

THIRD TERM E-LEARNING NOTES

SUBJECT: CHEMISTRY

CLASS: SS 2

SCHEME OF WORK

WEEK

TOPICS

1. **Revision of last term's work.**
2. **Oxidation-Reduction (Redox) Reactions:** Oxidation and Reduction Reactions, Oxidation Numbers of Central Element in Some Compounds, Oxidizing and Reducing Agents, Balancing redox equation
3. **Electrochemical cell:** Standard electrode potential, Drawing of cell diagram and writing cell notation, e.m.f of cells, Application of Electrochemical cells.
4. **Electrolysis: Ionic theory,** Meaning of Electrolysis and Terminologies – Electrodes, Electrolytes, Electrolytic Cell , comparison of electrolytic cells and electrochemical cells and principles of electrolysis, Factors affecting the discharges of ions/ product of electrolysis, Examples of electrolysis – electrolysis of acidified water, copper(II) tetraoxosulphate(VI) and brine.
5. **Electrolysis:** Faraday's laws of electrolysis and calculations based on Faraday's laws of electrolysis, Uses of Electrolysis, Corrosion of Metals: corrosion treated as a redox process, rusting of iron and its economic costs, Prevention of rusting of iron.

6. **Hydrocarbons:** Structure and Valency of Carbon, Hydrocarbon – Meaning and Examples, Homologous Series- Characteristics, functional group, Saturated Hydrocarbon -alkanes
7. **Mid-Term Break and Holiday Assignment.**
8. **Hydrocarbons: Isomerism, Unsaturation hydrocarbon- Alkene, Unsaturation in Alkyne, Aromatic Hydrocarbon**
9. **Alkanols: Types and classes of alkanols, laboratory and Industrial Production of Alkanols, physical propertie of alkanols, chemical properties of alkanols, laboratory test and uses of alkanols.**
10. **Revision**
11. **Examination.**

REFERENCE TEXTS

1. **New School Chemistry for Senior Secondary Schools by Osei Yaw Ababio; Africana First Publishers Plc.**
2. **Chemistry for Senior Secondary Schools 2 by Magbagbeola O, et al; Melrose Books and Publishers.**
3. **Comprehensive Certificate Chemistry by G N C Ohia, et al; University Press Plc.**

WEEK 1

Revision of last term's work.

WEEK 2

TOPIC: OXIDATION – REDUCTION (REDOX) REACTION

CONTENT: 1. OXIDATION AND REDUCTION

1. REDOX REACTIONS

2. OXIDATION NUMBERS OF CENTRAL ELEMENT IN SOME COMPOUNDS

3. CONNECTION OF OXIDATION NUMBER WITH IUPAC NAME

4. OXIDISING AND REDUCING AGENTS

5. REDOX EQUATION

PERIOD 1: OXIDATION AND REDUCTION

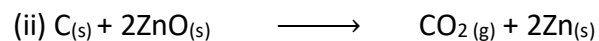
Redox is a short form for reduction and oxidation reactions. The two reactions are opposing and complementary and they occur simultaneously. Redox has the following definitions

1. Definition of Oxidation reactions with examples

(a) In terms of addition of oxygen: Oxidation is the addition of oxygen or removal of hydrogen from a substance. Any reaction where there is addition of oxygen to a reactant is regarded as Oxidation.

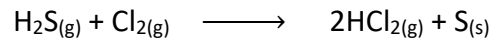
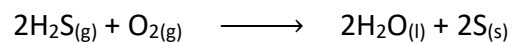
Some examples are:





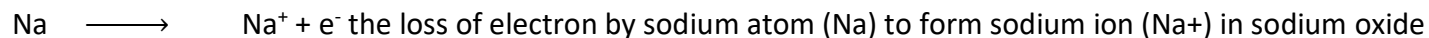
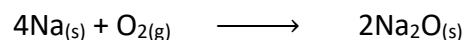
In this reaction, oxygen (O) was added to carbon (C) to form Carbon (iv) oxide (CO₂). That is carbon was oxidized to Carbon (iv) oxide.

(b) In terms of hydrogen removal: Oxidation is the removal of hydrogen from a compound e.g. (i) combustion of hydrogen sulphide



In both reactions, hydrogen sulphide was oxidized to Sulphur. The second reaction shows that oxidation can occur without the involvement of oxygen

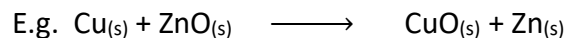
(c) In terms of loss of electron: Oxidation is the loss of electron.



2. Definition of reduction with examples

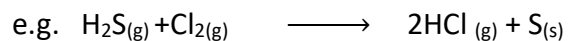
Definitions of reduction are the opposites of the definitions of oxidation.

(a) Removal of oxygen



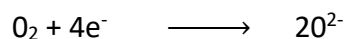
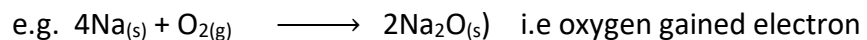
In this reaction zinc oxide was reduced to zinc

(b) Addition of hydrogen

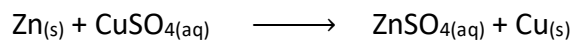


chlorine is reduced to hydrogen chloride.

(c) Gain of electrons



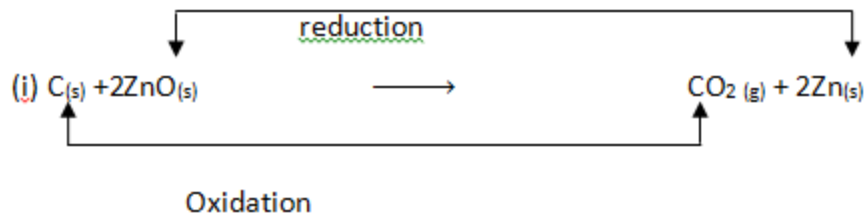
(d) In terms Of Decrease In Oxidation Number



The oxidation number of copper decreased from +2 in copper (ii) salt to zero in copper.

(e) Reduction occurs when electro negative element is added

3. Redox Reaction And Electron Transfer: considering all the reactions used as examples so far, it would be seen that both oxidation and reactions occur together see illustration below:

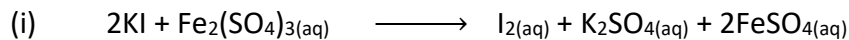


Because $2Na \longrightarrow 2Na^+ + 2e^-$ (loss of electron is oxidation)

$O_2 + 4e^- \longrightarrow 2O^{2-}$ (gain of electron is reduction)

The above electron transfer equations can be written for most redox reactions.

More examples are



Ionically

$2I^-_{(aq)} \longrightarrow I_{2(aq)} + 2e^-$ (loss of electrons in oxidation)

The above equation is the redox reaction .Each half reaction cannot occur alone because an atom cannot receive or donate electron without another atom around that must donate or receive the electron.

Evidence of electron transfer in redox reactions can be proved using a simple electro chemical cell.

EVALUATION

1. Define reduction in terms of electron transfer.
2. Explain the experiment which proves that there is electron transfer in redox reaction

PERIO 2: OXIDATION NUMBERS OF CENTRAL ELEMENTS SOME COMPOUNDS

Oxidation number is the electrical charge assigned to an atom in accordance with some prescribed set of rules

RULES FOR DETERMING OXIDATION NUMBER

The following sets of rules are used to determine the oxidation state or number of substances.

1. The oxidation number of an element in an un-combined state is zero, for example the oxidation number of Hydrogen, Oxygen and Sodium atom in a free state or un-combined with another element is zero.
2. In most compounds containing hydrogen, the oxidation number of hydrogen is +1 except in hybrids where it is -1
3. Electrons shared between two unlike atoms are counted with the more electro negative atom. For example, in water molecule the electron are regarded as being with the more electronegative oxygen. Thus in H_2O , each hydrogen atom is in +1 oxidation state, and the oxygen atom is in -2 oxidation state.

- In most compounds containing oxygen, the oxidation number of each oxygen atom is -2 except in peroxides where it is -1. E.g. hydrogen peroxides, (H₂O₂), sodium peroxide (Na₂O₂), barium peroxide (BaO₂)
- The oxidation number of each halogen is -1, except when bonded with fluorine which is the most electro negative. For instance in IF₇, each fluorine atom is in oxidation state of -1 and iodine is in oxidation state of +7
- The sum of all the oxidation numbers of elements in a compound is zero and with this simple relationship, the oxidation number of each element in a compound can be calculated.
- In simple ions, i.e. ions containing one atom, the oxidation number is equal to the charge on the ion. For example the ion Al³⁺ has an oxidation number of +3, the ion Cu²⁺ has an oxidation number of -2
- What about complex ion? In a complex ion (i.e. ion consisting of more than one element) the oxidation number is the algebraic sum of all the oxidation numbers of all the elements in the ion. This will be the sign on the ion and of the same size. For example, in tetraoxosulphate (vi) ion (SO₄²⁻) the overall charge is -2 which its oxidation number is. It is obtained as follows:

$$(Oxidation\ number\ of\ S) + 4(Oxidation\ number\ of\ O) = (Oxidation\ number\ of\ SO_4^{2-})$$

$$+6 \qquad +4 \times (-2) \qquad = \qquad -2$$

For other ions such as OH⁻, NO₃⁻, NO₂⁻, SO₃²⁻, PO₄³⁻, NH₄⁺ their oxidation numbers are -1, -1, -2, -3 and +1 respectively.

Calculation of oxidation number

With the rules for determining oxidation numbers in our memory, it is possible to calculate the oxidation number of any given element in an ion or a compound .

Worked examples

1. Calculate the oxidation number of chromium in CrO_4^{2-}

Solution

This can be solved by simple linear equation by making the unknown subject of formula. We have 4 oxygen and its oxidation is -2

Let the unknown (oxidation number of Cr be x)

But $CrO_4^{2-} = -2$ (because the sign on the ion is -2)

$$x + (-24) = -2$$

$$x - 8 = -2$$

$$x = -2 + 8$$

$$x = +6$$

Thus the oxidation number of Cr in CrO_4^{2-} is +6

Note: oxidation number is never written as a neutral number, i.e. it is either written as a neutral number, i.e. It is either written as a positive or negative number.

2. Calculate the oxidation number of sulphur in H_2SO_4

Solution

The oxidation number (O.N) of hydrogen is +1 and the two hydrogen atoms will give +2. The four oxygen atoms will give $(-2 \times 4) = -8$, since -2 is the oxidation number (O, N) of an atom of oxygen

$$\therefore 2 \times (+1) + x + (4 \times -2) = 0$$

$$+ 2 + x + -8 = 0$$

$$x = 0 + 8 - 2$$

$$x = +8 - 2$$

$$x = +6$$

CONNECTION OF OXIDATION NUMBERS WITH IUPAC NAMES

From all the examples considered so far, it can be seen that the names of most inorganic compounds reflect.

The oxidation numbers of the elements with variable oxidation numbers e.g. iron (ii) ion, in this, oxidation number of Fe is +2

1. the number of oxygen atoms in the compound e.g.

CO_3^{2-} is trioxo (i.e. three oxygen atoms)

SO_4^{2-} is tetraoxo (i.e. four oxygen atoms)

$\text{Cr}_2\text{O}_7^{2-}$ is heptaoxo (i.e. seven oxygen atoms)

Hence, if the formula of a substance is known, the IUPAC name can be written by calculating the oxidation number of the central atom using the fixed oxidation numbers of the other atoms like oxygen, hydrogen and halogen.

For example:

KMnO_4 has oxidation number $\text{K} = +1$ and $\text{O}_4 = 4 \times -2$

i.e. $+1 + x + -8 = 0$ where x is the oxidation number of Mn , therefore $x = 0 + 8 - 1 = +7$. Therefore, the name is potassium tetraoxomanganate(VII)

The Following Prefixes and Suffixes Should Be Noted

PREFIX	SUFFIX
1= mono-	SO_3^{2-} = sulphate (IV)
2= di-	SO_4^{2-} = sulphate (VI)
3.= tri -	MnO_4^- =manganate (VII)
4. =tetra-	PO_4^{3-} = phosphate(V)
5= penta	ClO_3^- =Chlorate(V)
6= hexa-	CO_3^{2-} = Carbonate(IV)
7= hepta –	$\text{Cr}_2\text{O}_7^{2-}$ = dichromate (vi)

8. =octa-

NO_3^- = nitrate (v)

9.= nona –

BrO^- = bromate (i)

10. = deca

IO^- =iodate (i)

IO_3^- - iodate (V)

. ClO^- = Chlorate (V)

Worked examples

1. Calculate the oxidation number of manganese in $KMnO_4$ and hence name the compound in IUPAC nomenclature

Solution

Let unknown (oxidation number of Mn) = x

K= +1, O = -2

$KMnO_4 = 0$

$1 + x + (-2 \times 4) = 0$

$1 + x - 8 = 0$

$x - 8 = -1$

$x = -1 + 8$

$$X = +7$$

The oxidation number of manganese is +7 and the name of the compound is potassium tetraoxomanganate(vii) . The tetraoxo means there are 4 oxygen atoms in the compound.

2. Calculate the oxidation state of chromium in $Cr_2O_7^{2-}$ and give its IUPAC name

Let the unknown = x

$$O = -2 \text{ and } Cr_2O_7^{2-} = -2$$

$$2x + (-2 \times 7) = -2$$

$$2x - 14 = -2$$

$$2x = -2 + 14$$

$$2x = +12$$

$$X = +6$$

The oxidation state of chromium is +6

There are two chromium atoms and there are seven oxygen atoms, there is a charge on the compound making it an ion. Thus, the name is heptaoxodichromate(VI) ion.

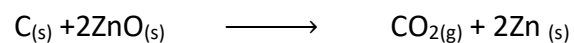
EVALUATION

1. What is the oxidation number of manganese in each of the following species? (i) MnCl_2 (b) MnO_2 (c) MnO_4^-
2. Determine the oxidation number of chromium in $\text{Cr}_2\text{O}_7^{2-}$, then name in IUPAC.

PERIOD 3: OXIDIZING AND REDUCING AGENTS

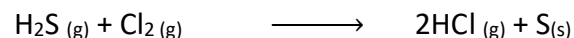
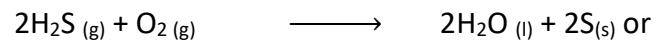
A. Oxidising agent: This is a substance which adds oxygen or removes hydrogen from a substance or a substance which accepts electron, i.e. electron acceptor. Consider the following reactions.

1. **Oxidation in terms of addition of oxygen:** oxidizing agent will add oxygen to a substance e.g.



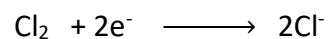
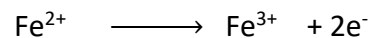
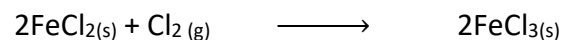
ZnO is the oxidising agent because it added oxygen to (i.e. oxidise) carbon to form CO_2

2. **Oxidation in term of removal of hydrogen:** oxidizing agent will add oxygen to a substance e.g



Oxygen and chlorine are the oxidizing agents because they removed hydrogen from H_2S to form sulphur

3. **Oxidation in term of loss of electrons:** oxidizing agent will gain electrons e.g



The oxidizing agent is chlorine (Cl_2) because it gained electrons lost by Fe^{2+}

Other means of identifying an oxidising agent is the substance.

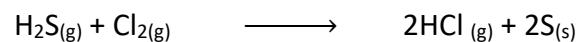
(i) Which is reduced

(ii) Whose oxidation number has decreased

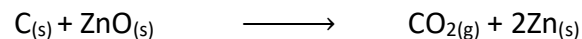
B. Reducing agent: This is a substance which adds hydrogen to another substance or removes oxygen from that substance or this substance which donates electron, i.e. an electron donor.

Consider the following reactions

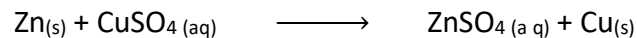
1. Adds hydrogen e.g hydrogen sulphide reduces chloride to hydrogen chloride in



2. Removes oxygen e.g. Carbon in



3. Has its oxidation number increases e.g. in



Zinc is the reducing agent because its oxidation number increased from zero in zinc to +2 in zinc tetraoxosulphate (vi) that is, it was oxidized.

The Table below give some common reducing agent and oxidizing agents.

OXIDIZING AGENTS

Oxygen	O ₂
Hydrogen peroxide	H ₂ O ₂
Chlorine (and other halogen)	Cl ₂

Acidified potassium tetraoxomanganate (viii)	KMnO_4
Acidified potassium heptaoxchromate (vi)	$\text{K}_2\text{Cr}_2\text{O}_7$
Hot concentrated tetraoxsulphate (vi) acid	H_2SO_4
Concentrated trioxonitrate(v) acid	HNO_3
Silver salt (and metals low in the series e.g cu)	Ag^+

REDUCING AGENTS

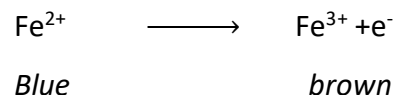
Hydrogen	H_2
Carbon	C
Potassium	KI
Hydrogen sulphide	H_2S
Sulphur(iv) oxide	SO_2
Iron(II) salt	SO_2
Sodium (and other reactive metals)	Na
Ammonia	NH_3

Generally, metals that are highly electro positive e.g. Na, K and Mg lose electrons easily and so are good reducing agents (they release their electrons to another thus reducing it). Non-metals that are highly electro negative e.g. O, Cl and F are good oxidizing agents (they easily accept electrons from another thus oxidizing it).

TEST FOR OXIDIZING AGENTS

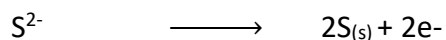
Strong reducing agents like iron (ii) salt and hydrogen sulphide are used.

1. Green aqueous solution of iron (ii) salt turns brown in the presence of an oxidising agent .



Note: Freshly prepared iron(II) salt is used.

- (ii) Hydrogen sulphide forms deposit of sulphur when passed through the solution of an oxidising agent.



TEST FOR REDUCING AGENTS

Strong oxidizing agents like acidified potassium tetraoxomagnate(viii) and acidified potassium heptaoxo chromate (vi) are used

- i. The purple colour of acidified potassium tetraoxomanganate (vii) is turned colourless in the presence of a reducing agent.

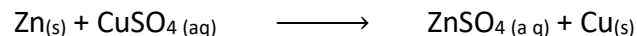


The orange colour of acidified potassium heptaoxodichromate(VI) is turned green in the presence of a reducing agent.



EVALUATION:

1. define reducing agent in terms of oxidation number.
2. Identify the oxidizing and reducing agent in the reaction below:



PERIOD 4: BALANCING IONIC (REDOX) EQUATION

In balancing ionic equations, the following steps can be taken.

- i. Examine the given equation carefully
- ii. Balance the number of atoms first by adding the correct number of H^+ and H_2O on the appropriate sides of the equation when an acid is involved.
- iii. Next, balance the number of charges by adding appropriate number of electrons on the right side of the equation.

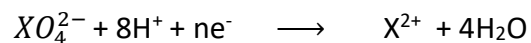
- iv. Lastly, count the atoms on both sides of the equation. If they are equal, then the equation is balanced in atoms (or molecules); also check the charge; if equal, then it is balanced ironically too.

WORKED EXAMPLES

1. What is the value of n in the following equation



Solution

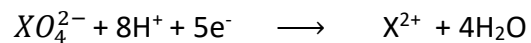


In the given equation, the atoms are balanced but the charges are not. On the left of the equation, we have a net charge of

-1 on $XO_4^{2-} + 8H^+ + ne^-$, i.e. $-1 + 8 = +7$

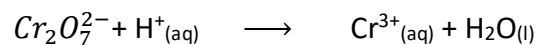
On the right hand side, we have just +2 which is on X^{2+} . Then, for these charges to be equal on both sides of the equation, we have to add 5 electrons which will neutralize 5 positive charges from the +7, thereby reducing it to +2, which balances the equation below

The balanced equation is



WORKED EXAMPLE 2

Balance the following Redox equation:



Solution

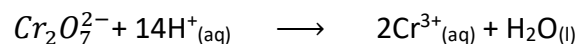
The first step is to balance the atoms before the charges.

On the left hand sides (LHS) there are 7 oxygen atoms on

CrO_7^{2-} , but on the right hand side (RHS), there is only one oxygen atom, the oxygen is balanced by writing 7 in front (coefficient) of H_2O

On the same LHS, there are 2 atoms of chromium, but one on RHS; write 2 in front of the Cr^{3+} to balance it.

The 7 in front of H_2O (i.e. $7\text{H}_2\text{O}$) now on RHS, gives 14 atoms of hydrogen, but only one on the LHS. Then write 14 as hydrogen, but only one on the LHS. Then write 14 as a coefficient of H^+ , this gives the equation below which is balanced in terms of atom, but not in charges yet



Now to balance the charges

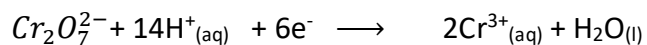
LHS: -2 on $\text{Cr}_2\text{O}_7^{2-}$, +14 on 14H^+

i.e. $-2 + 14 = +12$ net charge

RHS: $+3 \times 2 = +6$ net charge

Add 6 electrons to neutralize the 6 excess positive charges on LHS.

This gives the balanced equation below:



EVALUATION

1. Define reduction in terms of electron transfer
2. Using appropriate examples, explain (i) oxidizing agent (ii) reducing agent
3. What is the oxidation number of iron in FeO_3 ?
4. What is the oxidation number of nitrogen in $\text{Al}(\text{NO}_3)_3$?
5. What is the value of x in the following equation?



GENERAL EVALUATION

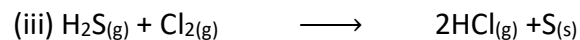
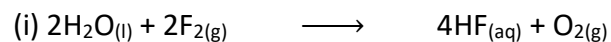
OBJECTIVE TEST:

1. The oxidation number of manganese in the reaction: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \longrightarrow \text{Mn}^{2+}_{(\text{aq})} + 4\text{H}_2\text{O}_{(\text{l})}$
(a) +3 to +5 (b) +7 to +2 (c) +3 to +2 (d) +7 to +5
2. $\text{Fe}_2\text{O}_3_{(\text{s})} + 3\text{CO}_{(\text{g})} \longrightarrow 2\text{Fe}_{(\text{s})} + 3\text{CO}_2_{(\text{g})}$ the oxidizing agent is (a) Fe_2O_3 (b) CO (c) Fe (d) CO_2
3. In which of the following is oxidation number of sulphur +2? (a) H_2SO_4 (b) H_2SO_3 (c) SO_2 (d) H_2S

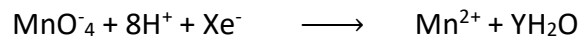
4. Which of the following can act as both oxidizing and reducing agent? (a) KMnO_4 (b) H_2O_2 (c) SO_2 (d) H_2S
5. Which of these does not define reduction? (a) it is loss of electron (b) it is addition of hydrogen (c) it is addition of electrons (d) decrease in oxidation number

ESSAY QUESTIONS

1. List the oxidising agent in the following

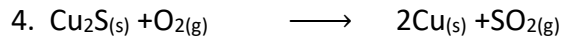
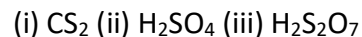


2. Consider the following equation:



State the (i) Values of x and y (ii) oxidation state of Mn in MnO_4^-

3. Calculate the oxidation number or state of sulphur in the following compounds and give their IUPAC names.



What is the change in oxidation number of copper in reaction?

5. What is the oxidation number of manganese in each of the following species? (i) MnCl_2 (ii) MnO_2 (iii) MnO_4^-

WEEKEND ASSIGNMENT

Read comprehensive chemistry by G N C. Ohia et al

PRE READING ASSIGNMENT

Read about ionic theory from Melrose chemistry for senior secondary 2 pages 111 to 116

WEEKEND ACTIVITY

Explain the experiment which proves that there is electron transfer in redox reactions

REFERENCE TEXTS

Comprehensive Chemistry by G N C Ohia et al

WEEK 3

TOPIC: ELECTROCHEMICAL CELL:

1. STANDARD ELECTRODE POTENTIAL,
2. DRAWING OF CELL DIAGRAM AND WRITING CELL NOTATION,
3. E.M.F OF CELLS,
4. APPLICATION OF ELECTROCHEMICAL CELLS.

PERIOD 1: STANDARD ELECTRODE POTENTIAL,

ELECTROCHEMICAL CELLS

Electrochemical cell is a device that converts chemical energy to electrical energy.

SOME PARAMETERS ASSOCIATED WITH ELECTROCHEMICAL CELL

1. Standard hydrogen electrode: The standard hydrogen electrode consists of hydrogen gas at 25°C and one atmospheric pressure bubbling around an inert platinum electrode in contact with one-molar solution of hydrogen ions i.e $H_{(aq)}^+/H_{2(g)}$ system. It is the reference standard by which the standard electrode potential of a metal ions/metal system is measured against. The standard hydrogen electrode potential has been given assigned a value of zero.

2. Standard electrode potential (E°): Standard electrode potential of a metal ions/metal system is the potential difference set up between the metal and one-molar solution of its ions at 25°C (arbitrarily taking the standard electrode potential of hydrogen ions/hydrogen gas system as zero volts.

3. Salt bridge: It is a piece of paper soaked in potassium chloride solution or ammonium chloride solution. Its functions are to complete the electric circuit and enable the movement of ions from one half-cell to another so as to maintain electrical neutrality in the solution.

MEASUREMENT OF STANDARD ELECTRODE POTENTIAL

In measuring the standard electrode potential of particular metal ion/metal system, the half-cell of the metal ions/metal system e.g $Zn_{(aq)}^{2+}/Zn_{(s)}$ is connected to that of hydrogen ions/hydrogen gas system by a salt bridge. The reading shown on the voltmeter is the electromotive force or the e.m.f of the cell. Since the standard electrode potential of hydrogen ions/hydrogen gas system is assigned the value zero, the value obtained in the standard electrode potential of the metal system.

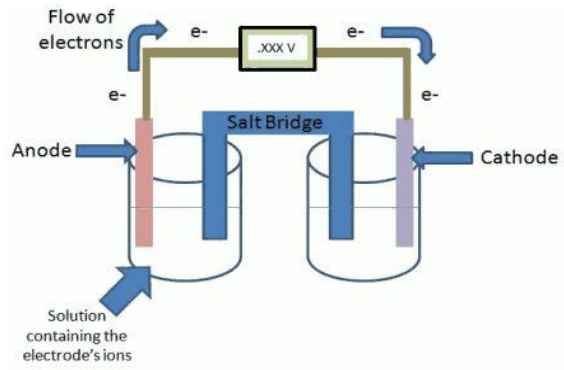
The sign of the standard electrode potential of a given metal system is dependent of the direction of the electron flow in the cell. If electron flows from hydrogen ions /hydrogen gas system to the the metal ions/metal system, the standard electrode potential of that metal system is assigned a negative value.If the flow of electron is in the reverse direction, it is assigned a positive value. For example, the standard electrode potential of zinc ions/zinc metal system is -0.76v because during the measurement, the flow of electrons is from hydrogen system to zinc system. $E^\circ Zn_{(aq)}^{2+}/Zn_{(s)} = -0.76\text{v}$

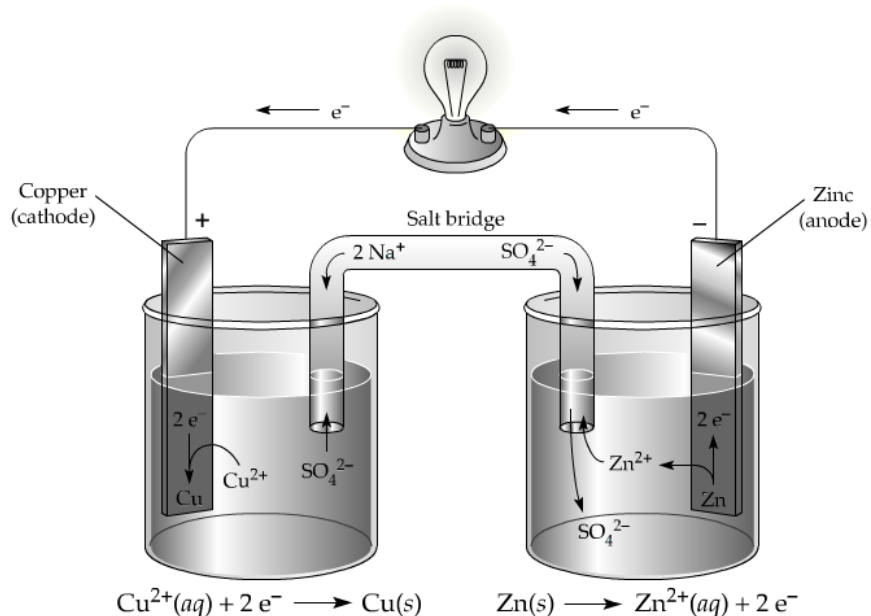
$$E^\circ Cu_{(aq)}^{2+}/Cu_{(s)} = +0.34$$

EVALUATION:

- 1.What is standard electrode potential?
2. What is the significance of salt bridge?
3. Explain how the standard electrode potential of silver system can be measured..

PERIOD 2: DRAWING OF CELL DIAGRAM AND WRITING CELL NOTATION,



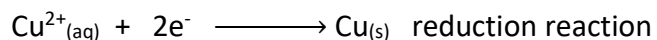
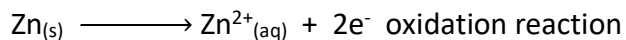


When two half – cells with different electrode potentials are connected through a salt-bridge, they form a cell called **Electrochemical cell**. The electrochemical is also called a **Voltaic cell or Galvanic cell**, and it is a device in which chemical energy is converted to electrical energy. In the electrochemical cell oxidation occurs in one –half –cell while reduction occurs in the other half –cell. The overall redox reaction results in flow of electrons or electric current. For example if the $\text{Zn}^{2+}(\text{aq}) / \text{Zn}(\text{s})$ half –cell is connected to $\text{Cu}^{2+}(\text{aq}) / \text{Cu}(\text{s})$ half cell through a salt bridge with the $\text{Cu}^{2+} / \text{Cu}(\text{s})$ made the right hand electrode, the cell produced is called a Daniel cell. In this cell, oxidation occurs at zinc electrode, this electrode is then negative electrode of anode. Reduction occurs at the copper electrode which functions as positive electrode or cathode as shown ion fig 7 presents the diagram of Daniel cell. The diagram is written in a short – hand form as $\text{Zn}(\text{s}) / \text{Zn}^{2+}(\text{aq}) // \text{Cu}^{2+}(\text{aq}) / \text{Cu}(\text{s})$. A n electric flows through the wire connecting the two metals or

electrons outside the solution. In this system, chemical energy is converted to electrical energy. In the fig7, Zinc atoms at the Zinc electrode lose two electrons each (oxidation) to form zinc ions which go into solution. In doing so, the zinc electrode becomes negatively charged and hence behaves as a negative electrode or anode. The excess electrons on the zinc electrode move along the wire to the electron deficient copper (II) ions. The copper (II) ions thus becomes reduced by gaining two electrons each to form metallic copper, which deposits itself on the copper electrode. The copper electrode becomes positively charged and behaves as positive electrode or anode. It is important to distinguish between the electrochemical cells and electric cell. In both, oxidation always occurs at the anode and reduction at the cathode. In an electrochemical cell, the negative electrode is the anode while the positive electrode is the cathode. In electrolytic cell on the other hand, the negative electrode is the cathode while the positive electrode is the anode.

CELL NOTATION

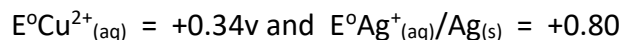
$Zn_{(s)} / Zn^{2+}_{(aq)} // Cu^{2+}_{(aq)} / Cu_{(s)}$. In the cell notation above, the single represent phase boundary between the metal ions(solution) and the metal (electrode) while the double lines represent the salt bridge.



ION: ion can be defined as element which has either gained or lost electron.

EVALUATION:

Draw an electrochemical cell to represent the following:



PERIOD 3 AND 4: E.M.F OF CELLS AND APPLICATION OF ELECTROCHEMICAL CELLS

The e.m.f of an electrochemical cell is the algebraic difference between between the standard electrode potentials of the two metallic electrodes i.e $e.m.f = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$. A big e,m,f means a powerful cell. if we need a powerful cell the metals which serve as electrode should be far apart. The sign of e.m.f tells whether the reaction is spontaneous or not. A positive e.m.f shows that the reaction in the cell is spontaneous and so such cell can work otherwise it is not workable.

RELATIONSHIP BETWEEN E.M.F AND ΔG

$$\Delta G = -nFE$$

Where n = number of moles of electron transfer, F = Faraday constant.

A negative value of ΔG shows the reaction is spontaneous

Solved examples

Find the e.m.f and ΔG of the cell represent by the half-cell below:

$$E^{\circ}_{\text{Cu}^{2+}(\text{aq})} = +0.34\text{v} \text{ and } E^{\circ}_{\text{Ag}^{+}(\text{aq})/\text{Ag}(\text{s})} = +0.80$$

Solution:

$$e.m.f = e.m.f = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

$$= +0.80 - 0.34$$

$$= +0.46\text{v}$$

$$\Delta G = -nFE$$

$$= 2 \times 96500 \times 0.46$$

$$= -88780 \text{ J}$$

APPLICATIONS OF ELECTROCHEMICAL CELLS

Examples of electrochemical cells are Daniell cell, Leclanche cell and lead acid accumulator (car battery). These cells are used in power generation..

Primary cells are not rechargeable. This because, the reactions in these cells are irreversible. Daniell and Leclanche cells are primary cells. Once they are discharged, they cannot be used again.

Secondary cells are rechargeable. The chemical reactions in these cells are reversible. They can be recharged after use. An example of this kind of cell is car battery. The batteries in our handset, laptops are secondary cells

EVALUATION:

1. Distinguish between a primary cell and a secondary cell.
2. Give four examples of a primary cell
3. Describe the working principle of a Leclanche cell.

WEEK 4

TOPIC: ELECTROLYSIS: IONIC THEORY

CONTENT:

1. MEANING OF ELECTROLYSIS AND TERMINOLOGIES – ELECTRODES, ELECTROLYTES, ELECTROLYTIC CELL
2. COMPARISON OF ELECTROLYTIC CELLS AND ELECTROCHEMICAL CELLS AND PRINCIPLES OF ELECTROLYSIS
3. FACTORS AFFECTING THE DISCHARGES OF IONS/ PRODUCT OF ELECTROLYSIS,
4. EXAMPLES OF ELECTROLYSIS – ELECTROLYSIS OF ACIDIFIED WATER, COPPER (II) TETRAOXOSULPHATE (VI) AND BRINE.

IONIC THEORY

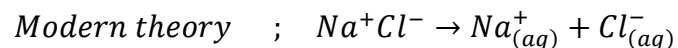
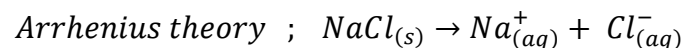
Ionic theory was used to explain the behaviour of electrolytes when electric current is passed through their solution.

A Swedish chemist, Swante Arrhenius (1887), was the first to present the ionic theory to describe electrolysis. The theory proposed that when an electrolyte is melted or dissolved in water, some if not all of the molecules of the substance dissociate into freely-moving charge ions i.e. positive ions (called cations) and negative ions (anions). The process of dissociation into ions is called ***ionization***.

When an electric current is passed through an electrolyte, the free ions lose their random movement. The positive ion becomes attracted to the negative electrode (called the cathode) while the negative ions move toward the positive electrode (called anode). It is this ionization and mobility of ions that constitute current flow.

Arrhenius version of the ionic theory has been modified since the X-ray diffraction studies show that salts and strong alkalis consist of oppositely charged ions even in solid state. The modern theory proposes that ions in such solid state are pulled away from

one another either as a result of heat energy applied when the solid melts or with the help of the solvent molecules when the solid dissolves. Thus,



PERIOD 1: MEANING OF ELECTROLYSIS

Electrolysis is the process whereby chemical decomposition occurs when an electric current is passed through an electrolyte.

TERMINOLOGIES

1. Electrode: Electrodes are conductors in the form of wires rods or plates through which an electric current enters or leaves the electrolyte. We have the cathode and the anode

Note: In an electrolytic cell, the negative electrode is known as the cathode while the positive electrode is known as the anode.

In an electrochemical cell, the negative electrode is the anode while the positive electrode is the cathode.

2. ELECTROLYTES

An electrolyte is a compound, in a molten or aqueous state which can conduct electricity and is decomposed in the process.

Electrolytes conduct electric currents by movement of ions. Examples of electrolytes include NaCl, CuCl₂, CuSO₄, HCl, H₂SO₄ etc.

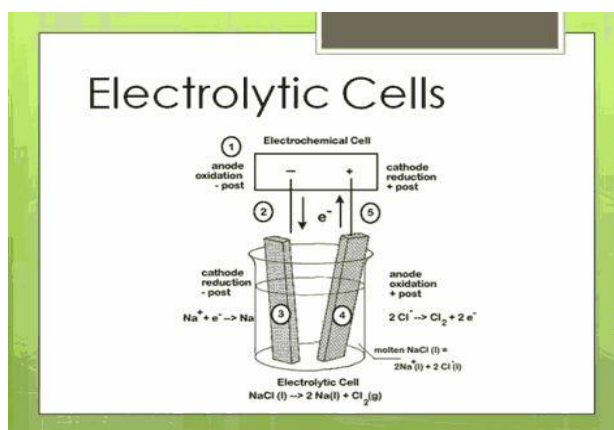
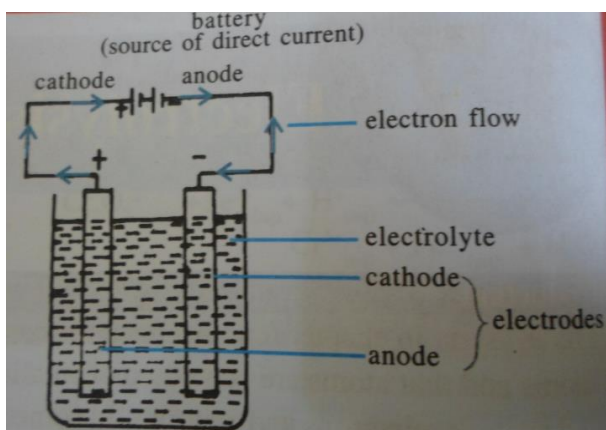
Electrolytes may be weak or strong:

(a) Strong electrolyte ionizes completely and can conduct large electric current since they are composed entirely of free moving ions. Examples are strong acids , strong bases and salts

(b) Weak electrolytes are partially ionized and conduct electricity poorly because they contain a low concentration of ions. Examples are weak acids and weak bases

3. Non-electrolytes do not conduct electricity since they do not ionize. **Examples** are urea, ethane, benzene, trichloromethane, cane sugar, ether, tetra chloromethane.

4. An electrolytic cell is an assembly of two electrodes in an electrolyte and is used for the electrolysis of a substance.



EVALUATION

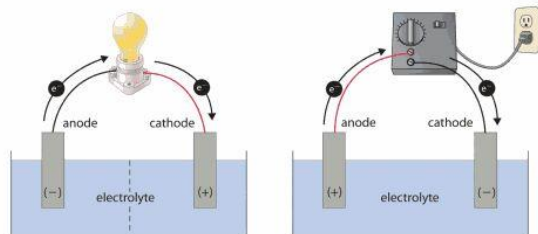
Write short note on the following terms:

1. Electrolysis
2. Electrolyte
3. Electrodes
4. Weak electrolyte

5. Strong electrolyte

6. Non-electrolyte

PERIOD 2: COMPARISON OF ELECTROLYTIC CELLS AND ELECTROCHEMICAL CELLS AND PRINCIPLES OF ELECTROLYSIS



GALVANIC CELL

Energy released by spontaneous redox reaction is converted to electrical energy.

Oxidation half-reaction:
 $Y \rightarrow Y^+ + e^-$

Reduction half-reaction:
 $Z + e^- \rightarrow Z^-$

Overall cell reaction:
 $Y + Z \rightarrow Y^+ + Z^- \quad (G < 0)$

ELECTROLYTIC CELL

Electrical energy is used to drive nonspontaneous redox reaction.

Oxidation half-reaction:
 $Z^- \rightarrow Z + e^-$

Reduction half-reaction:
 $Y^+ + e^- \rightarrow Y$

Overall cell reaction:
 $Y^+ + Z^- \rightarrow Y + Z \quad (G > 0)$

COMPARISON OF ELECTROLYTIC CELL AND ELECTROCHEMICAL CELL

	ELETROLYTIC CELL	ELETROCHEICAL CELL
1.	Convert electrical energy to chemical energy i.e. electric current is required for the reaction to occur.	Convert chemical energy to electrical energy i.e. chemical reaction produces electrical energy.
2.	Electrons are forced or pushed by an	Electrons are generated by oxidation at an

	outside source such as a battery. The process is non-spontaneous	electrode (anode). The process is spontaneous
3.	Cathode is the negative electrode while anode is the positive electrode	Cathode is the positive electrode while anode is the negative electrode
4.	Electrodes are in the same compartment	Electrodes are in separate compartments
5.	Salt bridge is not needed	Salt bridge is required
6.	Only one electrode is used	Two electrodes are used

The table above gives the salient differences between an electrolytic cell and electrochemical cell.

PRINCIPLES OF ELECTROLYSIS

Cations and anions

An electrolyte composed of positive and negative ions.

During electrolysis, electrons move from the negative pole of the battery to the cathode, making it to be electron-rich and negatively charged. Hence, positive ions migrate to the cathode; and are called cations. The anode becomes electron-poor and positively charged. Hence, negatively charged ions migrate to the anode, and are called anions.

PERIOD 3: FACTORS AFFECTING THE DISCHARGES OF IONS/ PRODUCT OF ELECTROLYSIS

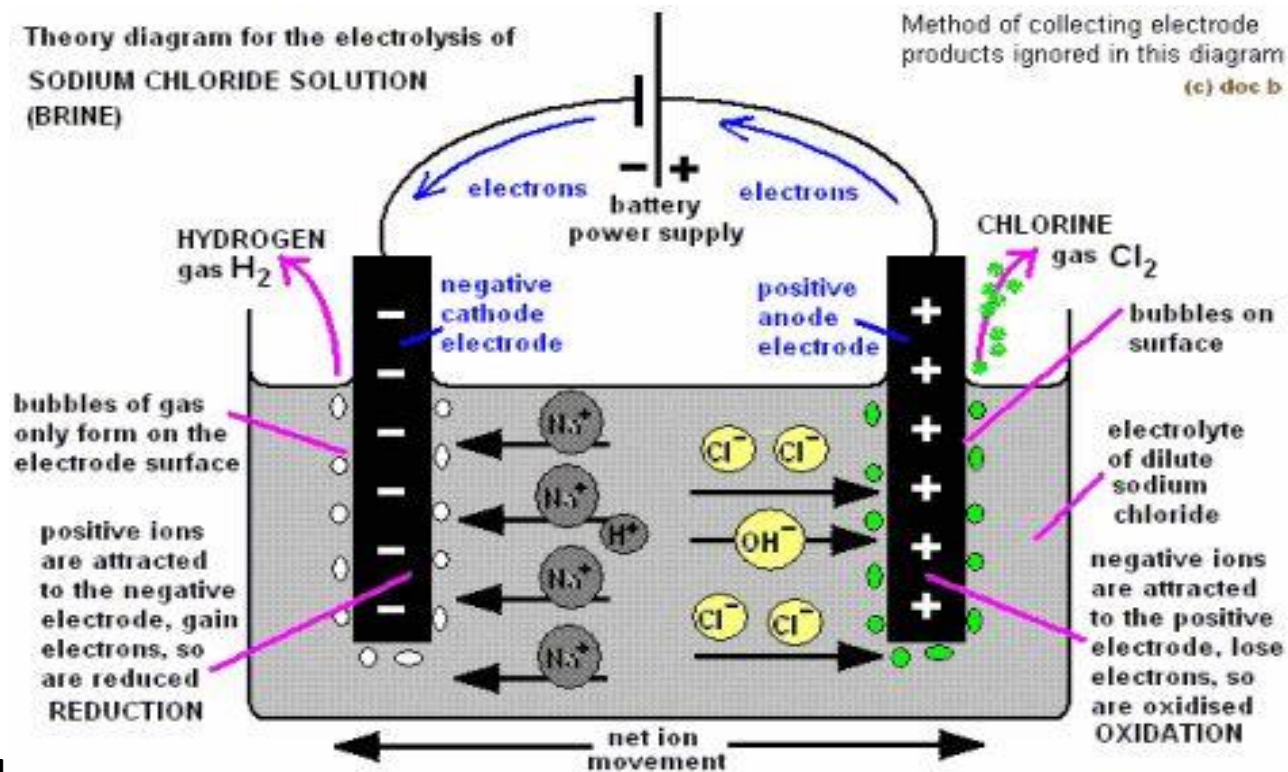
During electrolysis the ions in the electrolyte are positively and negatively charged. The negatively charge ion will migrate to the negative charged cathode. In order to know which of the ions will be preferentially discharged at the electrodes, the following factors must be taken into consideration.

1. The position of ions in the electrochemical series

CATIONS		ANIONES	
K ⁺	If two cations are present in an electrolyte the cation below the other one in the electrochemical series will be preferentially discharged.	If two anions are present in electrolyte, the anions that is above the other one in the electrochemical series will be preferentially discharged	OH ⁻
Na ⁺			I ⁻
Ca ²⁺			Br ⁻
Mg ²⁺			Cl ⁻
Al ³⁺			
Zn ²⁺			
Fe ²⁺			NO ₃
Sn ²⁺			
Pb ²⁺			SO ₄ ²⁻
H ⁺			
Cu ²⁺			F ⁻
Hg ²⁺			
Ag ⁺			
AU ⁺			
Pt ⁺			

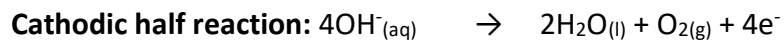
ELECTROLYSIS OF SODIUM CHLORIDE SOLUTION USING INERT ELECTRODES

Note: Inert electrodes are the electrodes that will tamper with the electrolysis. Examples are carbon (graphite) and platinum.

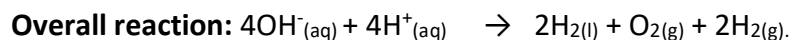
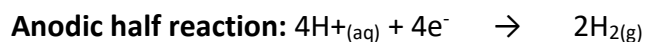


DIAGRAM

At the anode: Both OH^- and Cl^- will migrate to the anode where OH^- ions will be preferentially discharged because they are above Cl^- in the electrochemical.



At the cathode: Both Na^+ ions and H^+ ions migrate to the cathode where H^+ ions are preferentially discharged because they are below Na^+ ions in the electrochemical series.



Products of electrolysis: Hydrogen gas is produced at the cathode and oxygen gas is produced at the anode.

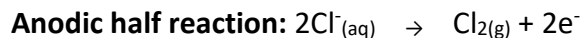
Effect of electrolyte on the electrolysis: At the end of the electrolysis, the electrolyte will more concentrated since OH^- and H^+ have been removed from the electrolyte.

2. CONCENTRATION OF IONS IN THE ELECTROLYTE e.g. electrolysis of concentrated $\text{NaCl}_{(\text{aq})}$ using inert electrodes. The electrolytic cell is the same as in number factor above, but the electrolyte is concentrated $\text{NaCl}_{(\text{aq})}$

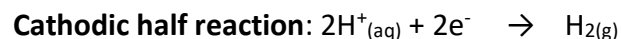
Ions in the electrolyte:

CATIONS	ANIONS
Na^+	OH^-
H^+	Cl^-

At the anode: Both, Cl^- and OH^- will migrate to the anode where Cl^- ions will be preferentially discharged because they are more concentrated. (i.e they are more in concentration inside the electrolyte).



At the cathode: Both the Na^+ and H^+ will migrate to the cathode where H^+ will be preferentially discharged. This is because even though Na^+ ions are more in concentration in the electrolyte, the distance between Na^+ and H^+ in the electrochemical series is very far.



Overall reaction of electrolysis: Chlorine gas is produced at the anode and hydrogen gas at the cathode.

Effect of electrolysis on the electrolyte

At the end of the electrolysis, the electrolyte will become alkaline. H^+ and Cl^- have been removed from the electrolyte remaining Na^+ and OH^- which forms NaOH solution.

3. EFFECT OF THE NATURE OF ELECTRODE e.g. Electrolysis of CuSO_4 solution using copper anode. The cathode can be any electrode e.g. platinum, graphite or even copper.

The electrolysis cell is the same as above but the anode is copper and cathode is graphite

Ions in the electrolyte:

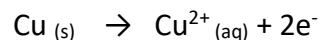
CATIONS



ANIONS



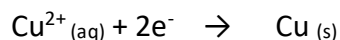
At the anode: Both SO_4^{2-} and OH^- will migrate to the anode but none of them will be discharged. The reason is that, the anode (which is a copper metal) has the nature as the metallic ions (copper (II) ions) in the electrolyte.



(From anode)

At the cathode: Both the Cu^{2+} ions and H^+ ions will migrate to the cathode where Cu^{2+} ions will be preferentially discharged.

Cathode half reaction



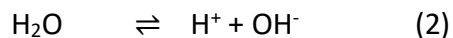
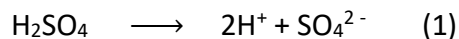
Product of electrolysis: Since the normal electrolysis did not take place, the anode ionizes into the electrolyte. Copper metal get deposited on the surface of the cathode and the cathode becomes bigger in size.

EVALUATION

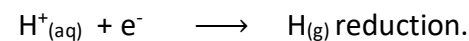
1. Differentiate the following terms; Electrolysis, electrodes, anode, cathode, ions, electrolytic cell, electrochemical cells.

PERIOD 4: ELECTROLYSIS OF ACIDIFIED WATER (DILUTE TETRAOXOSULPHATE (VI) ACID).

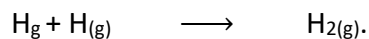
The equation of the reaction that is involved in this process is given as shown below:



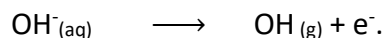
At the cathode: H^+ ions migrate to the cathode and are reduced by gaining electrons to become neutral hydrogen atoms.



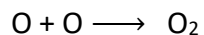
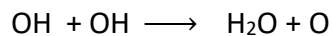
The hydrogen atoms then combine in pairs to form diatomic hydrogen gas molecules



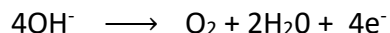
At the Anode: SO_4^{2-} and OH^- anode. The ions migrate to the anode. The OH^- ions being lower than SO_4^{2-} in the electrochemical series are preferentially discharged.



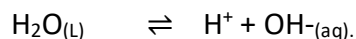
The OH groups interact to form water and oxygen molecules;



Overall half-equation



The net result of the electrolysis is that two volumes of hydrogen are produced at the cathode and one volume of oxygen is produced at the anode. The migration of SO_4^{2-} ions to the anode and the discharge of the H^+ ions cause a decrease in the concentration of H_2SO_4 acid around the anode disturbs the ionic equilibrium of water. To reverse this, more water ionizes



This produces an excess of H^+ ions and with the incoming SO_4^{2-} ions, an increase in the concentration of H_2SO_4 acid is obtained at the anode. The overall result of the electrolysis is that the total amounts of the acid in the solution remain unchanged at the end of the electrolysis. Since two volumes of hydrogen are obtained at the cathode and one volume of oxygen is obtained at the anode, the entire process is equivalent to the electrolysis of water.

ELECTROLYSIS OF COPPER (II) CHLORIDE USING COPPER – CARBON OR PLATINUM ELECTRODE

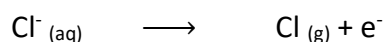
A copper rod is used as the cathode and carbon rod as the anode. The equations of the reaction involved were given as shown below:



When electric current is passed through the electrolytic solution, Cu^{2+} and H^+ ions migrate to the cathode, and Cl^- with OH^- migrate to the anode. At the cathode: Cu^{2+} being lower than H^+ in the electrochemical series is preferentially discharged. The Cu^{2+} gains two electrons from the cathode; which are in turn deposited as neutral metallic copper at the cathode.



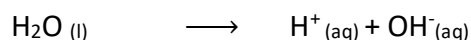
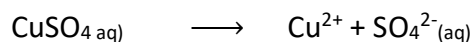
At the Anode: Cl^- for its higher concentration than OH^- is preferentially discharged as it loses an electron to the anode and becoming chlorine atoms.



The chlorine atoms then combine to form chlorine molecules. The discharge of copper and chlorine causes the solution to become progressively dilute as evidenced by the fading away of the light green colour of the electrolyte. If electric current is passed continuously without adding more electrolyte, at a stage, the hydroxyl ions begin to discharge, because at this stage, the influence of concentration of Cl^- ion is no longer the overall controlling factor. In the end, a mixture of chlorine and oxygen is obtained at the anode.

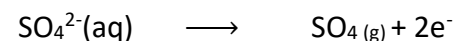
ELECTROLYSIS OF CuSO_4 USING COPPER –COPPER ELECTRODES (I.E COPPER AS CATHODE AND COPPER AS ANODE)

The equation of the reaction of the electrolysis is given as shown below.

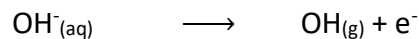


In this case, because both the cathode and anode are made up of copper metal, at anode three possible mechanisms occurs.

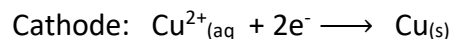
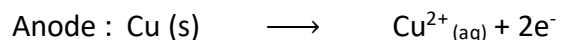
1. Discharge of SO_4^{2-} ion by loss of electrons



2. Discharge of OH^- ions by loss of electrons



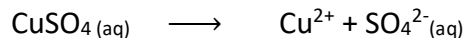
3. Copper metal loses electrons become $\text{Cu}^{2+}(\text{aq})$



Out of these possibilities, (3) occurs most readily because it requires the least energy. as a result, no ions are discharge at the anode. For each copper atom deposited, at the cathode, one atom of copper is dissolved from the anode to form Cu^{2+} ion in solution. Therefore, there is no change in the concentration of the electrolyte. The electrolysis merely produces copper ions at the anode and deposits same at the cathode. The colour of the electrolyte does not change.

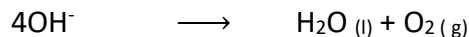
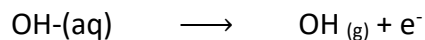
ELECTROLYSIS OF CUSO_4 SOLUTION USING COPPER – CARBON ELECTRODE

The electrolyte ionizes as follows



When electric current is passed through the electrolyte Cu^{2+} and H^{+} ion migrate to the cathode, and $\text{SO}_4^{2-}(\text{aq})$ and OH^{-} ions migrate to the anode. At cathode, $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \longrightarrow \text{Cu}(\text{s})$

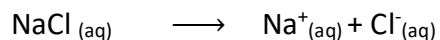
At anode: The OH^{-} ions is discharged preferentially because it is lower in the electrochemical series. Hence, oxygen is liberated at the anode.



The solution gradually becomes acidic due to the presence of H^{+} and SO_4^{2-} ions left behind in the solution. This is evident from the gradual fading away of the blue colour of the electrolyte.

ELECTROLYSIS OF BRINE: Concentrated sodium chloride solution (ionic NaCl solution) Note that in this electrolysis, the cathode can be either platinum carbon to the anode must be carbon to resist attack by chloride.

The electrolysis ionizes thus:

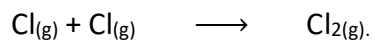
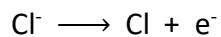


When electric current is passed through the electrolytic solution Na^+ and H^+ ions migrate to the cathode, while Cl^- and OH^- ions move to the anode.

At Cathode: Na^+ being lower than H^+ in the electrochemical series is preferentially discharged. The H^+ gains electrons from the cathode and is in turn liberated at the cathode



At Anode: Cl^- ions is preferentially discharged because its concentration is much higher than that of OH^- ions. The Cl^- ion loses an electron to the anode and becomes chloride atoms. The chlorine atoms then combine to form chlorine molecules.



The discharge of hydrogen ions and chloride ions causes the electrolyte to become basic.

ELECTROCHEMICAL SERIES

Electrochemical series also known as reactivity series is an arrangement of the standard electrolyte potentials of metal ions/metal half cells and those of non-metal/non-metal half cells in order of starting from the most negative to the most positive. Strongly electropositive series while strongly electronegative elements (usually non-metals and oxidizing agents) are at the bottom of the series.

Importance of electrochemical series

- 1. Predicting the preferential discharge of ions in electrolysis.** Metal down the series is preferentially discharged to the one above it and the non-metal higher in the series is discharged in preference to the ones below it.
- 2. Chemical reactivity of elements:** metals in the series are arranged in decreasing order of chemical reactivity except for calcium which is less reactive than sodium but is above sodium.
- 3. Natural occurrence:** metals above hydrogen in the series do not occur in the free state instead, they exist as compounds e.g. chlorides, sulphides, oxides etc.
- 4. Displacement and redox reactions:** the series shows at a glance the order in which element will displace another. Metals will displace metals which are below them in the series. For redox reactions, reducing power of elements decrease down the reactivity series while their oxidizing power increases down the series

EVALUATION

1. Distinguish between a conductor and an electrolyte
2. Distinguish between an electrolytic cell and an electrochemical cell.
3. How would you purify an impure copper metal
4. Distinguish between electrolyte and non-electrolyte.

5. Classify each of the following as strong/weak electrolyte: potassium chloride, sodium ethanoate, petrol, aqueous ammonia, cane sugar.
6. Explain why $\text{NaCl}_{(s)}$ does not conduct electricity, but $\text{NaCl}_{(aq)}$ does
7. (a) Define electrochemical series
(b) Explain the factors affecting preferential discharge of ion
State the importance of the electrochemical series

GENERAL EVALUATION

OBJECTIVE TEST

1. Which of the following methods cannot be used to distinguish between a strong acid and a weak acid? (a) conductivity measurement (b) action on starch iodide paper (c) measurement of pH (d) measurement of heat of reaction
2. Which of the following substances decomposes when an electric current is passed through it? (a) glucose solution (b) zinc rod (c) tetrachloromethane (d) hydrochloric acid
3. Which of the following is not an electrolyte? (a) silver trioxonitrate (V) solution (b) sugar solution (c) copper (II) tetraoxo sulphate (VI) solution (d) acidulated water
4. In electrolysis, reduction takes place at the ---- (a) anode (b) cathode (c) electrodes (d) cations
5. Electrolysis is not applied in ---- (a) tin plating (b) purification of copper (c) extraction of aluminium (d) extraction of iron (e) manufacture of sodium hydroxide

ESSAY QUESTIONS

1. (a) Define the term anode and cathode (b) Distinguish between weak electrolyte and strong electrolyte
(b) (a) What is an electrolyte? Give two examples (b) What is non-electrolyte? Give two examples
2. some metals are found to occur freely in nature while others occurs as compounds, explain each occurrence with examples of each types of metals.
3. Differentiate electrolyte from non-electrolyte. Give examples each
4. Briefly discuss ionic theory

WEEKEND ASSIGNMENT

Read New school chemistry for senior secondary schools by Osei Yau Ababio, pages 200-206 and 219-221

PRE-READING ASSIGNMENT

Read about electrolysis

WEEK 5

TOPIC: ELECTROLYSIS

CONTENT:

1. FARADAY'S LAW OF ELECTROLYSIS.
2. CALCULATIONS BASED ON FARADAY'S LAWS OF ELECTROLYSIS
3. USES OF ELECTROLYSIS AND CORROSION OF METALS: CORROSION TREATED AS A REDOX PROCESS
4. RUSTING OF IRON AND ITS ECONOMIC COSTS, PREVENTION OF RUSTING OF IRON

PERIOD 1: FARADAY'S LAWS OF ELECTROLYSIS

1. Faraday's first law of electrolysis states that *the mass (m) of a substance liberated or deposited at an electrode during electrolysis is directly proportional to the quantity (q) of electricity passing through the electrolyte.*

Mathematically;

$$M \propto Q \text{ ----- (1)}$$

The unit of quantity is coulombs (c). It is the product of current in Amperes and time in seconds

$$Q = It \text{ ----- (2)}$$

$$M \propto It$$

$$\therefore M = k It \text{ ----- (3)}$$

Where k = proportionality constant

I = current in ampere (A)

T = time in seconds (S)

It can be shown that the reciprocal of k, which is called charge –to-mass ratio, is Cf/M . It is constant for a particular element

$$\text{i.e. } m = \frac{Mit}{cF}$$

$$\text{Re-arranging } \frac{m}{M} = \frac{It}{cF}$$

Where m= mass f element deposited in grams

M = molar mass of the element in g/mol

I = current, in ampere

t = time, in seconds

c = charge on the element (number of Faradays)

F = Faraday (96500Cmol⁻¹)

The quantity of electricity required for the passage of one mole of electrons is 96500coulombs. It is called one Faraday. i.e. Faraday, F=96500C = 1 mole of electrons 6.02x10²³electrons

2. Faraday's second law of electrolysis states that when the same quantity of electricity is passed through different electrolytes the relative amounts in moles of the element deposited are inversely proportional to the charge on the ion of respective element.

Mathematically, $n \propto \frac{I}{c}$

i.e. $nc = \text{constant}$

Where n is the amount, in moles, of the element liberated, and c is the charge on its ion.

For two different elements: $n_1 c_1 = n_2 c_2$

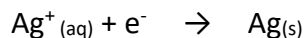
CALCULATIONS

Example 1

During the electrolysis of silver trioxonitrate (v) solution, 9650 coulombs of electricity was passed. Calculate (i) the amount of silver deposited in moles (ii) the mass of silver deposited in grains (Faraday = 9650 coulombs, Ag = 108g mol⁻¹)

Solutions

The equation of reaction is given as



(i) From the above equation, 1 mole of electron is required to deposited 1 mol of Ag but, 1 mol of electron = 1 Faraday 96500C.

i.e. 96500C will deposit 1 mol of Ag

$$\therefore 1/96500 \times 9650/1 = 0.1 \text{ mol}$$

\therefore Amount of silver deposited = 0.100 mol

(i) Since 1 mol of silver, (Ag) is 108g

\therefore 0.1mol of Ag will be $108/1 \times 0.1 = 10.8\text{g}$

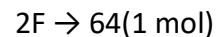
\therefore Mass of silver deposited = 10.8g

Example 2

When a solution of copper (II) salt was electrolyzed 3.20g of copper was deposited. Calculate the quantity of electricity required for electrolysis. (Cu=64; 1 F = 96500C).

Solutions

Equation of the reaction is: $\text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^{-} \rightarrow \text{Cu}_{(\text{s})}$



Amount in mole copper deposited is $n = \frac{\text{mass (g)}}{\text{Molar mass}}$

$$n = \frac{3.20 \text{ mol}}{64} = 0.050 \text{ mol}$$

From the equation, 2 mols of electrons will deposit 1 mol of Cu. But 1 mol of electrons = Faraday = 96500C \therefore 2 mol of electrons will require $2 \times 96500\text{C}$ to deposit 1 mol of Cu.

1 mol of Cu requires $2 \times 96500\text{C}$

\therefore 0.05mol of Cu will require $2 \times 96500 \times 0.05\text{C}$

Quantity of electricity required = 96500 coulombs.

EVALUATION

1. State the two faraday's laws of electrolysis
2. By passing an electric current for 1 hour through a copper and a silver coulometer connected in series, 1.28g of copper was deposited at the cathode of the copper coulometer calculate (a)the strength of the current used and (b) the mass of silver deposited on the cathode of the coulometer (faraday = 96500C mol⁻¹, Cu =63.5; Ag = 108).
3. Calculate the volume of gas that would be produced at the electrodes when 0.005 mole of electrons are passed through a dilute solution of tetraoxosulphate (VI) acid (gaseous molar volume = 22.4dm³ at stp.)

PERIOD 3: USES OF ELECTROLYSIS AND CORROSION OF METALS: CORROSION TREATED AS A REDOX PROCESS

Some of the main applications of electrolysis are as follows:

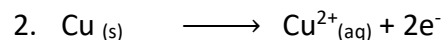
1. In the extraction of elements especially metals such as Na, K, Mg, Ca, Zn and non-metals (e.g.H₂, F₂, Cl₂) are obtained either from their fused compounds or their aqueous solutions.
2. Purification of metals (e.g. Cu, Hg, Ag, Au).
3. Electroplating of one metal by another.
4. Preparation of certain important compounds, such as sodium hydroxide and sodium trioxochlorate (v).

1. EXTRACTION OF ELEMENTS USING ELECTROLYSIS METHOD

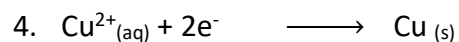
An example of extraction can be shown using extraction of aluminium from purified bauxite (Al₂O₃). Highly electropositive metals like sodium, potassium, and calcium are extracted by the electrolytic reduction of their fused chlorides. For example, electrolysis of molten sodium chloride liberates sodium at the cathode and chlorine gas at the anode. Elements like chlorine and hydrogen are extracted by electrolysis.

2. Purification of metals: The less electropositive metals e.g copper, mercury, silver and gold, tend to occur in the un-combined form in Ores. These ores are usually contaminated with impurities such as sand, limestone and iron compounds. An example is the extraction of copper from copper ore. In this process, an electrolysis cell is used in which the impure ore series as the anode and a pure copper plate series as the cathode. The electrolyte can be any soluble copper salt, such as copper(ii) tetraxoxsulphate (vi).

1. **At the Anode:** The metallic copper atoms in the ore give up two electrons each to go into solution as copper ions.



3. At Cathode: The copper ions in the solution are attracted to the cathode where they receive two electrons each from the cathode to become neutral copper atoms which are then deposited on the cathode.

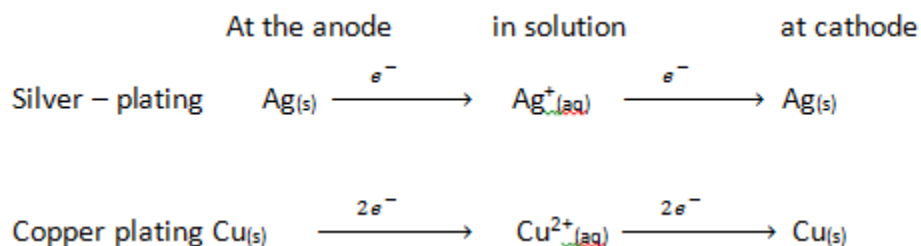


5. The net result is that pure atoms are transferred from the impure ore (anode) to the cathode. Similar arrangements are used in the extraction and purification of silver, mercury, and gold. In all cases, the electrolysis is soluble salt of the metal being extracted or purified, the anode is metallic ore or impure metal, while the cathode is a pure piece of metal on which the pure metallic atoms can be deposited.

3. ELECTROPLATING

Electroplating is a method of coating the surface of one metal with another metal, usually copper, silver, chromium, nickel, or gold by means of electrolysis, for decoration or protection against corrosion; Articles which are usually plated include table-ware, cutlery, jewellery and iron or steel objects such as the metallic parts of motor – cars, the arrangement of the electrolytic cell is essentially the same as that used in the extraction of metals. The cathode is the metallic object which is to be located or plated. The anode strips off the plating metals; and the electrolyte, a salt of the plating metal. Thus, in silver – plating a spoon, the cathode is the spoon, the anode is the silver rod and the electrolyte is a soluble silver salt, usually a silver trioxonitrate (v)

solution. As the current passes through the cell, the plating metal dissolves at the anode and the ions produced migrate to the cathode where they are discharged and deposited a layer on the object.



Many articles are plated at the same time. The thickness of the plating may be controlled by adjusting the magnitude of the current and the duration of the process. Where the first material used for coating could not adhere, (like in the case of chromium which does not adhere well to steel), first coat with copper which provides adhesion, then with a layer of nickel for protection against corrosion and finally, with a layer of chromium for a tarnish free finish.

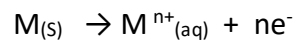
CORROSION OF METALS: CORROSION TREATED AS A REDOX PROCESS



Corrosion is the deterioration of materials by chemical interaction with their environment.

The term corrosion is sometimes also applied to the degradation of plastics, concrete and wood, but generally refers to metals.

Corrosion of metal is defined as the slow destruction of metals by chemical actions. For instance, all metals above hydrogen in the activity (electromotive) series corrode with mineral acids. In the process, a metal (M) is oxidized from zero oxidation state to a higher oxidation state (n^+):



During corrosion of metals by acid, the metals supply electrons which are taken up the H^+ ions of the acid to liberate hydrogen.

Example:

This net reaction would occur spontaneously since E_{Cell}^0 is positive.

The above example shows that corrosion is a redox reaction. The term “corrosion” is applicable mainly to other metal apart from iron. Actually, **the corrosion of iron is appropriately described as “rusting”**.

PERIOD 4: RUSTING OF IRON AND ITS ECONOMIC COSTS, PREVENTION OF RUSTING OF IRON

1. The most familiar chemical property of iron is rusting. Iron (or steel) combines with air and water to form rust, a reddish -brown substance.

Rust consists mainly of hydrated iron (III) oxide, $Fe_2O_3 \cdot 3H_2O$ and a little quantity of iron (II) trioxocarbonate (IV) $FeCO_3$. Rusting is an electrolytic process. The essential **conditions for rusting** are: **moisture and oxygen (air)** which acts together on iron. It is unfortunate that iron, which has very many uses, rusts easily on exposure to moist air.

2. THE ECONOMIC IMPLICATION of this is that very large sums of money are expended throughout the world on the various protective processes against rusting. Besides, a lot of money is spent every year for the replacement of rusted iron world-wide.

3. PREVENTION OF RUSTING OF IRON

There are many preventive methods against rusting of iron. The basic principle of the methods is prevention of **water and air** from coming into direct contact with iron. Examples of the methods are:

(a) Painting: This is good for iron or steel and equipment materials such as ships. Bridges, railings, agricultural machinery, cars and buses. It involves Coating layers of paints on the materials and equipment.

(b) Galvanizing: The method involves spraying iron materials (such as pipes, roofing sheets, etc with powdered zinc or dipping the materials (when properly cleaned) into molten zinc.

(c) Electroplating: it involves electrolytic deposition of metals such as tin, copper, chromium etc (tin-plating, copper plating, chromium-plating, etc) on iron materials such as car bumpers, spoons, knives, etc.

(d) Greasing or Oiling: The method involves rubbing iron or steel materials (such as tools or moving parts of machinery) with grease or oil.

EVALUATION

1. What is corrosion?

2. Write short note on the following:

(a) Painting (b) Galvanizing (b) Galvanizing (c) Electroplating (d) Greasing or Oiling

GENERAL EVALUATION

OBJECTIVE TEST:

1. One faraday is equal to (a) 9650 coulombs (b) 96500 coulombs (c) two moles of element (d) half a mole of electron

2. In the electrolysis of brine, the anode must be carbon because (a) carbon is a reducing agent (b) chlorine is a reducing agent (c) carbon induces the discharge of chlorine (d) chlorine attacks other elements but not carbon
3. Which of the electrodes does oxidation occur? (a) at the cathode (b) at the anode (c) at the electrolyte (d) at the half electrode.
4. How many moles of copper would be produced by passing one faraday of electricity through an aqueous solution of copper (ii) tetraoxosulphate (vi)? (a) 2 (b) 0.5 (c) 3 (d) 4

ESSAY QUESTIONS

1. List the factors that affect the discharge of ions in electrolysis.
2. State the two faraday's laws of electrolysis
3. How can you obtain pure copper from impure copper?

WEEKEND ASSIGNMENT

Read New school chemistry for secondary schools by Osei Yaw Ababio reversed by L.E.S. Akpanis H. I. Pages 504 – 513

PRE READING ASSIGNMENT

Read all about hydrocarbon and crude oil

WEEK END ACTIVITY

1. Describe how the following can be plated with equations only.
 - i. Silver
 - ii. Copper

iii. Spoon

2. Define the following (a) hydrocarbon (b) Isomerism (c) saturated and unsaturated hydrocarbon.

WEEK: 6

TOPIC: HYDROCARBONS

CONTENT:

1. STRUCTURE AND VALENCY OF CARBON, HYDROCARBON – MEANING AND EXAMPLES,

2. HOMOLOGOUS SERIES- CHARACTERISTICS, AND FUNCTIONAL GROUP

3. SATURATED HYDROCARBON -ALKANES

PERIOD 1: STRUCTURE AND VALENCY OF CARBON

Carbon atom belongs to group IV of the periodic table of element. The excited state electronic configuration of carbon is $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$. A carbon has four electrons in its outer shell and uses by sharing electrons to form covalent bonds. Carbon atoms can combine with one another to form chains and rings of different sizes. Carbon atom can catenate that is they can form single double or triple covalent bonds with themselves.

HYDROCARBONS MEANING AND EXAMPLES

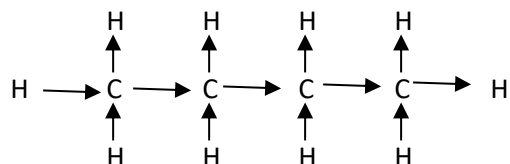
Hydro carbons are compounds composed of only carbon and hydrogen atoms.

Hydrocarbons are classified into the two classes (1) Aliphatic compounds and (2) aromatic compounds.

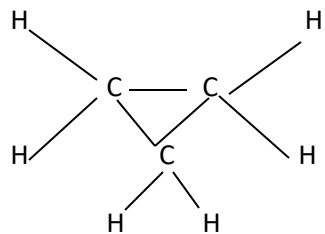
Aliphatic compounds are either acyclic or cyclic.

Acyclic compounds are straight chain or branched chain hydrocarbons. They are the alkanes, alkenes and alkynes. E.g

(a)



In cyclic compounds the carbon atoms form a ring e.g



EVALUATION

1. Explain the term 'catenation'
2. What is aliphatic hydrocarbon?

PERIOD 2 AND 3: HOMOLOGOUS SERIES AND FUNCTIONAL GROUP

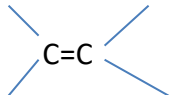
Definition: A homologous series is a family of organic compounds which follows a regular structural pattern, in which each successive member differs in its molecular formula by a $-CH_2$ group.

CHARACTERISTICS OF HOMOLOGOUS SERIES AND NAMING-IUPAC

1. All members of a series share a general molecular formula.
2. Each successive member of a series differs in its molecular formula by the addition of a $-CH_2$ group and in its relative molecular mass by an increase of 14.
3. The physical properties of the members in a series change gradually as the number of carbon atoms per molecule increase.
4. All members of a series can usually be prepared by using the same general method.
5. The members of a series show similar chemical properties.
6. All members of the same family have the same functional group.

Functional group

Functional group is an atom , bonds or group of atoms that identifies an organic compound and determines its chemical properties.

Homologous series	General molecular formula	General structural formula	Functional group
1. alkanes	C_nH_{2n+2}	R-H	
2. Alkene	C_nH_{2n}		
Alkyne	C_nH_{2n-2}		- C \equiv C-
Haloalkanes	$C_nH_{2n-1}X$	K-X	-X
Alkanol	$C_nH_{2n+1}OH$	R-O-H	-OH
Alkanoic acid	$C_nH_{2n+1}COOH$	R-COOH	-COOH
Alkanoate	$C_nH_{2n+1}COOC_nH_{2n+1}$	R-COOR ¹	-COOR
Alkanal	$C_nH_{2n+1}CHO$	RCHO	-CHO
Alkanone	$C_nH_{2n+1}COC_nH_{2n+1}$	RCOR	-COR
Amide	$C_nH_{2n+1}CONH_2$	RCONH ₂	-CONH ₂

Amine	$C_nH_{2n+1}NH_2$	RNH_2	$-NH_2$
Nitrile	$C_nH_{2n+1}CN$	RCN	$-CN$

PERIOD 4: SATURATED HYDROCARBON-ALKANES

ALKANES

The alkanes are hydrocarbons with the general molecular formula C_nH_{2n+2} , where n starts from 1. They are saturated hydrocarbons. There are only single covalent bonds in their structures. The sources of alkanes are crude oil, coal, marsh gas and natural gas. Every carbon atom in alkane is sp^3 hybridized

Examples of the first five members of the series are shown below.

FORMULA	NAME	STRUCTURAL FORMULA
CH_4	Methane	$ \begin{array}{c} H \\ \\ H - C - H \\ \\ H \end{array} $

C_2H_6	Ethane	$ \begin{array}{c} H & H \\ & \\ H - C & - C - H \\ & \\ H & H \end{array} $
C_3H_8	Propane	

THE ALKYL GROUPS: The general term alkyl group includes all groups derived from the alkanes by the loss of hydrogen atom.

They are given the general symbol, R. Example of the alkyl groups.

FORMULA	NAME
CH_3-	Methyl
C_2H_5-	Ethyl
C_3H_7	Propyl
C_4H_9	Butyl

IUPAC NOMENCLATURE

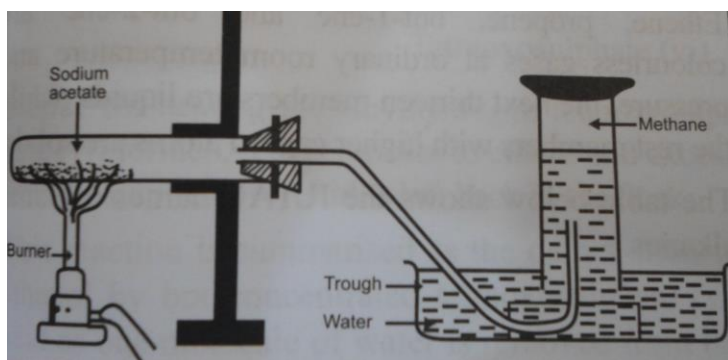
Rules for naming organic compounds

1. Take the longest continuous carbon chain as the root hydrocarbon and name according to the number of carbon atoms present.
2. Number the carbon atom in the root hydrocarbon from the end which gives the least number to the suffix and then the prefix.
3. Indicate any substituents by prefixes preceded by number to show their positions on the carbon chain.
4. When substituents are present, they should be named based on the one that comes first alphabetically.

METHANE

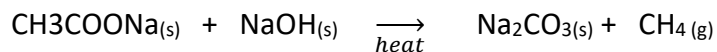
Occurrence: Methane is the major constituent of natural gas, petroleum gas and, is usually known as Marsh gas. It is also found in poorly ventilated coal mines as fire damp.

LABORATORY PREPARATION

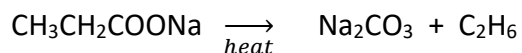


Procedure: A finely powdered mixture of equal amount of anhydrous sodium ethanoate and soda lime is heated in a hard glass test tube. The methane evolved is collected over water.

Equation of reaction is shown below:



To prepared other members, the same method is applied. For example the preparation of ethane required heating sodium propanoate according to the equation below:



PROPERTIES OF ALKANES

PHYSICAL PROPERTIES

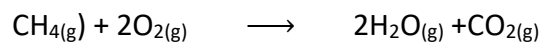
1. The first members (i.e. $\text{C}_1 - \text{C}_4$) are colourless gases. From ($\text{C}_5 - \text{C}_{17}$) are colourless liquids. Higher members greater than C_{17} are colourless, wax like solids.
2. They are insoluble in water
3. Branched alkanes have lower boiling and melting points than their corresponding straight chain alkanes. Braching makes the molecules spherical and therefore reduces the intermolecular forces between the molecules. For example, the boiling point of 2-methylpropane is lower than that of the straight chain isomer butane.

CHEMICAL PROPERTIES

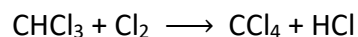
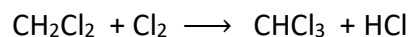
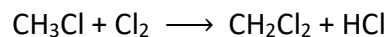
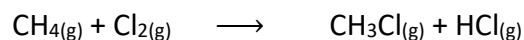
Alkanes are generally inert but they can undergo two types of reactions: combustion and substitution reactions

1. **Combustion reaction.** Alkanes burn in air (or oxygen) to form CO₂ and H₂O (steam). The reaction is exothermic e.g

Combustion of Methane, CH₄



2. **Substitution reaction.:** Alkanes undergo substitution reaction with halogens. The reaction of methane with chlorine takes place in the presence of ultra-violet light which serves as catalyst. The continues until all hydrogen atoms in methane are replaced with chlorine



Uses of methane

1. It is mainly used fuel
2. It is used to make carbon black, carbon (iv) sulphide, trichloromethane or tetrachloromethane.

EVALUATION

1. Give reasons why soda lime is used instead of pure sodium hydroxide in the laboratory preparation of methane
2. Explain why the boiling pint of 2-methylbutane is lower than that of pentane.
3. What is the product of complete chlorination of ethane.

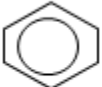
GENERAL EVALUATION

Objective test

1. Exceptionally large number of carbon compounds is essentially due to the ability of (a) carbon to catenate liberally (b) various groups to catenate (c) nitrogen, hydrogen, phosphorus and the halogens to catenate with themselves (d) hydrocarbons to dominate other groups

2. The name of $C(CH_3)_4$ is (a) Butane (b) tetramethylbutane (c) methylpropane (d) 2-methylpropane

3. The following are generally characteristics of carbon except (a) covalent nature and non polar (b) low melting and low boiling points (c) low reactivity with other elements except oxygen and the halogens (d) hydrogen bond in petrol

4. The name of  is (a) Benzene (b) Hexane (c) pentane (d) cyclohexane

5. Which of these compounds exhibit resonance? (a) Ethanol (b) Benzene (c) propylene (d) butyne

ESSAY QUESTIONS

1. Draw a labelled diagram to show how methane is prepared in the laboratory

2. State four characteristics of a homologous series
3. Draw the structures of (a) 2,3 –dimethylbutane (b) 2,2,3-trimethylhexane

WEEK 8

TOPIC: HYDROCARBONS

CONTENTS:

- 1. ISOMERISM-STRUCTURAL AND GEOMETRIC**
- 2. ALKENE-IUPAC NOMENCLATURE, METHOD OF PREPARATION, PROPERTIES AND USES**
- 3. ALKYNE- IUPAC NOMENCLATURE, METHOD OF PREPARATION, PROPERTIES AND USES**
- 4. AROMATIC HYDROCARBON: BENZENE-PROPERTIES AND USES**

ISOMERISM

Definition: Isomerism is the existence of two or more compounds (known as isomers) with the same molecular formulae but different molecular structures

ISOMERISM IN ALKANES

1. Structural isomerism in alkanes is obtained by branching the straight chain isomer.

(a) Draw the isomers of the compound with molecular formula

(i) C_4H_{10}

$CH_3CH_2CH_2CH_3$ (butane) $CH_3CH(CH_3)CH_3$ 2-methylpropane

(ii) C_5H_{12} $CH_3CH_2CH_2CH_2CH_3$ Pentane $CH_3CH(CH_3)CH_2CH_3$ 2-methylbutane $CH_3CH(CH_3)_2CH_3$ 2,2-dimethylpropane

GEOMETRIC ISOMERISM

This is the existence of organic compounds with the same molecular formula but different spatial arrangement of compound atoms. Geometric isomers have similar chemical properties but different physical properties. We have the cis-isomers and trans-isomers. Geometric isomerism is common in alkenes.

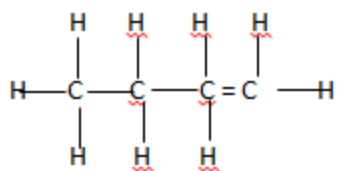
ALKENES

The alkenes are the homologous series with the general molecular formula C_nH_{2n} ; n starts from 2. They are unsaturated hydrocarbons due to the double bond in their structures.

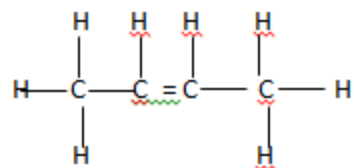
ISOMERISM IN ALKENES

1. Isomers of C_4H_8

1. Isomers of C₄H₈

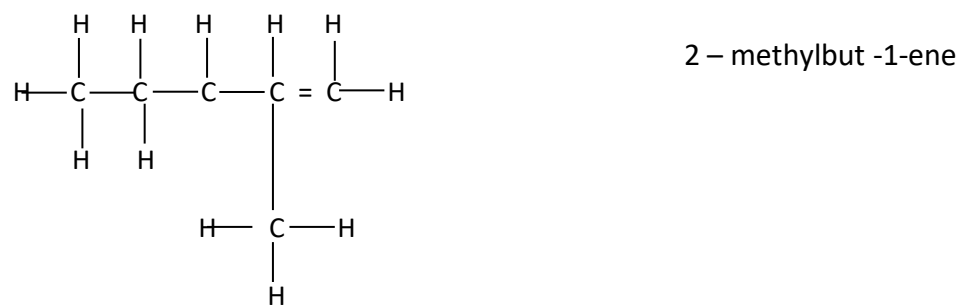
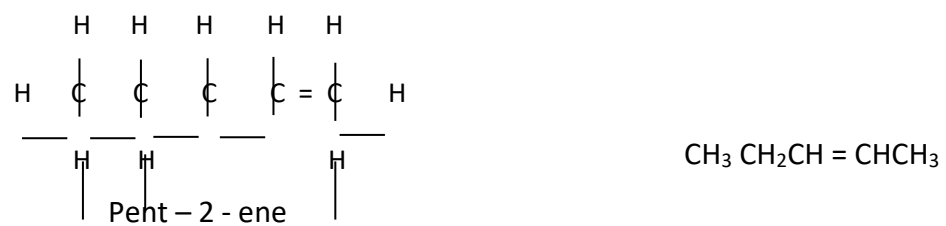
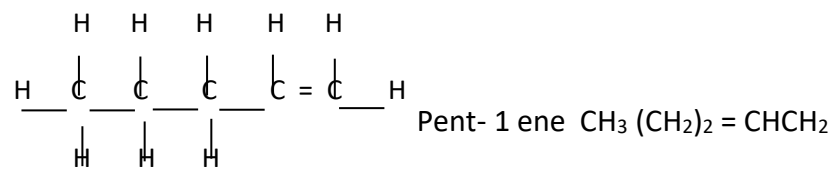


CH₃ CH₂ CH=CH₂ but-1-ene



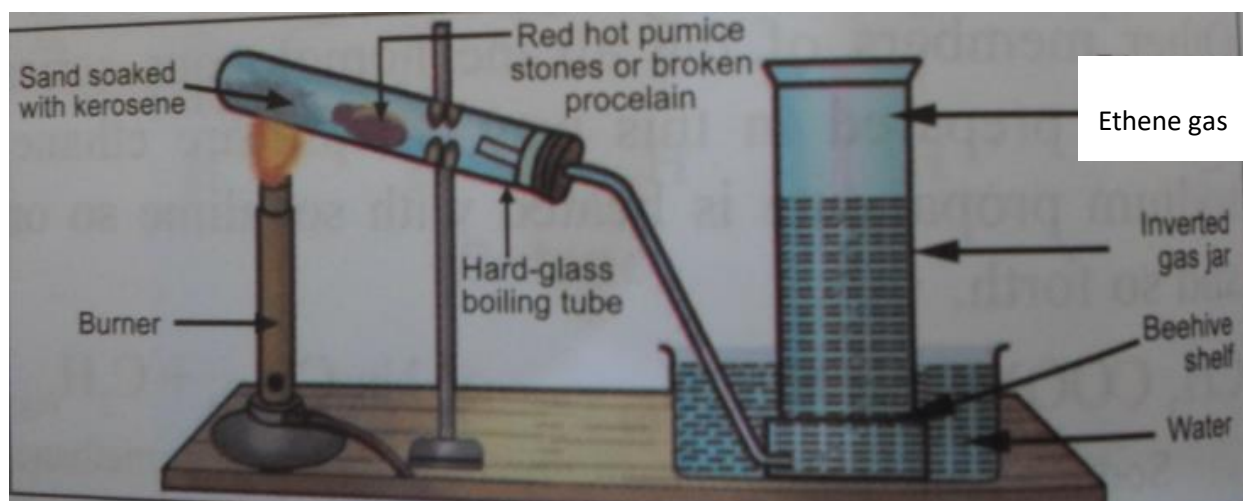
CH₃ CH = CH CH₃ but-2-ene

2. Isomers of C₅H₁₀

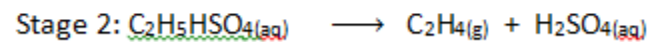
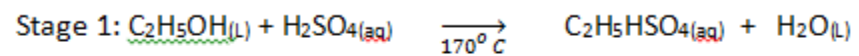


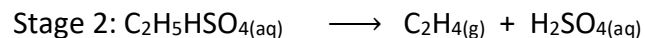
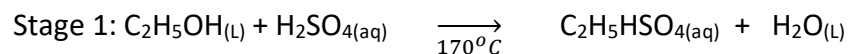
Note: The carbon atoms forming the straight chain is numbered from the carbon closest to the double bond.

Laboratory preparation of Ethene



Ethene is prepared in the laboratory by the dehydration of ethanol by concentrated tetraoxosulphate(vi). Equations for the reaction





PROPERTIES OF ETHENE

Physical properties

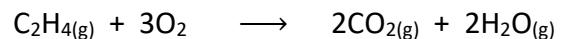
1. Ethane is a colourless gas with a faint sweetish smell.
2. It is sparingly soluble in water.
3. It is slightly less dense than air.
4. It has no action on litmus paper.

Chemical properties

1. Ethene undergoes polymerization reaction e.g



2. Combustion reaction: Ethene or any alkene burn in air (or oxygen) to form CO_2 and H_2O e.g



3. Addition reaction: An addition reaction involves the direct addition of an attacking reagent across the double bond or triple bond of an unsaturated compound to form a saturated compound or at least one in which the degree of saturation is increased.
Example of addition reaction between

(i) ethane and hydrogen.



(ii) Ethene and chlorine



Uses of Ethene

1. It is the main source of ethanol
2. It is a raw material for making polyethane.
3. Low concentration of ethane is used to ripen fruits.

Test for Unaturation

Alkenes and alkynes are unsaturated compounds.

Unsaturated compounds

1. Decolourized reddish brown bromine water or bromine.
2. Charge purple $\text{KMnO}_4(\text{aq})$ to colourless

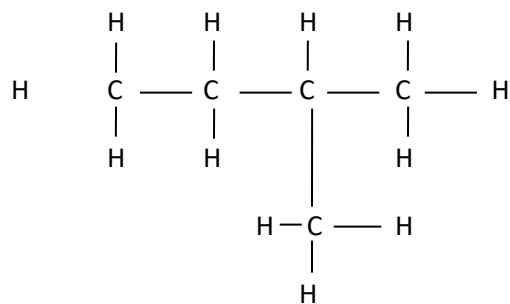
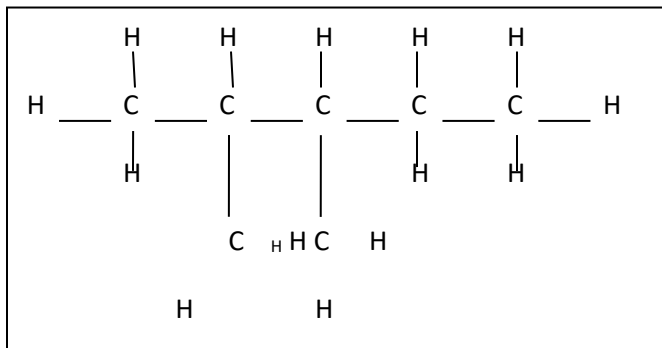
EVALUATION

1. What is the function of the empty flask during the preparation of ethane?

2. What is geometric isomerism? Give examples

3. Draw and name five possible isomers of hexane, C_6H_{14} .

1. Name these structures



PERIOD 3: THE ALKYNES

The alkynes are a homologous series of unsaturated hydrocarbon series of molecular formula of C_nH_{2n-2} where n starts from 2

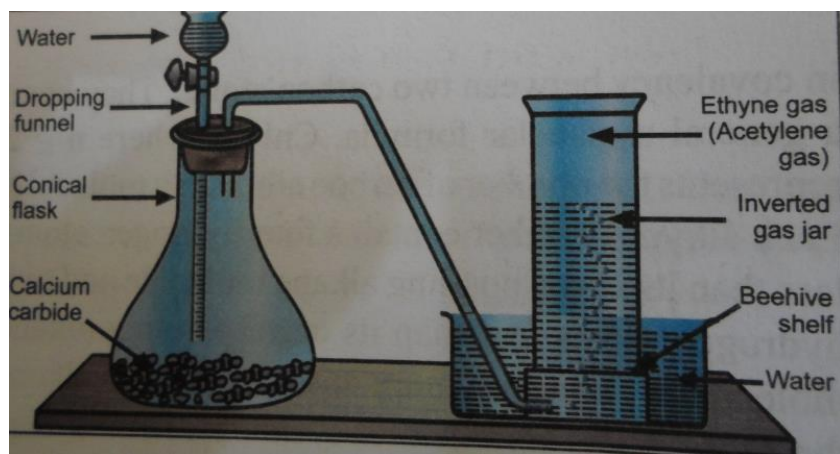
The examples of the first five members of the alkynes series are shown below.

Chemical Formula	Name	Structural Formula
C_2H_2	Ethyne	$H-C\equiv C-H$
C_3H_4	Propyne	$H-C\equiv C-CH_3$
C_4H_6	Butyne	$H-C\equiv C-CH_2-CH_3$
C_5H_8	Pentyne	$H_3-C-CH_2-CH_2-C\equiv C-H$

Note: The number of carbon atoms in alkynes starts from the last carbon closest to the triple bonds.

ETHYNE

LABORATORY PREPARATION OF ETHYNE



Procedure: Ethyne is prepared in the laboratory by the action of cold water on calcium carbide as shown in the diagram above. The process is exothermic

Equation for the Reaction



PROPERTIES OF ETHYNE

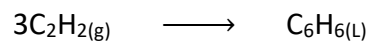
a. Physical Properties

1. It is a colourless gas with sweet smell

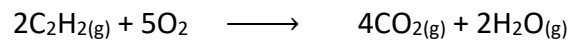
2. Sparingly soluble in water
3. Slightly less dense than air.
4. Unstable and may explode if compressed to liquid

b. Chemical Properties: Like any other alkyne, ethyne undergoes combustion, addition and reactions. Ethyne also undergoes polymerization reactions and substitute reactions.

1. Polymerization reaction Ethyne

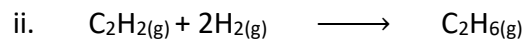
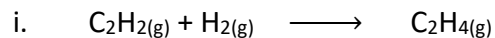


2. Combustion reaction of ethyne. Ethyne burns in oxygen to form CO_2 and steam.



3. Addition reaction of Ethyne

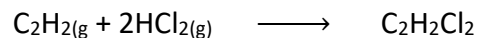
a. Ethyne reacts with hydrogen to form.



b. With halogens e.g with chlorine

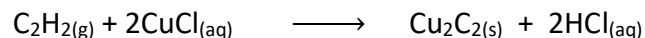


c. With hydrogen halides

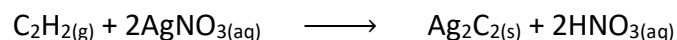


4. Substitution Reaction: Since ethyne is a terminal alkyne, it can react with ammoniacal solutions of copper(i)chloride and silver trioxonitrate(v) at room temperature as shown below:

a. With ammoniacal copper(i)chloride solution a reddish-brown precipitate is formed



b. With ammoniacal solution of silvertrioxonitrate(v) solution a yellowish white precipitate is formed



USES OF ETHYNE

1. As fuel in lamps such as miner's lamps
2. Ethyne is mixed with oxygen to produce a very hot flame called oxy-ethyne flame used for cutting and welding metals.
3. Ethyne is a source of ethanol, and solvents used in industry and in dry cleaning e.g 1,1,2 – trichloroethene and 1,1,2,2 – tetrachloroethene .

EVALUATION:

1. State three physical properties of ethyne
2. Show how you will distinguish between these gases: ethane, ethane and ethyne
3. Describe the method of preparation of ethyne

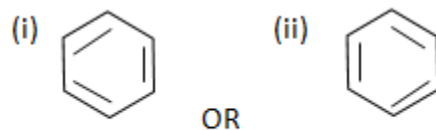
Read about types of alkanols from page 540 of new certificates chemistry by Osei Yaw Ababio.

PERIOD 4 : AROMATIC HYDROCARBON

BENZENE

Benzene is obtained from the destructive distillation of coal and also from Naptha.

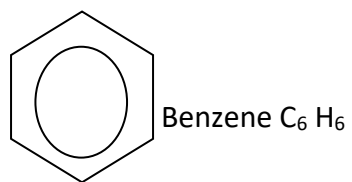
Structure: August Kekule proposed the structure of benzene in 1865 to be



In 1945, the concept of resonance was used to explain the structure of benzene

Resonance occurs when two equivalent forms of a compound are in equilibrium.

The real structure of benzene is represented as a regular hexagon with an inscribe circle, indicating that the double bonds are shifting at all times. The ring also shows that the six electrons involved in the double bonds are moving around the six carbon atoms in benzene.



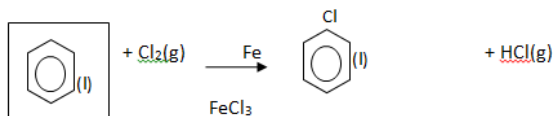
PROPERTIES OF BENZENE

A. PHYSICAL PROPERTIES

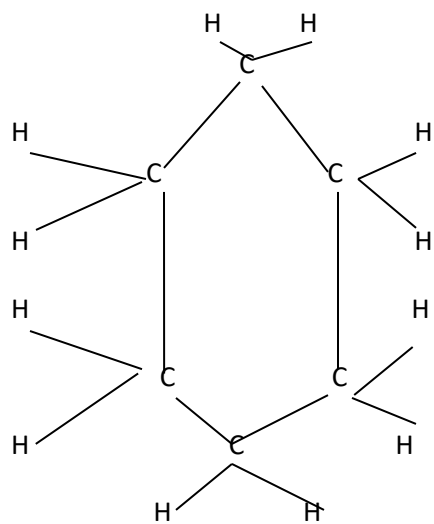
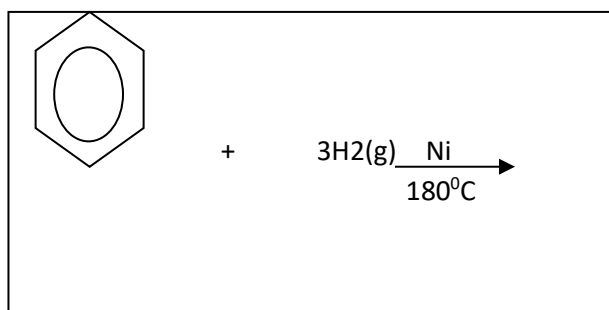
1. Benzene is a colourless liquid with a sweet smell.
2. It is insoluble in water.
3. It has a boiling point of 80°C

B. CHEMICAL PROPERTIES

1. Benzene undergoes substitution reaction e.g. when Benzene reacts with chlorine



2. Benzene also undergoes addition reaction e.g



Cyclohexane is formed when a mixture of benzene vapour and hydrogen are passed over a nickel catalyst 1899C.

EVALUATION

What is resonance? Explain using the structure of Benzene

GENERAL EVALUATION

OBJECTIVE TEST:

1. Which of the following decolourizes bromine water? (a) propane (b) ethanol (c) ethene (d) methanoic acid
2. Hydration of propene gives (a) propane (b) propan-2-ol (c) propan-1-ol (d) benzene
3. Which of the following gases is used in mariner's lamp? (a) ethane (b) ethane (c) ethyne (d) ethanol
4. Hydrogenation of benzene in the presence of nickel catalyst at 180°C produces (a) hexane (b) cyclohexene (c) cyclohexane (d) methylbenzene
5. Dehydration of propanol produces (a) propene (b) propane (c) propyne (d) propanoic acid

ESSAY QUESTIONS

1. Write an equation to show addition reaction between ethyne and hydrogen bromide to illustrate laboratory preparation of ethane.
2. Write an equation to show the reaction ethane with conc H_2SO_4
3. Write the general formula for the alkynes. How would you prepare ethyne in the laboratory.
4. Draw the possible isomers of the compound with molecular formula C_7H_{14}

WEEKEND ASSIGNMENT

Read about test for un-saturation on page 537 of New School chemistry by Osei Yaw Ababio.

PRE-READING ASSIGNMENT

Read about Alkanols on page 539 of new School chemistry by Osei Yaw Ababio

WEEKEND ACTIVITY

From the internet find out why terminal alkynes can form a precipitate with ammonical copper (i) chloride solution and non terminal alkynes cannot.

WEEK: 9

TOPIC: ALKANOLS

CONTENT:

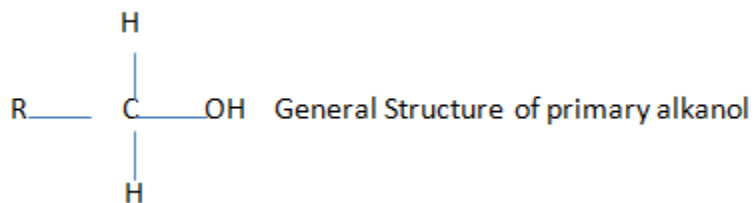
1. CLASSES AND TYPES OF ALKANOLS
2. LABORATORY AND INDUSTRIAL PRODUCTION OF ALKANOLS,
3. PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOLS
4. LABORATORY TEST AND USES OF ALKANOLS.

PERIOD 1: CLASSES AND TYPES OF ALKANOLS

CLASSES OF ALKANOLS

There are three classes of alkanols – primary, secondary and tertiary.

1. Primary (1°) alkanols: These have only one alkyls group attached to the carbon atom that carries the hydroxyl group e.g. C_2H_5OH , C_4H_9OH etc. Example: C_2H_5OH ,



2. Secondary (2°) alkanols: These have two alkyls attached to the carbon that carries the hydroxyl group e.g. Example: $CH_3CH_2CH(OH)CH_3$ (Butan - 2- ol)

WWW

3. **Tertiary (3°) Alkanols:** These have three alkyl group attached to the carbon atom that carries the hydroxyl group. Example:

C (CH₃)₃OH (2-methylpropan-2-ol).

2. TYPES OF ALKANOLS

1. **Trihydric alkanols:** These are alkanols with three (3) OH groups per molecule e.g. propane-1, 2, 3-triol

CH₂OH

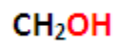
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CHOH

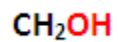
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CH₂OH

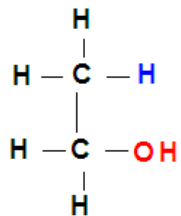
2. **Dihydric alkanols:** These contain 2 OH groups per molecule e.g. ethane-1, 2-diol



|



3. Monohydric alkanols: These contain only one (1) OH group per molecule e.g. ethanol, C₂H₅OH, etc.



They have the general molecular formula C_nH_{2n+1}OH. This is sometimes written as ROH. Where R stands for an alkyl group

Examples of the first five members of the group are:

Chemical Formula	Name	Structural Formula
CH ₃ OH	Methanol	$\begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{OH} \\ \\ \text{H} \end{array}$
C ₂ H ₅ OH	Ethanol	$\begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{H} \\ \\ \text{H} - \text{C} - \text{OH} \\ \\ \text{H} \end{array}$

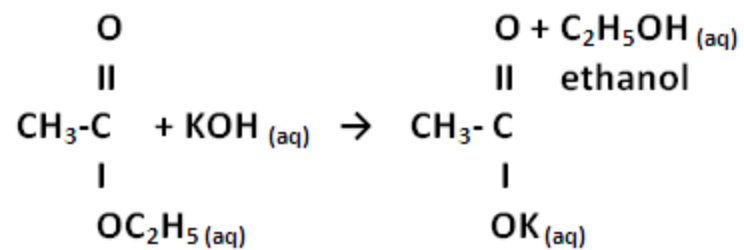
C_3H_7OH	Propanol	$ \begin{array}{c} H \\ \\ H-C-OH \\ \\ H-C-H \\ \\ H-C-H \\ \\ H \end{array} $
C_4H_9OH	Butanol	$ \begin{array}{c} H \\ \\ H-C-OH \\ \\ H-C-H \\ \\ H-C-H \\ \\ H-C-H \\ \\ H \end{array} $

$C_5H_{11}OH$	Pentanol	$ \begin{array}{c} H \\ \\ H-C-OH \\ \\ H-C-H \\ \\ H-C-H \\ \\ H-C-H \\ \\ H-C-H \\ \\ H \end{array} $
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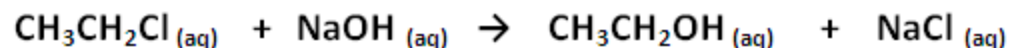
PERIOD 2: LABORATORY AND INDUSTRIAL PRODUCTION OF ALKANOLS

Ethanol is prepared by the following methods.

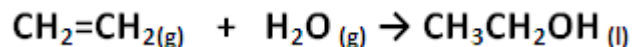
1. Hydrolysis of ethyl ethanoate in hot, dilute potassium hydroxide.



2. Hydrolysis of monochloroethane in the presence of dilute sodium hydroxide.



3. Hydration of ethene. A mixture of ethene and steam with tetraoxosulphate (VI) acid as a catalyst is heated to a temperature of 600°C and a Pressure of 100 atmospheres.



4. Reduction of ethanal with nascent hydrogen



Ethanal Ethanol

The reaction takes place in the presence of platinum catalyst.

INDUSTRIAL PRODUCTION OF ALKANOLS (ETHANOL) BY FERMENTATION

1. **Fermentation** is the slow decomposition by micro –organisms of large organic molecules (such as starch) into smaller molecules (such as ethanol).

Preparation of Ethanol by Fermentation

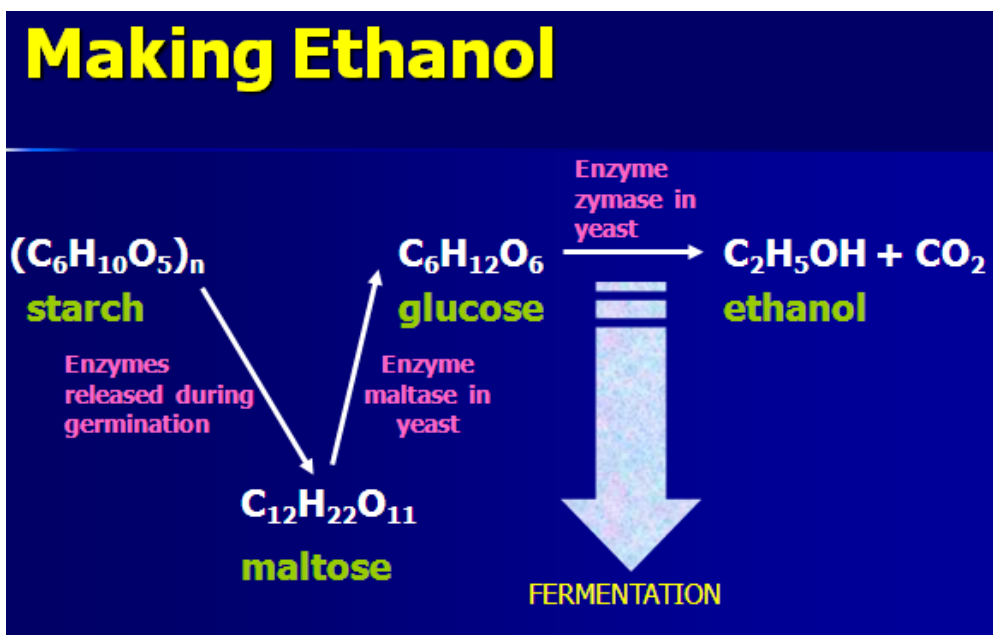
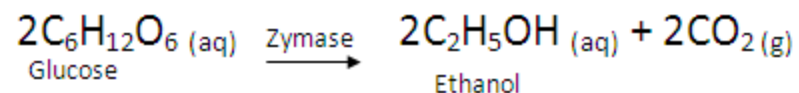
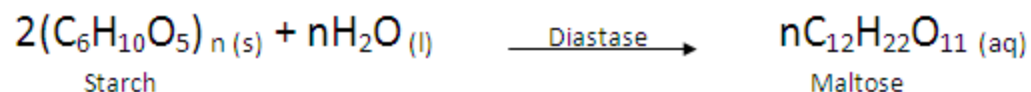
Process: The starch granules from any starchy food e.g. rice, maize, yam etc are extracted by crushing and pressure cooling.

1. They are then treated with malt at 50°C to 60°C for an hour.

2. Malt which contains enzymes DIASTASE is added. Diastase converts starch to maltose

3. Yeast which contains enzymes maltase and zymase is added at room temperature.

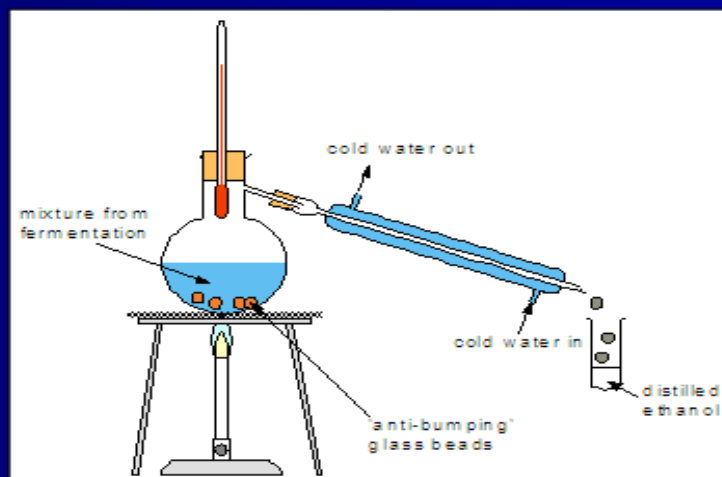
Enzymes maltase converts maltose to glucose. zymase decomposes glucose to ethanol and carbon (IV) oxide



Fractional distillation of the ethanol obtained by fermentation is carried out for further concentration and purification.

Distillation

To produce drinks with a high alcohol concentration, **DISTILLATION** must be carried out.



EVALUATION

Starting from potatoes explain how ethanol can be prepared by fermentation.

PERIOD 3: PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOLS

PHYSICAL PROPERTIES OF ETHANOL

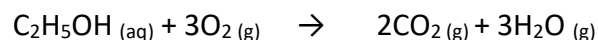
Ethanol

1. is readily soluble in water;
2. is has no effect on litmus;
3. is colourless and volatile;

4. is liquid at room temperature;
5. Smells characteristically and tastes sharply; and
6. boil at a temperature of 78°C

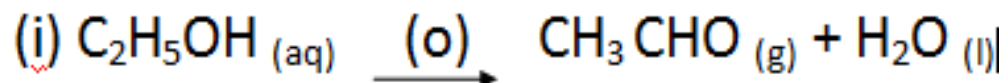
CHEMICAL PROPERTIES OF ALKANOLS

1. **Combustion reaction:** Alkanols undergo combustion reaction to form CO_2 and steam.

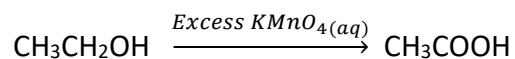


2. **Esterification reaction:** Alkanols react with alkanolic acid to produce esters and water. The process is reversible and is catalyzed with concentrated H_2SO_4 acid e.g.

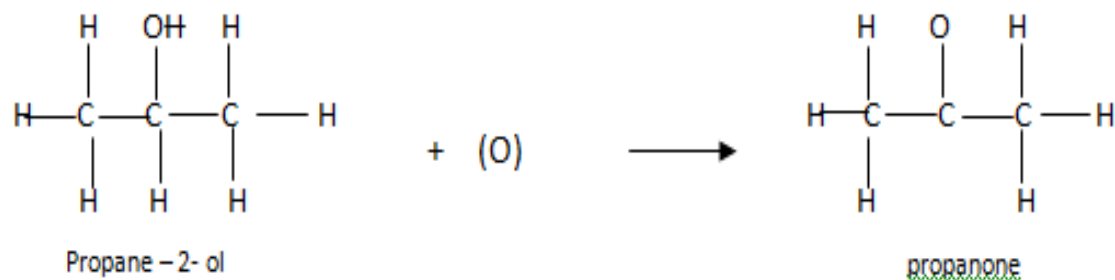
3. **Oxidation reaction:** (a) primary alkanol react with oxidizing agents to form alkanals oxidizing agents are acidified KMnO_4 (aq) or $\text{K}_2\text{Cr}_2\text{O}_7$ (aq)



Oxidation of primary alkanols with excess oxidizing agent gives alkanolic acid. Oxidation of ethanol in the presence of excess acidified KMnO_4 (aq) gives ethanoic acid.

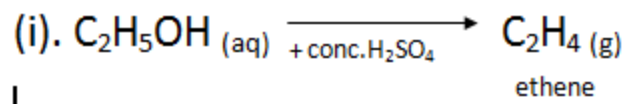


Generally primary alkanols are oxidized to alkanals (as in equation (1) above) and then finally to alkanolic acids (as in equation (ii) above) if excess oxidizing agent is used. Secondary alkanols are oxidized to Alkanones e.g.

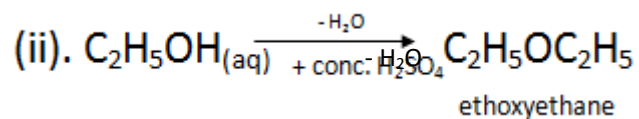


Tertiary alkanols are not oxidized at all.

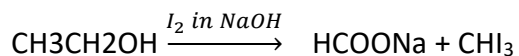
3. **Dehydration Reaction:** Alkanols undergo dehydration reactions as follows with concentrated H_2SO_4 acid. A dehydration agent removes the elements of water from a substance and changes the chemical nature of that substance. Dehydration of alkanol at 170°C gives alkene..So, ethanol is dehydrated to ethane.



But when alkanol is in excess and the reaction is carried out at 140°C , alkoxyalkane is formed. For example dehydration of excess ethanol at 140°C give ethoxyethane.



4. Ethanol and 2-ol alkanol(that is an alkanol that has methyl group bonded to the carbon atom that carries the hydroxyl group) react with mixture iodine in sodium hydroxide to give a yellow precipitate with antiseptic smell. This called iodoform test. The methyl group is converted to tri-iodomethane(iodoform



Propan-2-ol, butan-2-ol etc give positive iodoform test. But propan-1-ol does not give positive iodoform test because ethyl group is attached to the carbon that carries the hydroxyl group.

Period 4: LABORATORY TEST AND USES OF ALKANOLS

TEST FOR ALKANOLS

Add ethanoic acid to the unknown solution, then add few drops of concentrated tetraoxosulphate (vi) acid and heat gently. A fruity or fragrance smell indicates the presence of alkanol in the unknown solution.

Lucas reagent can also be used to distinguish between the classes of alkanols. When Lucas reagent (a solution of zinc chloride in hydrochloric acid) is added to alkanol, if tertiary alkanol is present, a white precipitate is formed immediately (the solution becomes cloudy), a secondary alkanol will give a white precipitate in five minutes while a primary alkanol will not give a white precipitate in five minutes.

1. Useful as an important solvent to dissolve perfumes, drugs, flavoring extracts etc.
2. As fuel either by itself or mixed with petrol in racing cars or rockets.
3. As anti-freeze in automobile radiators
4. Use in alcoholic beverages e.g. brandy etc.
5. As a preservative for biological specimens.

GENERAL EVALUATION

OBJECTIVE TEST:

1. Functional group for the alkanol is
(a) $C_nH_{2n+1}OH$ (b) C_nH_{2n-2} (c) carboxylic group (d) hydroxyl group
2. Primary alkanols are oxidized to carboxylic acid; secondary alkanols are oxidized to alkanones, while tertiary alkanols are (a) oxidized to alkanols (b) oxidized to mixture of alkanes (c) oxidized to alkanol (d) not oxidized.

3. The solubility of alkanols in water is due to (a) their covalent nature (b) hydrogen bonding (c) their low boiling point (d) their low freezing point
4. the general molecular formulae for the alkanols is (a) $C_nH_{2n+2}OH$ (b) $C_nH_{2n+1}OH$ (c) $C_nH_{2n+4}OH$ (d) $C_nH_{2n-2}OH$

ESSAY QUESTIONS:

1. Use equations only to explain the process of fermentation of starchy food to obtain ethanol
2. Write an equation to show esterification reaction of ethanol.

WEEKEND ASSIGNMENT

Read about fermentation of starchy food to prepare ethanol on page 541 of new School chemistry by Osei Yaw Ababio

WEEK END ACTIVITY

Take any bottled juice; place it under the sun for 1 or 2 days. Open it and taste it. What do you observe?

PREREADING ASSIGNMENT

Read the whole term's work.

WEEK 10 Revisions

WEEK 11 Examinations

