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CHEM 335

Laboratory of Chemical Kinetics

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Grading

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Attendance, experiment performance and reports	10
Final exam	20

Experiment (1): Catalytic Hydrogen Peroxide Decomposition

- **Aim of the experiment:**

  Determination of rate constant (k) and half-life ($t_{1/2}$).

- **Theoretical Principle:**

A solution of hydrogen peroxide decomposes automatically and slowly to give oxygen and water according to the following equation:

$$2H_2O_{2\,(aq)} \rightarrow 2H_2O_{(L)} + O_{2(g)}\uparrow$$

It has been observed that the rate of dissociation increases with the presence of catalytic solids such as manganese dioxide (MnO_2) or colloidal platinum that act as catalysts. The reaction can be followed either by titration of hydrogen peroxide by potassium permanganate in an acidic medium or by the aggregation of oxygen gas released from the dissociation reaction.

- **Procedure:**

1- Place 100 ml of (H_2O_2 0.5%) solution in a conical flask.

2- Add about 0.05 g of (MnO_2) to the solution, start the stopwatch and note the reaction start time.

3- After 5 minutes, transfer 10 ml of the mixture using a pipette and place it in a conical flask. Add 20 ml of (1M H_2SO_4), then quickly titrate with a solution of (0.02M $KMnO_4$) in the burette and record the titration volume ($KMnO_4$ acts as self-indicator).

4- Repeat step (3) at 10, 15, 20, and 25 minutes intervals.

- **Results :**

t (min)	V_{KMnO4} (mL)	log (a-x)
5		
10		
15		
20		
25		

• Calculations:

Since the reaction follows first-order, the kinetic equation is:

$$\log (a\text{-}x) = \log (a) - \frac{kt}{2.303}$$

The $KMnO_4$ volume represents the amount of the undissociated (residual) H_2O_2 solution (a-x) at a given time.

Organize the results in a table, plot a graph of log (a-x) on the Y-axis and (t) on the X-axis and determine the slope of the straight line, where:

$$\text{Slope} = - \frac{kt}{2.303}$$

The half-life for first-order reaction is:

$$t_{½} = \frac{\ln (2)}{k}$$

Determine the slope, calculate the rate constant (k), and then find the half-life ($t_1/_2$) of the reaction.

Experiment (2): Catalytic Effect of Salt

- **Aim of the experiment:**

Study of the catalytic effect of sodium chloride on the rate constant of reaction between potassium persulfate and potassium iodide.

$$S_2O_8{}^{2-} + 2I^- \rightharpoonup 2SO^{2-}{}_4 + I_2$$

- **Theoretical Principle:**

The reaction between potassium iodide and potassium persulfate is studied in solutions with different ionic strengths, using sodium chloride as an electrolyte salt that does not take part in the reaction.

$$K_2S_2O_8 + 2KI \rightharpoonup 2K_2SO_4 + I_2$$

In this experiment, the effect of salt addition on the rate of reaction will be studied according to the Bronsted-Bjerrum hypothesis. As proposed by the two scientists, the interaction between ions is supposed to pass through the formation of an intermediate complex that decompose unimoleculely to give the products. For example, the reaction between the ions A and B, whose charges are Z_a and Z_β, can be represented as follows:

$$A^{ZA} + B^{ZB} \rightarrow [X^{ZA\,ZB}] \rightarrow \text{Products}$$

Based on the Debye-Huckel equation, the two scientists were able to derive the following relationship:

$$Log\ k = log k_0 + Z_A Z_B \sqrt{I}$$

Where:

k_0 = constant rate of reaction velocity at infinite dilution.

k = rate constant.

I = Ionic strength of the reaction medium.

From the relationship, it is clear that the initial effect of adding a neutral salt to a reaction containing ions of the same charge is to increase the rate constant of this reaction, where the value of $Z_A Z_B$ becomes positive and the effect is reversible if different charged ions are found.

The rate of reaction speed can be followed by the addition of different concentrations of NaCl salt (neutral) to form different ionic strength. This experiment investigates the effect of salt, specifically sodium chloride (NaCl), on the rate of reaction between ($K_2S_2O_8$) and (KI). The hypothesis is based on the Debye-Hückel theory, which predicts that adding electrolytes affects the ionic strength and thus the reaction rate.

- **Procedure:**

1. In five conical flasks with lids, prepare the mixtures with the specified salt concentrations as shown in the table. Ensure that each flask is covered immediately after preparing the mixture.

2. Start the stopwatch immediately after adding the final component to each mixture to accurately measure the reaction time.

Mixture no. / Solutions	1	2	3	4	5
H_2O (distilled)	50 mL	40 mL	30 mL	20 mL	10 mL
4M NaCl	---	10 mL	20 mL	30 mL	40 mL
0.04 M KI	25 mL	25 mL	25 mL	25 mL	25 mL
0. 04 M $K_2S_2O_8$	25 mL	25 mL	25 mL	25 mL	25 mL

3- After 15 minutes, take 25 mL of the first reaction mixture using a pipette and transfer it to a clean conical flask. Titrate the iodine released with (0.01 M $Na_2S_2O_3$) solution in the burette using starch indicator until the blue color disappears, indicating the end point.

4- Repeat the previous process for the remaining mixtures once 15 minutes have passed for each.

Chemical Equation of the Titration Reaction of Sodium Thiosulfate with Iodine:

$$2Na_2S_2O_3 + I_2 \rightharpoonup 2NaI + Na_2S_4O_6$$

- **Results:**

Mixture number	1	2	3	4	5
Volume ($Na_2S_2O_3$) mL *(x)*					
K					
log k					
I					
$\sqrt{I}$					

- **Calculations:**

1. The molarity of $S_2O_8^{2-}$ in the mixture is equivalent to the molarity of $Na_2S_2O_3$ used in the titration (0.01 M). A 25 mL portion of the mixture is titrated. therefore, the initial concentration of the mixture can be considered equivalent to 25 mL of $Na_2S_2O_3$. If the volume of $Na_2S_2O_3$ required to titrate 25 mL of the mixture after a time (t) is equal to x ml, then the remaining concentration of the mixture for each component is equal to (25 - x) ml. By substituting these values into the second-order rate equation, the rate constant (k) can be determined at t = 15 minutes.

$$k = \frac{1}{t} \cdot \frac{x}{[0.01\,(25 - x)]}.$$

2. The ionic strength (I) for each mixture in the reaction can be calculated as follows:

$$I = \frac{1}{2} \sum MZ^2$$

Where:

M = molarity of each ion present in the reaction medium.

Z = ion charge.

3. Present the results for each mixture in a table showing the following values:

x, K, log k, I and $\sqrt{I}$

4. Plot the relationship between log k on the y-axis and $\sqrt{I}$ on the x-axis, then find the slope that represents the value of ($Z_A Z_B$).

Find Example:

a-For the method of calculating (***k***):

b-For the method of calculating (***I***):

Experiment (3): Study of the autocatalysis of the reaction of potassium permanganate with oxalic acid.

- ## Aim of the experiment:

Determination of the reaction order.

- ## Theoretical Principle:

Study the Autocatalytic Reaction between Potassium Permanganate and Oxalic acid .This reaction is self-catalyzed by the manganese ion that is formed from the products of the reaction, and more manganese ion can be added to increase the rate of the reaction.

$$2KMnO_4+5H_2C_2O_4+3H_2SO_4 \rightharpoonup K_2SO_4+ 2MnSO_4+10CO_2+8H_2O$$

- ## Procedure:

1- Prepare two mixtures of the given solutions as shown in the table using separate conical flasks. Start timing upon the addition of the potassium permanganate solution.

Mixture	0.1 M $H_2C_2O_4$	1M H_2SO_4	0.2M $MnSO_4$	Distilled Water	0.02M $KMnO_4$
1	100 ml	10 ml	5 ml	20 ml	30 ml
2	100 ml	5 ml	5 ml	40 ml	15 ml

2- After 5 minutes, use a pipette to withdraw 10 ml of the first mixture and transfer it into a conical flask. Add 15 ml of 10% KI solution to stop the reaction, then titrate the released iodine with 0.01 M sodium thiosulfate ($Na_2S_2O_3$) using starch as an indicator.

3- Repeat step (2) at time intervals of 10, 15, 20, and 25 minutes.

- **Results and Calculations:**

Time (min)	$Na_2S_2O_3$ Volume for first mixture (mL)	$Na_2S_2O_3$ Volume for second mixture (mL)
5		
10		
15		
20		
25		

The volume of sodium thiosulfate at a given time is equivalent to the volume of permanganate in the mixture at the same time.

1- Arrange your results in a table, then plot the relationship between the volume of sodium thiosulfate (representing the volume of permanganate) on the Y-axis and the corresponding time on the X-axis for each mixture.

2- Discuss the shape of the curve for each mixture, determine the initial concentration of ($KMnO_4$) for each by extrapolating the curve to the Y-axis, then find the half-life of the reaction for each mixture from the curve.

3- Determine the order of the reaction using the following relationship:

$$\frac{(t_{1/2})_1}{(t_{1/2})_2} = \left(\frac{a_2}{a_1}\right)^{n-1}$$

Where:

$(t_{1/2})_1$ = The half-life of the first mixture

$(t_{1/2})_2$ = The half-life of second mixture.

(a_1) and (a_2) = Initial concentrations of ($KMnO_4$) in the first and second mixtures respectively.

-Equation of the titration reaction between sodium thiosulfate with iodine.

$$2Na_2S_2O_3 + I_2 \rightharpoonup 2NaI + Na_2S_4O_6$$

-Reaction equation for adding KI to the reaction mixture (to stop the reaction):

$$10KI + 2KMnO_4 + 8H_2SO_4 \rightharpoonup 6\,K_2SO_4 + 2MnSO_4 + 5\,I_2 + 8H_2O$$

-

Experiment 4: Adsorption of Oxalic Acid on Activated Charcoal

- ## Aim of the experiment:

Determination of the value of the constant (n) of the adsorption of oxalic acid on the surface of the activated charcoal and the value of adsorption rate constant (k).

- ## Theoretical Principle:

An adsorption of Oxalic acid on surface of activated Charcoal when a quantity of activated charcoal is mixed with an oxalic acid solution, a quantity of this solution will be adsorbed on the surface of the charcoal The relationship that connects the adsorption process at a certain temperature is the Freundlich adsorption equation (Freundlich Adsorption Isotherm).

$$\frac{x}{m} = k\, C^{1/n}$$

Taking the logarithm of both sides:

$$\log \frac{x}{m} = \log k + \frac{1}{n} \log C$$

Where:

x = the amount of adsorbate per 1 g.

m = the mass of the adsorbent (g). in this experiment (activated charcoal).

C = solution concentration in units (g/L).

K = rate constant of adsorption. n =Constant expresses the quality of adsorption.

When plotting a graph of $\log \frac{x}{m}$ on the Y-axis against log (C) on the X-axis. A straight line is obtained, where the slope equal to $\frac{1}{n}$ and the intercept is equal to log(k).

- ## Procedure:

1- Inside six conical flasks with lids, prepare the following solutions:

Solution	0.25 M oxalic acid	Distilled Water
A	100 ml	-
B	80 ml	20 ml
C	60 ml	40 ml
D	50 ml	50 ml
E	40 ml	60 ml
F	25 ml	75 ml

2- Add 2 g of activated charcoal to each flask. Close the flask with its lid, place it on a mechanical shaker, and shake for 15 minutes.

3- Filter the solution using filter paper (the filtrate must be clear).

4- Transfer 10 ml from each filtrate and titrate with 0.1 M sodium hydroxide solution using phenolphthalein as an indicator. Record the volume of NaOH consumed.

- ## Results and Calculations:

1- The molar of solutions before adsorption can be calculated from the law of dilution, for example, the molar of solution B can be found as follows:

$$M_1 V_1 = M_2 V_2$$

The same applies to the rest of the solutions.

2- The molar of the solutions after adsorption can be calculated from the titration results where oxalic acid reacts with sodium hydroxide in a molar ratio of 2:1

$$\frac{(M \cdot V)}{n} = \frac{(M` \cdot V`)}{n`}$$

The equation for the calibration reaction is as follows:

$$H_2C_2O_4 + 2NaOH \rightharpoonup Na_2C_2O_4 + 2H_2O$$

3- To calculate the value of (x) use the following relationship:

$$x = [\ (M_2 - M`) \cdot (M_{wt}) \cdot V_{(ml)}] / 1000$$

Where:

M_2 = Molarity of oxalic acid solution before adsorption, M' = Molarity of the same solution after adsorption.

$M._{wt}$ = Molar mass of aqueous oxalic acid is 126 g/mol

4- To find the value of (C), use the following relationship:

$$C = \acute{M} \times 126 \quad g/L$$

5- For each solution, determine the value of $\frac{x}{m}$ then calculate (log $\frac{x}{m}$)

6- Arrange your results in a table, then plot a graph between log $\frac{x}{m}$ and log (C).

Determine the values of (n) and (k) from the graph

M	V_{NaOH}	M`	x	$\frac{x}{m}$	$log\frac{x}{m}$	C	log (C)

- Write examples for calculating methods of (M_2), ($\acute{M}$), (x) and (C).

Experiment 5: Effect of Concentration on Reaction Rate

- **Aim of the experiment:**

Determination rate constant (k) and half-life reaction time ($t_{1/2}$) for the reaction of potassium persulfate with potassium iodide.

- **Theoretical Principle:**

The reaction between potassium persulfate and potassium iodide is carried out according to the following equation:

$$K_2S_2O_8 + 2KI \rightarrow 2K_2SO_4 + I_2$$

The above reaction is considered to be of the second order, but since the concentration of (KI) is higher than the concentration of ($K_2S_2O_8$), any change in the concentration of (KI) during the reaction is considered very small compared to its original concentration, so the concentration of (KI) can be considered constant throughout the reaction, and therefore the rate of reaction depends only on the concentration of only one substance, which is ($K_2S_2O_8$). On this basis, the reaction can be considered kinetic of the first order for the reactant potassium persulfate, and reactions of this type that contain more than one reactant and achieve the kinetic first-order equation are called pseudo – first order reactions.

The kinetic equation for a first-order reaction:

$$In\ (a-x) = In\ (a) - kt$$

a: Initial concentration of potassium persulfate (mol/L).

x: the concentration of one of the products (I_2) (mol/L).

k: Rate constant.

t: Reaction time.

In this experiment, we will track the change in the concentration of (I_2) and calculate its concentration by titrating it with sodium thiosulfate in the presence of starch indicator.

- **Procedure:**

1-Mix 100 ml of (0.005 M $K_2S_2O_8$) with 100 ml of (0.02M KI) in conical flask fitted with lid and start the stopwatch.

2-After 5 minutes, take 10 ml of the reaction mixture and transfer it to a conical flask containing 25 ml of distilled water.

3- Titrate the iodine produced by the reaction with sodium thiosulfate (0.01 M $Na_2S_2O_3$) using starch as indicator and record of the titration volume in the table.

4- Repeat steps (2) and (3) at different times intervals (10,15,20 and 25 minutes) and record the corresponding titration volumes.

- **Results and Calculations:**

1- From the volume of sodium thiosulfate obtained from the titration, the concentration of iodine can be calculated according to the following equation:

$$I_2 + 2Na_2S_2O_3 \rightharpoonup 2NaI + Na_2S_4O_6$$

$$\frac{(M \cdot V)}{n} = \frac{(M` \cdot V`)}{n`}$$

Where ($\acute{M}$) is the concentration of iodine released (mole /L), which is equivalent to a concentration of potassium persulfate.

2- The initial concentration of ($K_2S_2O_8$) at the beginning of the reaction is calculated as follows:

$$M \cdot V = M` \cdot V`$$

3- Record your results in the following table:

t (min)	V $_{Na_2S_2O_3}$ (mL)	X	(a-x)	ln (a-x)
5				
10				
15				
20				
25				

4- plot a graph of ln (a-x) on the Y-axis and (t) on the X-axis. A straight line is obtained, where the slope equal to the negative rate constant of the reaction (-k) in (min^{-1}).

5- Calculate the half-life time using the following law:

$$t_{1/2} = \frac{\ln(2)}{k}$$

-

Experiment 6: Saponification reaction of ester

- **Aim of the experiment:**

Determination of rate constant and half-life for the reaction saponification of ethyl acetate.

- **Theoretical Principle:**

The kinetic study of the reaction of ethyl acetate in an alkaline medium is carried out according to the following equation:

$$CH_3COOC_2H_5 + NaOH \rightarrow CH_3COONa + C_2H_5OH$$

The reaction rate is directly proportional to the concentration of both the ester and base. Therefore, we find that this reaction belongs to the second order and can be expressed if the initial concentration of both the ester and base are equal (a=b). The differential kinetic equation is:

$$\frac{\mathrm{dx}}{\mathrm{dt}} = k(a-x)^2$$

Integrating the previous equation, we obtain:

$$\frac{1}{(a\ -\ \mathrm{x})} = kt + \frac{1}{(a)}$$

- **Procedure:**

1- Take 50 ml of (0.05 M NaOH) in beaker and 50 ml of (0.05 M ethyl acetate) in conical flask.
2- Quickly add the sodium hydroxide to the ester and mix well with turn on the stopwatch.
3- After 5 minutes, transfer 10 ml of the mixture to conical flask containing 50 ml of distilled water. Add exactly 10 ml of (0.025 M HCl) and two drops of phenolphthalein indicator.
4- Titrate by (0.025 M NaOH) then record the volume.
5- Repeat the following step at different times (25, 20, 15, 10) minutes, recording the volume of sodium hydroxide as (x) ml.

If (a = 10 ml) is the amount of hydrochloric acid equivalent to the initial concentration of both the base and ester, then (a - x) is the amount of base or easter remaining.

- **Results and Calculations:**

a =10 mL

t (min)	X= V_{NaOH} (mL)	a -x	$\frac{1}{(a-x)}$
5			
10			
15			
20			
25			

1- plot a graph of $\frac{1}{(a-x)}$ on the Y-axis and (t) on the X-axis. A straight line is obtained, where the slope equal to rate constant of the reaction (k) in (M^{-1} min^{-1}).

2- Calculate the half-life time using the following law:

$$t_{1/2} = \frac{1}{k \cdot a}$$

3- Why is this reaction called " Saponification Reaction"?

-

Experiment (7): The Corrosion Rate

- ## Aim of the experiment:

Study of the Corrosion Rate for some Metals & Alloys.

- ## Theoretical Principle:

Acids affect metals and alloys by causing corrosion (oxidation) of them, and this effect depends on type and concentration of the acid, and the duration of the metal's exposure to acid, as this leads to a decrease in the mass of the metal.

The relationship by which the Corrosion Rate can be estimated is:

$$\text{C.R} = \frac{534\,w}{DAT} = \;\dots\dots.(mpy)$$

Where:

W = Difference in sample weight (mg).

D = Sample density (g/cm^3).

A = Sample Area (in square inches). ($1\ inch^2 = 6.45\ cm^2$)

T = Time the sample in acid (in hours).

- **Procedure:**

1- Inside four beakers put 25 ml of the following acids in the mentioned concentrations:

a- 3M (HCl) b- 3M (HNO_3) c- 2M (HCl) d- 2M (HNO_3)

2- Take four copper metal samples or four aluminum metal samples and polish them to remove any oxides or impurities on their surface with sandpaper.

3- Wash the samples with acetone, then with water, and dry them with a tissue or blotting paper.

4- Weigh the samples individually accurately and then measure their dimensions with a ruler and record this in the attached table.

5- Soak the copper or aluminum slices in the solutions mentioned in the four glasses for half an hour.

6- Lift the samples after the end of the time and remove the abrasion products from them with a brush, then wash them with acetone, then with distilled water and dry them.

7- Weigh the samples thoroughly again.

8- Find the corrosion rate for each sample and make your observations on the effect of acid concentration and type on the rate of corrosion.

Note: Consider the thickness of the sample to be 0.025 cm for copper and 0.06 cm for aluminum.

Results and Calculations:

Explain how to calculate the volume, area, and density of samples only once in detail as follows:

Example: Sample No. (1):

Dimensions of Sample. Length = ... cm, Width = ...cm, Thickness = ...cm

Sample weight before corrosion = ... g = ... mg

Sample volume = ... cm^3

Sample Area = ... cm^2 = ... in^2

Sample Density = ... g/cm^3

Sample weight after corrosion = ... g = ... mg

Lost Weight = ... mg

Corrosion Rate = ... mpy

Write down your results in the table:

Beaker Symbol	Acid used and concentration	Length (cm)	Width (cm)	Area (inch^2)	Volume: (cm^3)	Density g/cm^3	Weight Before corrosion	Weight after corrosion	Lost mass (mg)	C.R (mpy)
a	3M HCl	....	....	....	....	....	....	....	....	....
b	3M HNO_3	....	....	....	....	....	....	....	....	....
c	2M HCl	....	....	....	....	....	....	....	....	....
d	2M HNO_3	....	....	....	....	....	....	....	....	....

From the results of your experiment: Which of the two acids under study causes the most corrosion of the metal? And what effect does the acid concentration have on the rate of corrosion?

Experiment (8): Hydrolysis of Ester (ethyl acetate)

Aim of the experiment:

Determination of the constant rate (k) and half-life ($t_{1/2}$) for the ethyl acetate reaction.

Theoretical Principle:

Ethyl acetate is hydrated in water according to the following equation:

$$CH_3COOCH_2CH_3 + H_2O \rightharpoonup CH_3COOH + CH_3CH_2OH$$

This reaction occurs very slowly at room temperature, so hydrochloric acid is added as a catalyst. Although the reaction is second-order, the large excess of water allows its concentration to be treated as constant. Therefore, the reaction is kinetically considered to be pseudo–first order, as its rate depends only on the concentration of ethyl acetate.

$$Rate = k\,[CH_3COOCH_2CH_3]$$

The kinetic equation of a first-order interaction:

$$In\,(a\text{-}x) = In(a) - k\,t$$

Where:

a= Initial concentration of ethyl acetate (mol/L)

x= Concentration of an product (acetic acid) formed (mol/L)

k= rate constant (min^{-1}).

t= The time (min)

The rate constant (k) can be determined by monitoring any physical property that varies linearly with concentration. In this experiment, the amount of acetic acid formed is determined by titration with a standard NaOH solution using phenolphthalein as an indicator.

- ## Procedure:

1- Place 100 mL of (0.1 M HCl) in a conical flask or beaker.

2- Add 5 mL of ethyl acetate to the conical flask and start the stopwatch immediately. Mix the contents thoroughly.

3- After 5 minutes, withdraw 10 mL of the reaction mixture and transfer it into another conical flask containing 20 mL of ice-cooled distilled water to quench the reaction.

4- Add two drops of phenolphthalein (ph.ph) indicator to the flask, then titrate the mixture using (0.1 M NaOH) and record the NaOH volume used.

5- Repeat the titration at different time intervals (10, 15, 20 and 25 minutes) record the volume of NaOH consumed for each time interval as (x) mL.

- ## Results and Calculations:

1- Calculate the values of x, (a − x), and ln (a − x) using the following table, where at the end of the reaction, the NaOH volume corresponding to the initial concentration(a) is is (**31.7 mL**).

t (min)	x	a-x	ln (a-x)
5			
10			
15			
20			
25			

2- plot a graph of ln (a-x) on the Y-axis and (t) on the X-axis. A straight line is obtained, where the slope equal to the rate constant of the reaction (k).

3- Calculate the half-life time using the following law:

$$t_{1/2} = \frac{\ln(2)}{k}$$

Experiment (9): Measurement of the rate constant by electrical conductivity method

- **Aim of the experiment:**

Measurement of Rate Constant by Conductimetric Method.

- **Theoretical Principle:**

To understand this method, the saponification reaction of ethyl acetate is studied

$$CH_3COOC_2H_5 + NaOH \rightharpoonup CH_3COONa + C_2H_5OH$$

This reaction is a second-order reaction and can be expressed if the initial concentration of both the ester and base are equal (a=b). So if the initial concentration of the reactants is a, and if the number of moles reacting after the time t is equal to x. then:

$$\frac{dx}{dx} = k(a-x)^2 \qquad (1)$$

By doing integration, the following:

$$k = \frac{1}{at} \cdot \frac{x}{a-x} \qquad (2)$$

As you can see from the reaction, the electrical conductivity value of the reaction solution decreases over time as the concentration of hydroxyl ions (with high conductivity) decreases and the concentration of acetate ions with lower conductivity increases, so the rate of the reaction can be tracked by tracking the change in the conductivity value over time. In other words, Equation (2) can be written in terms of conduction instead of concentration terms as follows:

$$k = \frac{1}{at} \cdot \frac{G_0 - G_t}{G_t - G_\infty} \qquad (3)$$

Where:

G_0 = Initial electrical conductivity of the solution (at the beginning of the reaction).

G_t = Conductivity of the solution after a time.

G_∞ = The electrical conductivity of the reaction solution at the end of the reaction.

By rearranging equation (3), we get:

$$G_t = \frac{1}{aK} \cdot \frac{G_0 - G_t}{t} + G_\infty \qquad (4)$$

This equation (4) represents the equation of a straight line, so when we plot the graphical relationship between the value of Gt on the Y-axis and the value of $\frac{G_0 - G_t}{t}$ on the x-axis, we get a straight line whose slope is equal to $\frac{1}{aK}$ and the intercept represents G_∞.

- **Procedure:**

1- Clean the delivery cell given with distilled water thoroughly.
2- Put 25 ml of (0.01 M NaOH) solution into a suitable beaker.
3- Connect the connecting cell to the electrical conductor and place the cell inside the beaker, turn on the device and adjust it according to the supervisor's instructions.
4- Add 25 ml of (0.01M ethyl acetate) to the NaOH beaker, start the stopwatch and quickly record the conductivity value at the start of the reaction (G_0).
5- Record the conductivity values after 5, 10, 15, 20 and 25 minutes.

- **Results and Calculations:**

1- Arrange the experiment results according to the following table:

G_0 =…………… S/cm

t (min)	G_t	$\frac{G_0 - G_t}{t}$
5		
10		
15		
20		
25		

2- plot a graph of (Gt) on the Y-axis and ($\frac{G_0 - G_t}{t}$) on the X-axis. A straight line is obtained, where the slope equal to rate constant of the reaction (k).

Experiment (10): Kinetic study of the reaction of sulfite ion with iodate ion

(Iodine Clock Reaction)

- **Aim of the experiment:**

Study of reaction kinetics of Sulfite with Iodate (Iodine Clock Reaction).

- **Theoretical Principle:**

Sulfite reacts with iodine in an acidic medium in two steps, a slow step (defined for the reaction) during which sulfite is oxidized into sulfate by iodate as shown by the equation:

$$3Na_2SO_3 + KIO_3 \rightarrow 3\ Na_2SO_4 + KI$$

When all sulfite is oxidized into sulfates, the product (KI) reacts in a very rapid step with iodate (KIO_3) in an acidic medium as shown in the equation:

$$5KI + KIO_3 + 3H_2SO_4 \rightarrow 3I_2 + 3K_2SO_4 + 3H_2O$$

Thus, the endpoint of the reaction, indicated by the formation of a dark blue complex, occurs due to the interaction of iodine with starch in the reaction medium.

At a constant concentration of iodine, the reaction rate equation can be written as follows:

$$\text{Rate} = k\ C^n$$

where:
k = rate constant.
C= Sulfite solution concentration (mol/L).
n = Reaction order for sulfite.

When all sulfite reacts, the reaction rate can be expressed by the equation.

$$R = C / t$$

- **Procedure:**

A- The Preparation of the acid sulfite solution (0.01 M) is as follows:

1- Weigh (2 g of starch) powder, then add it to 10 ml of distilled water, then add the starch suspension to 100 ml of boiling distilled water.

2- Chill the starch solution and then transfer it to a standard 250 ml flask containing 1ml of concentrated sulfuric acid added to 25 ml of distilled water (sulfuric acid is added to water in small quantities and not the other way around).

3- Weigh (0.315 g of Na_2SO_3) then dissolve it in the least amount of distilled water and add it to the solution prepared in the previous step and then complete the solution up to the mark with distilled water. (Explain how the molarity was calculated for Na_2SO_3 solution?)

B- The Preparation of KIO_3 (0.02M) potassium iodate solution is as follows:

Dissolve (0.428 g of KIO_3) in distilled water, transfer the solution to a 100 mL volumetric flask and complete to the mark with distilled water. (Explain how the molarity was calculated for KIO_3 solution?)

C- Different concentrations of sulfite solution are prepared as follows:

Blending quantity	
Distilled Water Volume (mL)	Acid sulfite solution volume (mL)
45	50
55	40
65	30
75	20

Add 5 mL of (KIO_3) solution to each flask, bringing the total volume to 100 mL, and start the stopwatch at the same time to measure the reaction time, indicated by the sudden appearance of dark blue color

- **Results and Calculations:**

Solution No.	Time at the end of the reaction (sec)
1	
2	
3	
4	

Taking the logarithm of the equation $Rate = k.\ C^{n}$ to be as follows:

$$log(R) = log\ (k) + n\ log(C)$$

1- Arrange the experiment results according to the following table:

C (Na_2SO_3) mol/L	Log(C)	t (sec)	R= C/ t	log (R)

2- plot a graph of log (R) on the Y-axis and log (C) on the X-axis. A straight line is obtained, where the slope equal to the reaction order (n) and the intercept is equal to (log k), from which the rate constant (k) can be calculated.