0.3986$$

$$\frac{0.1}{[R]} = \text{Anti log} (0.3986) = 2.503$$

$$[R] = \frac{0.1}{2.503} = 0.03995 \text{ M}$$

$$[R] = 0.04 \text{ M}$$

20.

It is given that a reaction is first order in A and second order in B.

i) The differential rate equation is as follows:

$$\text{Rate} = k[\text{A}][\text{B}]^2$$

(ii) On increasing the concentration of B three times, rate becomes 9 times.

$$(\text{d}[\text{B}])^2 = 3^2 = 9$$

(iii) When the concentrations of both A and B are doubled, rate becomes 8 times.

$$\text{d}[\text{A}](\text{d}[\text{B}])^2 = 2 \times 2^2 = 8$$

21.

$$t = \frac{2.303}{2.4 \times 10^{-3} \text{ s}^{-1}} \log 4$$

$$= \frac{2.303}{2.4 \times 10^{-3}} \times 0.6021 \text{ s}$$

$$t = 577.7 \text{ s} = 578 \text{ s}$$

22.

i) As step (1) is slow, it is the rate-determining step as it dictates the overall pace of the reaction, hence its $\text{Rate} = k[\text{H}_2\text{O}_2][\text{I}^-]$

ii) According to the rate law expressed in (i), Order = 2

iii) Step 1 is rate-determining, as it dictates the overall pace of the reaction.

23.

a.

Molecularity	Order
Molecularity is the number of ions or molecules that take part in the rate-determining step.	Number of molecules of the reactant whose concentration changes during the chemical change
It is always a whole number	It can either be a whole number or a fraction
The molecularity of the reaction is determined by looking at the reaction mechanism	The order of the reaction is determined by the experimental methods
It is theoretical concept	It is experimental concept
It is meaningful only for simple reactions or individual steps of a complex reaction. It is meaningless for an overall complex reaction.	It is meant for the reaction and not for its individual steps

b.

$$t = \frac{[R]_0 - [R]_t}{k}$$

$$= \frac{[0.1 - 0.064]}{4 \times 10^{-3}}$$

$$= 9s$$

24.

a.

$t_{1/2}$ for first order reaction :

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

$$\text{For } 300 \text{ K, } K_1 = \frac{0.693}{t_{1/2}} = \frac{0.693}{40} \text{ S}^{-1}$$

$$\text{For } 320 \text{ K, } K_2 = \frac{0.693}{t_{1/2}} = \frac{0.693}{20} \text{ S}^{-1}$$

$$\log\left(\frac{K_2}{K_1}\right) = \frac{E_a}{2.303R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\log\left(\frac{\frac{0.693}{20}}{\frac{0.693}{40}}\right) = \frac{E_a}{2.303 \times 8.314} \left[\frac{1}{300} - \frac{1}{200}\right]$$

$$\log(2) = \frac{E_a}{2.303 \times 8.314} \times \left(\frac{320 - 300}{300 \times 320}\right)$$

$$0.3010 = \frac{E_a}{2.303 \times 8.314} \times \frac{20}{200 \times 150 \times 320}$$

$$E_a = 0.3010 \times 2.303 \times 8.314 \times 150 \times 32$$

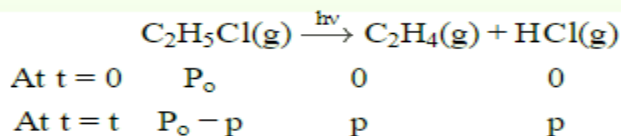
$$E_a = 27663.8 \text{ J/mol}$$

b. i. Zero order

ii. slope = -k

25.

a.



After time t, total pressure, $P_t = (P_o - p) + p + p$

$$\Rightarrow P_t = (P_o + p)$$

$$\Rightarrow p = P_t - P_o$$

Therefore, $P_o - p = P_o - P_t + P_o$

$$= 2P_o - P_t$$

$$= \frac{2.303}{t} \log \left(\frac{P_o}{2P_o - P_t} \right)$$

For a first order reaction,

$$k = \frac{2.303}{t} \log \left(\frac{P_o}{P_o - p} \right)$$

$$\text{At } t = 300, p_t = 0.5, p_o = 0.3$$

$$\text{So, } k = 3.6 \times 10^{-3} \text{sec}^{-1}$$

b.

$$k_2 / k_1 = 4, T_1 = 300\text{K}, T_2 = 320\text{K}.$$

$$\log \frac{k_2}{k_1} = \frac{E_a}{2.303 \times (8.314 \text{JK}^{-1}\text{mol}^{-1})} \left[\frac{320\text{K} - 300\text{K}}{320\text{K} \times 300\text{K}} \right]$$

$$0.6020 = \frac{E_a \times 20}{(2.303) \times (8.314 \text{Jmol}^{-1}) \times (320) \times (300)}$$

$$E_a = \frac{0.6020 \times 2.303 \times 8.314 \text{Jmol}^{-1} \times 320 \times 300}{20}$$

$$= 55327.6 \text{Jmol}^{-1} = 55.328 \text{kJmol}^{-1}$$

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