Al-Mustansiriyah University
College of Science
Chemistry Department
Practical physical chemistry

Electrochemistry

3th Grade

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Introduction of Electrical Chemistry

Introduction

Electrolytes: are substances that have ability to electrical conductivity.

* *There are two types of conductors involves:*

Electrolytic conductors	***Metallic conductors***
1) The current is caused by the positive and negative ions. 2) The current flows in opposite direction. 3) There is change in mass and concentration, therefore it does contain an energy gap.	1) The current is caused by the valence electrons. 2) The current flows in one direction. 3) There is no change in mass and concentration, therefore it does not contain an energy gap.

**There are varieties of electrolytes according to its equivalence :*

1) Single equivalence, example KCl, HCl, NaOH.

2) Double equivalence, example $BaCl_2$,H_2SO_4.

3) Tri equivalence, example $AlCl_3$, H_3PO_4.

* *There are two types of electrolyte conductor involves:*

1) ***Strongelectrolytes***, which undergo complete ionization, and therefore they have higher conductivity, such as salts, strong acids and strong bases, like KCl.

2) ***Weakelectrolytes***, which undergo only partial ionization, and therefore they have lower conductivity, such as weak acids and weak bases, like CH_3COOH.

** ***Resistance*** (R) is directly proportional with the cell length (ℓ cm) and inversely with the area (a cm^2).

$R \alpha \ell / a$ $R = r \times (\ell)\ cm / (a)\ cm^2 = r \times KK = \ell / a$

R = Resistance (Ω).

r = resistivity or specific resistance (Ω.cm).

K = cell constant = 1 (cm^{-1}).

** ***Conductance*** (G) is the inverse of the resistance.

G = 1/ R its unite (Ω^{-1}) or (S) k = 1 / r

$\therefore k = G \times K$

k = conductivity or specific conductance ($\Omega^{-1}.cm^{-1}$) or ($S.cm^{-1}$).

** ***Molar conductivity*** (Λ_m) is express delivery of one cm^3 from the solution containing one mole of solute.

$\Lambda_m = 1000\ k / C$ Λ_m = (Ω^{-1}. $mol^{-1}.cm^2$) or (S. $mol^{-1}.cm^2$)

C = concentration (mol /cm^3).

** ***Equivalent conductivity*** (Λ_{eq}) is express delivery of one cm^3 from the solution containing one equivalent of solute.

$\Lambda_{eq} = 1000\ k / C$ Λ_{eq} = (Ω^{-1}. $eq^{-1}.cm^2$) or (S. $eq^{-1}.cm^2$)

C = concentration (eq /cm^3).

** ***Specific or Limited molar conductivity*** ($\Lambda_\circ$)

$\Lambda_m = \Lambda_\circ - \beta \sqrt{C}$

$\Lambda_\circ$ = (S. $mol^{-1}.cm^2$)

β = kohlrausch's coefficient (S. $mol^{-1}.cm^2/\sqrt{M}$)

* Prepare Dilute Solution from Concentrate Solution:-

a) For liquid substance

Prepare (0.1M) in (100 mL) of HCl (s. g.=1.19 gm/L), % = 37% & M.Wt. = 36.5 gm/mol.

$$\mathbf{M} = \frac{\mathbf{s.\ g.} \text{ x percentage \% x } 1000}{\mathbf{M.Wt.}}$$

M= 1.19 × 0.37×1000/ 36.5 = 12 mol/L (initial concentration)

To prepare (0.1M) in (100mL)

$M_1 \times V_1 = M_2 \times V_2$

$12 \times V_1 = 0.1 \times 100$ $V_1 = 0.8$ mL

Prepare (0.02, 0.04, 0.06, 0.08 M) in (50 mL) from stock solution of (0.1M) of HCl.

$M_1 \times V_1 = M_2 \times V_2$

1) $0.1 \times V_{11} = 0.02 \times 50$ $V_{11} = 10$ mL

2) $0.1 \times V_{12} = 0.04 \times 50$ $V_{12} = 20$ mL

3) $0.1 \times V_{13} = 0.06 \times 50$ $V_{13} = 30$ mL

4) $0.1 \times V_{14} = 0.08 \times 50$ $V_{14} = 40$ mL

b) <u>For solidsubstance</u>

Prepare (0.1M) in (100 mL) of KCl (M.Wt. = 74.5 gm/mol).

$$\text{Molarity (M)} = \frac{\text{wt (g)}}{\text{M.wt}} \times \frac{1000}{\text{V mL}}$$

0.1 = wt. / 74.5 × 1000/ 100 wt.= 0.745 gm

Prepare (0.01, 0.03, 0.05, 0.07 M) in (50 mL) from stock solution of (0.1M) of KCl.

$M_1 \times V_1 = M_2 \times V_2$

1) $0.1 \times V_{11} = 0.01 \times 50$ $V_{11} = 5$ mL

2) $0.1 \times V_{12} = 0.03 \times 50$ $V_{12} = 15$ mL

3) $0.1 \times V_{13} = 0.05 \times 50$ $V_{13} = 25$ mL

4) $0.1 \times V_{14} = 0.07 \times 50$ $V_{14} = 35$ mL

Experiment (1): Determination the Limiting of Molar Conductivity for Strong and Weak Electrolytes

Introduction:

The conductance (G) of a solution is the inverse of its resistance.

$$G = \frac{1}{R} \quad \text{..... (1)}$$

As resistanceexpressed in ohm (Ω), the conductance of a sample expressed in (Ω^{-1}). The reciprocal ***ohm*** *used to be the* ***mho****, but now called Siemens (S). The conductance of a sample decreases with its length (ℓ) and increases with its cross-sectional area (a). We therefore write:*

$$G = k\frac{a}{\ell} \quad \text{.....(2)}$$

Where ***(ℓ /a)*** *is called cell constant,* ***(k)*** *is the conductivity of the solution. Can be defined conductivity is measurement of the ease with which electrical current flows through the solution. The units of* ***(k)*** *are (****mS.cm^{-1}, μS.cm^{-1}, S.cm^{-1}****). The conductivity depends upon:*

(i) *The molar concentration*

(ii) *The charge numbers*

(iii) *The mobility of the ionic species present.*

The molar conductivity (Λ_m) defined as:

$$\Lambda m = \frac{1000\ k}{C} \quad \text{.....(3)}$$

Where: (C) is concentration of ionic species present. The units of (Λ_m) are ($S.cm^2.mol^{-1}$, $\mu S.cm^2.\ mol^{-1}$, $mS.cm^2.\ mol^{-1}$). For strong electrolytes (Λ_m) varies linearly with square root of the concentration, this variation called kohlrausch's law:

$$\Lambda_m = \Lambda o - \beta\sqrt{C}$$

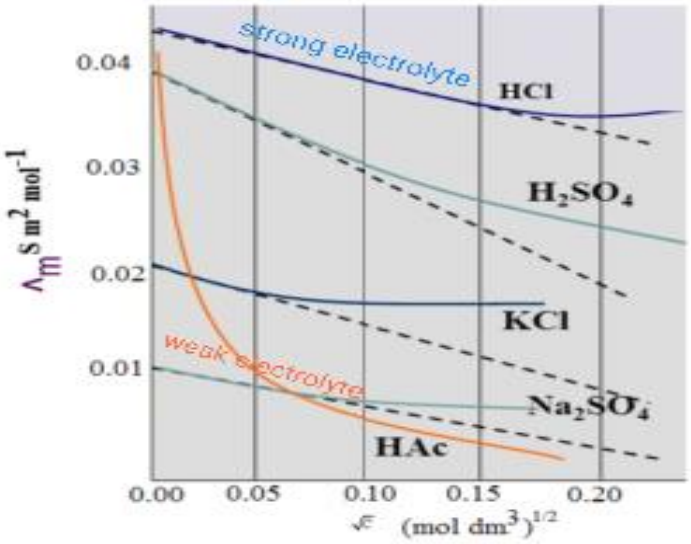
Where: (β) kohlrausch's coefficient.

$(\Lambda_o)$is the limiting molar conductivity; it presents the molar conductivity in the limit of zero concentration (i.e. infinite dilution).

Thus, a plot of (Λ_m) against $(\sqrt{c})$ then the intercept at (C≈0) will produce (Λ_o). The kohlrausch's law is can be used only for strong electrolytes because those substances are virtually fully ionized in solution. Weak electrolytes are not fully ionized in solution, therefore cannot be used the previous equation for calculate (Λ_o), for weak electrolytes (Λ_o) must be determined using un indirect approach, which relies upon the limiting molar conductivities of several strong electrolytes at infinite dilution (the kohlrausch method).for example the (Λo) of CH_3COOH can be calculated from following equation:

$$CH_3COONa + HCl \longrightarrow CH_3COOH + NaCl$$

$$\Lambda o(CH_3COOH) = \Lambda o(HCl) + \Lambda o(CH_3COONa) - \Lambda o(NaCl)$$



Procedure:

A- Preparation of required solutions:

1- *Prepare the following dilute solutions of **HCl** (2, 4, 6 and 8)×10^{-4} M in (50) mL from the stock solution (0.1)M.*

2- *Prepare the following dilute solutions of **CH_3COONa** (2, 4, 6 and 8)×10^{-4} Min (50) mL from the stock solution (0.1)M.*

3- *Prepare the following dilute solutions of **NaCl** (2, 4, 6 and 8)×10^{-4} M in (50) mL from the stock solution (0.1)M.*

B- Conductivity measurements:

1- *Making sure the conductivity probe has been correctly calibrated with standard solution of KCl its conductivity is known.*

2- *Use the conductivity meter to measure conductivity for each solution has been prepared during dip the conductivity probe in to the solution, which you want to measure its conductivity.*

3- *Once the reading has stabilized (10s), record the conductivity value on your data sheet.*

4- *Repeat steps (2,3) to measure conductivities of all solutions (3 sets of 4 solutions)and arrangedata in the table:*

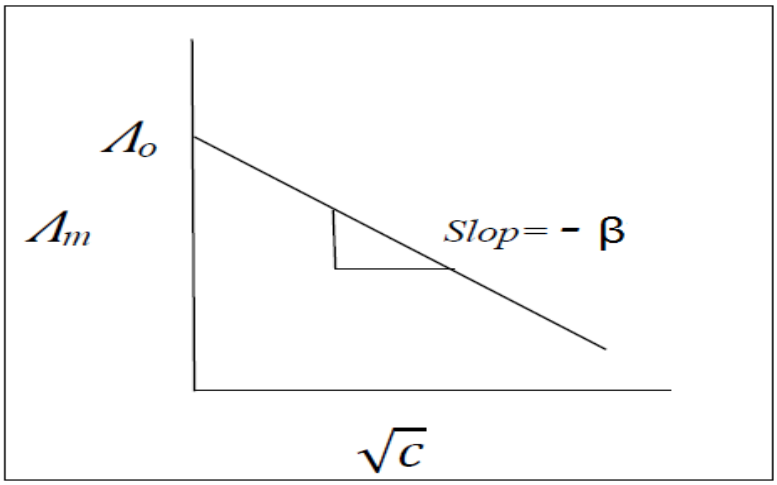
HCl				***CH_3COONa***				***NaCl***			
HCl (M) ×10^{-4}	k *(µs.cm^{-1})*	Λ_m *(µs. mol^{-1}. cm^2)*	$\sqrt{c}$	*CH_3COONa (M) ×10^{-4}*	k *(µs.cm^{-1})*	Λ_m *(µs. mol^{-1}. cm^2)*	$\sqrt{c}$	*NaCl (M) ×10^{-4}*	k *(µs.cm^{-1})*	Λ_m *(µs. mol^{-1}. cm^2)*	$\sqrt{c}$
2 4 6 8				2 4 6 8				2 4 6 8			

Calculations:

1- *Using equation* (Λ_m = 1000 k / C) *to determine* (Λ_m) *for all of the solutions that contain strong electrolyte.*

2- *Plot* (Λm) *versus* $\sqrt{C}$ *(eq. 4) and determine* (Λo) *for all the strong electrolytes*

3- *Using the kohlrausch method, evaluate* (Λo) *for acetic acid.*



Discussion:

1- Why do you can't use ($\Lambda_m = \Lambda o - \beta\sqrt{C}$) *to calculate*(Λo) *for weak electrolyte?*

2- Which has higher conductivity, the dilute solution or the concentrated solution? why?

3- What is the difference between the metallic conductivity and electrolytic conductivity?

Experiment (2): Determination of K_a Using the Conductivity Measurements

Introduction:

Weak electrolytes are not fully ionized in solution, they include weak acids and weak bases such as CH_3COOH and NH_3. The marked concentration dependence on their molar conductivities arises from the displacement of the equilibrium of reaction:

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

Towards products at low molar concentration there for Λ_m increase with dilution of solution and increasing degree of ionization (α)

$$\alpha = \frac{\Lambda_m}{\Lambda_o} \qquad \ldots\ldots\ldots\ldots\ldots\ldots (1)$$

*The equilibrium constant of acetic acid **(K_a)** can be represented by following equation:*

$$K_a = \frac{[H^+]\,[CH_3COO^-]}{[CH_3COOH]} \times \frac{\gamma_{H+}\,\gamma_{CH3COO-}}{\gamma_{CH3COOH}} \qquad \ldots\ldots(2)$$

At low concentrations the activity coefficients are approximately equal to (1) for H^+, CH_3COO- and CH_3COOH is an uncharged molecule has activitycoefficient is approximately unity. We can write equation (2) asfollowing:

$$K_a = \frac{[H^+]\,[CH_3COO^-]}{[CH_3COOH]} \qquad \ldots\ldots(3)$$

$[H^+] = \alpha C \qquad [CH_3COO^-] = \alpha C \qquad [CH_3COOH] = (1-\alpha)C$

Where C =primary concentration of CH_3COOH

Then:

$$K_a = \frac{\alpha^2 C^2}{(1-\alpha)C} \Longrightarrow \boxed{K_a = \frac{\alpha^2 C}{(1-\alpha)}} \qquad \text{.....(4)}$$

From equation (1) and (2) produce following equation:

$$\boxed{K_a = \frac{\Lambda_m{}^2 C}{\Lambda_o (\Lambda_o - \Lambda_m)}} \qquad \text{.....(5)}$$

We rearrange equation (5) to obtain:

$$\boxed{\frac{1}{\Lambda_m} = \frac{1}{\Lambda_o} + \frac{\Lambda_m C}{K_a (\Lambda_o)^2}} \qquad \text{.....(6)}$$

Procedure:

A- Preparation of required solutions:

Preparation of dilute solutions of acetic acid (0.002, 0.004, 0.006 and 0.008)M in (50)ml from the stock solution (0.1)M.

B- Conductivity measurements:

1- Making sure that the conductivity probe has been correctly calibrated with standard solution of KCl its conductivity is known.

2- Use conductivity meter to measure conductivity for each solution has been prepared during dip the conductivity probe in to the solution.

3- Once the reading has stabilized about (10s) record the conductivity value on your data sheet.

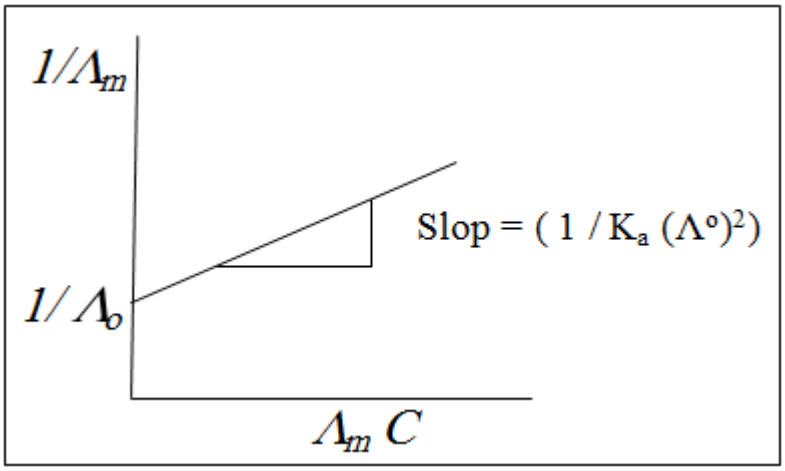
4- arrange data in the table:

$[CH_3COOH]M$	$K\ (\mu S.\ cm^{-1})$	Λ_m	$1/\Lambda_m$	$\Lambda_m C$
0.002 0.004 0.006 0.008				

Calculations:

1- using equation $\Lambda_m = \frac{1000\,k}{C}$ *to determine* Λ_m *for all solutions that contain acetic acid*

2- using equation (6) and plotting $(1/\Lambda_m)$ *against* $(\Lambda_m\ C)$ *from intercept youcan find* $(1/\Lambda_o)$ *and from the slop you can find* K_a *of acetic acid.*



Discussion:

1) what is the unite of K_a *?*

2- Why must temperature be determined in this experiment?

Experiment(3):Determination the Rate Constant of Saponification of Ethyl Acetate by the Conductometric Titration

Introduction:

Saponification express of the reaction between ethyl acetate and sodium hydroxide according to the following equation :

$$CH_3COOC_2H_5 + NaOH \rightleftharpoons CH_3COONa + C_2H_5OH$$

This reaction for second order, if the concentrations of the reactants are equal, can be used the reaction law in the following form:

$$\frac{x}{a-x} = a\,k_2\,t \qquad \ldots\ldots(1)$$

Thereaction involves strong electrolytes such as NaOH and CH_3COONa therefore we can follow this reaction by change of the conductivity with the concentration, this is possible because the conductivity is proportional to concentration. Wecan follow the reaction by observing the change of conductivity of NaOH. The conductivity at beginning of the reaction belong to NaOH and at end of the reaction it belongs to CH_3COONa. We can suppose the following:

$$\left.\begin{array}{l} a \propto k_o - k_\infty \\ x \propto k_o - k_t \end{array}\right\} \Longrightarrow a - x \propto k_t - k_\infty$$

Where (k_o) is the conductivity at beginning the reaction and(k_∞) is the conductivity at end the reaction. We can write the rate law equation (1) as following:

$$\frac{k_o - k_t}{k_t - k_\infty} = a\, k_2\, t \qquad \ldots\ldots\ldots (2)$$

We plot $(k_o - k_t / k_t - k_\infty)$ *against* (t) *we get straight line its slope is equal* (ak_2), (a) *represents the original concentration of NaOH is known, then we can find rate constant* (k_2). *It is considerable that the standard unit of rate constant* (k_2) *is* $(L.\ mole^{-1}.\ min^{-1})$ *or* $(L.\ mole^{-1}.\ sec^{-1})$ *depend on unit used in experiment.*

Procedure:

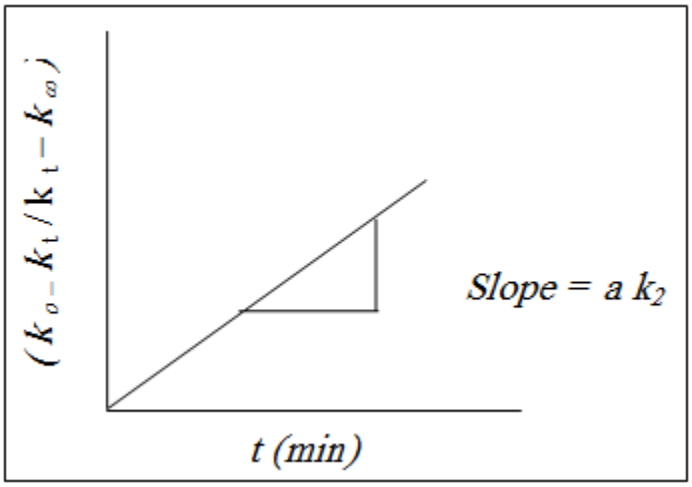
1) *Measure the conductivity of the distilled water.* ***(Note:- we can neglect the conductivity of distilled water).***
2) *Measure the conductivity of NaOH (0.005M) to calculate theconductivity at beginning the reaction (* k_o *).*
3) *Measure the conductivity of* CH_3COONa *(0.005M) it represents the conductivity at end the reaction (* k_∞*).*
4) *Add (25ml) of (0.01M) of NaOH in to (25ml) of (0.01M)* $CH_3COOC_2H_5$ *then record the time of beginning the reaction.*
5) *Measure the conductivity of the solution after (3,6,9,12,15) min to calculate the conductivity at time of the reaction (* k_t *).*

Calculations:

1) *Arrange the results as in table:*

time (min)	k_o	k_∞	k_t	$(k_o - k_t)$	$(k_t - k_\infty)$	$k_o - k_t / k_t - k_\infty$
3 *6* *9* *12* *15*						

2) Plot ($k_o - k_t / k_t - k_\infty$) against time (t) by using equation (2). From the slope determine the rate constant where (slope = a k_2).



Discussion:

1) *Is it important to make the original concentration of the reactants equal?*

2) *Does conductivity of the distilled water have effect on the results?*

Experiment (4): Conductometric Titration of Acids with Strong Base

Introduction:

Conductometric titrations used to investigate neutralization titrations by observing the change in electrical conductivity. The principle of conductometric titration is based on the fact that during the titration,one of the ions is replaced by the other and invariably these two ions differ in the ionic conductivity with the result that conductivity of the solution varies during the course of titration. The titration reaction involve the addition the reagent to a solution that contains a substance with unknown molarity. Just before and after the equivalence point, there is a marked difference in the rate of change of conductivity with addition of reagent. The curve of titration involve plotting conductivity of the solution against the volume of standard reagent. The use of an excess of the standard reagent establishes the correct location of the equivalence point therefore it is then possible to find the volume of reagent equivalent to the solution titrated by measuring the conductivity.

There aresome advantages of conductometric titration:

1) *No need toindicator.*

2) *The end point identified graphically with minimumerror.*

3) *Used with colored liquids for which ordinary indicators cannotwork.*

4) *Used for dilute solutions as well as for very weak acids.*

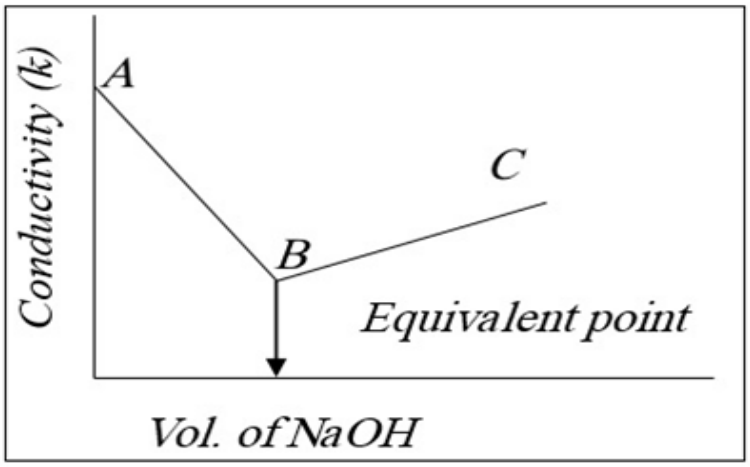
Examples about conductometric titrations:

1) ***Titration of strong acid (HCl) with strong base (NaOH).***

In first, the solution of HCl is unknown molarity; it is possible to find the concentration of HCl by titration with NaOH (known concentration) and using the titration curve. Before the addition of the base, the acid solution has a high conductivity (due to completely ionized in aqueous medium furthermore as H^+ ion has very large ionic mobility). As base added, the H^+ ion has react with OH^- ion to form water and being each H^+ ion replaced by Na^+ ion (where the conductivity of Na^+ less than conductivity of H^+). Consequently, the conductivity of the solution decreases and keeps on falling with addition of the base until reached to the equivalent point. At the equivalence point, the solution contains only NaCl.

$$HCl + NaOH \longrightarrow H_2O + Na^+ + Cl^-$$

*Further addition of alkali introduces now an excess of thefast OH^- ionsandtheycausestheconductivitytoriseagain. Thisvariation of solution conductivity plotted against the volume of alkali added. The result of the titration curve looks like the curve **ABC** shown in following figure:*



Where point (B) represents the minimum conductivity there is no excess present of either acid or base and hence it is theequivalentpoint.

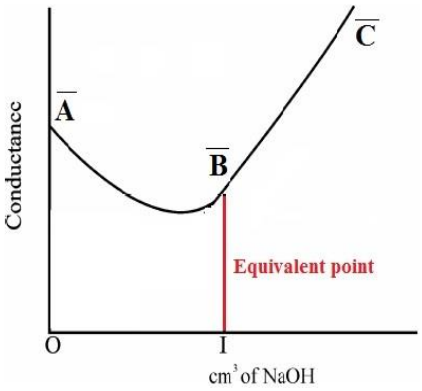
2) Titration of weak acid (CH_3COOH) with strong base (NaOH).

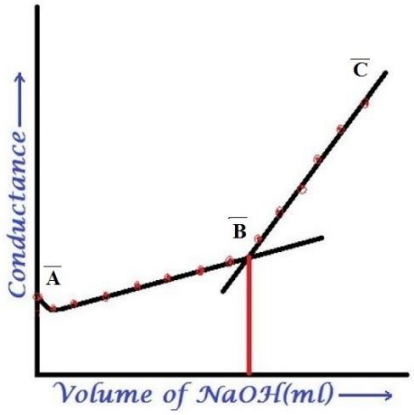
Since the acid is weak, its conductivity is relatively low. On the addition ofbase, there is decrease in conductance not only due to the replacement of $H^+$by Na^+but also suppresses the dissociation of acetic acid due to common ion acetate. But very soon, the conductance increases on adding NaOH as NaOH neutralizes the un-dissociated CH_3COOH to CH_3COONa which is the strong electrolyte.

This increase in conductance continues raise up to the equivalencepoint. The graph near the equivalence point is curved due the hydrolysis of salt CH_3COONa. Beyond the equivalence point, conductance increases more rapidly with the addition of NaOH due to the highly conducting OH^- ions.

$$CH_3COOH + NaOH \rightleftharpoons H_2O + CH_3COONa$$

The variation of conductivity with volume of NaOH gives curve looks like the curve $\overline{A}\,\overline{B}\overline{C}$.





Where point (B) (in cases titration HCl or CH_3COOH with NaOH) represents the equivalent point or the neutralization point. In this point, moles number of H^+ ions is equal to the moles number of OH^- ions, hence mole balance at the neutralization points yields:

$n H^+ = n OH^-$ ……. *(1)*

$(M \times V)_{acid} = (M \times V)_{base}$……..(2)

Procedure:

A- Titration HCl with NaOH:

(95mL) of distilled water added to (5mL) of HCl (unknown concentration) and its conductivity is measured, then it is measured after addition of (1mL) from solution (0.1M) of NaOH. This is repeated until (10mL) of NaOH is added.

B- Titration of CH_3COOH with NaOH:

Withdraw (5mL) of CH_3COOH from the stock solution (unknown concentration) and add (95mL) of distilled water, measure its conductivity, then add (1mL) of NaOH (0.1M) on solution of CH_3COOH and measure conductivity of solution. This is repeated until (10mL) of NaOH is added.

Calculations:

1) For each experiment, draw the conductivity (k) versus the volume of NaOH added.

2) Find the concentration of CH_3COOH and HCl.

Discussion:

1) What is the principle that the conductometric titration depend on it?
2)How can you know the neutralization point during the conductometric titration? Explain it.
3)Do not need indicator when we use the conductometric titration. Why?
4)When do you prefer the conductometric titration on the ordinary indicators?

Experiment (5):Potentiometric Titration of Strong Acid with Strong Base

Introduction:

Potentiometric titration is a laboratory method to determine the concentration of a given analyte (unknown). In this method, there is no use of a chemical indicator. Instead, thepH or electric potential across the substance is measured.By using the measurements of pH in this process can determine the equivalent point.

pH express to measure the hydrogen ions in solutions, or it's the negative logarithm for the activity of hydrogen ions in solution (pH = -log a_{H+}*)*

Potentiometric titration is provide more reliable data than data from titrations that use chemical indicators andare particularly useful with colored or turbid solutions and for detecting the presence ofunsuspected species

Difference between Equivalence point and Endpoint

Equivalence point	Endpoint
The point in the titration process where the chemical reaction in the titration mixture ends is called equivalence point.	The point in the titration process which is indicated by color change of the indicator is called endpoint.
It is the point where the analyte has completely reacted with the titrant.	It doesn't always give the point where the analyte has completely reacted with the titrant.
It is not always indicated by color change of the reaction mixture.	It is always indicated by the color change of the reaction mixture.
It gives the point where reaction ends.	It doesn't always give the point where reaction ends.
It comes either almost with endpoint or before the endpoint.	It comes either almost with the equivalence point or after the equivalence point.
Weak acids can show multiple equivalence points during titration.	Weak acids can show only one endpoint during titration.

Potentiometric titrations classified as precipitation titrations, complex formation titrations, oxidation/reduction titrations and acid-base titrations (neutralization titrations).

Procedure:

A- Titration by using indicators:

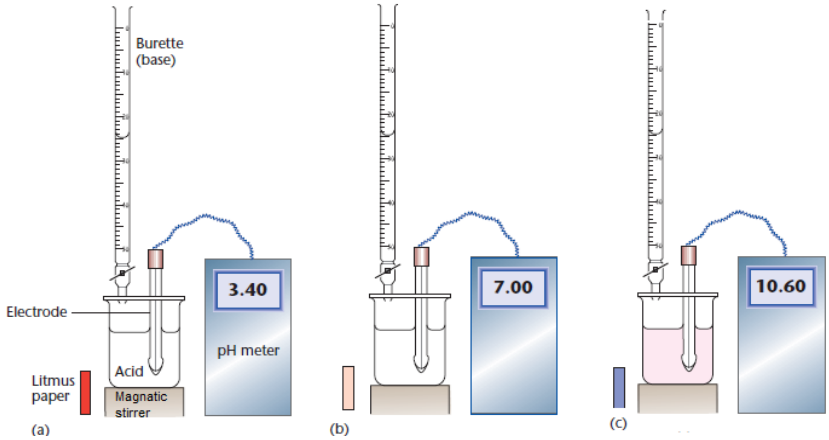
Take (25mL) of HCl (unknown) and put it in the beaker with capacity (100mL) then add two drops of indicator (phenolphthalein).Starttitration with the solution (0.1M) of NaOHuntil the color of the solution changed to pink, finally determine the volume of base which represents the volume at end point. Calculate the molarity of HCl solution.

$$(M\times V)_{acid} = (M\times V)_{base}$$

B- Titration by using pH –meter:

1) Wash the electrode with D. W. and dry it by filter paper then immerse it many times in beakers that contain the buffer solutions: pH =4, pH =9.

2) Wash the electrode again and dry it by filter paper then immerse it the beaker with capacity (250mL) which contain (25mL) of HCl (unknown) and record the pH value.



3) *begin the titration by adding the base of NaOH (0.1M) that is in burette, according to the following steps:*

a-At first, add (1mL) at time, towards the equivalence point, then reduce the amounts (about 0.1 mL) at the end.

b- Continue the titration, now adding larger volumes of NaOH until the pH stop to rise markedly.

Increments, mL	Volume of NaOH, mL
1	0-8
0.1	8-12
1	12-18
0.1	18-22
1	22-28

4- Record the values of pH for the mixture in the beaker after each addition.

5- In the end of experiment, wash the electrode and immerse it in the D.W.

Calculations:

1) Set table including the results of above method like below table.

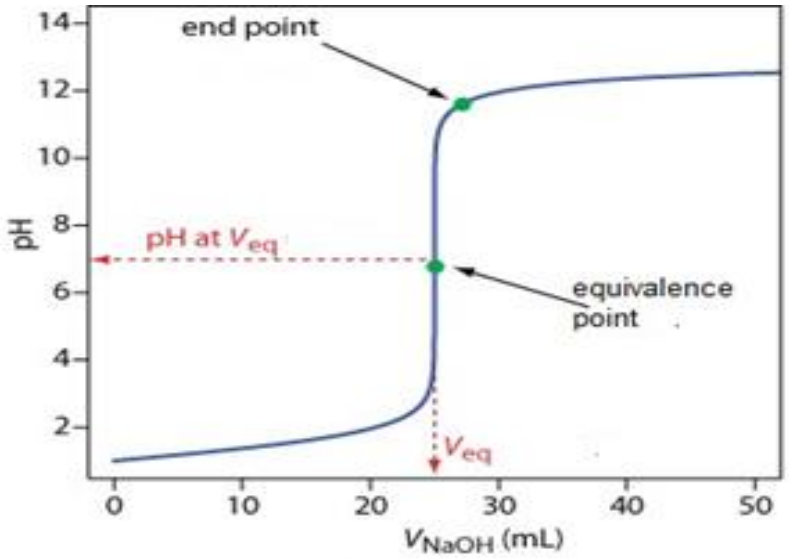
2) Plot a graph between the values of pH and volumes of NaOH.

3) Determine the equivalencepoint and calculate the molarity of HCl solution.

$(M\times V)_{acid} = (M\times V)_{base}$

V_{NaOH} (mL)	pH
0	
1	
2	

28	



Discussion:

1- Why is HCl consider a strong acid?

2- Classify the electrolytes in terms of equivalence.

Experiment (6): Application of (Lambert – Beer) Law

Introduction:

Spectrophotometric methods include measuring the intensity of incident light and transmitted light at aspecific wavelength. There are two types of devicearethe spectral: (1) Single beam (2) Double beam.

Law (Lambert – Beer):

Absorbs equal parts of the light beams by equal changes in the concentration of the absorbed material when the optical path length in the absorbing material constant.

$$A = \varepsilon\, C\, \ell$$

$\%T = (I / I_O) \times 100$

$A = \log I_O / I$

$= \log 1 / T$

A = absorbance

ε *= molar absorbance coefficient* ($mol^{-1}.L.cm^{-1}$)

ℓ *= long the light path inside the solution (1cm)*

T = transmittance

C = concentration ($mol.L^{-1}$)

I = the intensity of transmittedlight

I_o *= intensity of incidentlight*

Wavelength (λ):*The distance between adjacent peaks of awave packet, thathas the standard unit:nm=*10^{-9}*m,* $A^o = 10^{-10}$*m,* $\mu m = 10^{-6}$*m*

The greatest wavelength (λmax): *The wavelength, which has the highest absorption of the substance.*

Light intensity (I): *The number of photons absorbed per second.*

Photon: *Is units of energy.*

Beer-Lambert Law Applications*: (just view)*

This law finds applications in various fields such as:

- ***Analytical chemistry:*** *This analysis mainly concentrates on the separation, quantification and identification of matter by spectrophotometry. There is no involvement of extensive pre-processing of the sample to get the results. For example, bilirubin count in a blood sample can be determined by using a spectrophotometer.*
- ***Atomic Absorption:*** *The application of atomic absorption spectrometry (AAS) for the determination of metal concentrations.*
- ***In atmosphere:*** *Solar or stellar radiation in the atmosphere can be described using this law. The law in atmospheric applications has a modified equation.*

Beer-Lambert Law Limitations

Using this law it becomes easy to study the absorptivity coefficient of the sample when the concentration is low <10mM but as the concentration becomes high >10mM there is a deviation as the electrostatic interactions become more. The value of absorbance should be between (0-1).

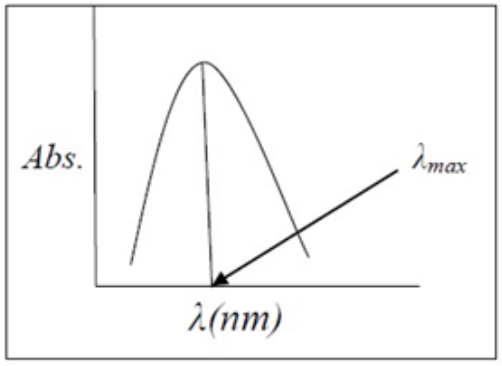
Measuring the absorbance of ($KMnO_4$)

a) Finding (λ_{max}):

1) Prepare dilute solutions of $KMnO_4$ (0.1M):(1,2, 3,4)$\times 10^{-4}$M in volumetric flask (50ml) according to law ($M_1 V_1 = M_2 V_2$).

2) Measure absorbance of the lower concentration (1×10^{-4})M versus water (blank) in the wave length range (400 – 600)nm and that each (10nm)

tonote the highest absorption value read by the device and install it to be (λmax).



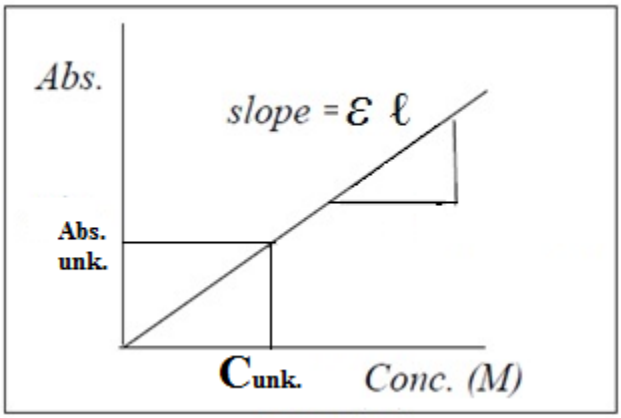
Wave length(nm)	Absorbance
400	
410	

600	

b) Finding (ε)& Concentration of unknown:

1- When you install (λmax) from step (a) was appointed absorbance of each solution.

*2- Plot a relationship between absorbance **(A)** and concentration (**C**) then determine the molar absorption coefficient **(ε)** from the slopand the concentration of unknown.*



Conc.(mol/L)	Absorbance
1×10^{-4}	
2×10^{-4}	
3×10^{-4}	
4×10^{-4}	
unknown	

Discussion:

1-Is it possible to measure the potassium permanganate in the ultraviolet region near the visible region?

2-What is the difference between ultraviolet and visible spectrum?