

**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:

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

As resistance is expressed in ohms( $\Omega$ ), the conductance of a sample is expressed in ( $\Omega^{-1}$ ). The reciprocal ohm used to be called the ohm, but now is called Siemens (S). The conductance of a sample decreases with its length(l) and increases with its cross-sectional area (A). We therefore write:

$$G = k \frac{A}{l} \quad \dots\dots\dots(2)$$

where  $\frac{A}{l}$  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 ( $\text{mScm}^{-1}$ ,  $\mu\text{Scm}^{-1}$ ,  $\text{Scm}^{-1}$ ). The conductivity depends upon :

- (i) the molar concentration
- (ii) the charge numbers
- (iii) the mobilities of the ionic species present.

The molar conductivity ( $\Lambda_m$ ) is defined as:

$$\Lambda_m = \frac{k}{1000 c} \quad \dots\dots\dots(3)$$

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

$$\Lambda_m = \Lambda_o - \epsilon \sqrt{C} \quad \dots\dots\dots(4)$$

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

Thus a plot of  $(\Lambda_m)$  against  $(\sqrt{C})$  then the intercept at  $(C=0)$  will produce  $(\Lambda_0)$ . The Kohlrausch's law can be used only for strong electrolytes because these substances are virtually fully in solution. Weak electrolytes are not fully ionized in solution, therefore cannot be used in the previous equation to calculate  $(\Lambda_0)$ , for weak electrolytes  $(\Lambda_0)$  must be determined using an indirect approach, which relies upon the limiting molar conductivities of several strong electrolytes at infinite dilution (the Kohlrausch method). For example, the  $(\Lambda_0)$  of  $\text{CH}_3\text{COOH}$  can be calculated from the following equation:

$$\Lambda_0(\text{CH}_3\text{COOH}) = \Lambda_0(\text{HCl}) + \Lambda_0(\text{CH}_3\text{COONa}) - \Lambda_0(\text{NaCl})$$

**Procedure:**

**A- Preparation of required solutions:**

- 1- Prepare the following dilute solutions of NaCl (0.1, 0.08, 0.04 and 0.02)M in (50)ml from the stock solution (0.5)M.
- 2- Repeat the step (1) to prepare the dilute solutions of  $\text{CH}_3\text{COONa}$  (0.1, 0.08, 0.04 and 0.02)M.
- 3- Prepare the following dilute solutions of HCl (0.01, 0.02, 0.04 and 0.005)M in (50)ml from the stock solution (0.05)M.

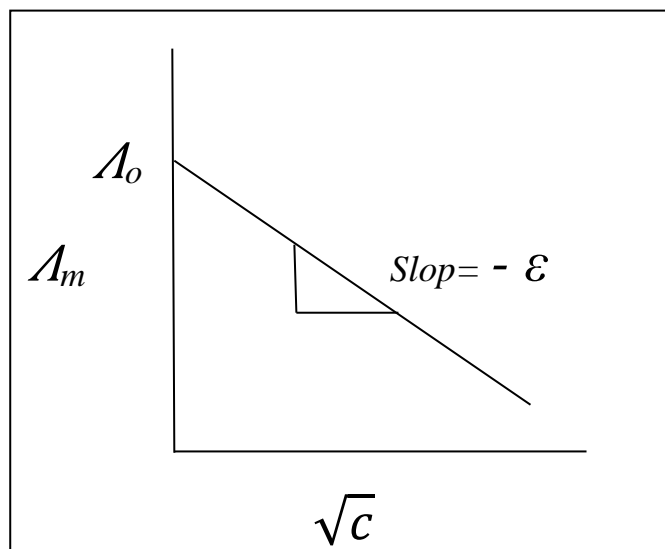
**B- Conductivity measurements:**

- 1- Making sure the conductivity probe has been correctly calibrated with a standard solution of KCl its conductivity is known.
- 2- Use the conductivity meter to measure conductivity for each solution that has been prepared during the dip the conductivity probe in to the solution, which you want to measure its conductivity.
- 3- Once the reading has stabilized (10 s), record the conductivity value on your data sheet.
- 4- Repeat steps (2, 3) to measure conductivities of all solutions (3 sets of 4 solutions).
- 5- Arrange data in the table:

NaCl		$\text{CH}_3\text{COONa}$		HCl	
$[\text{NaCl}]M$	$k(\mu\text{s cm}^{-1})$	$[\text{CH}_3\text{COONa}]M$	$k(\mu\text{s cm}^{-1})$	$[\text{HCl}]M$	$k(\mu\text{s cm}^{-1})$
0.02		0.02		0.005	
0.04		0.04		0.01	
0.08		0.08		0.02	
0.1		0.1		0.04	

**Calculations:**

- 1- Using equation (3) to determine ( $\Lambda_m$ ) for all of the solutions that contain strong electrolyte.
- 2- Plot  $\Lambda_m$  versus  $\sqrt{C}$  (eq. 4) and determine  $\Lambda_o$  for all the strong electrolytes
- 3- Using the Kohlrausch method, evaluate  $\Lambda_o$  for acetic acid.



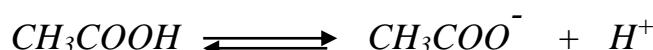
**Discussion:**

- 1- Why do you can't use  $\Lambda_m = \Lambda_o - \epsilon \sqrt{C}$  to calculate  $\Lambda_o$  for weak electrolytes?
- 2- in measurement of conductivity by using conductivity meter we must use alternating current (A. C ) and must not use direct current (D. C.).
- 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  $\text{CH}_3\text{COOH}$  and  $\text{NH}_3$ . The marked concentration dependence on their molar conductivities arises from the displacement of the equilibrium of reaction:



Towards products at low molar concentration there for  $\Lambda_m$  increase with dilution of solution and increasing degree of ionization ( $\alpha$ )

$$\alpha = \frac{\Lambda_m}{\Lambda^0} \quad \dots\dots(1)$$

The equilibrium constant of acetic acid ( $k_a$ ) can be represented by following equation:

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \times \frac{\gamma_{\text{H}^+} \gamma_{\text{CH}_3\text{COO}^-}}{\gamma_{\text{CH}_3\text{COOH}}} \quad \dots\dots(2)$$

At low concentrations the activity coefficients are approximately equal to (1) for  $\text{H}^+$ ,  $\text{CH}_3\text{COO}^-$  and  $\text{CH}_3\text{COOH}$  is an uncharged molecule has activity coefficient is approximately unity. We can write equation (2) as following:

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \quad \dots\dots(3)$$

$$[\text{H}^+] = \alpha C \quad [\text{CH}_3\text{COO}^-] = \alpha C \quad [\text{CH}_3\text{COOH}] = (1-\alpha) C$$

Where  $C$  = primary concentration of  $\text{CH}_3\text{COOH}$

Then:

$$K_a = \frac{\alpha^2 C^2}{(1-\alpha)C} \implies K_a = \frac{\alpha^2 C}{(1-\alpha)} \quad \dots\dots(4)$$

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

$$K_a = \frac{\Lambda_m^2 C}{\Lambda^0 (\Lambda^0 - \Lambda_m)} \quad \dots\dots(5)$$

We rearrange equation (5) to obtain:

$$\frac{1}{\Lambda_m} = \frac{1}{\Lambda^0} + \frac{\Lambda_m C}{K_a (\Lambda^0)^2} \quad \dots\dots(6)$$

**Procedure:**

**A- Preparation of required solutions:**

Preparation of dilute solutions of acetic acid (0.02, 0.01, 0.008 and 0.004)M in (50)ml from the stock solution (0.05)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 (10) sec record the conductivity value on your data sheet.

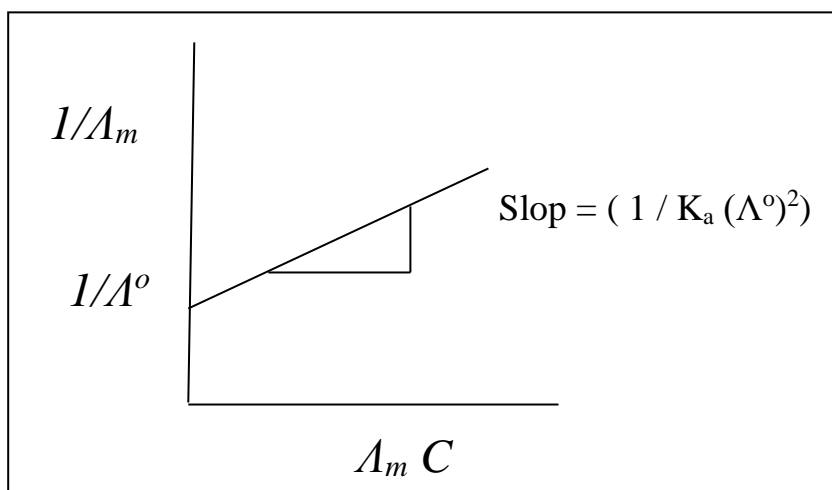
4- arrange data in the table:

$[CH_3COOH]M$	$K (\mu S cm^{-1})$
0.004	
0.008	
0.01	
0.02	

**Calculations:**

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

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



***Discussion:***

- 1- What are the sources of wrong in this experiment (theoretical and practical)?
- 2- Why must temperature be determined in this experiment?

**Experiment (3): Conductometric titrations of acids with strong base**

**Introduction:**

Conductometric titrations can be used to investigate neutralization titrations by observing the change in electrical conductivity. A reaction that causes change in the number or type of ions present in a solution can be observed by changing in conductivity of the solution. The titration reaction involves the addition of 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 involves 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 are some advantages of conductometric titration:

- 1) No need to indicator.
- 2) End point can be identified graphically with minimum error.
- 3) Can be used with colored liquids for which ordinary indicators cannot work.
- 4) Can be used for dilute solutions as well as for very weak acids.

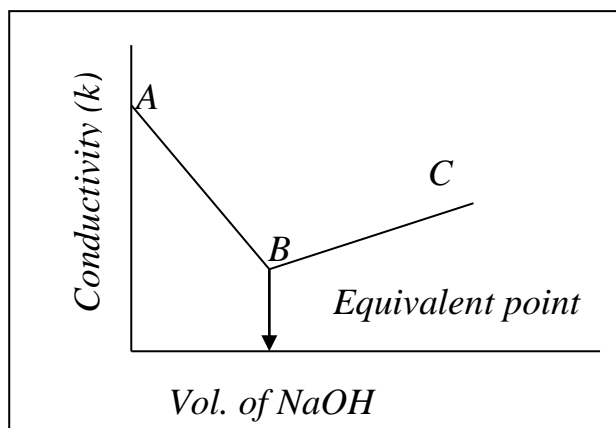
Examples about conductometric titrations:

1) The titration of a strong acid like HCl with a strong base like NaOH. Consider first the solution of HCl is unknown molarity it is possible to find the concentration of HCl by titration of HCl with NaOH has unknown concentration and using the titration curve. Before the addition of the base, the acid solution has a high conductivity. As base is added, the hydrogen has reacted with hydroxyl ions to form water.



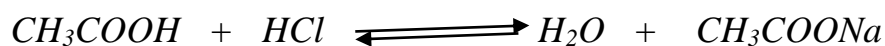
And they are replaced by the much slower cations ( $\text{Na}^+$ ) their conductivity is less than conductivity of ( $\text{H}^+$ ). Consequently the conductivity of the solution decreases and keeps on falling with addition of the base until the equivalent point is reached further addition of alkali introduces now an excess of the fast

hydroxyl ions and they causes the conductivity to rise again when this variation of conductivity of the solution is plotted against the volume of alkali added. the result of the titration curve looks like the curve ABC shown in following figure:

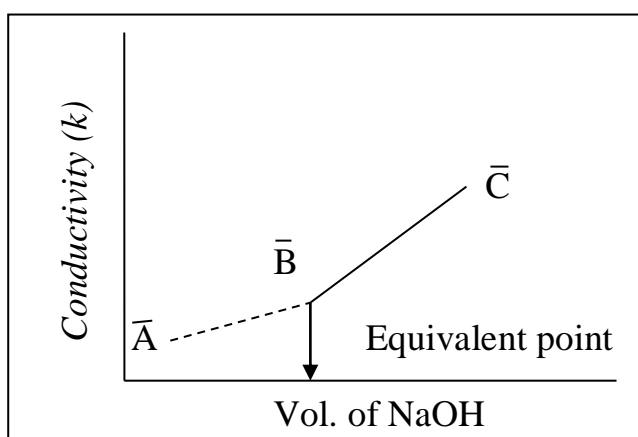


Where point (B) represents the minimum conductivity there is no excess present of either acid or base and hence it is the equivalent point.

2) The titration of a weak acid like  $\text{CH}_3\text{COOH}$  with a strong base like  $\text{NaOH}$ . since the acid is weak, its conductivity is relatively low. As the base is added. The poorly conductivity acid is converted to highly ionized salt



And consequently the conductivity goes up, once the acid is neutralized addition of excess base causes another sharp increasing in conductivity. The variation of conductivity with volume of  $\text{NaOH}$  gives curve looks like the curve  $\bar{A}\bar{B}\bar{C}$ :



Where point (B) represents the equivalent point in order to find the concentration of  $\text{HCl}$  or  $\text{CH}_3\text{COOH}$  during the titration, mole balance can be done at the neutralization point, moles number of hydrogen ions in the medium



is equal to the moles number of hydroxyl ions, hence mole balance at the neutralization points yields :

$$n H^+ = n OH^- \dots\dots\dots(1)$$

In cases like titration HCl with NaOH or titration CH<sub>3</sub>COOH with NaOH can be used equation (1) as at the following equation:

$$M_{acid} \times V_{acid} = M_{base} \times V_{base} \dots\dots(2)$$

**Procedure:**

**A- Titration HCl with NaOH:**

(95ml) of distilled water is 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<sub>3</sub>COOH with NaOH:**

Withdraw (5ml) of CH<sub>3</sub>COOH 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<sub>3</sub>COOH 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<sub>3</sub>COOH and HCl.

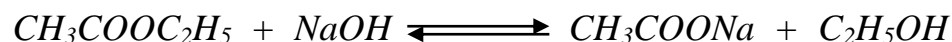
**Discussion:**

- 1- What is the principle that the conductometric titration depend on it?
- 2- Explain how can you know the neutralization point during the conductometric titration?
- 3- We don't need indicator when we use the conductometric titration?
- 4- When do you prefer the conductometric titration on the ordinary indicators?
- 5- At conductometric titration we prefer use concentration of reagent is greater than concentration of the solution titrated, why?

**Experiment (4):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 :



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

$$\frac{x}{a-x} = a k_2 t \quad \dots\dots(1)$$

The reaction involves strong electrolytes such as NaOH and CH<sub>3</sub>COONa therefore we can follow this reaction by change of the conductivity with the concentration, this is possible because the conductivity is proportional to concentration. We can 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<sub>3</sub>COONa. We can suppose the following :

$$a \propto k_o - k_\infty$$

$$x \propto k_o - k_t \quad \Longrightarrow \quad a - x \propto k_t - k_\infty$$

Where ( k<sub>o</sub> ) is the conductivity at beginning the reaction and ( k<sub>∞</sub> ) is the conductivity at end the reaction. We can write the rate law equation (1) as following:

$$\frac{k_o - k_t}{k_t - k_\infty} = a k_2 t \quad \dots\dots(2)$$

We plot  $(k_o - k_t / k_t - k_\infty)$  against  $(t)$  we get straight line its slope is equal  $(a k_2)$  are presents 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  $(l \cdot mole \cdot min)$  or  $(l \cdot mole \cdot sec)$  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 (0.025M) to calculate the conductivity at beginning the reaction  $(k_o)$  .

3) Measure the conductivity of  $CH_3COONa$  (0.025M) it represents the conductivity at end the reaction  $(k_\infty)$  .

4) Add (25ml) of (0.05M) of NaOH in to (25ml) of (0.05M)  $CH_3COOC_2H_5$  then record the time of beginning the reaction

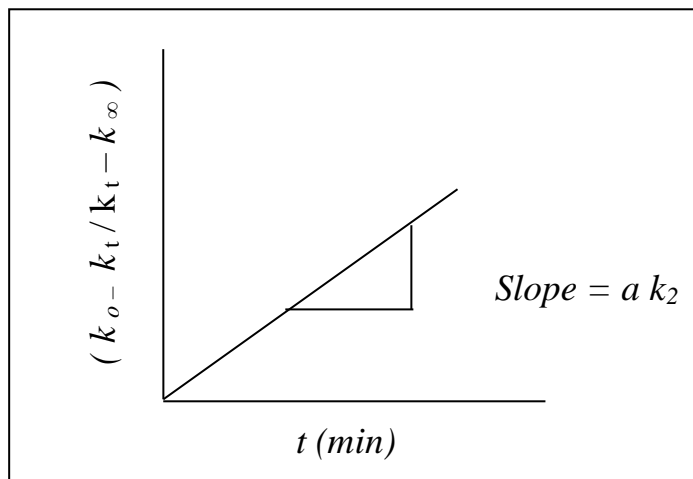
5) Measure the conductivity of the solution after  $(3, 6, 9, 12, 15)$  min to calculate the conductivity at time of the reaction  $(k_t)$

**Calculations:**

1) Arrange the results as in table:

Time (min)	$K_o$	$k_\infty$	$k_t$	$(k_o - k_t)$	$(k_t - k_\infty)$
3					
6					
9					
12					
15					

2) Plot  $(k_o - k_t / k_t - k_\infty)$  against time by using equation (2) from the slop Determine the rate constant where  $(S = a k_2)$ . Use value of  $(a)$  (0.025M) .



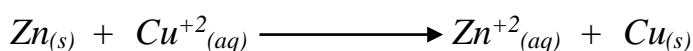
**Discussion:**

- 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 (5): Electrochemistry – Galvanic Cell**

**Introduction:**

*Oxidation – reduction reactions (redox – reactions) can be defined as chemical reactions involving the transfer of electrons from one reactant to another. An example the oxidation zinc atoms (Zn) by copper ion ( $\text{Cu}^{+2}$ ). The reaction is described by the equation:*



*In this reaction, two electrons transfer from Zn to  $\text{Cu}^{+2}$ .*

*In redox reactions two half – reactions occur. Oxidation in this reaction the substance donates electrons therefore its oxidation number increases.*



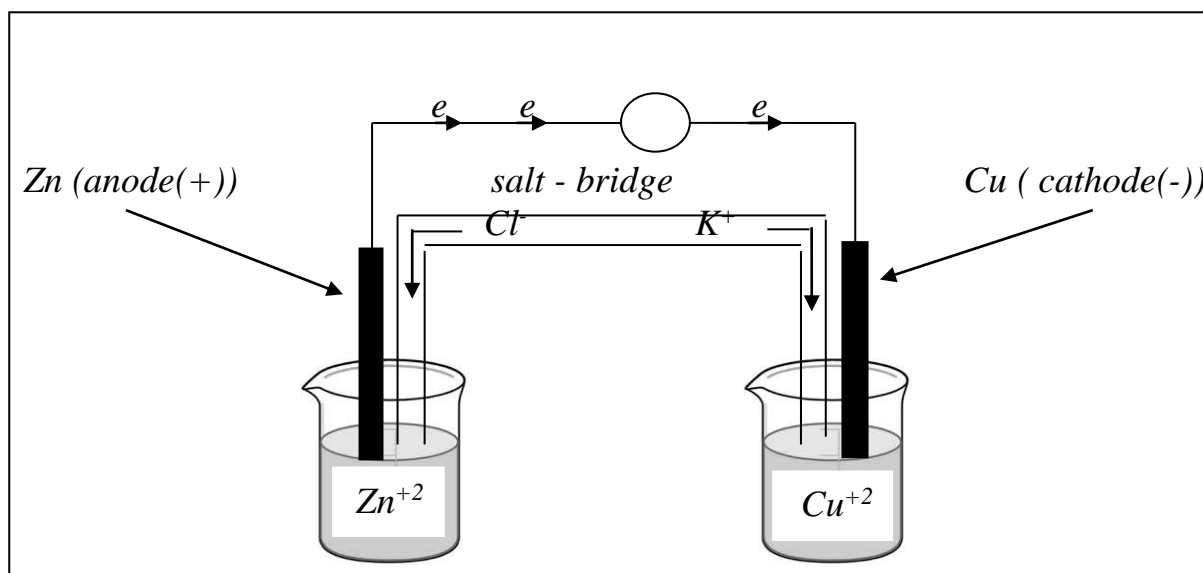
*A substance that under goes oxidation is called (reducing agent) because it causes reduction of another substance. Reduction in this reaction the substance gains electrons therefore its oxidation number decreases. A substance that under goes reduction is called (oxidizing reagent) because it causes oxidation another substance:*



*However, no half – reaction can occur by itself. A redox – reaction results when an oxidation and reduction half – reaction are combined to complete. A measuring of tendency for areduction to occur its reduction potential (E), measured in units of (volts). At standard conditions,  $25\text{C}^0$  and concentration of (1M) for the aqueous ions, the measured voltage of the reduction half – reaction. At standard conditions is defined as the standard reduction potential ( $E^0$ ). For zinc (-0.76V) and for copper (0.34V). the substance that has reduction potential greater than another under goes reduction. In the same time the substance that has reduction ion potential lesser under goes oxidation. For the oxidation half – reaction potential is called (oxidation potential), it is opposite in sign in the (reduction potential).*

**Galvanic cell or voltaic cell:**

A galvanic cell is a device in which a redox reaction spontaneously occurs and produces an electric current. A galvanic cell consists of two conductors (metals) called electrodes, each of which is immersed in an electrolyte solution and connected to an external wire. The purpose of using the galvanic cell is to produce an electric current where the electrons are made to pass through an external electrically conducting wire instead of being directly transferred between the oxidizing and reducing agents. The two reactions oxidation and reduction occur at cell electrodes, the electrode at which oxidation occurs is called (anode) and electrode at which reduction takes place is called (cathode). The electrons move from the anode (+) to the cathode (-) through an external electrically conducting wire producing an electric current. In many cells, the two electrodes are connected by a salt bridge that allows a current of ions from electrode to the other to complete the circuit of electron current in the external wires, in other words the purpose of using salt bridge is to isolate the reactants but maintains electrical contact between the two electrodes where can be used salt like KCl in salt bridge. Conduction of electricity from one electrode solution to the other then occurs by migration of potassium ions in the bridge in one direction and chloride ions in the other.



Galvanic cell consists of two electrode  $Cu//Cu^{+2}$  and  $Zn//Zn^{+2}$

At standard conditions indicated by the superscript (<sup>o</sup>), the standard cell potential ( $E^0_{cell}$ ) is based upon the standard reduction potentials as shown in equation (1):

$$E^0_{cell} = E^0_{red.(cathode)} - E^0_{red.(anode)} \dots\dots\dots(1)$$

When  $E^0_{cell}$  value is positive the reaction in the cell is spontaneous, but when  $E^0_{cell}$  value is negative in the cell the reaction is un spontaneous

**Nernst equation:**

The Nernst equation is equation used to calculate the potential of a galvanic cell when conditions are not standard as example when we used solutions their concentrations does not equal (1):

$$E_{cell} = E^0_{cell} - (RT / n F) \ln Q \dots\dots\dots(2)$$

Where ,  $E_{cell}$  : cell potential is un standard

$R=8.314 \text{ J / mole.K}$

$T$  : is temperature in (Kelvin)

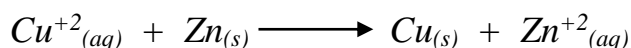
$n$  : is the number of electrons transferred in the redox reaction

$Q$  : is the reaction quotient for the ion products and ion reactants in the cell

Note: the solid electrodes have constant concentrations and so do not appear in  $Q$

$F$ : is Faraday constant its value is 96500 J / V.mole

for example, the cell consists of:  $\text{Cu} / \text{Cu}^{+2}(\text{cathode})$  and  $\text{Zn} / \text{Zn}^{+2} (\text{anode})$



$Q = ([\text{Zn}^{+2}] / [\text{Cu}^{+2}])$  then the Nernst equation becomes :

$$E_{cell} = E^0_{cell} - (R T / n F) \ln ([\text{Zn}^{+2}] / [\text{Cu}^{+2}]) \dots\dots(3)$$

Or we can write it as following:

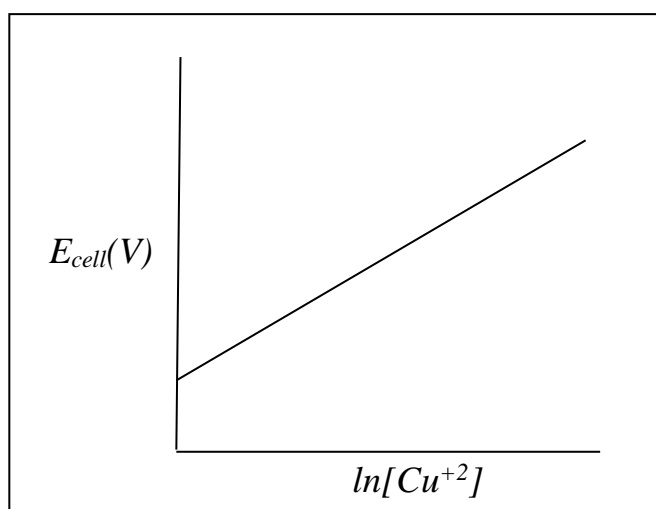
$$E_{cell} = E^0_{cell} - (R T / n F) \ln [Zn^{+2}] + (R T / n F) \ln [Cu^{+2}] \dots(4)$$

Where in cell of  $Cu / Cu^{+2}$ ,  $Zn / Zn^{+2}$

(  $n = 2$  ), the Nernst equation becomes :

$$E_{cell} = E^0_{cell} - (R T / 2 F) \ln [Zn^{+2}] + (R T / 2 F) \ln [Cu^{+2}] \dots(5)$$

From equation(5) when we use constant concentration of  $Zn^{+2}$  while  $[Cu^{+2}]$  is varied at constant temperature. The terms  $E^0_{cell}$  and  $-(R T / 2F) \ln[Zn^{+2}]$  are constant and together they equal the intercept in this case(  $E_{cell}$  ) will vary with  $[Cu^{+2}]$ . A plot of  $E_{cell}$  against  $\ln[Cu^{+2}]$  will result in a straight line:



The slope is constant because it equals  $(R T / 2 F)$  is provided the temperature is constant. The equation of the line in this plot can then be used to determine  $\ln[Cu^{+2}]$  in a solution of unknown concentration from its measured ( $E_{cell}$ ) as long as  $[Zn^{+2}]$  remains constant  $[1M]$  and the temperature is constant ( $25C^0$ ).

### Procedure:

#### **A- Determination of reduction potentials:**

- 1- Prepare solutions (1M) in (50ml) of  $Cu(NO_3)_2$ ,  $Zn(NO_3)_2$ ,  $Fe(NO_3)_2$  and  $NiSO_4 \cdot 9H_2O$ . Use beakers (50ml).
- 2- Prepare standard solution of  $KCl$  or  $KNO_3$ . Wet strip of filter paper with the standard solution, then fold the paper in to U shape, this will serve as your salt – bridge.



- 3- Immerse the each end of bridge in to two solutions in adjacent beakers.
- 4- Polish the electrodes (plates) of Cu and Zn by sand paper and connect the electrodes to voltmeter.
- 5- Immerse the two electrodes in the solution of  $\text{Cu}(\text{NO}_3)_2$  and  $\text{Zn}(\text{NO}_3)_2$
- 6- After (5-10)sec record the standard cell potential. If reading of voltmeter is positive, if not, re polishes and rinses the electrode and reverse the connection and record potential of cell.
- 7- repeat steps (3-6) to find standard cell potentials for cells  $\text{Zn} / \text{Zn}^{+2} // \text{Fe}(\text{NO}_3)_2 / \text{Fe}$  and  $\text{Zn} / \text{Zn}^{+2} // \text{NiSO}_4 / \text{Ni}$ . Use a freshly prepared filter paper for the salt – bridge for each cell.

**Calculations:**

- 1- Use standard reduction potential of electrode  $\text{Zn} / \text{Zn}^{+2}$  with value (-0.76V) and consider it as anode of cells and other electrodes as cathode
- 2- Using the equation :

$$E^0_{\text{cell}} = E^0_{\text{red.}(cathode)} - E^0_{\text{red.}(anode)}$$

To find the standard reduction potential for electrode  $\text{Cu} / \text{Cu}^{+2}$ ,  $\text{Ni} / \text{Ni}^{+2}$  and  $\text{Fe} / \text{Fe}^{+2}$

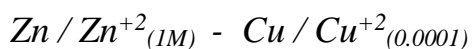
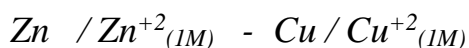
- 3- A range the data in a table as following:

Galvanic Cell	$E^0_{\text{cell}}$	Cathode half-reaction	Anode half-reaction	Reduction potential(V)
Cu - Zn				
Ni - Zn				
Fe - Zn				

**B- Nernst equation for varying  $\text{Cu}^{+2}$  concentration:**

- 1- Prepare three dilute solutions of  $\text{Cu}(\text{NO}_3)_2$  from (1M) stock solution as follows transfer (1ml) of (1M)  $\text{Cu}(\text{NO}_3)_2$  in to (100ml) volumetric flask and dilute to the mark to form (0.01M) solution then prepare ( $10^{-2}$ ,  $10^{-4}$  and  $10^{-6}$ )M
- 2- Prepare four electrodes from  $\text{Cu} / \text{Cu}^{+2}$  using the prepared solution in step(1)

3- A range four galvanic cells as follows:



In the same way is part (A)

4- Measure potential of the cells by using voltmeter, measure  $E_{\text{cell}}$  for the unknown  $\text{Cu}^{+2}$  solution . record the measured values in a table as shown bellow.

$[\text{Cu}^{+2}]M$	$10^{-6}$	$10^{-4}$	$10^{-2}$	1	unknown
$E_{\text{cell}}$					

**Calculations:**

1- using Nernst equation (5) , plot  $E_{\text{cell}}$  versus  $\ln[\text{Cu}^{+2}]$

2- from the equation for the line , determine  $[\text{Cu}^{+2}]$  of the unknown solution.

**C- Determination the  $E^0$  of a voltaic cell using Cu and unknown metal:**

1- Prepare (1M) of the unknown electrolyte solution and use the corresponding metal strip (X) in order that you get an electrode of  $X / X_{(1M)}$  ion

2- Use the standard electrode of  $\text{Cu} / \text{Cu}(\text{NO}_3)_2(1M)$  that have been prepared in part (A)

3- Make a new salt bridge by soaking a short length of filter paper in  $\text{KNO}_3$  solution.

4- Connect the two electrodes of X and Cu with the filter paper.

5- Measure the positive potential of the X – Cu voltaic cell using the same technique as in part (A). Record the potential of the cell after (5-10)sec then remove both electrodes from the solution and clean each electrode, set up the galvanic cell again and record the potential again. If the two measured potentials do not agree with measuring the potential of the cell a third time but after removing electrodes from the solution and clean them, then calculate the average of measured potentials

6- Using the equation:

$$E^0_{\text{cell}} = E^0_{\text{red.}(cathode)} - E^0_{\text{red.}(anode)}$$

To find  $E^0_{\text{red.}}$  Of the unknown metal.

**Discussion:**

1- You may find that the voltage record is less than what you would expect based on standard reduction potentials for example Zn – Cu cell may be less than (1.1V).

2- We must separate the solution of the electrodes of the cell, why?

3- Explain the part of salt – bridge?

**Experiment (6): Potentiometric titration of strong acid with strong base**

**Introduction:**

*It's the process through it can determine the equivalent point by using the measurements of pH. The neutralization or equivalence point can be determined by optical indicators.*

*Types of electrodes:*

- 1) Glass electrode*
- 2) Hydrogen electrode*
- 3) Quinhydrone electrode*
- 4) Antimony electrode*

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

**Procedure:**

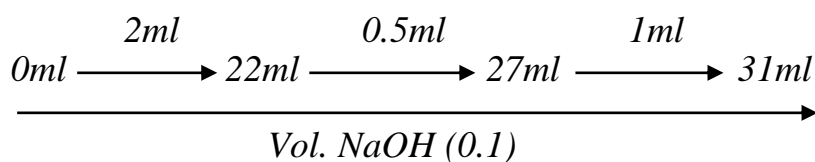
***A- Titration by using indicators:***

*Take (25ml) of HCl (0.1M) and put it in the beaker with capacity (100ml) then add two drops of indicator (phenolphthalein), finally determine the volume of base when the color of the solution changed to pink, which represents the volume at equivalent point.*

***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:  $\text{pH} = 4$ ,  $\text{pH} = 7$*
- 2- wash the electrode again and dry it by filter paper then immerse it the beaker with capacity (250ml) which contain (25ml) of HCl(0.1M) and record the pH value.*
- 3- begin the titration by adding the base of sodium hydroxide (0.1M) that is in burette, according to the following steps:*
  - a- At first, add(2ml) at time, towards the equivalence point, then reduce the amounts (about 0.5 ml) at the end.*

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



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:**

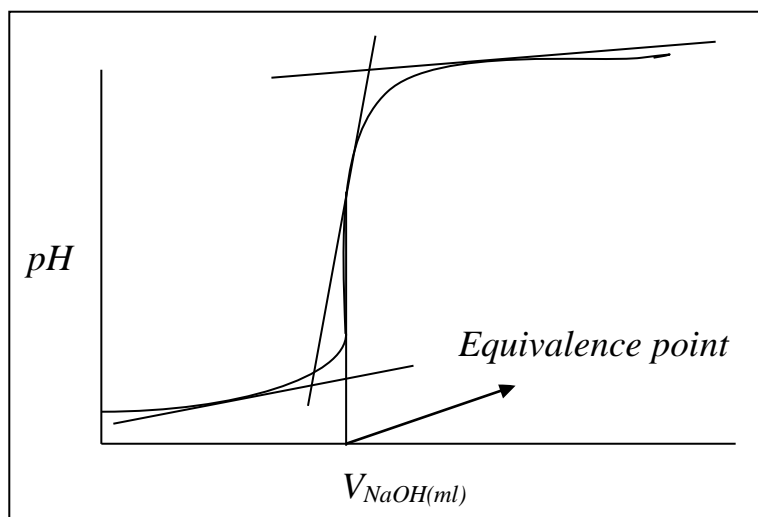
1- Set table including the results of above method like table (1):

2- Plot a graph between the values of pH and volumes of NaOH like shape (A)

3- Determine the equivalence point

$V_{\text{NaOH}} \text{ (ml)}$	pH
0	
2	
4	
-----	
-----	
-----	
31	

Table (1)



Shape (A)

**Discussion:**

1- Why is HCl consider a strong acid?

2- Class types of electrolytes in terms of equivalence?

**Experiment (7): Application of (Lambert – Beer) law**

**Introduction:**

Spectrophotometric methods include measuring the intensity of incident light and transmitted light at a specific wavelength. There are two types of devices are the 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.

$$A = \epsilon C l$$

$$\%T = (I / I_0) * 100$$

$$A = \log I_0 / I \\ = \log 1 / T$$

$A$  = absorbance

$\epsilon$  = molar absorbance coefficient ( $\text{mol}^{-1} \cdot \text{L} \cdot \text{cm}^{-1}$ )

$l$  = long the light path inside the solution (cm)

$T$  = transmittance

$C$  = concentration ( $\text{mol} \cdot \text{L}^{-1}$ )

$I$  = the intensity of transmitted light

$I_0$  = intensity of incident light

Wave length( $\lambda$ ): The distance between adjacent peaks of a wave packet, that has the standard unit:  $\text{nm} = 10^{-9} \text{m}$ ,  $\text{Å} = 10^{-10} \text{m}$ ,  $\mu\text{m} = 10^{-6} \text{m}$

**The greatest wave length ( $\lambda_{\text{max}}$ ):** The wave length which has the highest absorption of the substance.

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

**Photon:** Is units of energy

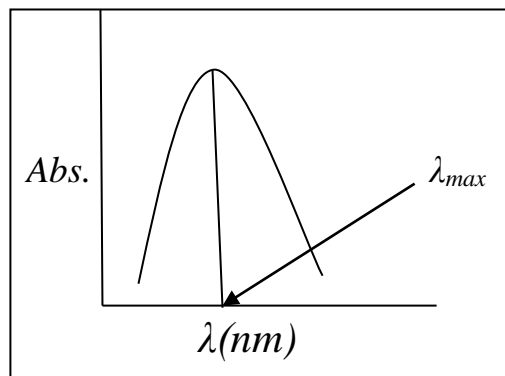
**Measuring the absorbance of potassium permanganate ( $\text{KMnO}_4$ ):**

**A- Finding ( $\lambda_{\text{max}}$ ):**

1- prepare dilute solutions of  $\text{KMnO}_4(0.1\text{M})$ : ( 1 , 3 , 5 , 7 )  $\times 10^{-4}\text{M}$  using volumetric flask (50ml) according to law (  $N_1 V_1 = N_2 V_2$  )

2- measure absorbance of the lower concentration ( $1 \times 10^{-4}$ )M versus water (blank) in the wave length range (400 – 600 )nm and that each (10nm) to note the highest absorption value read by the device and install it to be ( $\lambda_{max}$ ).

Wave length(nm)	Absorbance
400	
410	
-----	
-----	
600	

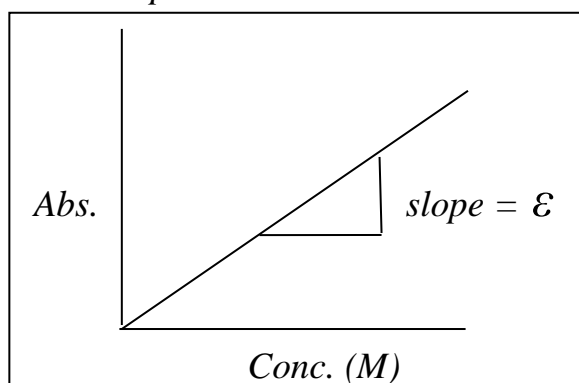


**B- Finding ( $\epsilon$ ):**

1- When you install ( $\lambda_{max}$ ) from step (A) was appointed absorbance of each solution.

2- Plot a relationship between absorbance (A) and concentration, determine the molar absorption coefficient ( $\epsilon$ ) from the slop.

Conc.(mol/L)	Absorbance
$1 \times 10^{-4}$	
$3 \times 10^{-4}$	
$5 \times 10^{-4}$	
$7 \times 10^{-4}$	



**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?