

# CHAPTER ONE: OXIDATION AND REDUCTION REACTIONS



## Competency:

*You should understand oxidation and reduction in terms of gain or loss of oxygen and in terms of electron transfer, and appreciate that the two processes always occur together.*

## Learning outcomes

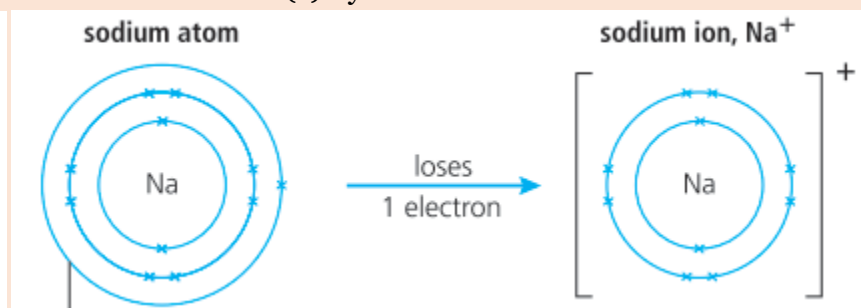
You should be able to:

- a) Understand the processes of oxidation and reduction and their importance in the chemical industry (u, s)
- b) Explain redox reactions in terms of electron transfer (u)
- c) Understand the changes that take place during the electrolysis of some compounds (u, s)

## INTRODUCTION

**Oxidation** can be defined as:

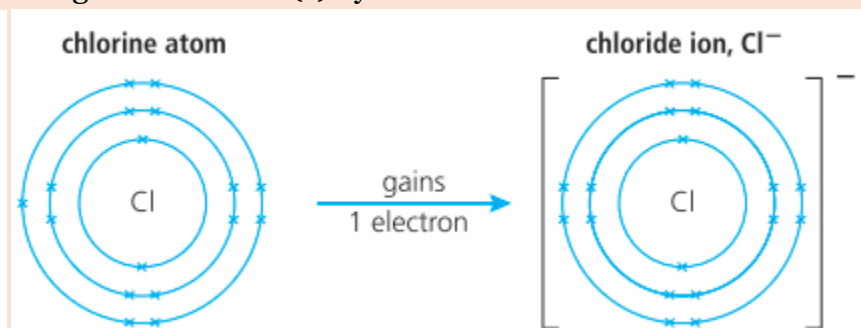
- i) The gain of oxygen by a chemical substance.
- ii) The loss of hydrogen by a chemical substance.
- iii) The loss of electron(s) by a chemical substance.



- iv) Increase in oxidation number of a chemical substance

**Reduction** can be defined as:

- i) The loss of oxygen by a chemical substance.
- ii) The of gain hydrogen by a chemical substance.
- iii) The gain of electron(s) by a chemical substance.



- iv) Decrease in oxidation number of a chemical substance

### Summary

Oxidation	Reduction
1. gain of oxygen	1. Loss of oxygen
2. loss of hydrogen	2. Gain of hydrogen
3. loss of electrons	3. Gain of electrons
4. increase in oxidation number	4. Decrease in oxidation number



**Activity:** Explain what happens when dry hydrogen gas is passed over heated copper(II) oxide.

## Redox reaction

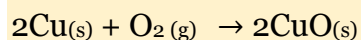
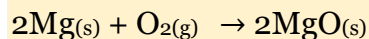
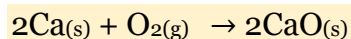
**Redox reaction** is a reaction in which oxidation and reduction occur simultaneously. Alternatively, an oxidation-reduction (redox) reaction is a type of chemical reaction that involves a transfer of electrons between two species. The word **redox** is an acronym of two words **reduction** and **oxidation**. *Redox reactions are common and vital to some of the basic functions of life, including photosynthesis, respiration, combustion, and corrosion or rusting. Most reactions are redox.*

**Oxidizing agents** are substances that accept electrons, allowing the other substance to be oxidized. By accepting electrons, the oxidizing agent is reduced.

**Reducing agents** are substances that donate electrons, allowing the substance to be reduced. By donating electrons, the element(s) in the reducing agent are oxidized.

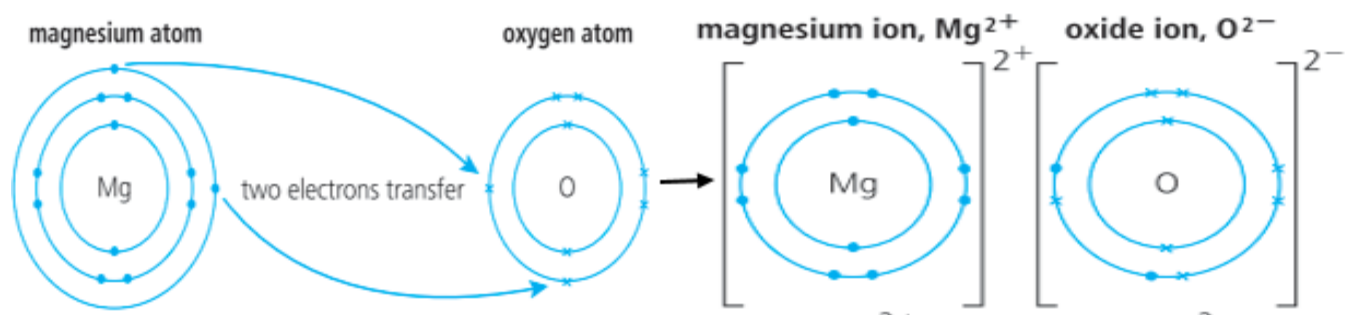
**Examples of simple redox reactions in which the metal is being oxidised by losing electrons in the outermost energy level and oxygen is being reduced by accepting electrons**

Burning of metals in dry oxygen.



## Formation of magnesium oxide

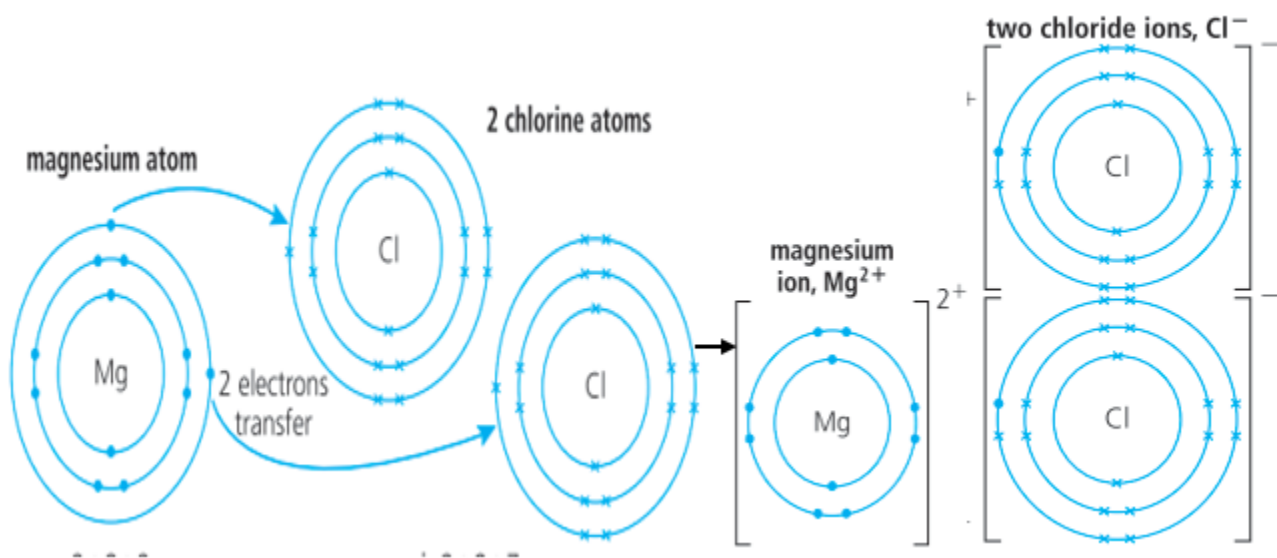
A magnesium atom has 2 outer most electrons and an oxygen atom has 6. When magnesium burns in oxygen, each magnesium atom loses its 2 outermost electrons to an oxygen atom. Magnesium ion and oxide ions are formed.



*The magnesium atom is oxidised to magnesium ion after losing two electrons and the oxygen atom is reduced to oxide ion after accepting two electrons from magnesium. The reaction is redox since oxidation and reduction are occurring simultaneously.*

### **Formation of magnesium chloride**

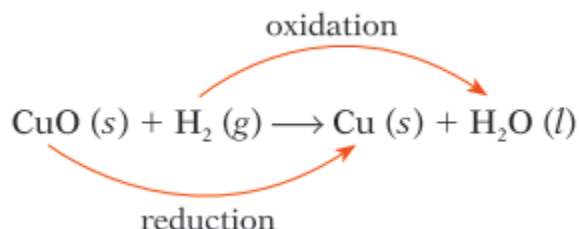
When heated magnesium reacts with dry chlorine, each magnesium atom reacts with two chlorine atoms, to form magnesium chloride. Each ion has 8 outermost electrons.



*The magnesium atom is oxidised to magnesium ion after losing two electrons and each of chlorine atom is reduced to chloride ion after accepting one electron from magnesium. The reaction is redox since oxidation and reduction are occurring simultaneously.*

## Reaction between copper(II) oxide and hydrogen

Look again at the reaction between copper(II) oxide and hydrogen. Copper(II) oxide loses oxygen, and hydrogen gains oxygen.



*So the copper(II) oxide is reduced, and the hydrogen is oxidised. Oxidation and reduction always take place together. So the reaction is called a redox reaction.*

## The reaction between chlorine and potassium bromide

When chlorine gas is bubbled through a colourless solution of potassium bromide, the colorless solution turns reddish brown due to this reaction.



The half-equations for the reaction are



The overall ionic equation after combining the half equations



## OXIDATION NUMBER (OXIDATION STATE)

The oxidation state, or oxidation number, is **the hypothetical charge of an atom if all of its bonds to other atoms were fully ionic**. It describes the degree of oxidation (loss of electrons) of an atom in a chemical compound. **Oxidation number** is defined as the total number of electrons that an atom either gains or loses to form a chemical bond with another atom. **Alternatively an oxidation number is a number assigned to an element to show its state of oxidation or reduction.** If an element's oxidation number increases

from reactant to product, it is being oxidized. If an element's oxidation number decreases from reactant to product, it is being reduced. Conceptually, the oxidation state may be positive, negative or zero. The oxidation state of an atom does not represent the "real" charge on that atom..

How do we know whether or not an element has gained or lost electrons?

To determine which elements have been oxidized or reduced, we look at changes in the **oxidation number** of the element.

### **RULES FOR ASSIGNING OXIDATION STATES**

In determining the oxidation state of an atom, there are some guidelines to follow:

- a) The oxidation state of an individual atom is 0.
- b) The total oxidation state of all atoms in: a *neutral species* (element, molecules or compound) is 0 and in an *ion* is equal to the ionic charge.
- c) Group I metals have an oxidation state of +1 and Group II an oxidation state of +2
- d) The oxidation state of halogens(Group VII elements) is -1 in compounds
- e) Hydrogen generally has an oxidation state of +1 in compounds except in metal hydrides where it is -1
- f) Oxygen generally has an oxidation state of -2 in compounds except in metal peroxides
- g) In binary metal compounds, Group VII elements have an oxidation state of -1 except when combined with oxygen, Group VI elements of -2, and Group V elements of -3.
- h) For a complex ion, the sum of the positive and negative oxidation numbers of all elements in the ion equals the charge on the complex ion.



Element	Usual oxidation state	Exceptions
Group 1 metals	Always +1	
Group 2 metals	Always +2	
Oxygen	Usually -2	Peroxides and $F_2O$ (see below)
Hydrogen	Usually +1	Metal hydrides (-1) (see below)
Fluorine	Always -1	
Chlorine	usually -1	Compounds with O or F (see below)

## Exceptions

**Hydrogen in the metal hydrides:** Metal hydrides include compounds like sodium hydride,  $NaH$ . Here the hydrogen exists as a hydride ion,  $H^-$ . The oxidation state of a simple ion like hydride is equal to the charge on the ion in this case, -1. Alternatively, the sum of the oxidation states in a neutral compound is zero. Because Group I metals always have an oxidation state of +1 in their compounds, it follows that the hydrogen must have an oxidation state of -1 (+1 -1 = 0).

**Oxygen in peroxides:** Peroxides include hydrogen peroxide,  $H_2O_2$ . This is an electrically neutral compound, so the sum of the oxidation states of the hydrogen and oxygen must be zero. Because each hydrogen has an oxidation state of +1, each oxygen must have an oxidation state of -1 to balance it.



**Oxygen in  $F_2O$ :** The deviation here stems from the fact that oxygen is less electronegative than fluorine; the fluorine takes priority with an oxidation state of -1. Because the compound is neutral, the oxygen has an oxidation state of +2.

**Chlorine in compounds with fluorine or oxygen:** Because chlorine adopts such a wide variety of oxidation states in these compounds, it is safer to simply remember that its oxidation state is not -1, and work the correct state out using fluorine or oxygen as a reference.

### Activity

Assign correct oxidation numbers to all elements in the following substances.

(i)  $H_2O$  (ii)  $CaO$  (iii)  $Na_2S$  (iv)  $HCl$  (v)  $Na_2O_2$  (vi)  $P_2O_5$  (vii)  $CaF_2$ , (viii)  $SO_4^{2-}$

### Activity two: Read the articles below

#### Fortifying Food with Iron

Iron is an essential mineral in our diet; iron-containing compounds like the heme protein in hemoglobin could not function without it. Most biological iron has the form of the  $Fe^{2+}$  ion; iron with other oxidation numbers is almost inconsequential in human biology (although the body does contain an enzyme to reduce  $Fe^{3+}$  to  $Fe^{2+}$ , so  $Fe^{3+}$  must have some biological significance. To ensure that we ingest enough iron, many foods are enriched with iron. Although  $Fe^{2+}$  compounds are the most logical substances to use, some foods like bread and breakfast cereals are the most well-known examples use “reduced iron” as an ingredient. Reduced iron is simply iron metal; iron is added as a fine metallic powder. The metallic iron is oxidized to  $Fe^{2+}$  in the digestive system and then absorbed by the body, but the question remains: Why are we ingesting metallic iron? Why not just use  $Fe^{2+}$  salts as an additive?

### Response

Although it is difficult to establish conclusive reasons, a search of scientific and medical literature suggests a few reasons. One reason is that fine iron filings do not

affect the taste of the product. The size of the iron powder (several dozen micrometers) is not noticeable when chewing iron-supplemented foods, and the tongue does not detect any changes in flavour that can be detected when using  $\text{Fe}^{2+}$  salts.  $\text{Fe}^{2+}$  compounds can affect other properties of foodstuffs during preparation and cooking, like dough pliability, yeast growth, and color. Finally, of the common iron substances that might be used, metallic iron is the least expensive. These factors appear to be among the reasons why metallic iron is the supplement of choice in some foods.

## CALCULATING OXIDATION STATE (OXIDATION NUMBER) OF CHEMICAL SUBSTANCES

The oxidation number of an atom is defined as **the charge acquired by the atom when they form ionic bonds with atoms of another element**. We always assume the compounds to be ionic for the calculation of oxidation number. Since every molecule possesses partial ionic and partially covalent character we assume ionic character for ease of calculation.

To find the oxidation number of an atom we need to follow the following steps:

**Step 1:** Assume the oxidation number of the atom to be  $x$  which you need to calculate.

**Step 2:** Mention the oxidation state of other bonded atoms and multiply it with the number of such atoms present in one molecule, compound or any chemical species.

**Step 3:** Write the oxidation number of all the atoms in the chemical substance in a linear sum format and equate it to the overall charge of the chemical substance such as a compound

**Step 4:** Solve for  $X$ .

Worked examples

(a) Carbon monoxide (CO).

**Determine the oxidation number of C in CO**

- ❖ In Carbon monoxide we know that bond involved between the Carbon atom and the Oxygen atom is covalent
- ❖ Since oxygen can gain a maximum of two electrons so its oxidation number will be  $-2$ .
- ❖ Now assume the oxidation state of Carbon in CO to be  $X$ .
- ❖ Since CO is an electrically neutral molecule i.e. its overall charge is zero hence the sum of oxidation numbers of C and O will be zero i.e.

$$X + (-2) = 0$$

$$X - 2 = 0$$

$$X = +2$$

Hence the oxidation state of the carbon atom in carbon monoxide is **+2**.

**(b) Oxidation Number of Sulphur in  $\text{H}_2\text{SO}_4$**

**Solution:**

**Step 1:** Assume the oxidation number of sulphur to be x

**Step 2:** The oxidation number for Hydrogen is +1 and for O is -2.

**Step 3:** Since the overall charge on the molecule is 0, therefore  $2(+1) + X + 4(-2) = 0$

**Step 4:**  $2 + X - 8 = 0 \Rightarrow X - 6 = 0 \Rightarrow X = +6$

Hence, the oxidation number of Sulphur in  $\text{H}_2\text{SO}_4$  is **+6**

**(c) Oxidation Number of Chromium in  $\text{Cr}_2\text{O}_7^{2-}$  (Dichromate(VI) ion)**

**Solution:**

**Step 1:** Assume the oxidation number for Chromium be X

**Step 2:** The oxidation number for oxygen is -2

**Step 3:** Since the molecule has an overall charge of -2 the equation can be written as

$$2X + 7(-2) = -2$$

**Step 4:**  $2X - 14 = -2 \Rightarrow 2X = +12 \Rightarrow X = +6$

Hence the oxidation number of chromium in  $\text{Cr}_2\text{O}_7^{2-}$  is **+6**.

**(d) Oxidation Number of Nitrogen in Ammonium nitrate i.e.  $\text{NH}_4\text{NO}_3$**

**Solution:**

**Step 1:** In this case, we will split the molecule into two ions Ammonium ion ( $\text{NH}_4^+$ ) and Nitrate ion ( $\text{NO}_3^-$ )

**Step 2:** We will find the oxidation number of Nitrogen in each ion. Assume the oxidation number for Nitrogen to be  $X$  in each case

**Step 3:** For the Ammonium ion, since the molecule has an overall +1 charge, hence equation can be written as  $X + 4(+1) = +1$

**Step 4:** Solving the equation  $X + 4 = 1 \Rightarrow X = -3$

**Step 5:** In the case of nitrate ion, the oxidation number for oxygen is -2 and the molecule has an overall charge of -1.

**Step 6:** Hence equation, in this case, will be  $X + 3(-2) = -1$

**Step 7:** Solving the equation we get  $X - 6 = -1 \Rightarrow X = +5$

Hence the oxidation number of Nitrogen in ammonium nitrate is -3 and +5

Reminder: With a few exceptions

- a) Oxidation number of I group elements = +1.
- b) Oxidation number of II group elements = +2.
- c) Oxidation number of III group elements = +3.
- d) Oxidation number of IV group elements = +4.
- e) Oxidation number of V group elements = -3.
- f) Oxidation number of VI group elements = -2.
- g) Oxidation number of VII group elements = -1.
- h) Oxidation number of VIII group elements = 0.

### (e) Oxidation Number of Manganese in $\text{KMnO}_4$

**Solution:**

Let the oxidation number of Manganese in  $\text{KMnO}_4$  be equal to  $X$

The oxidation number of Potassium = +1

The oxidation number of oxygen is = -2

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

$$\Rightarrow 1 + X - 8 = 0$$

$$\Rightarrow X - 7 = 0$$

$$\Rightarrow X = +7$$

The oxidation number of Manganese in  $\text{KMnO}_4$  is +7.

**(f) Oxidation Number of Carbon in  $\text{Na}_2\text{CO}_3$**

**Solution:**

Oxidation number of oxygen is = -2

Oxidation number of Sodium = +1

Let the oxidation number of carbon be X

$$\Rightarrow 2(+1) + X + 3(-2) = 0$$

$$\Rightarrow 2 + X - 6 = 0$$

$$\Rightarrow X - 4 = 0$$

$$\Rightarrow X = +4$$

Therefore X = +4

Oxidation Number of Carbon in  $\text{Na}_2\text{CO}_3$  is +4

**(g) Oxidation Number of Nitrogen in Ammonium nitrite**

**Solution:**

Ammonium nitrite is an ionic compound containing  $\text{NH}_4^+$  and  $\text{NO}_2^-$  ions.

**Oxidation number of nitrogen in  $\text{NH}_4^+$**

Let oxidation number of N atom is x

$$\Rightarrow x + 4 = +1$$

$$\Rightarrow x = 1 - 4 = -3$$

### **Oxidation number of nitrogen in $\text{NO}_2^-$**

Let oxidation number of Nitrogen atom be Y

$$Y + 2(-2) = -1$$

$$\Rightarrow Y - 4 = -1$$

$$\Rightarrow Y = -1 + 4 = +3$$

Thus, one atom of Nitrogen in Ammonium Nitrite is in the -3 oxidation state, while the other nitrogen atom is in the +3 oxidation state.

### **(h) Oxidation Number of N in $\text{NH}_3$**

#### **Solution:**

Let Oxidation Number of N in  $\text{NH}_3$  be x and we know that Oxidation Number of H is +1.

$$\Rightarrow x + 3(+1) = 0$$

$$\Rightarrow x + 3 = 0$$

$$\Rightarrow x = -3$$

Hence, the Oxidation Number of N atom in  $\text{NH}_3$  is -3.

### **EXPLAINING THE CONCEPT OF REDUCTION AND OXIDATION USING IONIC EQUATIONS**

#### **Ionic equations**

Most of the equations that you have studied so far are molecular (condensed) equations. In molecular equation, all reactants and products are written as electrically neutral chemical substances. In ionic equations, we express dissolved (soluble) ionic compounds as their dissociated ions, like  $\text{Na}^+(\text{aq})$  and  $\text{Cl}^-(\text{aq})$ . This means that we have to know what is soluble and insoluble.

#### **General guidelines for writing ionic equations**

- Write a well-balanced molecular equation with correct physical states of reactants and products
- Soluble ionic compounds ionize.
- Pure liquids like water, gases do not ionize.

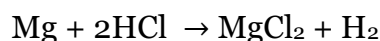


- d) Solids do not usually ionize except in few special cases
- e) Cancel out the spectator ions. Spectator ions are those ions that appear on both sides of the equation
- f) Write the net ionic equation.

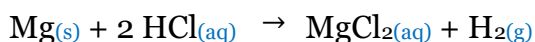
Consider the examples below

### 1. Reaction between magnesium and dilute hydrochloric acid

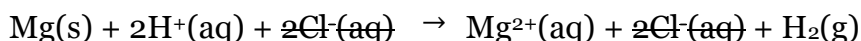
(a) Write the molecular equation and balance it



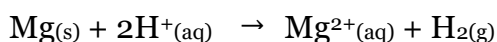
(b) Determine the state of each substance (gas, liquid, solid, aqueous). Use the [solubility rules](#) to determine which of the ionic compounds are soluble in water. Soluble ionic compounds are identified with an (aq), insoluble ones with an (s). Most elements and covalent compounds (covalent compounds are formed when two or more nonmetallic atoms of nonmetals are chemically bonded to each other) are insoluble in water and should be shown with an (s), (l) or (g). Acids ionize when dissolved in water.



(c) Write the ionic equation by breaking all the soluble ionic compounds (those marked with an (aq)) into their respective ions. Each ion should be shown with its charge and an (aq) to show that it is present in solution. Use coefficients to show the number of each ion present. Rewrite the elements and covalent compounds as they appeared in the preceding step.

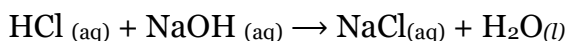


(d) Write the net ionic equation by removing the spectator ions. Spectator ions are those ions that appear exactly the same on each side of the ionic equation.

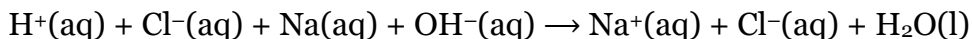


### 2. Reaction between sodium hydroxide and dilute hydrochloric acid

**Write a well balanced molecular equation with state symbols**

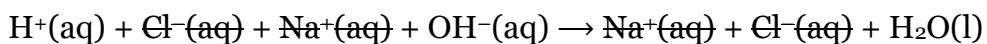


**Rewrite by separating the soluble ionic compounds into their dissociated ions**



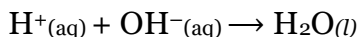
Note that water is not separated because it is not an ionic compound. As a simple molecule, the atoms do not separate and remain bonded as a discrete molecule of  $\text{H}_2\text{O}$ .  $\text{HCl}$  is separated. While pure  $\text{HCl}(\text{g})$  is a simple molecule, it is an **acid** that undergoes **ionisation in water** to form hydrogen ions. Therefore, dissolved  $\text{HCl}(\text{aq})$  should be treated as a soluble ionic compound, with dissociated ions.

**Cancel out common ions, which are the spectator ions**



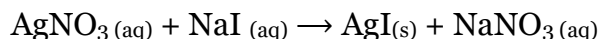
Sodium and chloride ions remain dissolved in the solution. They are the spectator ions, that watch and do *not* undergo the chemical reaction themselves.

**Write the net ionic equation for neutralisation**



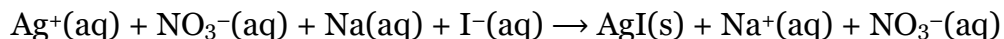
**3. Reaction between sodium iodide and silver nitrate**

Write the molecular equation



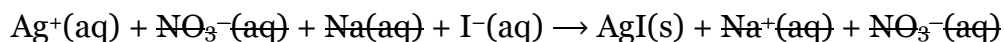
As silver iodide is insoluble, it precipitates as a solid. Therefore, its state symbol is (s) instead of (aq).

**Rewrite by separating the soluble ionic compounds into their dissociated ions**



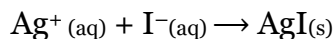
Note that  $\text{AgI}$  is not separated because it is an **insoluble ionic compound**.

**Cancel out common ions, which are the spectator ions**



Sodium and nitrate ions remain dissolved in the solution. They are the spectator ions, that watch and do *not* undergo the chemical reaction themselves.

### Write the net ionic equation for precipitation



#### Activity

- Write the ionic equation for the **acid-metal reaction** between zinc and dilute sulphuric acid to form zinc sulphate and hydrogen gas.
- Write the ionic equation for the **acid-carbonate reaction** between dilute hydrochloric acid and sodium carbonate to form sodium chloride, water, and carbon dioxide.
- Write the ionic equation for the **acid-carbonate reaction** between hydrochloric acid and insoluble magnesium carbonate to form magnesium chloride, water, and carbon dioxide.
- Write the ionic equation for the **neutralisation reaction** between sodium hydroxide and sulphuric acid to form sodium sulphate and water.
- Write the ionic equation for the **precipitation reaction** between calcium chloride solution and sodium carbonate solution.
- Write the ionic equation for the **redox reaction** between copper and concentrated nitric acid to form copper nitrate, nitrogen dioxide, and water.

## Explaining oxidation and reduction using ionic equations

To help us distinguish between oxidation and reduction, we can use a couple of mnemonic devices. One phrase to remember is “OIL RIG,” or “Oxidation Involves Loss of electrons, Reduction Involves Gain of electrons.” Another option is the phrase “LEO the lion says GER,” or “*Lose Electrons Oxidation, Gain Electrons Reduction.*”

When an atom or ion gains electrons, its charge becomes more negative. When a neutral atom turns into a negative ion, it is gaining electrons and being reduced. In addition, when a positive ion's charge decreases, for example, from 2+ to 1+ or from 1+ to 0, it is gaining electrons and being reduced.

When an atom or ion loses electrons, its charge becomes more positive. When a neutral atom turns into a positive ion, it is losing electrons and being oxidized. Furthermore, when a negative ion's charge increases, for example, from 2- to 1- or zero, it is losing electrons and being oxidized.

## ORES IN UGANDA

### a) Various ores in Uganda

An **ore** is a naturally occurring substance from which an element can be extracted.

#### Iron ores (haematite and magnetite)



Iron ores occurs mainly as haematite and magnetite. High quality iron ores occur in several parts of Uganda. The biggest deposit is haematite found at Muko in Kabale and Kisoro districts with total resources of over 50 million tons. Similar haematite with a resource of 2 million tons occurs at Mugabuzi in Mbarara district.

Magnetite ore occurs around the carbonatite complexes of Sukulu in Tororo and Bukusu in Manafwa district. At Sukulu, magnetite occurs in residual soils with apatite (phosphate). A resource of 45 million tons has been estimated. Twenty three (23) million tons have been estimated at Nakhupa, Nangalwe and Surumbusa. Namekhara contains an estimated resource of 18 million tons. Nangalwe is the most promising deposit.

There has been only limited production of iron ores in Uganda mainly for use as an additive in the steel scrap smelting and for special cement by Hima Cement. The Madhvani Group of Uganda planned to restart production of rebar at its rolling mill Steel Corporation of East Africa Ltd.

*Uses: The principal use of iron ore is in making of steel*

### Lead ore (Galena)

Galena, a mineral containing lead (with minor zinc and gold), occurs in quartz veins at Kampono, Kanyambogo and Kitaka in Kitomi Forest, Ibanda district. It also occurs associated with tin (cassiterite) at Kikagati. These deposits are all very small. Galena was mined only at Kitaka and production totaled only 750 tons over a 13- year period to 1960 when mining stopped.

*Uses: In making motor vehicle batteries and heavy metal shield for nuclear radiation protection.*

### Lithium ore (Amblygonite)

Amblygonite the ore that contains lithium mineral is found to be associated with pegmatites at Ruhuma in Kabale district; Mwerasandu, Lwamwire and Nyabushenyi in Ntungamo district; Lunya in Mukono district; Nampeyo and Mbale Estate in Mubende district.

Lithium has been exploited only from the Nyabushenyi (Ntungamo) and Mbale Estate (Mubende) pegmatites. Most of the pegmatites are small and of irregular bodies, which mitigates against large-scale exploitation, but is well suited to small-scale production by local entrepreneurs.

*Uses: Amblygonite the ore of lithium is used mainly as a non-metallic mineral, especially in chemicals.*

**Tin ore (cassiterite):** Several tin deposits occur throughout southwest Uganda in quartz-mica veins in contact with granitic bodies intruded into shales and sandstone host rocks of Karagwe- Ankolean System confirmed to be a tin-field province that extends southwest into Rwanda and Congo and northern Tanzania. The individual veins are thin (rarely more than a metre in width) irregular and of small tonnage potential.

Tin occurs in single veins at Mwerasandu, Kaina, Nyinamaherere in Ntungamo district; Kikagati and Ndaniyankoko in Isingiro district and Burama Ridge on the Kabale/Ntungamo border. Stockworks and sheeted vein swarms occur at Rwaminyinya in Kisoro district and Kitezo in Mbarara district and these may have large tonnage potential.

Uganda's tin concentrate production 1927 to 2001 totaled about 13,000 tons. The bulk of this production came from hard rock deposits, with minor eluvial production and no alluvial production. The largest deposit was Mwerasandu in Ntungamo district and substantial production also came from Kikagati in Isingiro district. Other producers were Rwaminyinya, Burama Ridge, Ndaniyankoko, Kaina and Nyinamaherere.

*Uses: used mainly for coating iron/steel to minimize rusting and also making cans for the food industry.*

## **Magnesite**





Magnesite ( $\text{MgCO}_3$ ) occurs at Lolung north of Moroto town. The quality is moderate but it is a small deposit. Another similar but smaller deposit occurs at Lolukei at the Karamoja-South Sudan boarder.

### Manganese ore

Manganese ore occurrences have been found in Isandara, Gulika, Kirongo and Gweitengya in western Uganda. No deposits of commercial value have been found.

### Nickel ore

Nickel ore occurs at the pyrite-pyrrhotite concentrate stock pile at Kasese. The proven reserve is 1.1 million tonnes.

### Chromite

It is found about 6 km long north of Mt. Moroto at Nakiloro and Lolung. The chromite is rarely pure and usually contains either talc or antigorite. Analyses in the past show that  $\text{Cr}_2\text{O}_3$  generally forms between 45% to 55% of the ore. There has not been any chromite production in Uganda.

### Titanium ores

The two ores ilmenite and rutile are well distributed throughout the country but are not concentrated in workable deposits. Rutile occurs in sands of the Kalere

River in small amounts. Iron ores at Bukusu and Sukulu contain a high proportion of titanium and magnetite.

*Uses: The main use of titanium is currently as a non-metallic mineral, especially as titanium oxide pigment in paint, paper, rubber, etc. rather than as a metal. The metal is used in the manufacture of corrosive resistant steel.*

#### **Tungsten ores (wolframite/scheelite):**

Numerous tungsten deposits occur in quartz vein type closely associated with granitoid intrusions. Tungsten occurs in quartz veins closely associated with granitoid intrusions at Nyamuliro (also called Bjordal Mine), Rushunga and Ruhija in Kabale district; Kirwa, Mutolere, Rwamanyinya, Mpororo and Bahati in Kisoro district; Kyasampawo in Mubende district, Buyaga in Lyantonde district and Nakaseta in Mityana district.

The main deposits that have been mined are Nyamuliro (also called Bjordal Mine), Kirwa, Ruhija, Mutolere, Rwamanyinya and Bahati in Kabale and Kisoro districts. Others are Kyasampawo and Mbale Estate in Mubende district and Buyaga in Rakai district.

Uganda's wolframite concentrate production from 1935 to 2001 totaled over 5,000 tons and this came from the various low-grade deposits. Kirwa Mine one of the large producer from late 1940's to 1979 has a resource estimated at 1.25 million tons averaging 0.19%  $\text{WO}_3$ . Bjordal Mine is currently being re-developed by M/S Krone Uganda Ltd. and production is up to 15 tons/month.

*Uses: Tungsten is mainly used in making armour plate in military equipment, manufacture of filaments for electric bulbs and in making tungsten-carbide for drilling bits.*

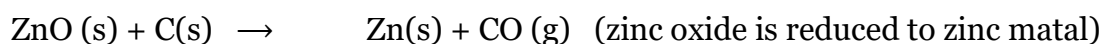
#### **Rock salt**

Salt for human and animal consumption has been extracted on a small scale from hot springs at Kibiro in Hoima district and on a larger scale from the floor of crater lakes Katwe and Kasenyi in Kasese district for centuries. The rock salt is a mixture of sodium and potassium chlorides. Currently, production is based on solar evaporation in ponds and the product is crude due to mixing of the salts during fractional crystallisation as well as with mud at the bottom. The reserves is estimated to be about 22.3 million tons.

## THE RELEVANCE OF REDUCTION/ OXIDATION TO METAL EXTRACTION

**Extraction of metal** refers to the process of obtaining metal from its ore. **Ore** is a naturally occurring substance from which an element can be extracted. **The electrochemical series** is a valuable guide to study of metals and their compounds. Gold, silver, mercury and copper are the only unreactive metals near the bottom of the **electrochemical series that can occur naturally**. The **method of metal** extraction depends on the **stability of the ore**. **Mercury** near the bottom of the electrochemical series is obtained by simply by **heating its sulphide ore**. Moderately reactive metals in the middle of the electrochemical series that is zinc, tin, iron, lead are obtained by reducing their oxides with coke or carbon monoxide. Coke is cheap but contaminates the metal. The **sulphide and the carbonate ores** are first converted to oxides. **Aluminium oxide is very stable** that it has to be reduced electrolytically (electrolytic reduction). Very reactive metals on top of the electrochemical series that is potassium, sodium, calcium, magnesium are obtained by electrolysis of their fused or molten metal chloride (chloride ore). Thus potassium, sodium, calcium, magnesium and aluminium are obtained by **electrolysis**.

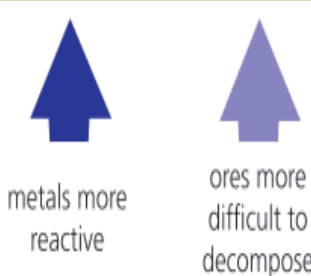

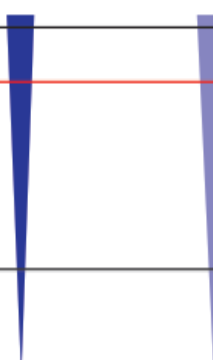

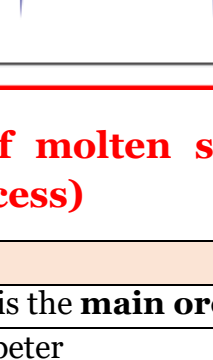
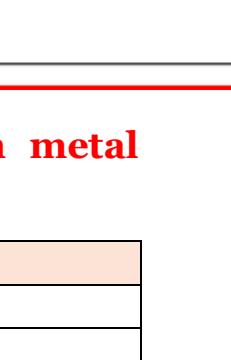
**Extraction of metals is a reduction process. During electrolysis metal ions accept electrons to form a metal (Reduction)**  $\text{Al}^{3+}(\text{l}) + 3\text{e}^- \rightarrow \text{Al}(\text{l})$  and for the oxides of moderately reactive metals the coke or carbon monoxide removes the oxide ions from the oxides



Therefore, **very reactive metals** in the electrochemical series (K, Na, Ca, Mg and Al) are extracted by electrolysis of their fused or molten salt and (oxide for aluminium).

**Moderately reactive metals** found as oxides, carbonates, or sulphides are extracted by **reduction** of their oxides by carbon or carbon monoxide.

Metals lower or at the bottom of the electrochemical series (mercury, silver, gold and times copper) mainly occur as free metals in the earth's crust. **These metals are not very reactive.**

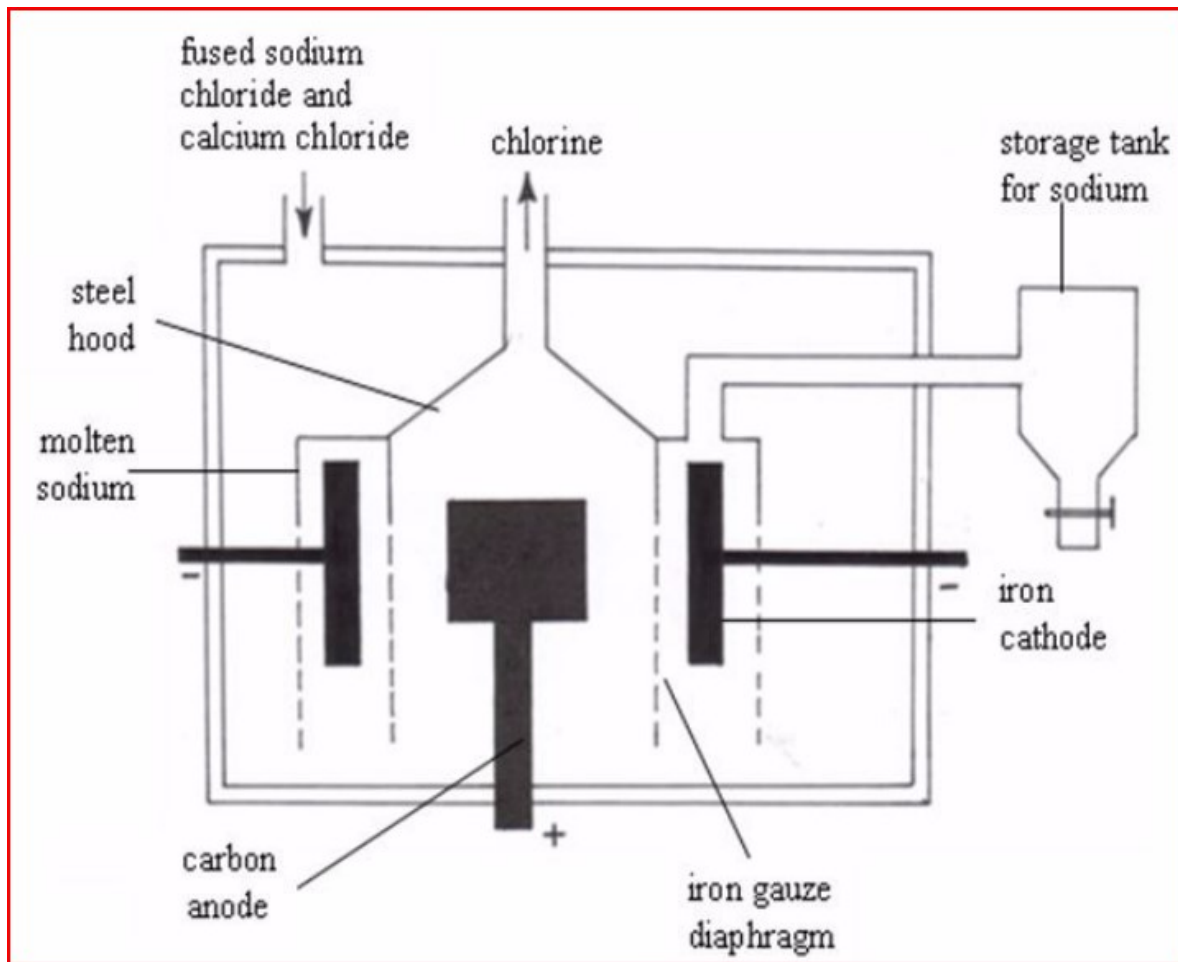
Metal		Method of extraction from ore	
potassium		electrolysis	
sodium			
calcium			
magnesium			
aluminium			
carbon		heating with a reducing agent – carbon or carbon monoxide	
zinc			
iron			
lead			
silver		occur naturally as elements so no chemical reactions needed	
gold			

### Reduction of molten sodium ion to form sodium metal (Downs' process)

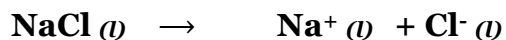
Ores.	Chemical name
Rock salt is the <b>main ore</b>	Sodium chloride
Chile saltpeter	Sodium nitrate
Soda ash	Sodium carbonate

The metal occurs chiefly as **rock salt and it is the main ore**. Sodium metal is **very reactive and therefore extracted by electrolytic reduction (electrolysis)**. Sodium metal is extracted by electrolysis of fused or molten sodium chloride (rock salt). The sodium chloride as obtained from rock salt or sea water. This is called **Downs' process** and the entire set up used is called **Downs' cell**.

## The Downs' cell of sodium extraction.



During this electrolysis, sodium chloride dissociates into sodium ions and chloride ions.



Sodium ions migrate to the iron cathode while the chloride ions migrate to the graphite anode where they are discharged forming sodium and chlorine respectively. Chlorine is a valuable by-product of the process.

Ions that migrate to graphite anode	Ions that migrate to the circular iron cathode.
$\text{Cl}^-$	$\text{Na}^+$

**Equation at the anode.**  $2\text{Cl}^- (l) \rightarrow \text{Cl}_2(g) + 2e$

*Chloride ions are being oxidised to chlorine molecules as they lose electrons*

**Equation at the cathode.**  $\text{Na}^+ (\text{l}) + \text{e}^- \rightarrow \text{Na}(\text{l})$ .

*Sodium ions are being reduced to sodium metal as they accept electrons.*

### Uses of sodium metal

- a) Used in the manufacture of sodium cyanide which is employed in the extraction of gold.
- b) Alloy of sodium and lead has been employed in the manufacture of tetraethyl lead which is an anti-knock additive used in petrol.
- c) Sodium is used for the withdrawal of heat from some nuclear reactors.
- d) Sodium has also found use in the reduction of titanium (IV) chloride to metal by heat.
- e) Manufacture of sodium vapour lamps well known for intensely yellow illumination they give.

### Reduction of aluminium ion to aluminium metal

The principle ore is called **Bauxite**.

#### Ores of aluminium.

Bauxite ( $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$ ) and the chemical name is aluminum oxide dihydrate. Cryolite ( $\text{Na}_3\text{AlF}_6$ ) chemical name is sodium hexafluoroaluminate. Corundum ( $\text{Al}_2\text{O}_3$ ) chemical name is aluminium oxide. Kaolin (china clay) ( $\text{Al}_2\text{Si}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$ ) used for making porcelain. Mica/ feldspar  $\text{K}_2\text{Al}_2\text{Si}_6\text{O}_{16}$ .

#### Electrolytic cell used to extract aluminium

um, copper, and manganese; **Magnalium** contain aluminium and magnesium; **Aluminium bronze** contains aluminium and copper.

### Uses of duralumin

Construction of air craft. Making of car bodies example piston rings and cylinder head. Making bicycle parts.

### Reduction zinc oxide to zinc metal



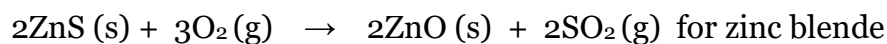
Zinc is extracted by **reduction of hot zinc oxide** using carbon (coke) or carbon monoxide.

### Ores of zinc.

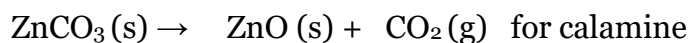
Ore of zinc	Chemical name	Chemical formula
Zinc blende	Zinc sulphide	ZnS
Calamine.	Zinc carbonate	ZnCO <sub>3</sub>

### Roasting

The concentrated ore is then roasted in air to convert it into the oxide:

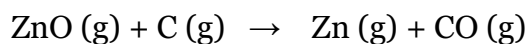


For the case of sulphur dioxide formed when zinc blende is roasted in air, it is frequently used for the manufacture of sulphuric acid in contact process.



### Reduction of zinc oxide to zinc metal

The **zinc oxide is reduced to metallic zinc by coke (carbon) and the coke is oxidized to carbon monoxide.** The zinc is led off as a vapour, which is then cooled by lead spray.



**Impure zinc** is obtained in this case and it is **purified by electrolysis.** Zinc is frequently granulated by running the molten metal into water.

### Alloy of zinc

**Brass** contains **zinc and copper**

### Uses of zinc

Used to coat iron forming galvanized iron to delay rusting.

Zinc is used as negative electrode of Leclanche batteries, example the outer case of dry batteries

Used to make alloys such as brass (copper and zinc).

### Uses of galvanized iron

Used for making buckets, dustbins, water tanks, bath tubs, wire fences, barbed wire and galvanized iron sheets.

### Reduction of iron(III) oxide, $\text{Fe}_2\text{O}_3$ or iron(II,III) oxide $\text{Fe}_3\text{O}_4$ to iron metal

The ores from which iron can be extracted.

Name of the ore.	Chemical name	Chemical formula
Haematite	Iron(III) oxide	$\text{Fe}_2\text{O}_3$
Magnetite or magnetic iron ore	IRON(II, III) oxide	$\text{Fe}_3\text{O}_4$
Spathic iron ore (siderite when relatively pure)	Iron(II) carbonate	$\text{FeCO}_3$
Iron pyrites	Iron(II) disulphide	$\text{FeS}_2$

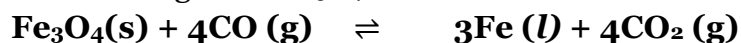
The carbon monoxide at high temperature (of about  $1000^\circ\text{C}$ ) reduces the haematite or **magnetite to metallic iron and itself to carbon dioxide.**

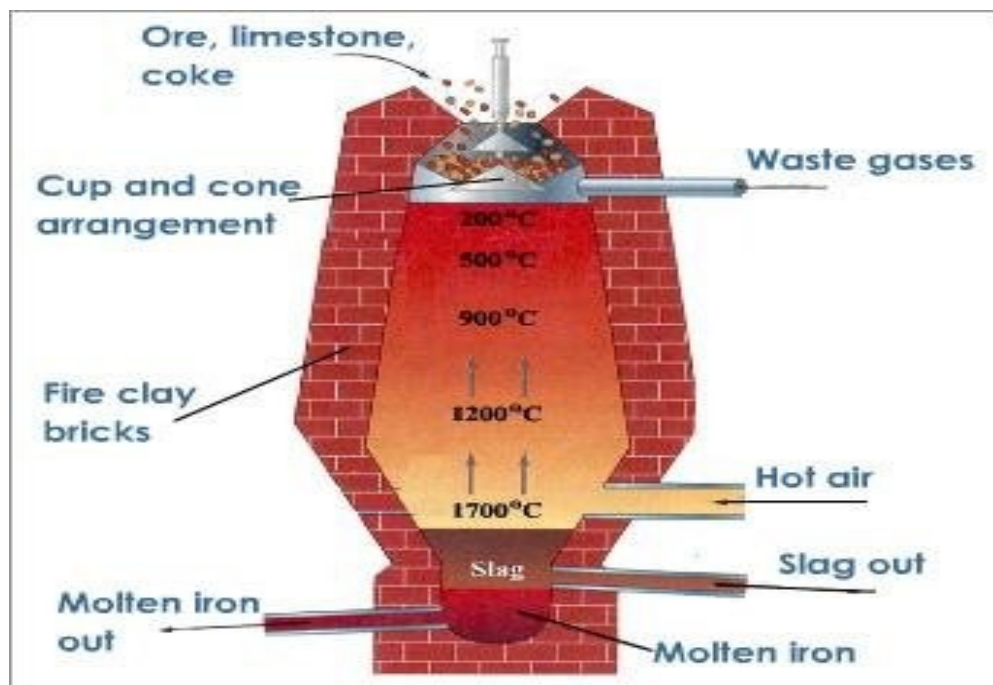
Chemical equation

For hematite: ( $\text{Fe}_2\text{O}_3$ )



OR. For magnetite  $\text{Fe}_3\text{O}_4$





## Reduction of copper(I) sulphide to copper metal

**Ores from which copper can be extracted.**

Copper pyrites ( $\text{CuFeS}_2$ )

Cuprite ( $\text{Cu}_2\text{O}$ ) chemical name is copper(I) oxide

Copper(I) sulphide ( $\text{Cu}_2\text{S}$ )

Malachite ( $\text{CuCO}_3 \cdot \text{Cu(OH)}_2$ )

It is also mined as the free metal in **parts of Canada and the USA**

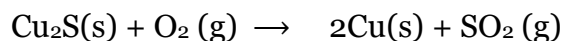
The **copper pyrites** which is the **main ore** is first concentrated by **froth floatation** method to remove **some earthly impurities**.

The ore is **then roasted in air** to form **copper (I) sulphide**



**Reduction process**

The **copper (I) sulphide** is reduced to **copper metal** by heating it in **regulated supply of air**



The **copper obtained in above is impure and is usually called blister copper.**

#### **Uses of copper.**

Used to make copper alloys such as brass used to make screws, nuts because it is stronger and malleable than copper.

Used to make boilers, kettles, electricity wires and cables electric motors and dynamos because it is a very good conductor of electricity.

Used to make coins, ear rings.

Used to make soldering equipment because of its high thermal conductivity.

### The contribution of metal extraction to the Ugandan economy

Extraction of metals has served as a viable route to economic transformation in resource rich countries. Some of the economic gains from mining and extraction include foreign direct investment (FDI), employment creation, infrastructure improvement, and essential services such as adult literacy education and primary healthcare.

While it is clear that mining and extraction has transformed Ugandan economy, it has also impacted negatively on the environment and, to some extent, society. Some of the negative impacts of mining and extraction of metals are loss of vegetation cover, mass destruction of water bodies, loss of biodiversity, land-use changes and food insecurity, increased social vices and conflicts, high cost of living, and air pollution.

The iron and steel industry is valued at over USD one billion and is growing. Much of the potential is still untapped. The industry employs an estimated 10,000

Ugandans directly, and many more indirectly. The iron and steel industry will continue to, and increasingly, anchor Uganda's industrialization.

## **OTHER INDUSTRIAL PROCESSES THAT INVOLVE OXIDATION OR REDUCTION**

### **Haber process (Manufacture of ammonia)**

Haber process is an industrial process for the catalyzed production of ammonia from nitrogen and hydrogen at temperature between 450 – 500 °C and high pressure between 200 – 500 atmospheres.

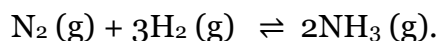
Ammonia is obtained on an industrial scale by the **Haber process** which is based on the direct combination of nitrogen and hydrogen under suitable conditions.

Dry nitrogen and dry hydrogen mixed in the ratio of one to three respectively at a relatively high pressure usually between 200 to 500 atmospheres are passed over a catalyst finely reduced iron impregnated with aluminum oxide at a temperature of 450 – 500 °C.

The iron catalyst is finely divided to increase surface area over which the reaction occurs since the reaction occurs at the surface of the catalyst.

The aluminum oxide improves the performance of the catalyst by making it more porous thus providing a higher surface area for the reaction.

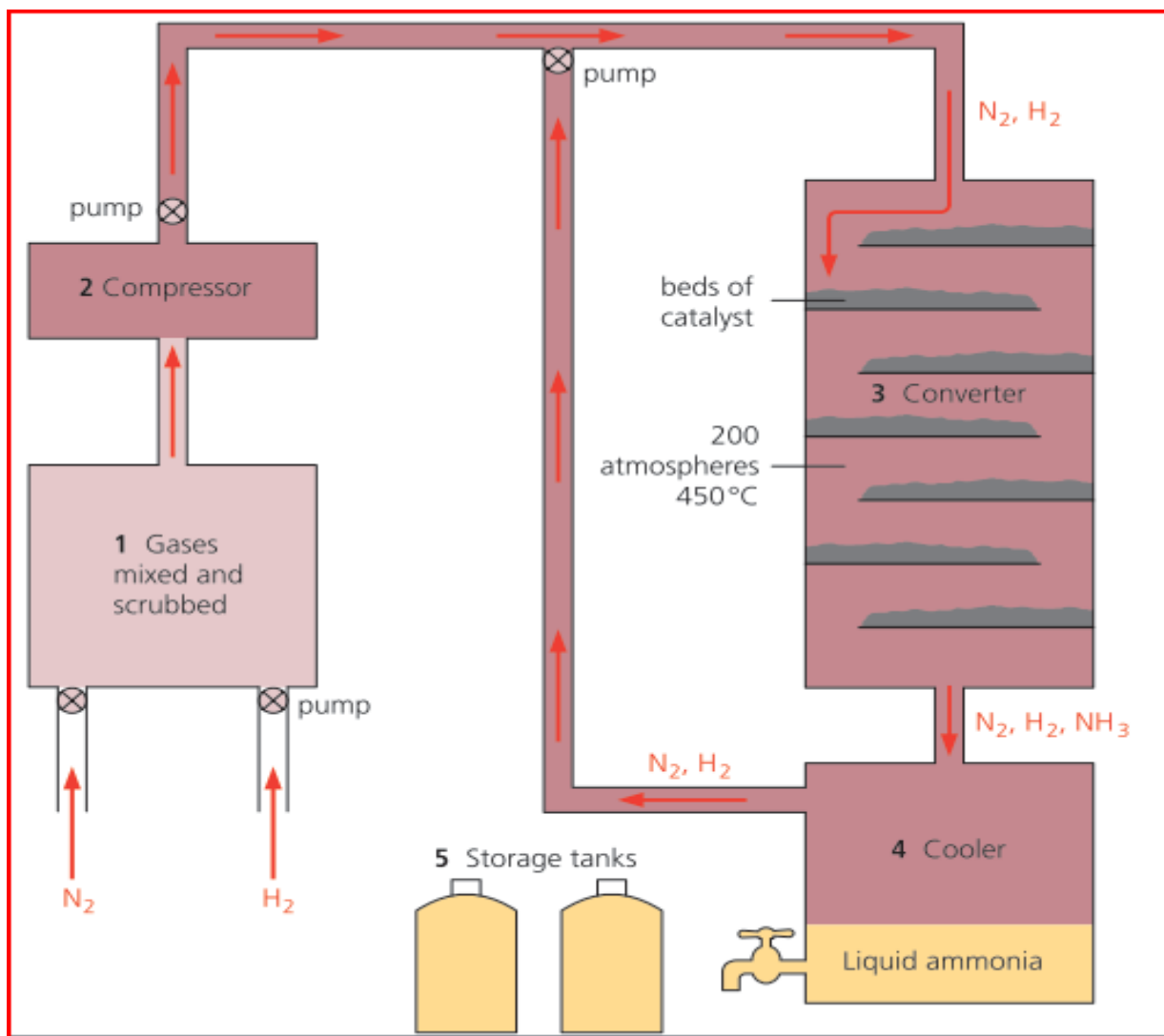
Ammonia is produced according to the chemical equation below and the reaction is reversible.



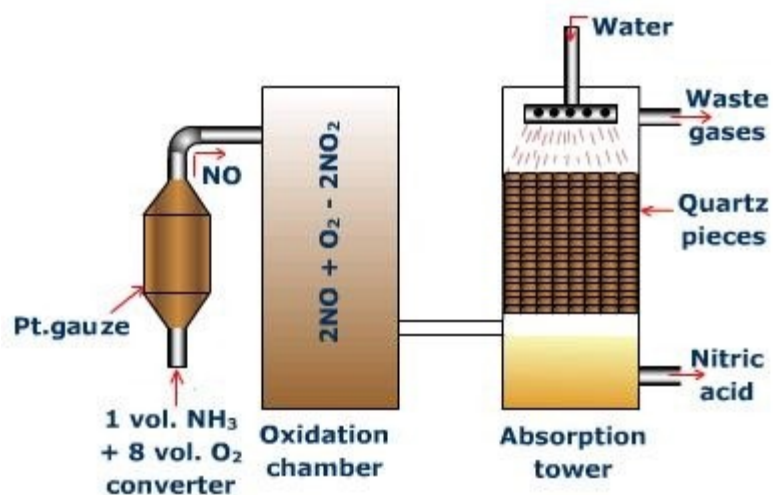
About 10% of the gases react to form ammonia. The formed ammonia is either dissolved in water or liquefied.

The unreacted gases (Hydrogen and nitrogen) are again passed over the catalyst until they react. The reaction between nitrogen and hydrogen is reversible and exothermic.

Ammonia is removed from the mixture of gases by cooling the mixture with freezing mixture. It is only ammonia that liquefies and can be removed from the mixture of unreacted nitrogen and hydrogen.



## Industrial manufacture of nitric acid (the Ostwald's method)



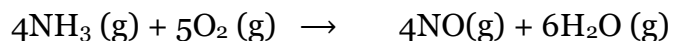


Nitric acid is industrially manufactured by catalytic oxidation of dry ammonia.

Ammonia obtained by direct synthesis using nitrogen and hydrogen in the Haber process is dried and mixed with excess dry air (oxygen) and passed over catalyst platinum -rhodium gauze at about 700 °C and 9 atmospheres pressure.

Dry ammonia is oxidized to nitrogen monoxide and water.

Chemical equation

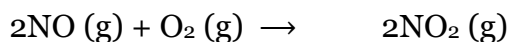


The pressure of 9 atmosphere increases the rate of reaction and the yield of the acid.

The platinum-rhodium catalyst is heated to a temperature of 700 °C (red heat) to start the reaction. The reaction is exothermic and once it starts, maintains the temperature of the gauze.

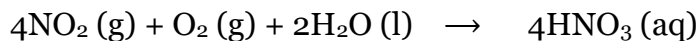
Nitrogen monoxide formed is rapidly cooled to a temperature of about 30 °C and allowed to react with oxygen from excess air to form nitrogen dioxide.

Chemical equation



The nitrogen dioxide in the presence of more oxygen from air is then absorbed in hot water. At this stage nitrogen dioxide and excess air are still under high pressure. Nitric acid is formed as the product.

Chemical equation

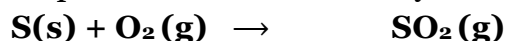




- Purification of sulphur dioxide to remove impurities like Arsenic compounds that may poison the catalyst.
- Conversion of sulphur dioxide into sulphur trioxide
- Dissolution of sulphur trioxide in concentrated sulphuric acid forming oleum.
- Dilution of oleum (done carefully) with correct amount of water forming ordinary concentrated sulphuric acid.

### Description of contact process

Sulphur is burnt in excess dry air (oxygen) to produce sulphur dioxide

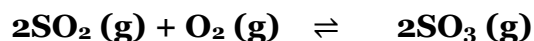


Alternatively, iron pyrites is burnt in air to produce sulphur dioxide



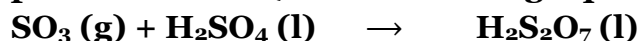
Sulphur dioxide may **contain some impurities such as arsenic compounds** which may poison the catalyst, this makes the **catalyst ineffective**, **there sulphur dioxide is cleaned to remove impurities and then dried**.

The **dried sulphur dioxide is mixed with dry air (oxygen)** and passed along heated pipes containing pellets of Vanadium pentoxide ( $\text{V}_2\text{O}_5$ ) at a temperature of **450 – 500 °C** under a **pressure of 2 – 3 atmospheres** and sulphur trioxide is formed.

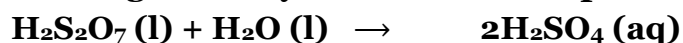


Note: The close contact of the reacting gases in presence of vanadium(V) oxide as catalyst gives the **name contact process**.

Sulphur trioxide formed is dissolved in **concentrated sulphuric acid to produce oleum (which is fuming liquid)**.



The **oleum formed is carefully diluted with correct amount of water forming ordinary concentrated sulphuric acid**



### Uses of sulphuric acid

Used in car batteries and lead accumulators as an electrolyte

Used in manufacture of fertilizers such as ammonium sulphate

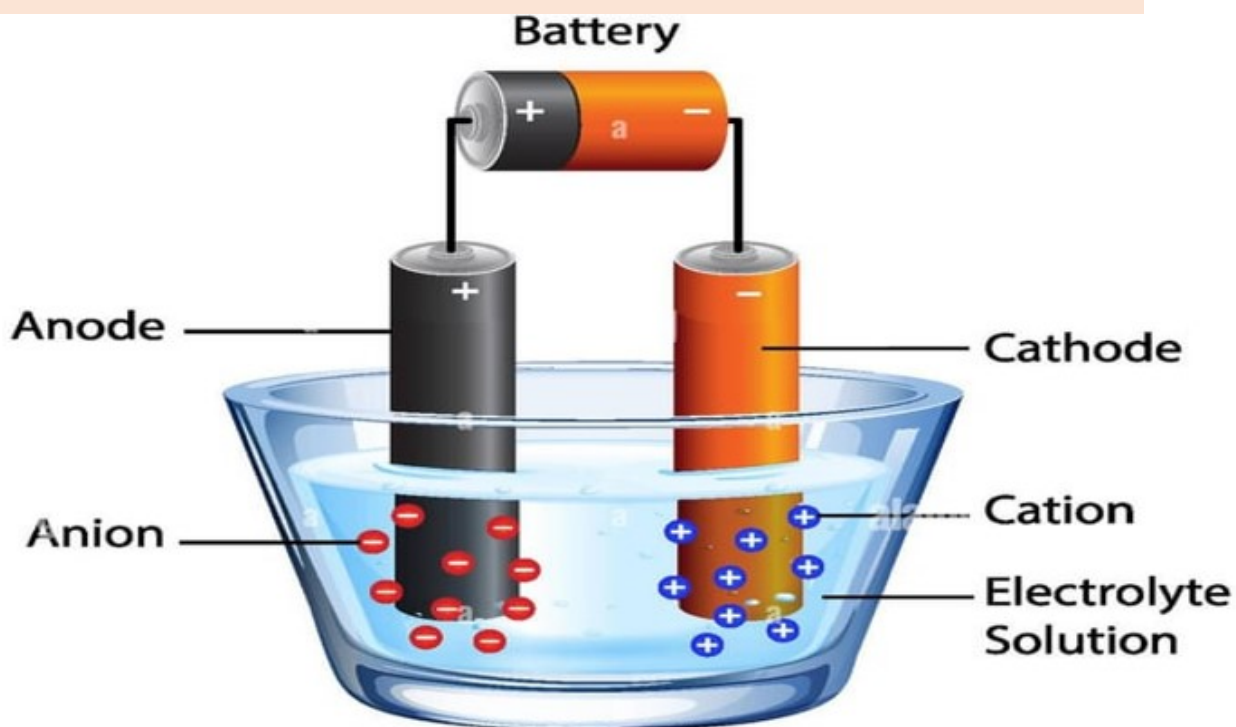
Used in extraction of metals and cleaning of metals prior to painting

Used in manufacture of paints

Manufacture of soap and detergents.

'Pickling' of metal: Is the cleaning of a metal surface before it is plated.

## ELECTROLYSIS



Electrolysis is the decomposition of an electrolyte by passing an electric current through it.

Alternatively, electrolysis is the decomposition of a compound in **molten or aqueous form** by passing electric current through it.

**Important terms to note**

**Electrolyte** is a compound in molten or aqueous form that conducts electric current and decomposed by it.

### Types of electrolytes

**Strong electrolyte** is a compound in molten or aqueous (solution) form that ionizes completely. **Alternatively, electrolysis** is a compound which is completely ionized in dilute solution and in the molten state.

#### Examples of strong electrolytes.

Mineral acids for examples dilute sulphuric acid, nitric acid and hydrochloric acid.  
Salts in aqueous or molten form for examples aqueous sodium chloride, molten lead(II) chloride

Molten or aqueous sodium hydroxide and potassium hydroxide (strong alkalis)

Molten metal oxides for examples aluminium oxide, magnesium oxide.

#### Equations for complete ionization

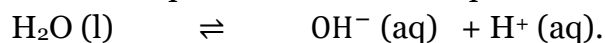
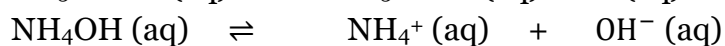


**Weak electrolyte** is a compound in molten or aqueous form that ionizes **only slightly** and most of it exists in form of unionized molecules.

#### Examples of weak electrolytes.

Organic acids for examples methanoic acid, ethanoic acid, propanoic acid, citric acid. Aqueous ammonia, inorganic weak acids for examples sulphurous acid, carbonic acids, nitrous acid. Water. Electricity does not pass **readily** through water, but a small current flows if a **large voltage** is used thus water is a **very weak electrolyte**

#### Equations showing slight ionization of weak electrolytes.



**Non electrolyte (Not a type of electrolyte)** is a solution or molten compound which **cannot** be decomposed by an electric current.

**Non electrolytes** exist only in the form of molecules and are **incapable of ionization**.

#### Examples of non-electrolytes.

Glucose, sucrose(Cane sugar), trichloromethane, tetrachloromethane, alcohols (methanol, ethanol, propan-1-ol etc), urea, molten candle wax.

**Electrode** is a positive or negative pole at which electrons enter or leave an electrolyte.

**An electrode** is usually a metal or graphite which is an allotrope of carbon.

**Anode** is a positive electrode at which the electrons enter the external circuit.

*Alternatively:*

Is a positive electrode at which the electrons leave the electrolyte.

Is an electrode where oxidation takes place.

Is an electrode surrounded by anions.

**Cathode** is a negative electrode at which the electrons enter the internal circuit.

*Alternatively:*

Is a positive electrode at which the electrons enter the electrolyte.

Is **an electrode** where reduction takes place.

Is an electrode surrounded by cations.

**Common examples of electrodes are:** Graphite electrode (carbon electrode), copper electrode, platinum electrode and mercury electrode.

**Ions** are electrically charged particles.

**Cations** are positively charged ions. Cations move to the cathode during electrolysis

**Anions** are negatively charged ions. Anions move the anode during electrolysis.

### **Quick check**

**Explain why solid sodium chloride does not conduct electricity while molten sodium chloride conducts.**

Sodium ions and chloride ions in solid sodium chloride are strongly held together by electrostatic force of attraction. The ions are not free to move (not mobile), thus incapable of carrying electric charge. Sodium ions and chloride ions in molten

sodium chloride are free and able to move because the electrostatic force is broken by heat energy thus the ions are free to carry electric charges

*Note molten compound contains no water thus the constituent ions are those from the compound.*

**Explain why solid sodium chloride does not conduct electricity while sodium chloride solution conducts.**

Sodium ions and chloride ions in solid sodium chloride are strongly held together by electrostatic force of attraction. The ions are not free to move (not mobile), thus incapable of carrying electric charge. Sodium ions and chloride ions in sodium chloride solution are free and able to move because the electrostatic force of attraction is broken off by water molecules thus setting the ions free and able to move thus conducting electricity.

*Note; in solutions of ionic compounds, water is usually the solvent.*





### Differences between an electrolyte and a conductor.

Electrolyte	Conductor
Mobile ions conduct electricity	Mobile electrons conduct electricity.
Conducts electricity when in aqueous solution or molten state only	Conducts electricity when in solid form except mercury.
It is compound.	An element usually a metal or graphite.
It is decomposed by electricity.	Remains chemically unchanged when electric current flows through it.

#### Reminder

Graphite is an allotrope of carbon and carbon is a non-metal. Conductors are solid substances and mercury which is a liquid which allow electricity to pass through but remain chemically unchanged. Conductors allow electricity to flow through them because they have free mobile electrons.

#### IONIC THEORY OF ELECTROLYSIS.

**Ionic theory** states that electrolytes consist of ions which are positively and negatively charged atoms or radicals. **An anion** is a negatively charged ion that moves to the anode during electrolysis. **A cation** is a positively charged ion that moves to the cathode during electrolysis.

Compound	Formula	Cations	Anions
Sodium chloride	NaCl	Na <sup>+</sup>	Cl <sup>-</sup>
Sodium hydroxide	NaOH	Na <sup>+</sup>	OH <sup>-</sup>
Copper(II) sulphate	CuSO <sub>4</sub>	Cu <sup>2+</sup>	SO <sub>4</sub> <sup>2-</sup>

Sulphuric acid	$\text{H}_2\text{SO}_4$	$2\text{H}^+$	$\text{SO}_4^{2-}$
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Properties of an ion differ from those of its element or radical. Ions are inert and stable because they have electronic configuration of noble gases in which the **outermost energy level is fully filled** with electrons. Metals and hydrogen form positive ions(cations) that move to the cathode during electrolysis. Non-metals and acid radicals form negative ions(anions) that move to the anode during electrolysis. The charge of an ion is the same as the valency of its atom or radical

Ions	Valency
$\text{Na}^+$	<b>1</b>
$\text{Cu}^{2+}$	<b>2</b>
$\text{Fe}^{3+}$	<b>3</b>
$\text{SO}_4^{2-}$	<b>2</b>

A solution or molten compound is neutral and the total charge of the positive ions **equals** that of the negative ions. When electrolytes are dissolved in water, they split into positive and negative ions. This is **called ionization**.

### SELECTIVE DISCHARGE OF IONS

When two or more ions of similar charge are present under similar conditions in a solution for example hydrogen ion ( $\text{H}^+$ ) and sodium ion ( $\text{Na}^+$ ) or  $\text{OH}^-$  and  $\text{SO}_4^{2-}$  one is preferentially selected for discharge at the electrode and the selection depends on the following factors.

#### Factors that affect selective discharge of the ion at the electrodes.

- Position of the ions in the electrochemical series.
- Concentration of the ions in the solution.
- Nature of electrode used during electrolysis.

#### (a) **Position of the ions in the electrochemical series**

The **electrochemical series for cations and anions**.

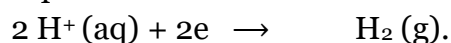
Cation	Anions
--------	--------

K <sup>+</sup> (most reactive)	SO <sub>4</sub> <sup>2-</sup> (most reactive)
Ca <sup>2+</sup>	NO <sub>3</sub> <sup>-</sup>
Na <sup>+</sup>	Cl <sup>-</sup>
Mg <sup>2+</sup>	Br <sup>-</sup>
Zn <sup>2+</sup>	I <sup>-</sup>
Fe <sup>2+</sup>	OH <sup>-</sup>
Pb <sup>2+</sup>	
H <sup>+</sup>	
Cu <sup>2+</sup>	
Ag <sup>+</sup>	

*Provided other factors are kept constant, an ion which is lower in the electrochemical series will be selected and discharged from solution at the electrode in preference to one above it in the electrochemical series.*

For example if sodium ions and hydrogen ions are at the cathode, the hydrogen ion will be selected and discharged in preference to the sodium ion.

Equation

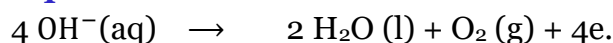


### Observation.

**Effervescence of a colourless gas.**

If the sulphate ions and hydroxide ions are present at the anode, **the hydroxide ions will be discharged in preference to sulphate ions.**

### Equation



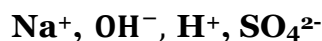
### Observation.

**Effervescence of a colourless gas.**

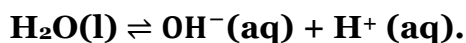
### Quick check

Dilute sodium sulphate solution was electrolyzed using graphite anode and cathode.

(a) Identify the ions present in the solution



*The hydroxide ions and hydrogen ions are coming from partial ionization of water which is a weak electrolyte*



(b) Identify the ions that migrate to the cathode and anode respectively

Anode	Cathode
$\text{OH}^-$	$\text{Na}^+$
$\text{SO}_4^{2-}$	$\text{H}^+$

(c) Identify the ion that will be discharged at the cathode and anode respectively.

**$\text{H}^+$  ions will be discharged at the cathode.**

**$\text{OH}^-$  ions will be discharged at the anode**

(d) State what will be observed at the cathode and anode respectively

**At the anode: Effervescence of a colourless gas.**

**At the cathode : Effervescence of a colourless gas.**

(e) Write half equation for reaction that take place at the anode and cathode respectively.

**Cathode equation**  $2 \text{H}^+ (\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2 (\text{g})$  **(reduction)**

**Anode equation**  $4 \text{OH}^- (\text{aq}) \rightarrow 2 \text{H}_2\text{O} (\text{l}) + \text{O}_2 (\text{g}) + 4\text{e}^-$   
**(oxidation)**

(f). Write the overall equation for reactions that took place at the electrodes.

**$4 \text{H}^+ (\text{aq}) + 4\text{OH}^- (\text{aq}) \rightarrow 2 \text{H}_2 (\text{g}) + 2 \text{H}_2\text{O} (\text{l}) + \text{O}_2 (\text{g})$**

(g) The resultant solution after electrolysis was tested using litmus paper, state what was observed at the respective electrodes.

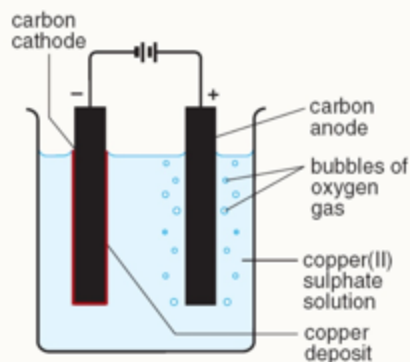
**The solution has no effect on litmus paper because sodium sulphate is formed which is neutral.**

## Quick check

## Electrolysis of Copper(II) Sulphate Solution using Carbon Electrodes

- Ions present:

	cations	anions
From copper(II) sulphate	$\text{Cu}^{2+}(\text{aq})$	$\text{SO}_4^{2-}(\text{aq})$
From water	$\text{H}^+(\text{aq})$	$\text{OH}^-(\text{aq})$



Electric current was passed through copper(II) sulphate solution for one hour using graphite electrodes

(a) Identify the ions present in the solution.

**$\text{Cu}^{2+}$ ,  $\text{SO}_4^{2-}$ ,  $\text{H}^+$  and  $\text{OH}^-$**

**Reason:** Because sulphuric acid is formed.

**Quick check**

Dilute sulphuric acid was electrolyzed using graphite electrodes.

(a) Identify the ions present in the solution.

**H<sup>+</sup>, SO<sub>4</sub><sup>2-</sup>, and OH<sup>-</sup>**

(b). Identify the ions that will be discharged at the anode and cathode respectively.

Anode	Cathode
OH <sup>-</sup>	H <sup>+</sup>

(c). State the observation at the cathode and anode respectively.

**At the anode: Effervescence of a colourless gas**

**At the cathode: Effervescence of a colourless gas**

(d) Write equations for reactions that took place at the cathode and anode respectively.

**At the cathode:  $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$  (Reduction)**

**At anode:  $4\text{OH}^-(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) + 4\text{e}^-$  (Oxidation)**

(d). Write the overall equation for reaction at the electrodes.

**$4\text{H}^+(\text{aq}) + 4\text{OH}^-(\text{aq}) \rightarrow 2\text{H}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$  (Redox)**

(e). State how the alkalinity and acidity vary at the anode and cathode.

**At the cathode the alkalinity increases and the acidity decreases.**

**At the anode the alkalinity decreases and acidity increases.**

(f). Describe how the gaseous products at the cathode and anode are tested in the laboratory.

**The gaseous product at the anode is oxygen. If a glowing splint is lowered into a gas jar containing dry sample of oxygen gas, the glowing splint is relit or relighted.**

*Dinitrogen oxide also relights a glowing splint however dinitrogen oxide has faint sweet smell unlike oxygen which is odourless. The gaseous product at the cathode is hydrogen. If burning splint is introduced into inverted gas jar*

*containing dry sample of hydrogen gas, the burning splint is extinguished as the gas burns with pop sound. The gas at times explode if large amounts are used instead of small ones.*

THE CHAPTER CONTNUES .....

*The book is already on market with sample questions and answers as well as AOI*