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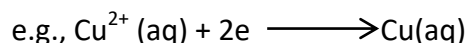
SENIOR SIX TERM 2

TOPIC 1/2: Electrochemistry

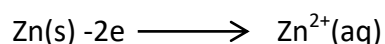
Topic competency: The learner analyses redox reactions, electrochemical processes and electrochemical cells, applies the principles of electrode potentials to predict reaction spontaneity and electrolysis outcomes.

Redox Equations

Redox reaction is a reaction in which both reduction and oxidation take place simultaneously. A reduction reaction is where there is addition of electrons to an ion species;



whereas, oxidation is loss of electrons from a compound/ion species .



The species which accepts electrons such as Cu^{2+} is reduced but is referred to as an oxidizing agent.

The species which donates electrons such as Zn^{2+} is oxidized but is referred to as a reducing agent

Oxidation Number

An oxidation number is a hypothetical charge assigned to an atom in a chemical compound to track electron movement; an increase in this number indicates oxidation (loss of electrons), and a decrease indicates reduction (gain of electrons). These numbers can be positive, negative, or zero, and are used to determine the oxidation states of individual elements in a reaction. For a neutral compound, the sum of the oxidation numbers is zero, while for an ion, it equals the ion's charge.

Rules governing oxidation numbers

- (i) **Free elements:** The oxidation number is always zero (e.g., O₂, Fe)
- (ii) **Monatomic ions:** The oxidation number is equal to the charge of the ion (e.g., Na⁺) has oxidation number +1
- (iii) **Group 1 elements:** Alkali metals (like Na and K) are always +1 in a compound.
- (iv) **Group 2 elements:** Alkaline earth metals (like Ca and Mg) are always +2 in a compound.
- (v) **Hydrogen:** It is +1 when bonded to a nonmetal but -1 when bonded to a metal (in metal hydrides).
- (vi) **Oxygen:** It is typically -2, except in peroxides (like H₂O₂), where it is -1, or when bonded to fluorine, where it can be positive.
- (vii) **Fluorine:** It is always -1 in a compound.
- (viii) **Overall charge:** The sum of all oxidation numbers in a neutral compound equals zero. The sum of all oxidation numbers in a polyatomic ion equals the ion's charge.

Calculating oxidation number of ions of an element in complex ions

Unless otherwise stated or implied in these calculations, the oxidation number of a metal is equal to group number, e.g., Na is 1, Mg is 2, Al is 3. Oxidation number of O is -2 and that of any group 7 element is -1.

Example 1

Calculate the oxidation number of manganese ion in KMnO₄.

Here the oxidation number of K = 1 and that of oxygen is implied to be -2. If the oxidation number of Mn is X

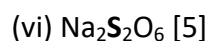
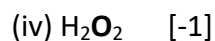
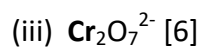
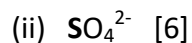
Then, $1 + X + -2 \times 4 = 0$ (overall charge on the molecule)

$$X = 7$$

Trial 1

Calculate the oxidation numbers of the bolded atoms in molecules or atoms below

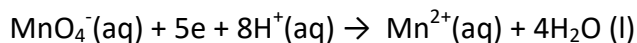
1. (i) **K**IO₃ [5]



Half equations

These are balanced based on the number of electrons lost or accepted by a central atom.

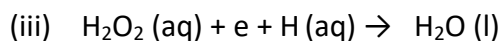
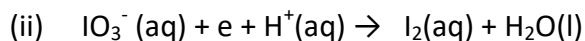
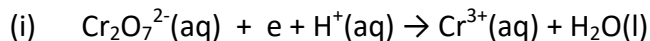
Example, equation for reduction of manganate (VII) in acidic medium is written as follows:



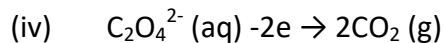
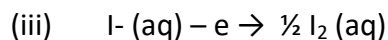
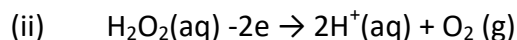
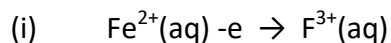
Five electrons are added because the oxidation number of Mn is reduced from 7 to 2.

Trial 2

Balance the following reduction equations



The following are common oxidation half equation reactions:

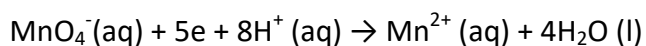


Overall equation

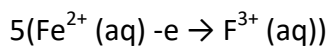
These are constructed from half equation in such a way that the number electron received by an oxidizing agent is equal to the number of electron lost from a reducing agent.

For example a reaction between MnO_4^- and Fe^{2+} in acidic medium, the half equation of Fe^{2+} is multiplied by 5 to provide the 5 electron necessary to reduce MnO_4^- .

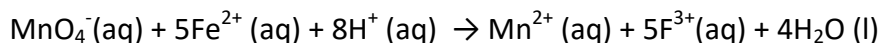
That is,



Plus



Equals



Trial 3

1. Write overall equations between the following pairs of ions or compounds, assume acidic medium where necessary
 - (a) MnO_4^- and I⁻
 - (b) MnO_4^- and H_2O_2
 - (c) KIO_3 and KI
 - (d) MnO_4^- and $\text{C}_2\text{O}_4^{2-}$
 - (e) $\text{K}_2\text{Cr}_2\text{O}_7$ and FeSO_4
2. 0.9875g of an impure potassium manganate (VII) was dissolved in water to make 250cm³ of solution. When 20.0cm³ of this solution was acidified with dilute sulphuric acid, warmed and

titrated against sodium ethanedioate (oxalate) solution, made by dissolving 1.675g of anhydrous sodium ethanedioate to make 250 cm³, 24.40cm³ of sodium ethanedioate solution was used. (Na₂C₂O₄ =134 and KMnO₄ = 158).

Calculate percentage purity of potassium permanganate (VII) [97.6%]

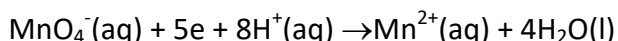
3. 10cm³ of concentrated hydrogen peroxide was diluted to 250cm³. 20cm³ of the dilute solution required 28.30cm³ of 0.02M potassium manganate (VII) solution for complete reaction. Find the volume strength of the concentrated hydrogen peroxide solution. [1.77M]
4. Potassium permanganate is commonly used as a reagent in volumetric analysis. However it is not a primary standard.

(a) State two disadvantage of using potassium manganate (VII) as a reagent in volumetric analysis.

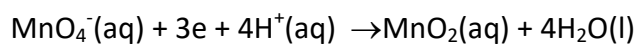
(b) Give two reasons why potassium manganate (VII) is not used as a primary standard

(c) Write a half equation for the reduction of manganate (VII) ions in

(i) acid solution



(ii) alkaline solution



5. In volumetric estimation of reducing agent, potassium dichromate (VI) is preferred to potassium manganate (VII) as an oxidant

(a) Explain why potassium dichromate (VI) is preferred as an oxidant. (2 marks)

(b) 3.8g of solder containing tin was dissolved in excess hydrochloric acid. The solution was made up to 250cm³. 25.0cm³ of this solution required 23.5cm³ of a 0.01M potassium dichromate (VI) solution for complete reaction.

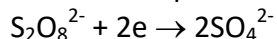
Calculate the percentage, by mass, of tin in the solder. [22%]

(c) Explain why chlorine is more soluble in dilute sodium hydroxide than in water (4 marks)

6. (a) Determine the oxidation number of

- i. Nitrogen in HNO_3 [5]
- ii. Sulphur in S_2O_8 (1marks) [8]

(b) Write half equation for the conversion of $\text{S}_2\text{O}_8^{2-}$ to SO_4^{2-} ions



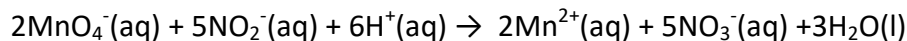
(c) Complete the following and balance the equations

- i. $\text{S}_2\text{O}_8^{2-} + \text{I}^- \rightarrow$
- ii. $\text{Sn}^{2+}(\text{aq}) + \text{S}_2\text{O}_8^{2-}(\text{aq}) \rightarrow$

7. Iron (II) sulphate is normally used to standardize a solution of potassium manganate (VII) acidified with sulphuric acid.

- (i) write equation for the reaction between potassium manganate (VII) and iron (II) sulphate
- (ii) State why hydrochloric acid is not used to acidify potassium manganate (VII) solution
 Cl^- are oxidized by MnO_4^- as well

(b) 25 cm^3 of acidified solution of 0.02M potassium manganate (II) reacted exactly with 24.95 cm^3 of sodium nitrite. Potassium manganate (VII) reacts with sodium nitrite according to the following equation



Calculate the concentration of sodium nitrite in moles per litre. [0.05M]

Application of redox reactions in Biological systems

- (i) **Cellular respiration:** This process uses a series of redox reactions to convert chemical energy stored in glucose into adenosine triphosphate (ATP), the energy currency of the cell.
- (ii) **Photosynthesis:** The complementary process to respiration in plants also relies on redox reactions to convert light energy into chemical energy.
- (iii) **Metabolic pathways:** Redox reactions are the backbone of many metabolic pathways, where enzymes control the transfer of electrons in controlled steps.
- (iv) **Oxidative stress:** An imbalance between pro-oxidants and antioxidants can lead to oxidative stress, which is linked to various diseases.

Application of redox reactions in Industrial systems

(i) Energy production:

- **Batteries and fuel cells:** These devices generate electricity through controlled redox reactions, where electrons are transferred from one chemical to another.
- **Metal extraction:** Redox reactions are used to extract pure metals from their ores.

(ii) Corrosion: This is an unwanted redox process where metals oxidize, leading to degradation.

- **Rusting of iron:** A common example where iron is oxidized by oxygen in the presence of water.
- **Prevention:** Industries use techniques like cathodic protection, which makes the metal the cathode in an electrochemical cell, to prevent it from oxidizing.

(iii) Chemical manufacturing: Redox reactions are used in the synthesis of numerous chemicals, such as the production of nitric acid from ammonia or chlorine from sodium chloride.

(iv) Environmental remediation: Redox reactions can be used to treat pollutants, such as reducing toxic hexavalent chromium to a less harmful form.

Electro-chemical conductivity

Resistance R of electrolytic solution

This is the drag force experienced by ions as they migrate towards the respective electrodes. Resistance of an electrolyte is proportional to the length l of an electrolyte between electrodes and inversely proportional to the square root of the cross section area A , of each electrode.

$$\text{Thus, } R \propto \frac{l}{A}$$

$$\text{Or } R = \rho \frac{l}{A}$$

ρ is called resistivity of an electrolyte and is the resistance of an electrolyte placed between electrodes each 1cm^2 in area and 1cm apart.

In chemistry reciprocal terms are preferred.

$$\text{That is, conductance, } \frac{1}{R} = \frac{1}{\rho \frac{l}{A}} = \frac{1}{\rho} \times \frac{A}{l}$$

The reciprocal of resistivity, $\frac{1}{\rho}$ is called conductivity, K .

It implies that conductance, $\frac{1}{R} = K \times \frac{A}{l}$

And $K = \frac{1}{R} \times \frac{l}{A} \Omega^{-1}\text{cm}^{-1}$

Factors that affect the conductivity of electrolytes

- 1. Concentration of Ions in Solution:** The higher the concentration of ions in solution, the higher its conductivity due to high number of conducting ions.
- 2. Weak/Strong Electrolyte:** Strong electrolyte have higher conductivity than weak electrolytes because they are completely ionized in water to produce many ions
- 3. Temperature:** The solubility of the solute increases with the increase in the temperature. Therefore, the conductivity of the electrolyte increases. Secondly temperature increase the speed of ions in solution

Molar conductivity, Λ_c

This is the conductivity of a solution that contains 1 mole of an electrolyte.

$$\Lambda_c = \frac{K}{C} \Omega^{-1}\text{cm}^2\text{mol}^{-1}$$

where C = moles of electrolyte in the solution between electrodes.

Example 2

The resistance of 0.1M KCl solution placed between electrodes each 1cm^2 in area and 0.5 cm apart is 400Ω . Calculate the molar conductivity of the solution.

$$K = \frac{1}{R} \times \frac{l}{A} = \frac{1}{400} \times \frac{0.5}{1} = 0.00125 \Omega^{-1}\text{cm}^{-1}$$

Volume of electrolyte between the electrodes = $1 \times 0.5 = 0.5\text{cm}^3$

Moles of the electrolyte between the electrodes

1000cm^3 contain 0.1mole

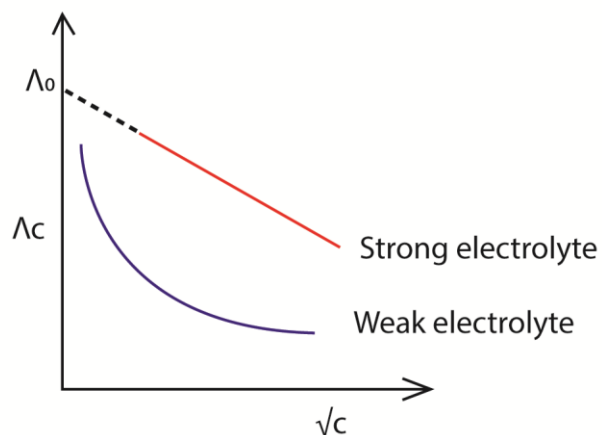
$\therefore 0.5\text{cm}^3$ contain $\frac{0.5 \times 0.1}{1000} = 0.00005\text{moles}$

$$\Lambda_c = \frac{K}{C} = \frac{0.00125}{0.00005} = 25 \Omega^{-1}\text{cm}^2\text{mol}^{-1}$$

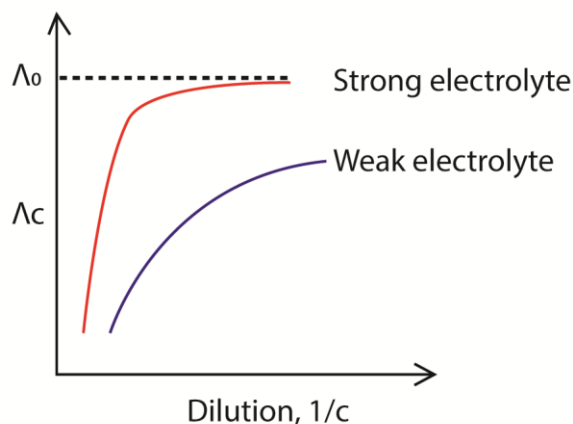
Variation of molar conductivity with concentration

Molar conductivity for both strong and weak electrolyte decrease as the concentration increase as shown by the graphs below

(a) A graph of molar conductivity, Λ_0 against root concentration, \sqrt{c}



(b) A graph of molar conductivity, Λ_0 against dilution, $1/c$



Explanation

(a) Strong electrolyte

- (i) **Strong electrolyte** are completely ionized in solution;
- (ii) at high concentration the solution contains high density of ions
- (iii) high density of ions lead ion interaction which reduce mobility of ions which lowers molar conductivity.
- (iv) On dilution the tendency to form these interactions reduce lowering the drag force thus molar conductivity increases.

(b) Weak electrolyte

- (i) Weak electrolytes are poorly ionized in solution which provides low concentration conducting ions which leads to low molar conductivity compared to strong electrolytes
- (ii) At high concentration, the percentage ionization of weak electrolytes is very low leading to very few conducting ions per mole of electrolytes and low molar conductivity.
- (iii) Percentage ionization of weak electrolytes increases with dilution leading to increase in the number of ions per mole of an electrolyte and molar conductivity.

Molar conductivity at infinite dilution, Λ_0

This is the conductivity of 1 mole of electrolyte when the solution is very dilute that the ions experience no interaction from other ions.

For **strong electrolyte**, molar conductivity at infinite dilution is obtained by extra plotting of molar conductivity against root concentration to zero concentration.

For weak electrolyte, molar conductivity at infinite dilution is obtained by application of Kohlrausch's law of independent migration of ions.

It states: "the molar conductivity of an electrolyte at infinite dilution is the sum of molar conductivity of the constituent ions at infinite dilution.

For example

$$\Lambda_0 \text{NaCl} = \lambda_0 \text{Na}^+ + \lambda_0 \text{Cl}^-$$

$$\Lambda_0 \text{AB} = \Lambda_0 \text{AC} + \Lambda_0 \text{BD} - \Lambda_0 \text{CD}$$

Example 3

The molar conductivity of nitric acid, potassium nitrate and potassium fluoride are 421, 145 and 129 $\Omega^{-1}\text{cm}^2\text{mol}^{-1}$ respectively at infinite dilution. Molar conductivity of hydrofluoric acid at infinite dilution.

$$\begin{aligned}\Lambda_0 \text{HF} &= \Lambda_0 \text{HNO}_3 + \Lambda_0 \text{KF} - \Lambda_0 \text{KNO}_3 \\ &= 421 + 129 - 145 \\ &= 405 \Omega^{-1}\text{cm}^2\text{mol}^{-1}\end{aligned}$$

Factors affecting conductivity of an ion at infinite dilution

1. **Charge on the ion**: the bigger the charge the higher the conductivity because the ion carries bigger charge per ion
2. **The size of an ion**: small ions have high speed of movement leading to high conductivity. However, small ions with high charge have high density that may attract a big cloud of water of hydration that its effective mass may be bigger than that of a big ion. That is why the small cations may have lower conductivity than big cation.

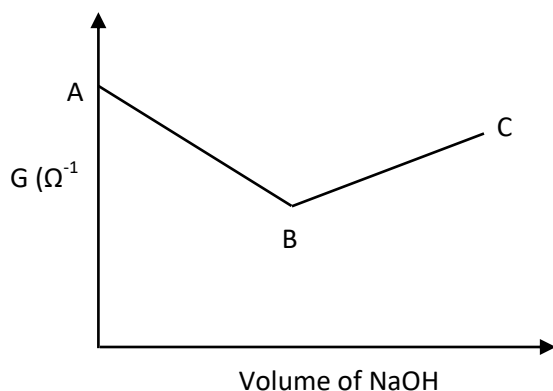
Example of conductivities of common ions at infinite dilution

ion	λ_0 ($\Omega^{-1}\text{cm}^2\text{mol}^{-1}$)
Na^+	50.1
OH^-	198.6
H^+	349.8
Cl^-	76.4

Conductometric titrations

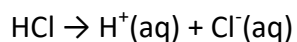
These are titrations in which ions are replaced by others of different conductivity, the titration followed by conductivity measurement.

(a) Titration of a strong acid with strong base

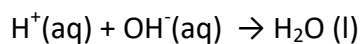


Explanation

Initially at A, conductivity is high due to presence of highly conducting H^+ from ionization of hydrogen chloride

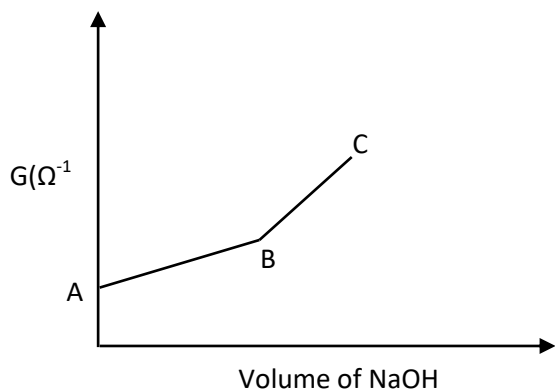


Along AB conductivity decreases up to the end point due to removal of highly conducting hydrogen ions.



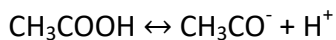
Along BC after the end point, conductivity increases due to excess OH⁻ ions that are relatively highly conducting.

(b) Titration of a weak acid e.g. CH₃COOH with strong base (e.g. NaOH)

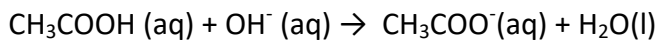


Explanation

Initially at A conductivity is low due to small concentration of H⁺ ions since the acid is partially ionized.

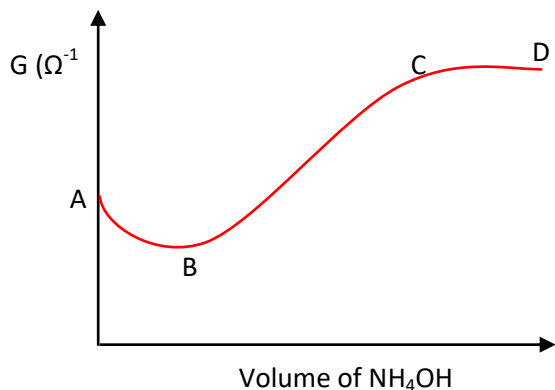


Along AB conductivity increases up to end point due to addition of salt ions and further ionization of the acid due to dilution.

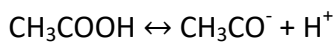


Along BC conductivity increase to due to excess OH⁻.

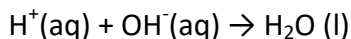
(c) Titration of a weak acid (e.g. CH₃COOH) with weak base (e.g. NH₄OH)



Initially at A, conductivity is high due to presence of H^+ from partial ionization of the acid



Along AB conductivity decreases due to removal of H^+ .



Along BC conductivity increases due to addition of salt ion



Along CD conductivity remains almost constant due to attainment of equilibrium.

Trial 4

(a) Define the term

(i) Molar conductivity (1mark)

(ii) Electrolytic conductivity (1mark)

(b) The table below shows molar conductivities, Λ , of potassium chloride and ethanoic acid at different concentrations.

(concentration) ^{1/2} (mol dm ⁻³) ^{1/2}	0.01	0.02	0.03	0.04	0.074	0.106
Λ , for KCl ($\Omega^{-1}cm^2mol^{-1}$)	147.0	145.0	142.0	140.0	132.0	128.0
Λ , for ethanoic acid ($\Omega^{-1}cm^2mol^{-1}$)	70.0	49.0	33.0	24.0	10.0	6.0

(i) On the same axes draw graphs of molar conductivity against square root of concentration, for both potassium chloride and ethanoic acid. (3marks)

(ii) State the shapes of the graphs in (b)(i) and explain your answer. (5marks)

(c) The electrolytic conductivity of a solution containing 6.66g of ethanoic acid in one litre is $5.21 \times 10^{-2} \Omega^{-1}m^{-1}$ at 25°C.

Calculate the dissociation constant K_a of ethanoic acid at 25°C.

(the molar conductivities of hydrochloric acid, sodium chloride and sodium ethanoate are 426×10^{-4} , 126×10^{-4} and $91 \times 10^{-4} \Omega^{-1}m^2mol^{-1}$ respectively at infinite dilution). (06marks)

(d) To a solution of hydrochloric acid was added small amounts of sodium hydroxide solution at a time until in excess. State how the conductivity of the solution varied and explain your answer. (4marks)

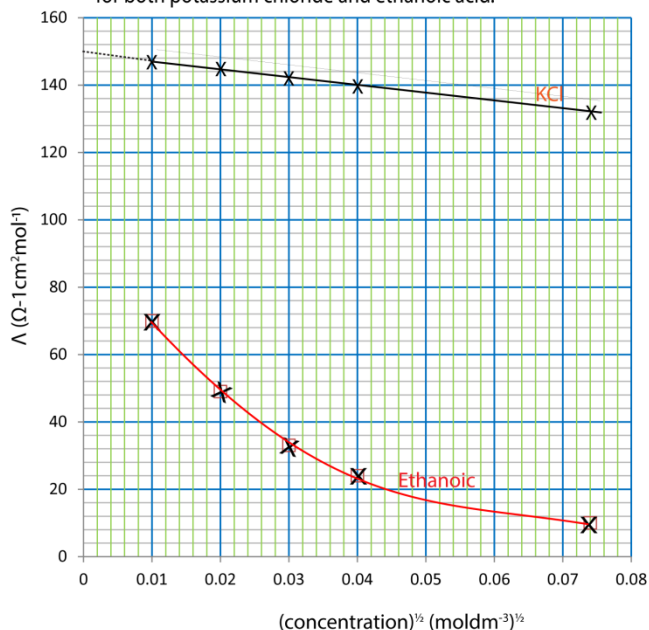
Solution

(a) (i) Molar conductivity is a conductivity of solution that contains 1 mole of electrolyte.

(ii) Electrolytic conductivity is the reciprocal of resistance of a solution between electrodes each 1cm^2 in area and 1cm apart.

(b) (i)

Molar conductivity against square root of concentration, for both potassium chloride and ethanoic acid.



KCl is Strong electrolyte

- **It is** are completely ionized in solution;
- at **high** concentration the solution contains high density of K^+ and Cl^- ions
- high density of ions lead ion interaction that increases the drag force of ions which lowers molar conductivity.
- On dilution the tendency to form these interactions reduce lowering the drag force thus molar **conductivity** increases.

Ethanoic acid is weak electrolyte

- It is poorly ionized in solution which provides low concentration conducting CH_3COO^- and H^+ ions which leads to low molar conductivity compared to strong electrolytes
- At high concentration, the percentage ionization of ethanoic is very low leading to very few conducting ions per mole of electrolytes and low molar conductivity.
- Percentage ionization of ethanoic acid increases with dilution leading to increase in the number of ions per mole of an electrolyte and molar conductivity.

(c) Formula mass of ethanoic acid, $\text{CH}_3\text{COOH} = 60$

$$\text{Moles of ethanoic acid per dm}^3 = \frac{6.66}{60} = 0.111 \text{ mol dm}^{-3} \equiv 111 \text{ mol m}^{-3}.$$

$$\text{Molar conductivity, } \Lambda_c = \frac{K}{C} = \frac{5.21 \times 10^{-2}}{111} = 0.000469 \text{ } \Omega^{-1} \text{m}^2 \text{mol}^{-1}$$

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

$$= 429 \times 10^{-4} + 91 \times 10^{-4} - 126 \times 10^{-4}$$

$$= 391 \times 10^{-4} \text{ } \Omega^{-1} \text{m}^2 \text{mol}^{-1}$$

$$\text{The degree of ionization, } \alpha = \frac{\Lambda_c}{\Lambda_0} = \frac{0.000469}{391 \times 10^{-4}} = 0.012$$

$$\text{Acid constant, } K_a = \alpha^2 C = (0.012)^2 \times 0.111 = 1.6 \times 10^{-5} \text{ mol dm}^{-3}$$

- (d) Conductivity initially decreases with increasing volume of sodium hydroxide due to removal of highly conducting hydrogen ions up to the end point. It then increases due to excess OH⁻ ions.

Trial 5

- (a) State **two** factors that can affect the molar conductivity of an electrolyte.
 (b) The molar conductivity of potassium hydroxide at various concentrations are given in the table below.

Molar conductivity ($\Omega^{-1} \text{cm}^2 \text{mol}^{-1}$)	238	230	224	217	210
Concentration of potassium hydroxide (mol/dm ³)	0.01	0.04	0.09	0.16	0.25
Square root of concentration of potassium hydroxide (mol/dm ³) ^½					

- (i) Complete the table above. (1mark)
 (ii) Draw a graph of molar conductivity against the square root of the concentration of potassium hydroxide. (3marks)
 (iii) Explain the shape of the graph you have drawn in (b)(ii).
 (iv) Determine the value of molar conductivity of potassium hydroxide at infinite dilution

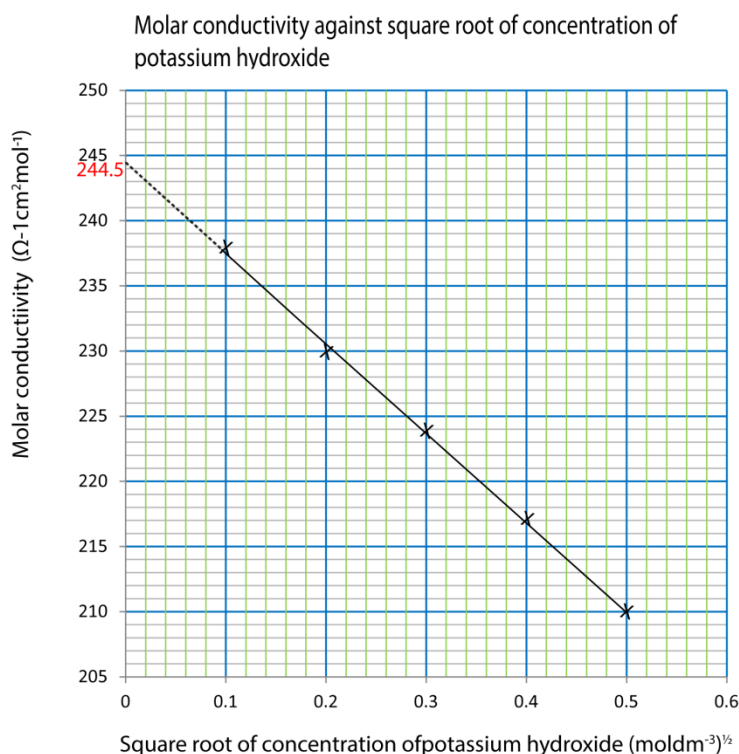
Solution

- (a) – concentration and temperature

(b) (i)

Molar conductivity ($\Omega^{-1}\text{cm}^2\text{mol}^{-1}$)	238	230	224	217	210
Concentration of potassium hydroxide (mol/dm^3)	0.01	0.04	0.09	0.16	0.25
Square root of concentration of potassium hydroxide (mol/dm^3) ^{1/2}	0.1	0.2	0.3	0.4	0.5

(ii)



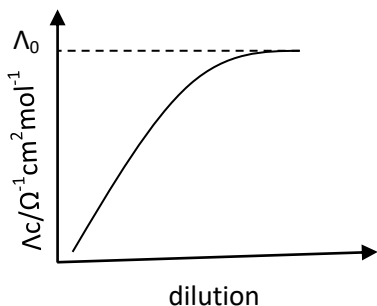
(iii) **KOH is Strong electrolyte:** molar conductivity is decreases as concentration increase because

- **It is** are completely ionized in solution;
- at **high** concentration the solution contains high density of K^+ and Cl^- ions
- high density of ions lead ion interaction that increases the drag force of ions which lowers molar conductivity.
- On dilution the tendency to form these interactions reduce lowering the drag force thus molar **conductivity** increases.

(i) $244.5 \Omega^{-1}\text{cm}^2\text{mol}^{-1}$

Trial 6

- (a) State three factors that can affect molar conductivity of electrolytes (1 ½ marks)
- (b) The graph below shows the variation of molar conductivity of a strong electrolyte with dilution



Briefly explain the shape of the graph

- (c) The molar conductivity of nitric acid, potassium nitrate and potassium fluoride are 421, 145 and $129 \Omega^{-1}\text{cm}^2\text{mol}^{-1}$ respectively at infinite dilution
- (i) Molar conductivity of hydrofluoric acid at infinite dilution (02marks)
- (ii) Dissociation constant, K_a , of a 0.1M hydrofluoric acid solution. (The electrolytic conductivity of hydrofluoric acid is $3.15 \times 10^{-5} \Omega^{-1}\text{cm}^{-1}$)

Solution

- concentration
- temperature
- whether weak or strong electrolyte

(b) molar conductivity increases with dilution because opposite ions get far apart which reduces ionic interference.

$$\begin{aligned} \text{(c) (i) } \Lambda_0(\text{HF}) &= \Lambda_0(\text{HNO}_3) + \Lambda_0(\text{KF}) - \Lambda_0(\text{KNO}_3) \\ &= 421 + 129 - 145 \\ &= 405 \Omega^{-1}\text{cm}^2\text{mol}^{-1} \end{aligned}$$

$$\text{(ii) Molar conductivity, } \Lambda_c = \frac{K}{C} = \frac{3.15 \times 10^{-5}}{0.1 \times 10^{-3}} = 0.001 \Omega^{-1}\text{m}^2\text{mol}^{-1}$$

$$\text{Acid constant, } K_a = \alpha^2 C = (0.001)^2 \times 0.1 = 1 \times 10^{-7} \text{ moldm}^{-3}.$$

Trial 7

(a) Define 'electrolytic conductivity'.

(b) The molar ionic conductivity at infinite dilution of some ionic species are shown below

ion	λ_0 ($\Omega^{-1}\text{cm}^2\text{mol}^{-1}$)
Na^+	50.1
OH^-	198.6
H^+	349.8
Cl^-	76.4

Calculate the electrolytic conductivities for:

(i) 0.01M sodium hydroxide solution

(ii) A solution made by mixing 50cm^3 of 0.01M sodium hydroxide and 50cm^3 of 0.02M hydrochloric acid

Solution

(a) Electro conductivity is the reciprocal of resistance of an electrolyte placed between electrodes each 1cm^2 and 1cm apart.

$$\begin{aligned} \text{(b)(i) Electrolytic conductivity} &= C(\lambda_0\text{Na}^+ + \lambda_0\text{Cl}^-) \\ &= 0.01(50.1 + 76.1) \\ &= 1.265 \Omega^{-1}\text{cm}^2\text{mol}^{-1} \end{aligned}$$

$$\begin{aligned} \text{(ii) Moles of sodium hydroxide} &= \frac{50 \times 0.01}{1000} = 0.0005 \text{ moles} \\ \text{Moles of hydrochloric acid} &= \frac{50 \times 0.02}{1000} = 0.001 \text{ moles} \\ \text{Moles of unreacted hydrochloric acid} &= 0.001 - 0.0005 = 0.0005 \text{ moles} \\ \text{Moles of sodium chloride formed} &= \text{moles of sodium hydroxide that reacted} \\ &= 0.0005 \text{ moles} \\ \text{Electrolytic conductivity} &= C(\lambda_0\text{Na}^+ + \lambda_0\text{Cl}^-) + C(\lambda_0\text{H}^+ + \lambda_0\text{Cl}^-) \\ &= 0.0005(50.1 + 76.1) + 0.0005(349.8 + 76.1) \\ &= 0.06235 + 0.2122 = 0.27455 \Omega^{-1}\text{cm}^2\text{mol}^{-1} \end{aligned}$$

Trial 8

- (a) The molar conductivity of sodium hydroxide solutions of different concentrations are shown in the table below:

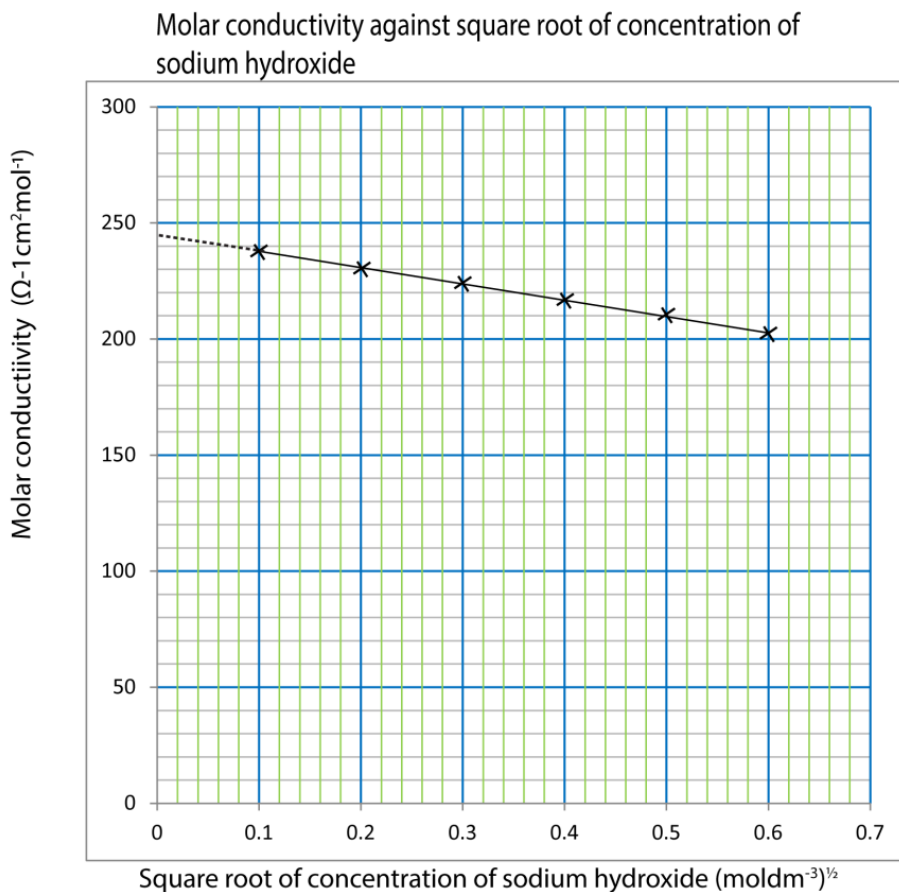
Concentration/ mol dm^{-3}	0.01	0.04	0.09	0.16	0.25	0.36
Molar conductivity ($\Omega^{-1}\text{cm}^2\text{mol}^{-1}$)	238	230	224	217	210	202

- (i) Draw a graph of molar conductivity against the square root of concentration (4marks)
- (ii) Explain the shape of the graph (4 ½ marks)
- (iii) Determine the value of molar conductivity at infinite dilution of sodium hydroxide and indicate its units (2marks)
- (b) Using the same conductivity cell, the resistance of 0.1M potassium chloride and 0.1M bromoethanoic acid solution were found to be 24.96 and 66.50 ohms respectively at 25°C when determined using the same conductivity cell. [the conductivity of potassium chloride at 25°C is $0.01164 \Omega^{-1}\text{cm}^{-1}$ and the molar conductivity of bromoethanoic acid at infinite dilution is $389 \Omega^{-1}\text{cm}^2\text{mol}^{-1}$]
- (i) Calculate the cell constant (2marks)
- (ii) Calculate the molar conductivity of the 0.1M bromoethanoic acid (3marks)
- (iii) Determine the pH of 0.1M bromoethanoic acid (4 ½ marks)

Solution

(a)(i)

Concentration/ mol dm^{-3}	0.01	0.04	0.09	0.16	0.25	0.36
$\text{Concentration}^{\frac{1}{2}}/(\text{mol dm}^{-3})^{\frac{1}{2}}$	0.1	0.2	0.3	0.4	0.5	0.6
Molar conductivity ($\Omega^{-1}\text{cm}^2\text{mol}^{-1}$)	238	230	224	217	210	202



- (ii) **NaOH is Strong electrolyte:** molar conductivity is decreases as concentration increase because
- **It is** are completely ionized in solution;
 - at **high** concentration the solution contains high density of K⁺ and Cl⁻ ions
 - high density of ions lead ion interaction that increases the drag force of ions which lowers molar conductivity.
 - On dilution the tendency to form these interactions reduce lowering the drag force thus molar **conductivity** increases.

(iii) $245 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$

(c) (i) $K = \frac{1}{R} \times \text{cell constant}$

$$0.0116 = \frac{1}{24.96} \times \text{cell constant}$$

$$\text{Cell constant} = 0.29 \text{cm}^{-1}$$

(ii) K for bromoethanoic acid = $\frac{1}{66.5} \times 0.29 = 4.36 \times 10^{-3} \Omega^{-1} \text{cm}^{-1}$

$$(iii) \Lambda_c = \frac{K}{C} = \frac{4.36 \times 10^{-3}}{0.1 \times 10^{-3}} = 43.62 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$(iii) \alpha = \frac{\Lambda_c}{\Lambda_0} = \frac{43.6}{389} = 0.112$$

$$[\text{H}^+] = \alpha C = 0.112 \times 0.1 = 0.0112 \text{mol dm}^{-3}$$

$$\text{pH} = -\log 0.0112 = 1.95$$

Trial 9

The molarity of a sample of hydrochloric acid about 0.1M was determined accurately by measuring the conductivity of solution as 1.0M sodium hydroxide was added to 50cm³ of the acid. The results were as follow:

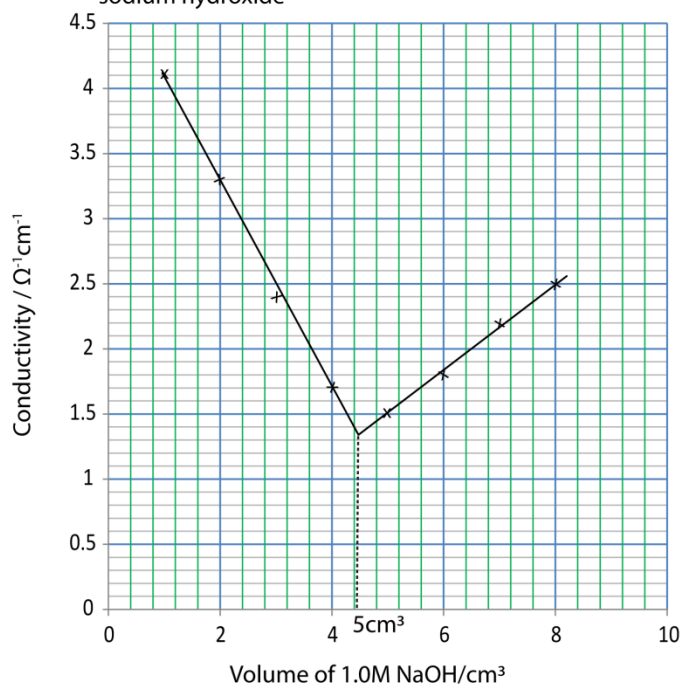
Conductivity/ $\Omega^{-1} \text{cm}^{-1}$	4.1	3.3	2.4	1.7	1.5	1.8	2.2	2.5
Volume of 1.0M NaOH/cm ³	1	2	3	4	5	6	7	8

- (a) (i) Plot a graph of conductivity against the volume of 1.0M sodium hydroxide (3marks)
 (ii) Determine from the graph the volume of 1.0M sodium hydroxide used to reach the end point (1mark)
 (iii) Calculate the molarity of hydrochloric acid (02marks)
 (iv) Explain the shape of the graph (3marks)
- (b) Name one other application of conductivity measurement (1mark)

Solution

6 (a)(i)

A graph of conductivity against the volume of 1.0M sodium hydroxide



(ii) 5cm³

(iii) moles of hydrochloric acid = moles of sodium hydroxide = $\frac{1 \times 5}{1000} = 0.005 \text{ moles}$

$$\text{Molarity of hydrochloric acid} = \frac{1000 \times 0.005}{50} = 0.1M$$

(iv) Conductivity initially decreases with increasing volume of sodium hydroxide due to removal of highly conducting hydrogen ions up to the end point. It then increases due to excess OH⁻ ions.

(b) study of complexes

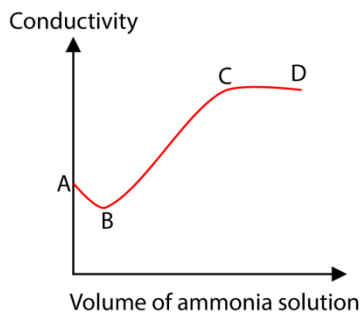
Trial 10

(a) Define the following terms

(i) Conductivity

(ii) Molar conductivity (3marks)

(b) The graph below shows the change in conductivity when 0.01M methanoic acid is titrated with 0.1M ammonium solution



Explain the shape of the graph.

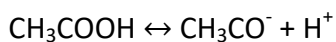
- (c) At 25⁰C, the molar conductivity of silver nitrate, potassium nitrate and potassium chloride are 133.4, 145.0 and 149.9 $\Omega^{-1}\text{cm}^2\text{mol}^{-1}$ respectively.

At the same temperature, the conductivity of a saturated solution of silver chloride is $3.41 \times 10^{-6} \Omega^{-1}\text{cm}^{-1}$ while that of pure water is $1.6 \times 10^{-6} \Omega^{-1}\text{cm}^{-1}$

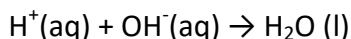
- (i) Calculate the solubility of silver chloride in moles per dm^3 at 25⁰C. (4marks)
 (ii) Determine the solubility product of silver chloride at 25⁰C. (3marks)
 (d) The ionic conductivities of rubidium and sodium ions are 73.3 and 50.1 $\Omega^{-1}\text{cm}^{-1}$ respectively. Explain why the ionic conductivity of rubidium ion is higher than that of sodium ions (3marks)

Solution

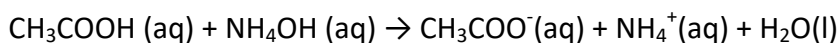
- (a) (i) Electrolytic conductivity is the reciprocal of resistance of a solution between electrodes each 1cm^2 in area and 1cm apart.
 (ii) Molar conductivity is a conductivity of solution that contains 1 mole of electrolyte.
 (b) Initially at A, conductivity is high due to presence of H^+ from partial ionization of the acid



Along AB conductivity decreases due to removal of H^+ .



Along BC conductivity increases due to addition of salt ion



Along CD conductivity remains almost constant due to attainment of equilibrium.

- (c) $\Lambda_0 \text{AgCl} = \Lambda_0 \text{AgNO}_3 + \Lambda_0 \text{KCl} - \Lambda_0 \text{KNO}_3$

$$\begin{aligned}
&= 133.4 + 149.9 - 145 \\
&= 138.3 \Omega^{-1}\text{cm}^2\text{mol}^{-1} \\
\text{KAgCl} &= 3.41 \times 10^{-6} - 1.6 \times 10^{-6} \\
&= 1.81 \times 10^{-6} \Omega^{-1}\text{cm}^{-1} \\
\Lambda_0 \text{AgCl} &= \frac{K}{C}
\end{aligned}$$

$$C = \frac{1.6 \times 10^{-6}}{138.3} = 1.3 \times 10^{-8} \text{ molcm}^{-3} \text{ or } 1.3 \times 10^{-5} \text{ moldm}^{-3}$$

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (1.3 \times 10^{-5})^2 = 1.7 \times 10^{-10} \text{ mol}^2\text{dm}^{-6}$$

Trial 11

(a) (i) state Kohlrausch's law of ionic conductivity at infinite dilution (1mark)

(ii) Calculate the molar conductivity of methanoic acid at infinite dilution

($\lambda_0(\text{HCOONa}) = 9.5 \times 10^{-2} \text{ Sm}^2\text{mol}^{-1}$, $\lambda_0(\text{NaCl}) = 1.26 \times 10^{-1} \text{ Sm}^2\text{mol}^{-1}$, $\lambda_0(\text{HCl}) = 4.26 \times 10^{-1} \text{ Sm}^2\text{mol}^{-1}$,)
(2marks)

(b) Ionic conductivities of Ag^+ and Cl^- ions at infinite dilution are 6.2×10^{-2} and $7.6 \times 10^{-2} \text{ Sm}^2\text{mol}^{-1}$ respectively at 298K. The electrolytic conductivity of silver chloride at 298K is $1.22 \times 10^{-4} \text{ Sm}^{-1}$. Calculate the solubility, in mol dm^{-3} , of silver chloride at 298K. (5marks)

(c) Name one practical application of ionic conductivity apart from the determination of solubility of electrolytes.

Solution

(a) (i) It states: "the molar conductivity of an electrolyte at infinite dilution is the sum of molar conductivity of the constituent ions at infinite dilution."

$$\begin{aligned}
\text{(ii) } \Lambda_0 \text{HCOOH} &= \Lambda_0 \text{HCOONa} + \Lambda_0 \text{HCl} - \Lambda_0 \text{NaCl} \\
&= 0.095 + 0.426 - 0.126 \\
&= 0.395 \Omega^{-1}\text{cm}^2\text{mol}^{-1}
\end{aligned}$$

$$\begin{aligned}
\text{(b) } \Lambda_0 \text{AgCl} &= \lambda_0 \text{Ag}^+ + \lambda_0 \text{Cl}^- \\
&= 0.062 + 0.076 = 0.138 \Omega^{-1}\text{cm}^2\text{mol}^{-1}
\end{aligned}$$

$$\Lambda_0 \text{AgCl} = \frac{K}{C}$$

$$C = \frac{1.22 \times 10^{-4}}{0.138} = 8.84 \times 10^{-4} \text{ molm}^{-3} \text{ or } 8.8 \times 10^{-7} \text{ moldm}^{-3}$$

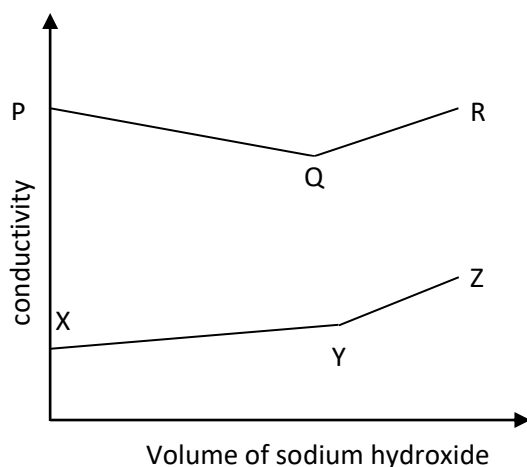
(c) – study complexes

- Determine end point in titration

Trial 12

- (a) (i) State Faraday's laws of electrolysis. (4marks)
- (ii) A current of 2A was passed for 30 minutes through a cell containing dilute sulphuric acid and the hydrogen produced at the cathode collected, calculate the volume of hydrogen, in cm^3 , that was produced at 23°C and 100kPa. (6marks)

(b)



Graphs PQR and XYZ show variation of conductivities of solutions formed when equal volumes of 0.1M hydrochloric acid and 0.1M ethanoic acid respectively were titrated with 1M sodium hydroxide. Account for the difference in the shapes of the graphs.

Solution

- (a) (i) the mass of a substance liberated at an electrode is proportional to the quantity of electricity used.

the number of moles of electricity liberated by 1 mole of electricity is inversely proportional to the charge on its ions

- (ii) $Q = it = 2 \times 30 \times 60 = 3600\text{C}$
 $2\text{H}^+(\text{aq}) + 2\text{e} \rightarrow \text{H}_2(\text{g})$
Volume of hydrogen liberated at stp.
 96500×2 liberates 22400cm^3 of hydrogen

$$3600 \text{ liberate } \frac{3600 \times 22400}{96500 \times 2} = 417.8 \text{ cm}^3$$

Volume of hydrogen at 23°C and 100KPa

From $\frac{PV}{T} = \text{constant}$

$$\text{Then, } \frac{100000 \times V}{273+23} = \frac{101325 \times 417}{273}$$

$$V = 459 \text{ cm}^3$$

(b) Graph PQR

- At P the conductivity is high due to presence of hydrogen ions from ionization of HCl.
- Along PQ conductivity decrease due to removal of hydrogen ions by OH⁻ ions.
- After the end point at Q, conductivity increase due to excess OH⁻ ions.

Graph XYZ

- At X conductivity is low due to presence of few hydrogen ions from partial ionization of ethanoic acid.
- Along XY conductivity increases due to addition of salt ions and further ionization of the acid on dilution
- After the end point at Y conductivity increases due to excess hydroxyl ions from sodium hydroxide

Trial 13

The ionic radii and molar ionic conductivities at infinite dilution at 18°C of lithium and cesium ions are given in the table below

	Li ⁺	Cs ⁺
Ionic radius (nm)	0.06	0.17
Molar ionic conductivity (Ω ⁻¹ cm ² mol ⁻¹)	33.5	68.0

Explain why the molar conductivity of lithium ions is lower than that cesium ions.

Solution

- Lithium ions have a large charge density.
- In aqueous solution they attract a big cloud of water molecules around themselves.
- This makes their effective size bigger than that of cesium.
- The big size of lithium ions the drag force that reduces its conductivity.

Trial 14

During the electrolysis of dilute sulphuric acid using platinum electrodes, a current of 2A was passed for 6.7 minutes at room temperature.

Calculate the volume of the gas evolved at

- (ii) The anode
(iii) The cathode (6 marks)

[1 Faraday = 96500C, 1 mole of a gas occupies 24dm³ at room conditions]

- (iv) Name one application of
- Electrolysis
- Standard electrode potential.

Solution

$$Q = it = 2 \times 6.7 \times 60 = 804C$$

- (i) At the anode oxygen is liberated
 $4OH^-(aq) + 4e \rightarrow 2H_2O(l) + O_2(gas)$
(4 x 96500)C liberate 24000cm³
 $\therefore 804 C$ liberate 50cm³
- (ii) At cathode hydrogen is liberated.
 $2H^+(aq) + 2e \rightarrow H_2(g)$
(2 x 96500)C liberate 24000cm³
 $\therefore 804 C$ liberate 10cm³

Electrochemical cells

When a strip of metal is placed in a solution of its ions, the metal atoms tend to ionize; cations dissolve in solutions while electrons remain on the metal surface. For instance, a divalent metal, M, ionizes as follows

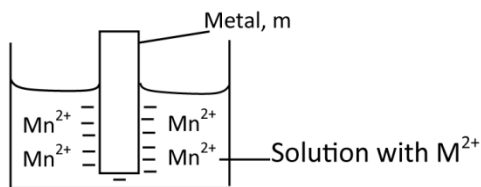


Fig. 4.1

The attraction of the metal rod (containing negative charge) and the solution (containing positive charges) causes a potential difference called **electrode potential**.

The metal dipped into its solution containing metal ions is called electrode half-cell represented symbolically as M^{n+}/M .

Definition

Electrode potential is a potential difference that is set up between a metal and its solution containing metal ions.

Factors affecting the magnitude of the potential difference

1. **Concentration of metal ions already in solution.** Since the ionization equation (1) is a reversible reaction, the higher the concentration of metal ions in solution, the further to the left will be the state of equilibrium and hence the smaller the electrode potential.
2. **Temperature.** Increasing temperature increases the ionization of metal atoms and thus increases the electrode potential.
3. **Pressure: applicable to gas systems,** the higher the pressure the higher the electrode potential because increased pressure increases the gas concentration at gas-liquid junction.
4. **Position of metal in electrochemical series.**
Metals above hydrogen in electrochemical series, have ionization reactions that lie more to the right than left, thus, the metal tends to be negatively charged with respect to their solutions. Thus have negative electrode potentials.

Metals below hydrogen in electrochemical series, their ionization reaction lie more to the left and their electrode potentials are positive.

Standard electrode potential (SEP)

This is the electrode potential of a metal dipped in a solution containing one mole of metal ions in a dm^3 /litre of solution at 298K and 1 atmosphere. It is denoted by E^0 .

Factors affecting standard electrode potential

Electrode potential involves formation of hydrated ions. It is a sum of atomization energy (which is endothermic) + ionization energy (which is endothermic) + ionic hydration energy (which is exothermic).

When atomization and/or ionization energy is high, the standard electrode potential becomes more positive.

On the other hand, when the hydration energy is higher than the sum of ionization and atomization energy, standard electrode potential is negative.

Short-hand Notation for an electrode half cell

A metal rod dipped in a solution of metal ions constitutes a half cell. The accepted short form of writing an electrode is to write 'oxidized form/reduced form' of metal. For example, the zinc half-cell is written as Zn^{2+}/Zn and copper electrode as Cu^{2+}/Cu .

Standard Hydrogen electrode (fig. 4.2)

It consists of hydrogen gas at 298K and pressure of 1 atmospheres bubbling over a strip of platinized foil (i.e. platinum coated) in a solution which is 1M with respect to H^+ ions

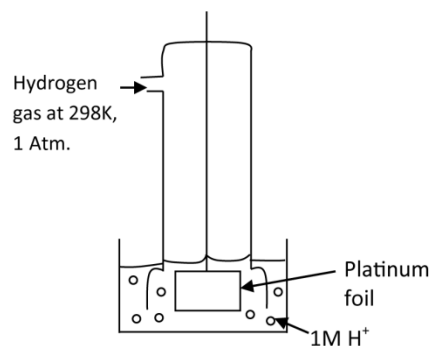
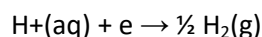


Fig. 4.2

The short hand notation for a hydrogen electrode is

$H^+, M/ \frac{1}{2} H_2, 1 \text{ atm.}$ Or $H^+ / \frac{1}{2} H_2$.

And electrode reaction is



A potential develops on the surface of the platinum; by convention, it is assigned an arbitrary value of **zero volts**.

Hydrogen electrode is used as a reference electrode and the electrode potential of all other electrodes are measured relative to this.

Measurement of electrode potential of metal (fig. 3)

A standard metal electrode is combined with a standard hydrogen electrode as shown in Fig.4.3 below

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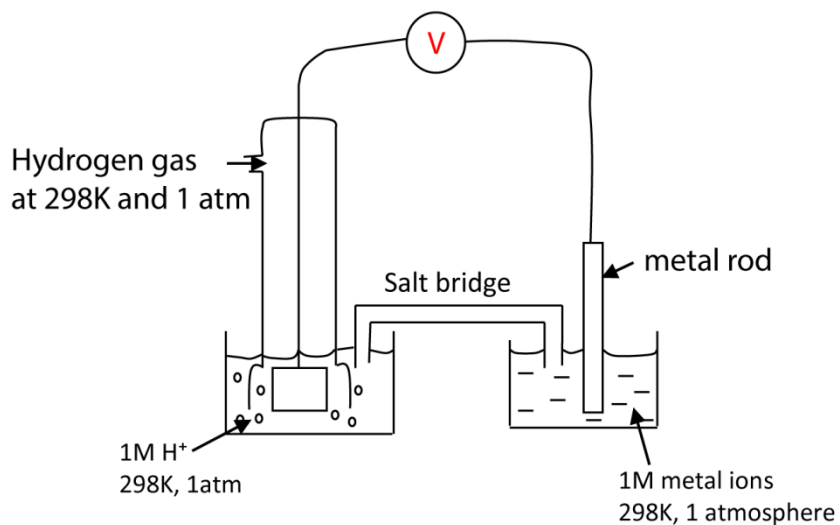


Fig. 4.3

The two compartments in the figure are connected by a **salt bridge**. This contains an electrolyte such as potassium chloride, which conducts electricity but does not allow mixing of two solutions in half cells. The measurement on the high resistance voltmeter is the standard electrode of the metal since that of hydrogen half-cell is zero.

Definition

Standard electrode potential is the electrode potential value of an electrode measure with respect to a standard hydrogen electrode of 0 volts. When an electrode is immersed or dipped into a solution of 1M concentration of its ions at a standard temperature of 298K and pressure of 1 atmosphere.

Some electrode and their standard electrode potentials are given below:

Electrode	Electrode reaction	E^0 /volts
Zn^{2+}/Zn	$Zn^{2+}(aq) + 2e \rightarrow Zn(s)$	-0.76
Cu^{2+}/Cu	$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$	+0.34
Fe^{2+}/Fe	$Fe^{2+}(aq) + 2e \rightarrow Fe(s)$	-0.44
Ag^+/Ag	$Ag^+(aq) + e \rightarrow Ag(s)$	+0.80

Galvanic or Voltaic cells or electrochemical cells

This is a type of cell in which a chemical reaction results in production of an electric current.

By combining two suitable electrodes (half cells) a cell of a particular emf may be obtained. One electrode acts as electron supplier and the other as electron acceptor.

The Daniel cell (fig. 4.4)

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It consists of the zinc electrode dipped into 1M zinc sulphate solution and copper electrode dipped into 1M copper sulphate solution, the two solution separated by a porous partition. The cell develops an electromotive force (emf) of 1.10V.

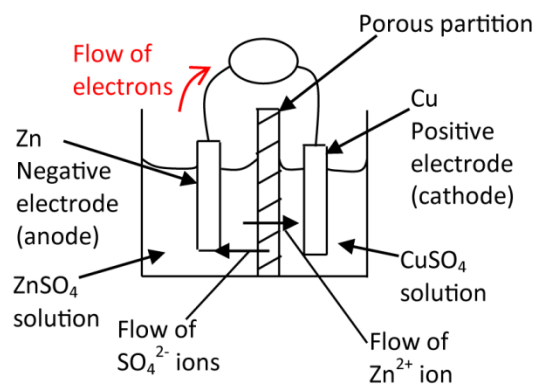
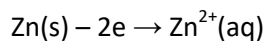


Fig.4.4

Half cell reaction

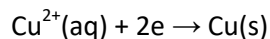
At zinc electrode (**anode**)

Oxidation occurs and the electrode dissolves.

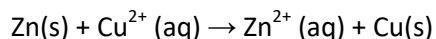


At the copper electrode

Reduction occurs (**cathode**)

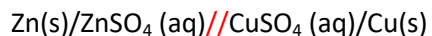


Overall equation

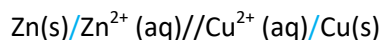


Cell notation

The cell in fig. 4 can be represented as



Or simply



The single vertical line (/) indicates a phase boundary and double vertical lines (//) indicate a porous partition between the solutions.

Emf of a cell

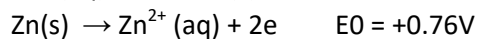
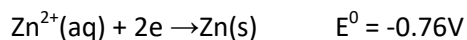
The emf of the above cell is given by

$$E = E_{\text{Cu}} - E_{\text{Zn}}$$

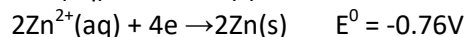
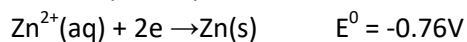
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Rules regarding electrode potentials

1. If the direction of an electrode reaction is reversed, the sign of its electrode potential must also be reversed.



2. If an electrode reaction equation is multiplied by a positive factor. The electrode potential must **not** be multiplied by that factor. It remains unchanged.



Not 1.52V

Generally the emf of a cell is defined as follow

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{R.H.E}} - E^{\circ}_{\text{L.H.E}}$$

NB. For a positive cell emf, R.H.E is for the half cell with most positive electrode potential.

Example 4

Calculate the standard emf of $\text{Zn}/\text{Zn}^{2+}/\text{Cu}^{2+}/\text{Cu}$ from the data given below. Write down the half cell reactions and cell reaction.

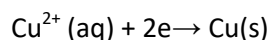
Electrode	E°/V
Zn^{2+}/Zn	-0.76
Cu^{2+}/Cu	+0.34

Answer

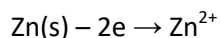
$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{R.H.E}} - E^{\circ}_{\text{L.H.E}} \\ &= 0.34 - (-0.76) \\ &= +1.10\text{V} \end{aligned}$$

Electrode reactions

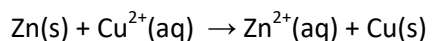
At R.H.E (**cathode**)



L.H.E (**anode**)



Cell reaction



Implication of the emf of a cell

For electrode; since, standard electrode potentials are reduction potentials, a negative value for E^0 implies that the reduction reaction is non-spontaneous while a positive value implies that the reduction reaction is spontaneous.

For a cell a negative value of E^0_{cell} indicates that the cell reaction as written is non-spontaneous and instead is spontaneous in the reverse direction.

Therefore, the value of standard electrode potential may be used to predict the direction of spontaneous chemical reaction. Then when writing down or constructing a cell given two electrodes together with their standard electrode potentials, the electrode with the less negative E^0 (if both are negative or electrode with more positive value if both are positive) must be the right hand electrode of the cell and the other the left hand electrode.

Example 5

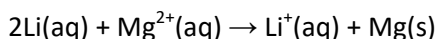
- Construct a cell using Li^+/Li and Mg^{2+}/Mg electrodes, given that $E^0_{\text{Li}^+/\text{Li}} = -3.04\text{V}$ and $E^0_{\text{Mg}^{2+}/\text{Mg}} = -2.37\text{V}$.
- Calculate the E^0_{cell} .
- Give the cell reaction

Solution



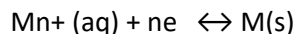
$$\begin{aligned} \text{(b) } E^0_{\text{cell}} &= E^0_{\text{R.H.E}} - E^0_{\text{L.H.E}} \\ &= -2.37 - (-3.04) \\ &= +0.67\text{V} \end{aligned}$$

- (c) Cell reaction



Standard electrode potential and standard free energy change

There is a relationship between electrochemistry and thermochemistry. In the electrode reaction

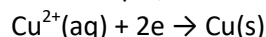


If the standard free energy change, ΔG^0 then the value of the standard electrode potential, E^0 , at the same temperature is given by

$$\Delta G^0 = -nFE^0$$

Where n is the number of electrons transferred in the electrode reaction and F is the Faraday constant.

For example, $E^0 = +0.34\text{V}$ at 298K for the reaction



The value of ΔG^0 is given by

$$\begin{aligned} \Delta G_0 &= -2 \times 96500 \times (-0.34) \\ &= 65.6\text{kJmol}^{-1}. \end{aligned}$$

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Liquid junction potential

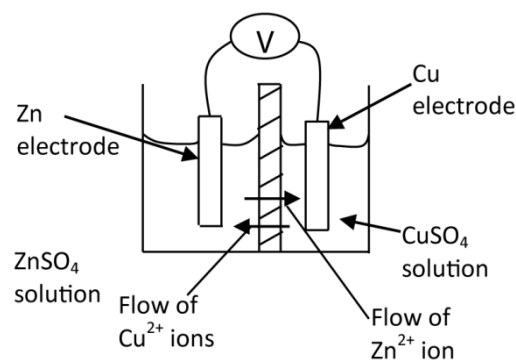
Is a potential difference set up across a phase boundary between two solutions in a cell. It tends to oppose the cell potential i.e. it reduces it and hence the observed cell emf is less than the actual emf of the cell

$$E^{\circ}_{\text{cell}} \quad \text{Observed} = E^{\circ}_{\text{cell}} \quad \text{actual} - E^{\circ}_{\text{j}}$$

the liquid junction
potential

Causes of liquid junction potential

It is caused by the differential diffusion of charges across the porous partition between the two solutions of the cell. For example, in the Daniel cell, the Zn^{2+} ions diffuse into the CuSO_4 solution and Cu^{2+} diffuse into the ZnSO_4 solution.



The speed of migration of Zn^{2+} and Cu^{2+} cations across the phase boundary is different due to a number of factors. Some of which are either the ionic size and/or the extent to which they are hydrated. Due to this, a p.d. is set up across the boundary called Liquid Junction Potential.

This liquid junction potential may be eliminated by connecting the solutions in the two half cells by means of a salt bridge which is a glass tube containing a saturated solution of either KCl or NH_4NO_3 .

NB. In the cell diagram/notation, the double lines either represent a porous partition or a salt bridge.

Types of electrodes

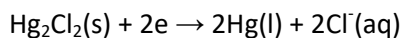
1. The metal/metalloid e.g. Zn^{2+}/Zn
2. The gas electrode e.g. $\text{Pt}, \text{H}^+ / \frac{1}{2} \text{H}_2$
3. The metal/insoluble electrode: this type of a metal coated with one of its insoluble salt surrounded by a solution containing the anion of the salt e.g.
 - (i) $\text{Ag}/\text{AgCl}(\text{s}), \text{Cl}^-$
Possible electrode reactions
As R.H.E (reduction)
 $\text{AgCl}(\text{s}) + \text{e} \rightarrow \text{Ag}(\text{s}) + \text{Cl}^-(\text{aq})$
As L.H.E (oxidation)
 $\text{Ag}(\text{s}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{e}$

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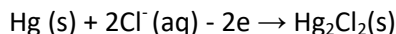
- (ii) $\text{Hg}/\text{Hg}_2\text{Cl}_2, \text{KCl}(\text{aq})$ or $\text{Hg}/\text{Hg}_2\text{Cl}_2, \text{Cl}^-$ This electrode is commonly known as a calomel electrode.

Possible electrode reactions

As R.H.E (reduction)



As L.H.E (oxidation)



The two electrodes $\text{Ag}/\text{AgCl}(\text{s}), \text{Cl}^-$ and $\text{Hg}/\text{Hg}_2\text{Cl}_2, \text{KCl}(\text{aq})$ have stable potential and are not easily affected by oxidants or reducing agents, and because of their high stability, they are sometimes used as reference electrodes since hydrogen electrode in practice is easily poisoned by oxidants and reducing agents.

4. Oxidation—reducing (redox) electrodes

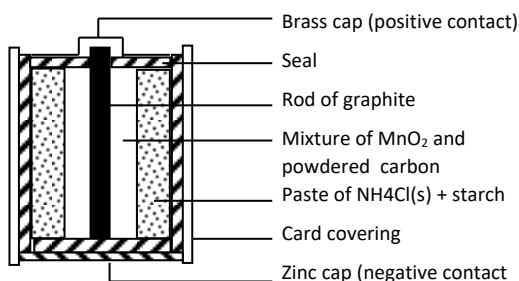
This type comprises of a reduced species and oxidized species in the same solution, e.g. $\text{Pt}/\text{Fe}^{3+}, \text{Fe}^{2+}$ or $\text{Pt}/\text{Cr}^{3+}, \text{Cr}^{2+}$.

Differences between electrolytic and e.m.f cells

Electrolytic cell	E.m.f. cells
Uses single electrolyte	Uses two electrolytes
Changes electrical energy into chemical energy	Changes chemical energy into electrical energy
Cathode is negative electrode	Cathode is positive electrode
Anode is positive electrode	Anode is negative electrode

Dry cells

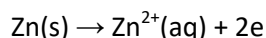
These were invented to overcome the difficulty of electrolytic solution leaking out of cell such as Daniel cell. In dry cell, the electrolyte is made into a paste. An example is shown below



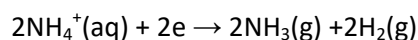
This type of cell is used in radio, flash light and clocks as it is portable.

The initial electrode processes are

Anode:



Cathode:



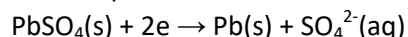
The lead-acid accumulator

The cell stores or accumulates electric charge. It consists of two lead plates dipping into a 30% solution of sulphuric acid. Both plate become covered with insoluble film of lead II sulphate

First the cell must be charged. A direct current is passed through the cell. The processes which take place are:

Charge

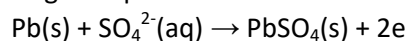
Positive plate



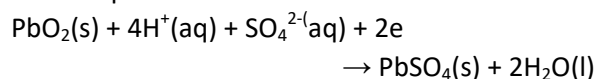
Negative plate

The plates are now different and therefore have different potentials, so that, when they are connected, an electric current will flow between them. When the cell supplies an electric current, i.e. discharge the process which take place are:

Negative plate



Positive plates



NB: the plate which is positive during the charge becomes negative during the discharge.

Example 6

Explain why distilled water is preferred to dilute sulphuric acid when changing an accumulator cell.

During recharging, there can be some production of hydrogen and oxygen caused by splitting of the water part of the acid. This needs to be replaced with distilled water. The sulphuric acid level remains the same, but since water is being lost, the concentration goes up, so adding water restores it to how it was when the battery was new.

Example 7

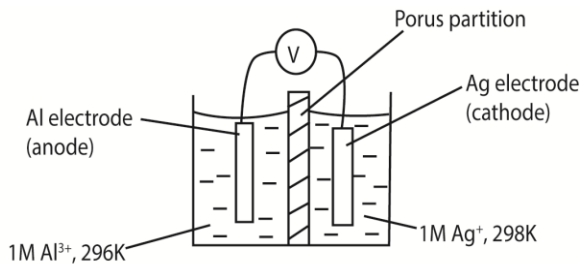
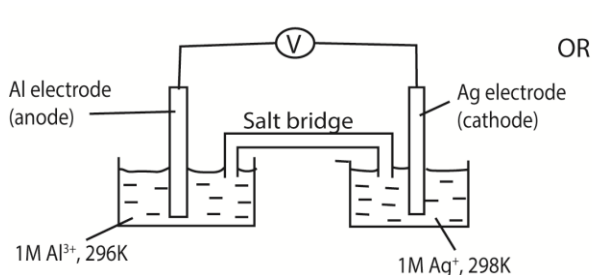
1(a.) What is meant by the term voltaic cell?

2 marks

A voltaic cell is an electrochemical cell that uses a chemical reaction to produce electrical energy.

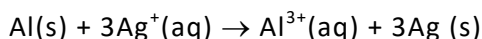
b.(i) With aid of labeled diagram describe briefly how the standard electrode potential of a Voltaic cell made up of silver electrode and aluminium electrode can be determined.

8 marks



(ii) Write the overall equation for the cell reaction.

1 ½ marks



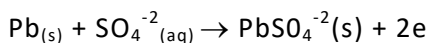
(iii) If the cell liberated 708.9 kJ of heat, determine the e.m.f of the cell. 3 ½ marks

$$\Delta G = -nFE^{\circ}$$

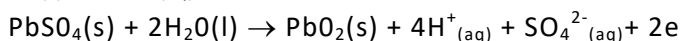
$$708.9 \times 1000 = 3 \times 96500 \times E^{\circ}$$

$$E^{\circ} = 2.45\text{V}$$

(c) Some electrode potentials of the electrode reactions of an accumulator cell are shown below;

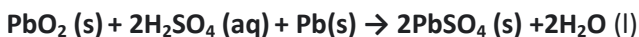


$$E^{\circ} = +0.13\text{V}$$



$$E^{\circ} = -1.47\text{V}$$

(i) Write the overall cell reaction and hence determine the e.m.f of the cell. 2 ½ marks

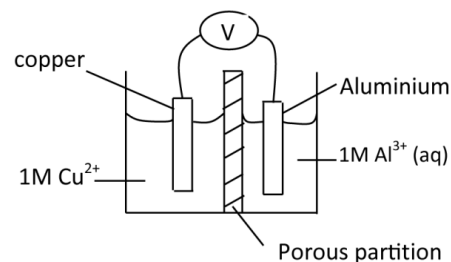


(ii) Explain why distilled water is preferred to dilute sulphuric acid when changing an accumulator cell. 2 ½ marks

During recharging, there can be some production of hydrogen and oxygen caused by splitting of the water part of the acid. This needs to be replaced with distilled water. The sulphuric acid level remains the same, but since water is being lost, the concentration goes up, so adding water restores it to how it was when the battery was new.

Trial 15

The diagram below shows an electrochemical cell



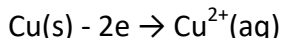
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- (i) State the purpose of porous partition (1mark)
- (ii) Give the ionic equation for the reaction taking place at each electrode (2marks)
- (a) (i) the standard electrode potentials (E^0) at 25°C are $+0.34\text{Volts}$ for copper electrode and -1.66V for aluminium electrode. Calculate the e.m.f of the cell.
- (ii) What does the value of the e.m.f calculated in (b)(i) signify?

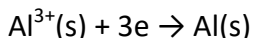
Solution

(a)(i) the porous partition prevents mixing of solution while allowing exchange of charges.

(a)(ii) At copper electrode



At aluminium electrode



$$\begin{aligned} \text{(b)} \quad E^0_{\text{cell}} &= E^0_{\text{R.H.E}} - E^0_{\text{L.H.E}} \\ &= -1.66 - 0.34 \\ &= -2.00\text{V} \end{aligned}$$

(i) A negative e.m.f indicates that the cell as set up is non-spontaneous.

Trial 16

- (a) Explain what is meant by the terms
- (i) Standard electrode potential (2mks)
- (ii) First ionization energy (2marks)
- (b) The first ionization energies (I.E/kJ) and standard electrode potentials (E^0/V) of group I elements of the Periodic Table are given in the table below

	Li	Na	K	Rb	Cs
1 st I.E/ kJmol ⁻¹	518.8	497.9	422.6	401.7	380.7
E^0/V	-3.03	-2.71	-2.93	-2.99	-3.02

- (i) Why is the standard electrode potential of lithium is anomalously high? (3mks)
- (ii) What would you expect to be the trend in ionic radii of univalent cations formed by group I elements? Explain your answer. (3mks)

Solution

(a)(i) Standard electrode potential is a reduction potential between a metal and a solution containing 1mole of metal ions, at 298K and 1 atmosphere.

(a)(ii) First ionization energy is the minimum energy required to remove an electron in its ground state completely from a gaseous atom

(b)(i) The small Li^+ ion has a high charge density leading to high negative hydration energy that offsets its high positive ionization and its high atomization energies. Other elements in the group have ionization, atomization energies and hydration energies.

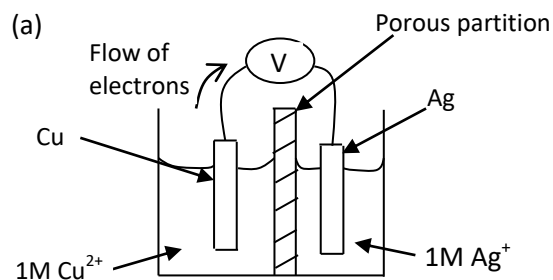
(ii) ionic radii increase from Li to Cs because

- There are additional filled electron shells from one element to another one element to the next.
- There is a reduction in the nuclear charge and an increase in the screening effect down the group

Trail 17

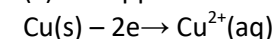
- (a) Draw a fully labeled diagram to show the arrangement of apparatus you would use to measure the emf of a cell consisting of copper and silver electrode. Indicate the direction of electron flow (5mks)
- (b) Write
- The notation for the cell you have drawn in (a) and state the convention used (3mks)
 - The equation taking place at each electrode (2mks)
 - The overall cell reaction (1mk)
- (c) If the standard electrode potential for copper ions/copper and silver ions/silver are +0.34V and 0.81V respectively, calculate the e.m.f of the cell consisting of copper and silver electrodes.
- (d) Describe briefly how the standard electrode potential of copper can be determined and write the cell notation (7mks)

Solution

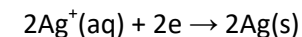


(b) (i) $\text{Cu(s)}/\text{Cu}^{2+}(\text{aq})//\text{Ag}^{+}(\text{aq})/\text{Ag(s)}$
 (Convention used?)

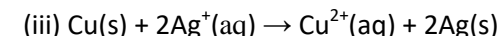
(ii) At copper electrode (anode)



At silver electrode (cathode)



Overall cell reaction



(c) 0.47V

(d) Arrange the apparatus as in fig.5 below; since the standard electrode potential of hydrogen gas is zero, the reading on the voltmeter is the standard electrode potential of copper electrode.

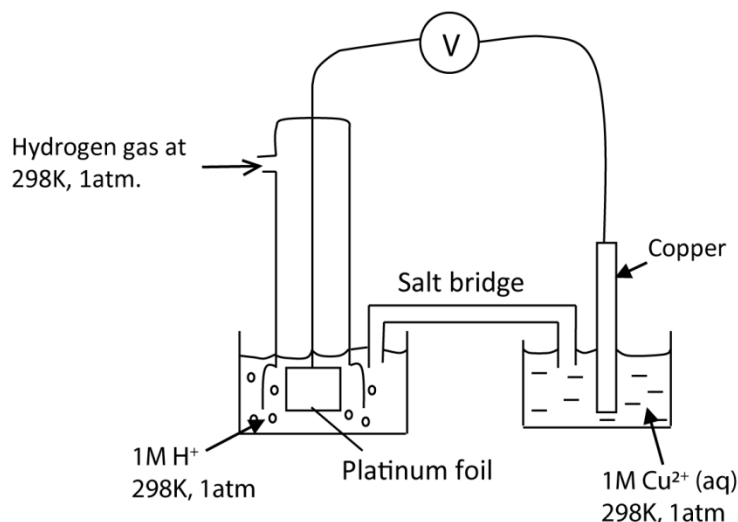
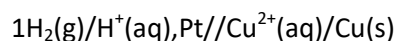


Fig. 5

Cell notation



Trial 18

The standard electrode potentials E^0 for four half-cell reactions are given below

	E^0/V
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.52
$\text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+ 0.20
$\text{Br}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$	+1.06
$\text{Cl}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	+1.36

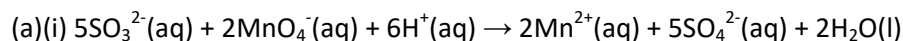
(a) (i) write an equation between sulphate ions and acidified solution of potassium manganate (VII)

(ii) Calculate the E^0 for the reaction in (i)

(b) Explain why hydrochloric acid is not used to acidify titrants in volumetric analysis involving manganate (VII)

(c) Write the cell convention for the reaction between bromine water and acidified potassium manganate(VII)

Solution



(a)(ii) $E^0_{\text{cell}} = E^0_{\text{Reduction}} - E^0_{\text{oxidation}}$
 $= 1.52 - 0.20$
 $= 1.32\text{V}$

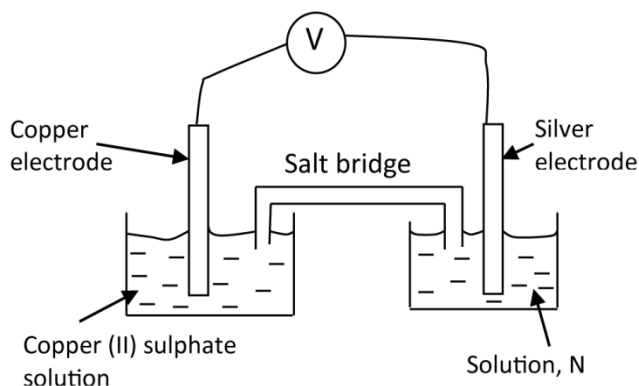
(b) Because chloride ions are strong oxidizing agents and will be oxidized by manganate (VII) in the reaction.



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Trial 19

A cell consisting of copper and silver electrode is shown in the diagram below



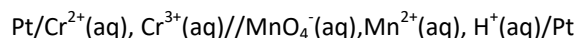
- (a) What is solution, N? (1mark)
- (b) Write
- (i) The notation for the cell (2mks)
- (ii) Overall equation for the cell (2mks)
- (c) (i) Write a substance that can be used in the salt bridge (1mk)
- (ii) State the purpose of the salt bridge (1mk)
- (d) Calculate the overall electrode potential, E , of the cell (2mk)

Solution

- (a) 1M Silver nitrate solution (or any soluble silver salt)
- (b) (i) $\text{Cu(s)}/\text{Cu}^{2+}(\text{aq})//\text{Ag}^{+}(\text{aq})/\text{Ag(s)}$
- (ii) $\text{Cu(s)} + 2\text{Ag}^{+}(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{Ag(s)}$
- (c) (i) KCl or NH_4NO_3
- (d) 0.46V

Trial 20

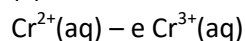
A cell was constructed as shown below



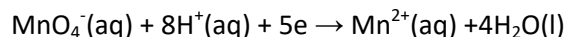
- (a) Write equation for the reaction that occur at each electrode (2½mks)
- (b) Write the equation for overall reaction
- (c) The electrode potential for system $\text{Cr}^{3+}/\text{Cr}^{2+}$ and $\text{MnO}_4^{-}/\text{Mn}^{2+}$ are -0.402 and +1.52volts respectively. Calculate the cell voltage (1mk)

Solution

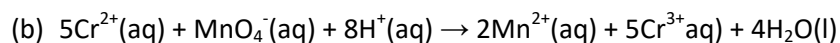
(a) L.H.E



R.H.E



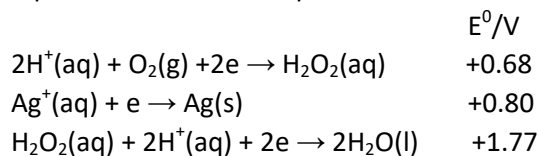
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(c) 1.922V

Trial 21

Equation and electrode potentials for some reaction are given below

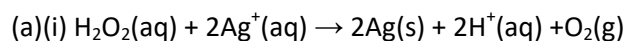


(a) (i) write equation for the reaction between silver nitrate and hydrogen peroxide (2mk)

(ii) Calculate E° cell for the reaction (i)

(b) State the property shown by hydrogen peroxide in the reaction in (a)(i) (1mk)

Solution

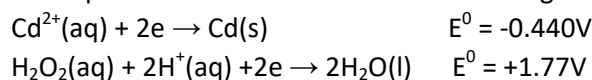


(a)(ii) +0.12V

(b) it is a reducing agent

Trial 22

The equation for some half-cell reaction are given below



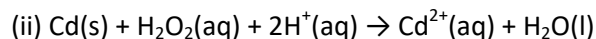
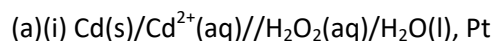
(a) Write

(i) The cell notation

(ii) The equation for overall cell reaction.

(b) The electrodes were connected externally by a conductor. State the direction of flow of electrons.

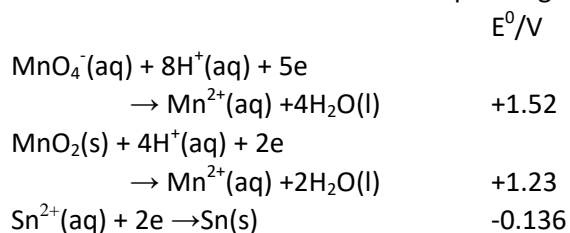
Solution



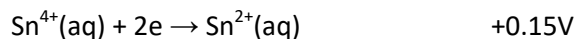
(b) from Cd (anode) to H_2O_2 , Pt

Trial 23

Some half reactions and their corresponding reduction potentials are given below:



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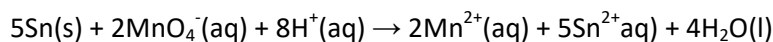


- (a) (i) state what would be observed when excess tin powder is added to acidified solution of potassium manganate (VII). Explain your answer
 (ii) Write equation for the reaction that takes place.
- (b) Solid manganese (IV) oxide was added to acidified solution of tin (II) ions
 (i) State what was observed and explain your answer (3mks)
 (ii) Write equation for the reaction (1½mks)

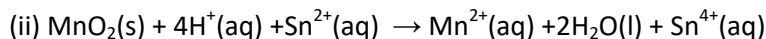
Solution

(a)(i) The purple color of acidified potassium manganate (VII) decolorize; because manganate (VII) is reduced to colorless Mn^{2+} ions

(a)(ii) Note that in excess tin powder, tin is oxidized to tin II rather Tin (IV)



(b) (i) Solid manganese (IV) oxide dissolved forming colorless solution because it is reduced to soluble colorless Mn^{2+} ions.



Trial 24

Given the following electrode potentials

	E^0/V
$\text{Na}^+(\text{aq}) + \text{e} \rightarrow \text{Na}(\text{s})$	-2.71
$\text{Cu}^{2+}(\text{aq}) + 2\text{e} \rightarrow \text{Cu}(\text{s})$	+0.337
$\text{Zn}^{2+}(\text{aq}) + 2\text{e} \rightarrow \text{Zn}(\text{s})$	-0.763
$\text{Ag}^+(\text{aq}) + \text{e} \rightarrow \text{Ag}(\text{s})$	+0.799

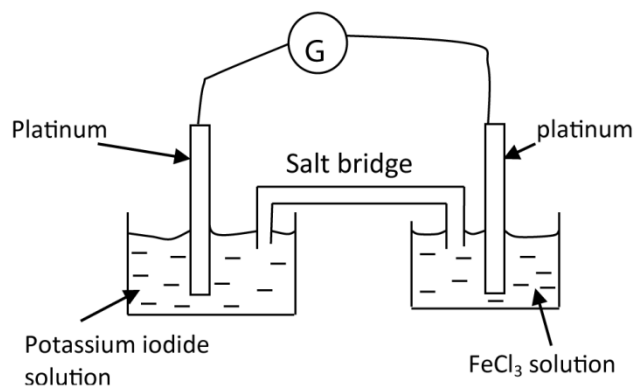
- (a) Arrange the electrodes in order of reducing power starting with the least reducing (1mk)
- (b) (i) Which two electrodes will form a cell of maximum emf? (½mk)
 (ii) Write the cell notation for the cell you gave in (i) above (1½mks)
 (iii) Write the overall cell reaction for the cell in (i). (1½mks)

Solution

- (a) Ag, Cu, Zn, Na (i.e. the more negative the E^0 , the more reducing or the less oxidizing the electrode is)
- (b) (i) Ag and Na
 (ii) $\text{Na}(\text{s})/\text{Na}^+(\text{aq}) // \text{Ag}^+(\text{aq})/\text{Ag}(\text{s})$
 (iii) $\text{Na}(\text{s}) + \text{Ag}^+(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{Ag}(\text{s})$

Trial 25

The diagram below shows emf cell for the reaction between iron (III) chloride and potassium iodide solution



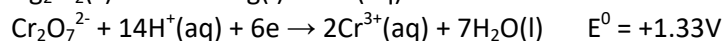
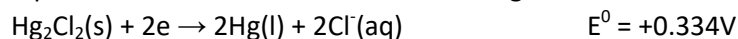
- (a) Indicate on the diagram the direction of flow of current (½mk)
- (b) State what is observed at the
- Anode
 - cathode
- (c) Write equation for the cell reaction

Solution

- (a) From a platinum electrode dipped into iron (III) chloride to platinum dipped in KI. Note current and electron flow are in opposite direction)
- (b) (i) brown solution
(ii) Brown solution turns green
- (c) $2I^-(aq) + 2Fe^{3+}(aq) \rightarrow I_2(aq) + 2Fe^{2+}(aq)$

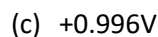
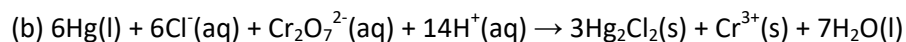
Trial 26

Equations for some half cell reactions are given below



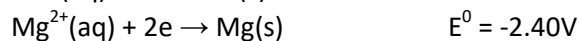
- (a) Write the convention for the cell (1mk)
- (b) Write equation for the overall reaction (1½)
- (c) Calculate the cell voltage

Solution



Trial 27

The electrode potentials, E^0 , of lead and magnesium are shown below



- (a) Write the cell convention for the cell that can be formed

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- (b) Write the equation for the
- Reaction taking place at each electrode
 - Overall reaction
- (c) Calculate the emf of the cell

Solution

- (a) $\text{Mg(s)}/\text{Mg}^{2+}(\text{aq})//\text{Pb}^{2+}(\text{aq})/\text{Pb(s)}$
- (b) (i) anode: $\text{Mg(s)} \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^-$
 Cathode: $\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb(s)}$
- (ii) $\text{Mg(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Pb(s)}$
- $$-0.13 - (-2.40) = +2.27\text{V}$$

Trial 28

The standard electrode potentials for some system are given below:

	E^0/V
$\text{Mn}^{3+}(\text{aq})/\text{Mn}^{2+}(\text{aq})$	+1.5
$\text{V}^{3+}(\text{aq})/\text{V}^{2+}(\text{aq})$	-0.26

- Write the convention for the cell formed by combining two systems (1mk)
- Write the overall equation for the cell reaction (1½mk)
- Calculate the emf of the cell. State whether the reaction is feasible or not and give a reason for your answer (2½mk)

Solution

- (i) $\text{V}^{2+}(\text{aq})/\text{V}^{3+}(\text{aq})//\text{Mn}^{3+}(\text{aq})/\text{Mn}^{2+}(\text{aq})$
- (ii) $\text{V}^{2+}(\text{aq}) + \text{Mn}^{3+}(\text{aq}) \rightarrow \text{V}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq})$
- (iii) $1.5 - (-0.27) = +1.77\text{V}$

Trial 29

- Define 'standard electrode potential'
 - Why is it not possible to measure standard electrode potential absolutely?
 - Discuss the factors which affect the value of standard electrode potential
- Describe a standard hydrogen half cell
- How would you measure standard electrode potential of a metal in a solution of its ions.
- Given that

$$\text{Ca}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ca(s)} \quad E^0 = -2.87\text{V}$$

$$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg(s)} \quad E^0 = -2.37\text{V}$$
 A cell was set up as below:

$$\text{Mg(s)}/\text{Mg}^{2+}(\text{aq})//\text{Ca}^{2+}(\text{aq})/\text{Ca(s)}$$
 - Calculate the emf of the cell
 - What conclusion can you draw from your emf value in (d)(i) above?

Solution

(a)(i) it is the electrode potential value of an electrode measure with respect to a standard hydrogen electrode of 0 volts. When an electrode is immersed or dipped into a solution of 1M concentration of its ions at a standard temperature of 298K and pressure of 1atmospheres.

(a)(ii) it requires a second electrode to be introduced since it's a difference in potential

However, the second electrode also produces its own electrode potential making such a difference relative rather than absolute

(a)(iii) Electrode potential = sublimation energy

+ ionization energy

+ Hydration energy

- Sublimation energy: the higher the sublimation energy the more positive the electrode potential. This is because it becomes difficult to convert a solid into gaseous atoms (i.e. sublimation is endothermic)
 - Ionization energy: if ionization energy is high; electrode potential becomes more positive because ionization energy is endothermic.
 - Hydration energy: a high hydration energy give a more negative electrode potential since hydration is exothermic reaction.
- (b) It consists of a molar solution of H⁺ ions having platinized titanium or platinum coated with finely divided titanium, around which pure hydrogen gas at 1 atmosphere and 298K (25^o) is bubbled.
- (c) The metal is placed in one molar solution of its ions and connected to the standard hydrogen electrode by a salt bridge

A voltmeter connected in parallel will show the emf of the cell

The standard electrode potential of the metal is equal to the emf shown by the voltmeter, measured relative to the electrode potential of hydrogen electrode considered to be zero

$$\begin{aligned} \text{(d) (i) } E_{\text{cell}}^{\circ} &= E_{\text{Reduction}}^{\circ} - E_{\text{oxidation}}^{\circ} \\ &= -2.87 - (-2.37) \\ &= -0.5\text{V} \end{aligned}$$

(ii) The cell as set up is non-spontaneous because emf is negative. It is spontaneous in the opposite direction.

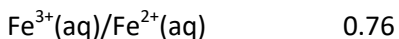
Trial 30

(c) Explain is meant by the term 'standard electrode potential' (2mks)

(d) Same half cells and their emf are given below:

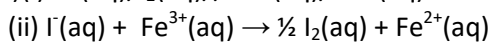
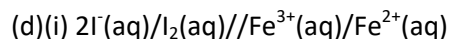
Half-cell	emf/volts
I ₂ (aq)/2I ⁻ (aq)	0.54

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- (i) Write the cell convention for the cell made up of the half cells (01mk)
- (ii) Write equation for the cell reaction
- (iii) Determine emf of the cell
- (iv) Calculate the standard free energy for the reaction and indicate its units
- (v) State whether the reaction (ii) is possible or not give reason for your answer (2mks)

Solution



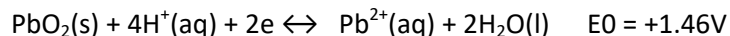
(ii) 0.22V

(iii) $\Delta G^{\circ} = -1 \times 96500 \times (-0.22)$
 $= 42.46\text{kJmol}^{-1}$

(iv) The reaction is feasible because emf is positive.

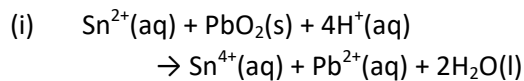
Trial 31

Equations for some half cell reaction are shown below



- (i) Write the overall equation for the spontaneously feasible reaction
- (ii) Calculate the emf of the cell in (e)(i)

Solution



(ii) 1.31V

Electrolysis

This is the decomposition of an ionic compound in molten or solution form to its constituent elements.

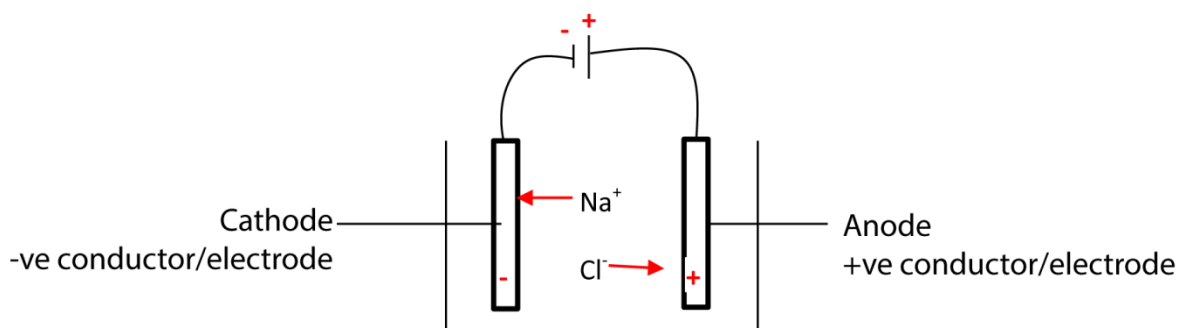
Mechanism of electrolysis

Consider electrolysis of molten sodium

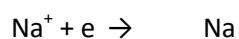
In molten form sodium chloride ionizes as follows



In an electrolytic cell, unlike charges attract; Na^+ migrate to the cathode while Cl^- migrate to the anode.



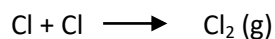
At the cathode, Na^+ ion acquires an electron to become sodium atom.



At the anode the Cl^- loses an electron to form a Cl atom



The chlorine atoms combine to form chlorine gas.

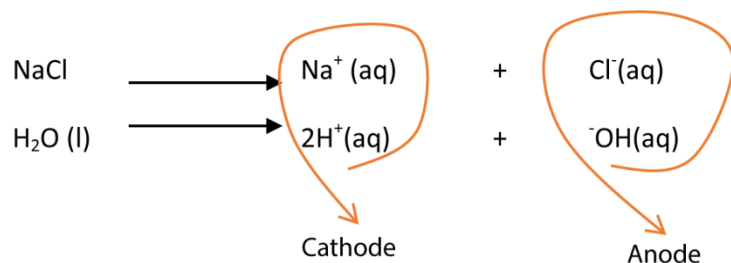


Ultimately, sodium chloride in molten form is decomposed by electrolysis to sodium metal and chlorine gas.

Selective discharge

Consider electrolysis of sodium chloride solution

Both sodium chloride and water ionize to form ions



The positively charged ion migrate to the cathode while negatively charged ions migrate to the anode.

Factors that decide the ion to be eliminated or discharged first.

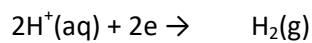
(i) **Position in electro-chemical series:**

Series for cations (**high to low**) K⁺, Ca²⁺, Na⁺, Mg²⁺, Al³⁺, [C], Zn²⁺, Fe²⁺, Pb²⁺, Cu²⁺, Ag⁺

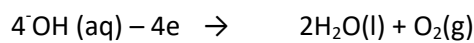
Series for anions (**high to low**) SO₄²⁻, NO₃⁻, Cl⁻, Br⁻, I⁻, -OH

The ion lower in electro-chemical series is discharged first. Consequently, during electrolysis of dilute sodium chloride, H⁺ which is low in the reactivity series than Na⁺ is discharged at the cathode whereas, OH⁻ ions are discharged at the anode.

At the cathode



At the anode

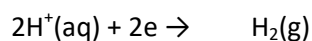


(ii) **Concentration:**

When Cl⁻, Br⁻ or I⁻ are concentrated, then they will be discharged with respect to OH⁻.

In this case electrolysis of concentrated sodium chloride solution liberates chlorine at the anode and hydrogen gas at the cathode (concentration does not affect Na⁺)

At the cathode



At the anode

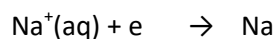


(iii) **Nature of electrode**

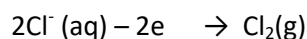
(a) When mercury cathode is used Na⁺ is discharged with preference to H⁺.

Therefore, electrolysis concentrated sodium chloride using mercury cathode liberates Na at the cathode and chlorine gas at the anode (due to high concentration of Cl⁻).

At the cathode



At the anode



Application of electrolysis

(i) **Industrial preparation of gases; Cl₂, H₂ and O₂.**

(ii) **Extraction of metal: metals above carbon in the reactivity series are almost all extracted by electrolysis.**

Example Extraction of sodium.

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Ore: molten NaCl

Anode: iron cylinder is cheap and has a melting point above the melting point of NaCl.

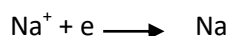
Cathode: carbon because it does not react with chloride.

Calcium chloride added to the mixture to

- lower the melting point of sodium chloride from 800°C to 600°C,
- reduce solubility of sodium in molten sodium chloride,
- Lower the corrosive vapor of sodium chloride.

Sodium is collected in dry nitrogen to protecting it from reacting with air.

At the cathode

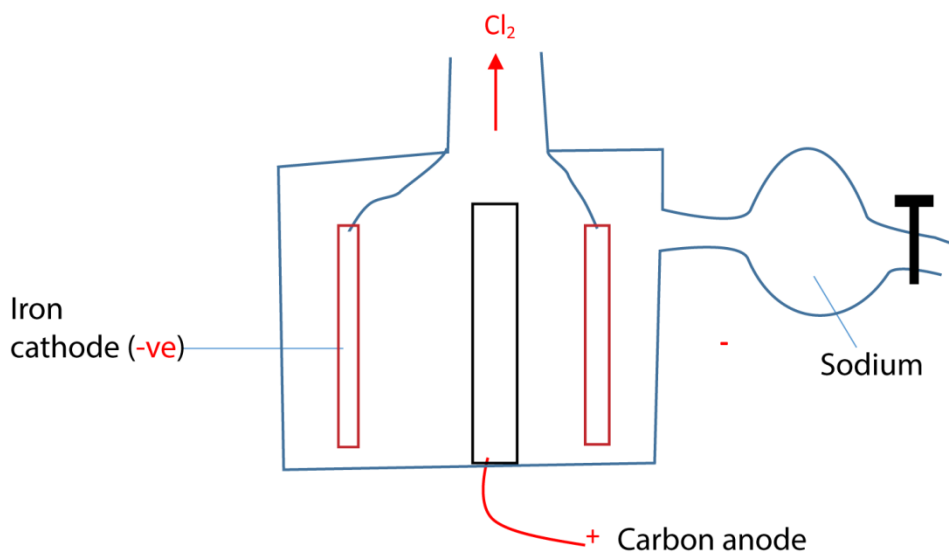


At the anode



Setup

The setup is such chlorine produced does not react with sodium



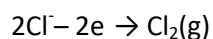
(iii) Preparation of sodium hydroxide

By electrolysis of concentrated solution of sodium chloride using carbon anode and mercury cathode. At the anode chlorine is liberated and at mercury cathode Na^+ instead of H^+ is discharged and dissolved in mercury to form mercury amalgam.

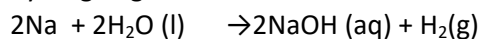
At the cathode



At the anode



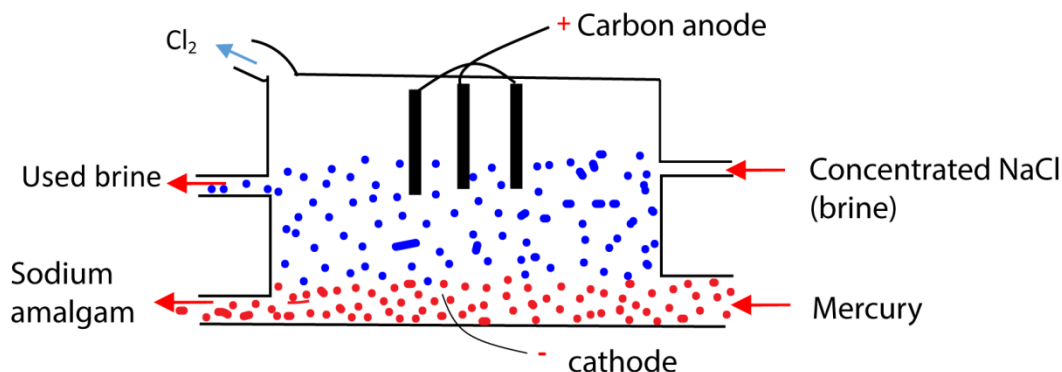
When sodium amalgam is dropped in water, sodium reacts to form sodium hydroxide and hydrogen gas.



Sodium hydroxide is concentrated to form pellets.

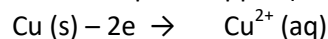
This process is disadvantageous because it releases poisonous mercury into the environment

Setup

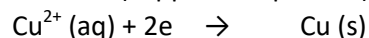


(iv) Purification of copper

Anode: impure copper (dissolves)



Cathode (copper is deposited)



Electrolyte: copper sulphate solution

Copper migrated from the anode to the cathode.

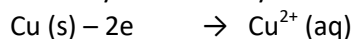
Faraday's Laws of electrolysis

1. The mass of a substance liberated at an electrode is proportional to the quantity of electricity used.
The quantity of electricity Q in coulombs = It
(I = current of electricity in amperes, t = time in second)
2. The moles of electricity required to liberate one mole of an element is proportion to the charge on its ions. (1mol of electricity = 1Faraday = $1F = 96500 \text{ C}$)

Example

Calculate the mass of copper liberated by a current of 1A for 1hour. ($\text{Cu} = 63.5$)

Quantity of electricity $Q = It = 1 \times (1 \times 60 \times 60) = 3600\text{C}$



It implies that $(96500 \times 2) \text{ C}$ is required to liberate 63.5 g of copper

Therefore, 3600C will liberate $\frac{63.5 \times 3600}{2 \times 96500} = 1.1$

Thank you
Dr. Bbosa Science