



ALPHA CHOICE INNOVATIVE ACADEMY
(International Secondary School)

4/6 Richard Okoroike Close, Praise Hill Estate, Arepo, Ogun State.

GRADE 8
SECOND TERM, 2022/2023 SESSION

CHEMISTRY
MASTERNOTE

NAME: _____

CHEMISTRY SYNOPSIS FOR GRADE 8

SECOND TERM SKILL

WEEK

- 1 Physical and chemical changes
- 2 Metals and non-metals
- 3 Hydrogen
- 4 Oxygen
- 5 Oxidation and reduction
- 6 Endothermic and exothermic reactions
- 7 Chemical Reaction of Metals; Rusting, combustion and respiration as oxidation processes.
- 8 Chemical Reaction of Metals; Factors that increase rate of rusting, Economic importance of rusting
- 9&10. Chemical Reaction of Metals; Ways of preventing rusting.
11. Revision.
12. Examination

WEEK 1

PHYSICAL AND CHEMICAL CHANGES

ESSENTIAL QUESTIONS: By the end of the lesson, students should be able to:

- Define physical change and chemical change.
- Highlight FOUR differences between physical changes and chemical changes.
- Give FOUR examples EACH of physical changes and chemical changes.

PHYSICAL AND CHEMICAL CHANGES

Matter is anything that has mass and occupies space. Matter can exist in three states, namely the gaseous state, the liquid state and the solid state. Matter could be transformed from one state to another. The transformation may involve physical changes in which the matter only changes in form and still retain its composition and structure, or the transformation may involve chemical changes in which the composition and structure changes to give a new substance/matter. Thus we have TWO types of changes in states of matter, namely:

- Physical changes
- Chemical changes.



- Physical changes:** These are changes which are easily reversible and in which no new substance(s) is/are formed. Examples are:
 - All cases of melting, boiling, vaporization, freezing, sublimation and condensation.

- ii. All cases involving separation of mixture of substances e.g. filtration, magnetization, crystallization, distillation etc.
 - iii. All cases involving expansion of metals through heating.
 - iv. All cases of dissolution of salt.
- b. Chemical changes:** These are changes that are not easily reversible and in which new substances are formed. Examples are:
- i. All cases of burning and heating of substances in air.
 - ii. Reactions of reactive metals with water/acids/alkalis.
 - iii. Dissolution and reaction of salts with acids.
 - iv. Corrosion of metals in moist air. E.g. rusting of iron.
 - v. Fermentation and decay of organic substances.
 - vi. Dehydration of substances.
 - vii. All electrolytic and electrochemical processes.

Comparing the characteristics of physical and chemical changes

Physical change	Chemical change
It is easily reversible	It is not easily reversible
No new substances are formed	Entirely new substances are always produced
There is no change in the mass of a substance involved in such a change	There is a change in the mass of a substance that undergoes such a change
It does not involve any great heat changes except the latent heat changes which occur during changes of state, e.g. the heat of dissolution and the latent heats of fusion and vaporization	A considerable amount of heat change is usually involved.

EXERCISE:

- Define physical change
- Are the following physical or chemical change ?, Melting, boiling, vaporization, freezing, sublimation and condensation, give a reason to back up your answers
- Define chemical change
- Identify physical and chemical changes in their environments
- Differentiate between physical and chemical change
- Classify each of the following changes as either a physical or a chemical change
 - The melting of candle wax
 - The dissolution of common salt in water
 - The addition of water to quicklime
 - The hardening of cement by the absorption of carbon (iv) oxide
 - Change in colour of zinc oxide from white to yellow and vice versa during heating and after cooling respectively.
- Highlight (3) other examples of physical changes
- Highlight (5) other examples of chemical changes
- What do you understand by latent heat?
- Mention the three-state matter can exist

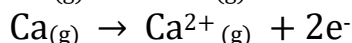
WEEK 2 METALS AND NON-METALS

Objectives:

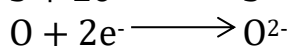
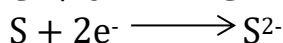
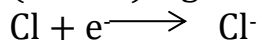
1. Define metals and non-metals
2. Identify metals and non-metals around them
3. Differentiate between metals and non-metals
4. Highlight the examples of metals and non-metals

Elements are classified as metals, metalloids and non-metals.

Metals are elements, which ionize by losing electron(s) to form positive ions (cations) e. g. calcium, magnesium, iron, etc. They are usually solids (except mercury which is a liquid at room temperature) and generally have less than four electrons in the outer shells of their atoms.









Non-metals are elements that ionize/react by gaining electrons to form negative ions (anions) e.g. chlorine, oxygen, nitrogen, iodine, etc.



Characteristics Properties of Metals

Physical properties of metals.

- i. They are sonorous i.e. can produce logical sound when struck.
- ii. They are malleable because can be hammered into shapes.
- iii. They are ductile because they can be drawn into (thin) wire.
- iv. They have high tensile strength.
- v. They are lustrous (i.e. have shiny appearance) and can be polished.
- vi. They are good conductors of heat and current/electricity because they consist of delocalize or mobile electrons in their crystal structure.
- vii. They have high relative density.
- viii. They have high melting points and boiling points because of the strong metallic bond that bind their atoms together.

Metals are silver grey in colour	Non-metals are of different colours	Metals like gold-yellow, copper - reddish brown.
<p>Al</p> 	<p>Si - SILICON</p> 	<p>Au</p> 
<p>Fe</p> 	<p>S</p> 	<p>Cu</p> 

Chemical Properties of Metals

- They ionize by losing electrons during reactions.
- They are reducing agents.
- They form basic oxides.
- They form ionic chlorides
- Reactive metals displace hydrogen of dilute acid to give off hydrogen gas.

Characteristics Properties of Non-metals

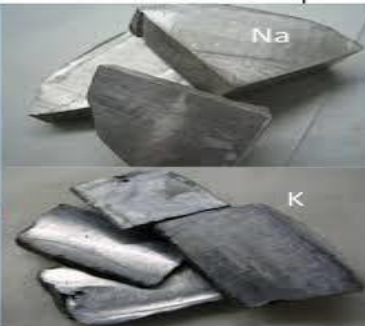
Physical properties of non-metals.

- They are not sonorous.
- They are brittle.
- They are not ductile.
- They are either soft solids or liquids or gases except diamond.
- They have no metallic lustre and cannot be polished.
- They are poor conductors of heat and current/electricity except graphite, a form of carbon.
- They have low densities except iodine.
- They have low melting points and boiling points except carbon.

Metals are generally hard.

Non-metals are brittle not hard.

Metals like lithium , sodium and potassium are soft.
Non-metals like carbon (diamond) is the hardest substance on Earth.



Chemical Properties of Non-metals

- They ionize by gaining electrons during reactions.
- They are oxidising agents.
- They form acidic oxides.
- They form covalent chlorides
- Non-metals cannot displace hydrogen from solutions of dilute acid to give off hydrogen gas.

METALLOIDS.



Silicon



Boron



Antimony



Germanium

Difference between the Physical Properties of Metals and Non-metals

Metals	Non-metals
Metals are malleable, ductile and sonorous.	Non-metals are not malleable, not ductile or not sonorous.
Metals have great tensile strength.	Non-metals are brittle or soft or not hard.
Metals are lustrous.	Non-metals are non-lustrous.
Metals are good conductors of heat and electricity.	Generally they are non-conductors or poor conductors except graphite.
Metals have high melting point and boiling point.	Non-metals except carbon have low melting and boiling point.
Metals have relatively high densities.	Non-metals have low densities.

Difference between the Chemical Properties of Metals and Non-metals

Metals	Non-metals
Metals are reducing agents or they have tendency to ionize by losing electrons to form cations.	Non-metals are oxidizing agents or they tend to ionize by gaining electrons to form anions.
Metals form basic oxides with oxygen.	Non-metals form acidic oxides with oxygen.
Some reactive metals can displace hydrogen from acids.	Non-metals cannot displace hydrogen from acids.
They generally tend to form ionic compounds.	Generally, tend to form covalent compounds.

EXERCISE:

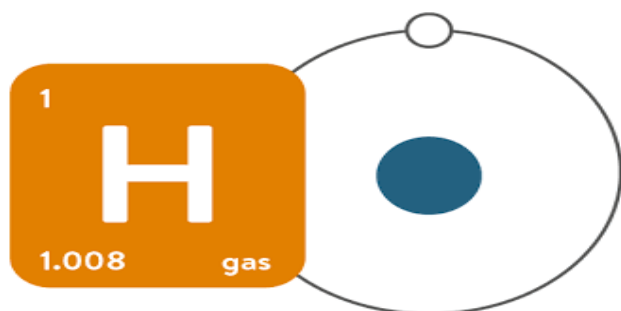
1. Define metals and non-metals
 2. Identify metals and non-metals around them
 3. Differentiate between metals and non-metals
 4. Highlight the examples of metals and non-metals
- Explain the following terms
5. Sonorous
 6. Malleable
 7. Boiling point
 8. Melting point
 9. Ionization
 10. Highlight the chemical properties of metals

WEEKS 3 HYDROGEN

ESSENTIAL QUESTIONS:

- Describe a method of preparing hydrogen in the laboratory.
- State the physical properties of hydrogen
- Highlight the uses of hydrogen

WHAT IS HYDROGEN?



LIGHTEST AND MOST ABUNDANT

Hydrogen is the first element in the periodic table. It is the lightest, most abundant and one of the oldest chemical elements in the universe.

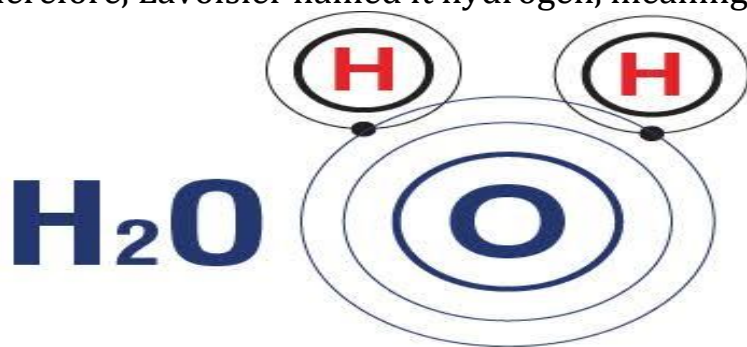
NEVER ALONE

On Earth, hydrogen is found in more complex molecules, such as water or hydrocarbons. To be used in its pure form, it has to be extracted.

FUEL OF STARS

Hydrogen fuels stars through nuclear fusion reaction. This creates energy and all the other chemical elements which are found on Earth.

Cavendish is regarded as the discoverer of hydrogen because he was the first (in 1766) to prepare it in the pure state, describes its properties and recognizes it as an element. He found that the gas was flammable and that it burned in air to produce water. Therefore, Lavoisier named it hydrogen, meaning *water-former*.



Hydrogen can be found in water molecules (H₂O), consisting of 2 hydrogens and 1 oxygen atom

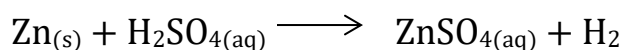
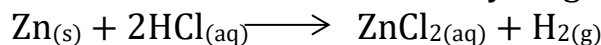
PREPARATION OF HYDROGEN

Laboratory Preparation:

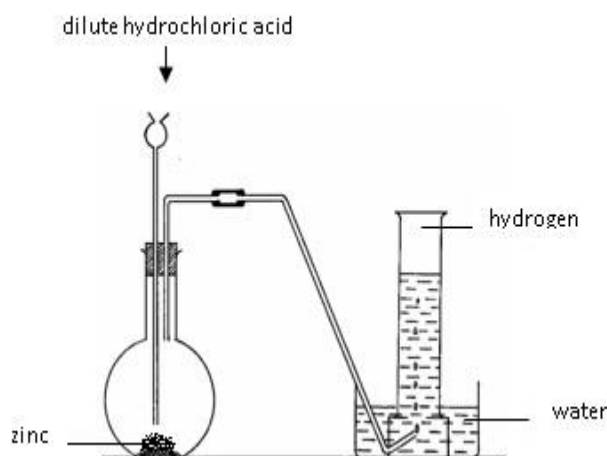
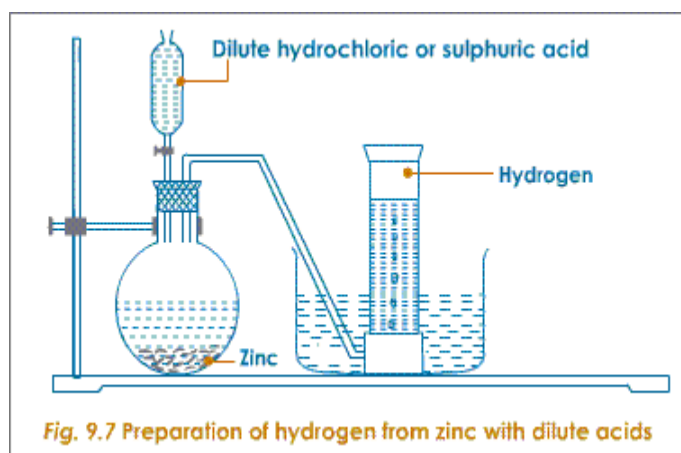
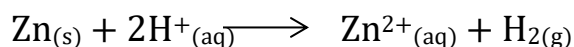
Hydrogen is liberated when certain metals react with dilute mineral acids, water or steam. It is also given off when tin, zinc or aluminium reacts with hot concentrated solutions of sodium or potassium hydroxide.

The three methods commonly used for the preparation of hydrogen in the laboratory are as follows:

Action of Zinc on an acid: dilute hydrochloric or tetraoxosulphate (VI) acid attacks metallic zinc with the liberation of hydrogen gas. No heating is necessary.

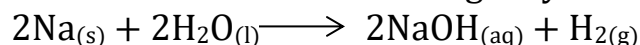


Ionically, both reactions can be represented as follows:



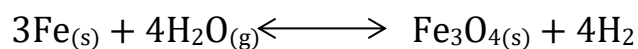
Action of Sodium on cold water:

Sodium liberates hydrogen from cold water. This reaction is very vigorous and should be carried out with extreme care using only a small piece of sodium.



Action of iron on steam

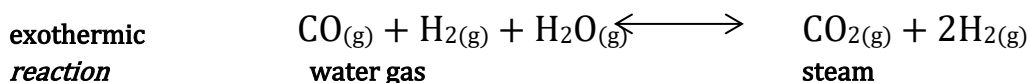
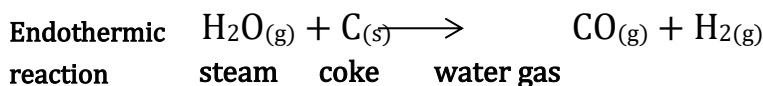
Iron, at red heat liberates hydrogen from steam. Iron (II) diiron (III); oxide, Fe_3O_4 , is formed at the same time. This reaction is reversible.



INDUSTRIAL PREPARATION:

From water gas (Bosch process):

In this process, large quantities of hydrogen are produced from cheap raw materials, namely water and coke. When steam is passed over red-hot coke (carbon) at about 1200°C , a mixture of carbon (II) oxide and hydrogen known as water gas is produced. Excess steam is then mixed with the water gas and passed over a catalyst, iron (III) oxide or chromium (III) oxide, at 450°C . As a result, the carbon (II) oxide in the water gas is converted to carbon (IV) oxide with a further yield of hydrogen.



The carbon (IV) oxide is then removed from the mixture by dissolving it in water (under a pressure of 30 atmospheres) or other solvents such as caustic soda solution. Any unreacted carbon (II) oxide is absorbed in an ammoniacal solution of copper (I) ethanoate.

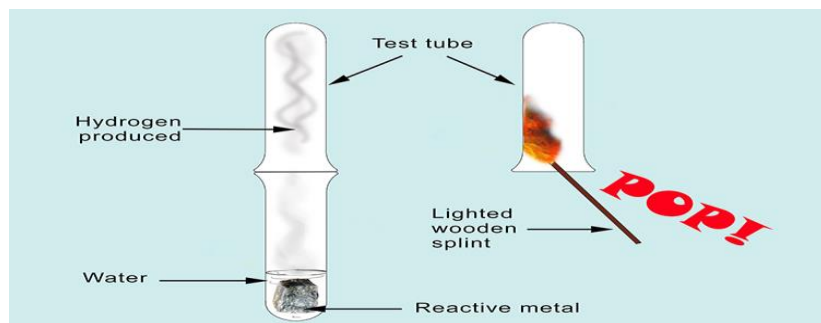
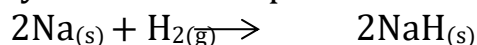
Physical Properties of Hydrogen

1. Hydrogen is a colourless, odourless and tasteless gas
2. It is neutral to moist litmus paper
3. It is relatively insoluble in water. Only 2 cm³ of it will dissolve in 100 cm³ of water at s.t.p.
4. Hydrogen is the lightest known substance. It is 14.4 times less dense than air.
5. It has a very low boiling point of -253°C.

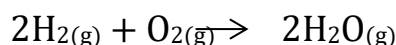
Chemical Properties:

1. Hydrogen combines with certain metals and non-metals to form electrovalent and covalent hydrides respectively.

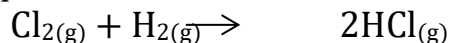
With metals: hydrogen combines directly with several of the more active metals to form ionic hydrides i.e. compounds which contain the hydride ion, H⁻. For example



With oxygen: pure hydrogen burns with a pale blue flame as it combines with oxygen to produce steam.

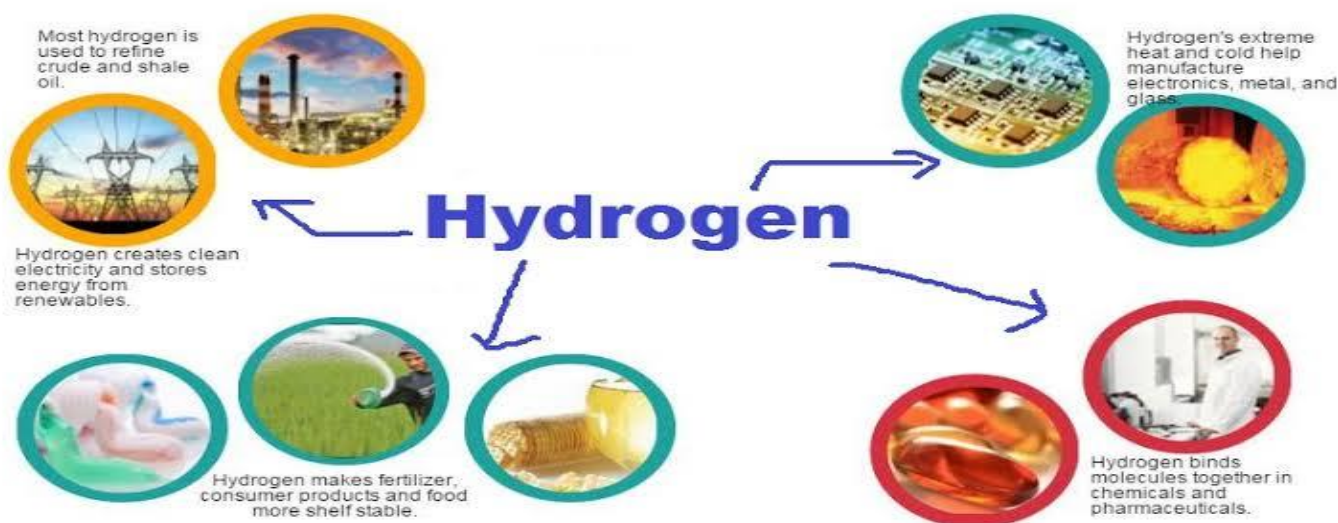


With halogens: hydrogen combines directly with the halogens to produce halides. For example,



Uses of Hydrogen

1. Hydrogen is used in the manufacture of ammonia, hydrochloric acid and methanol.
2. Hydrogen under high pressure is passed through vegetable oils (e.g. palm oil, corn oil, cotton seed oil or soybean oil) in the presence of a nickel catalyst to give solid fats, which are used as margarine or as lard substitutes, and in the manufacture of soap and candles.
3. Since hydrogen has a low density, it is used for filling balloons. Its highly flammable nature, however, limits this use to meteorological studies and other scientific purposes.
4. Hydrogen is passed through a mixture of oil and powdered coal at high temperatures and pressures to yield a mixture of hydrocarbons from which synthetic petrol can be extracted by fractional distillation.
5. Hydrogen is a constituent of many gaseous fuels such as water gas and coal gas. Liquid hydrogen is also used as a rocket fuel.
6. Hydrogen gives out a lot of heat when it burns. Thus it is used in oxy-hydrogen flames to produce high temperatures (over 2 000°C) that can melt metals.



EXERCISE:

1. Describe the discovery of hydrogen gas
2. How can you prepare hydrogen gas in the laboratory?
3. How can you identify hydrogen gas in the laboratory?
4. Highlight (4) chemical properties of hydrogen gas.
5. Explain the industrial production of hydrogen
6. State different uses of hydrogen gas
7. Who discovered hydrogen gas?
8. Highlight (4) physical properties of hydrogen gas.
9. List (4) uses of hydrogen
10. Is hydrogen a monoatomic, diatomic, tri atomic or polyatomic element

WEEK 4

OXYGEN

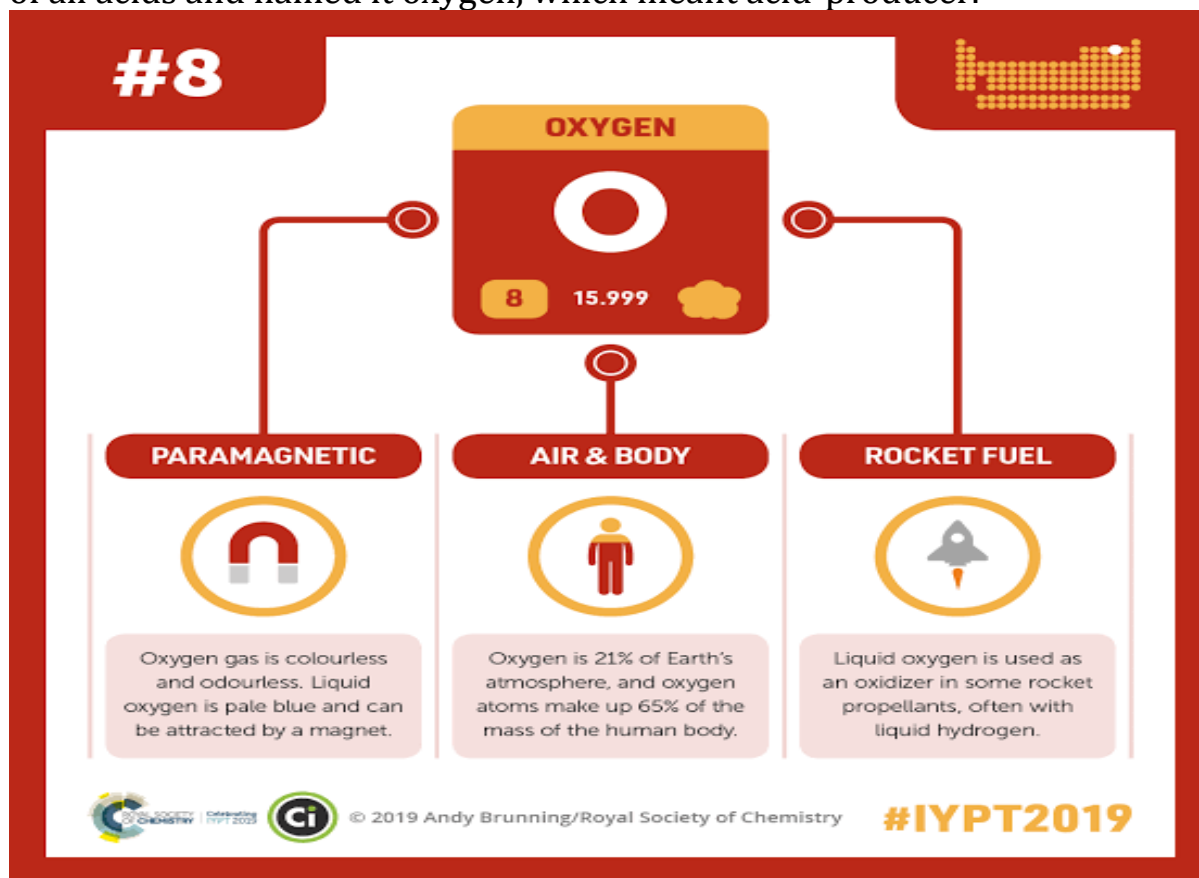
Objectives

1. Preparation of oxygen
2. Properties: Physical and Chemical properties of oxygen

Uses of oxygen.

Water as a product of hydrogen and oxygen- electrolysis of water.

Oxygen was discovered independently by Scheele in 1772 and Priestly in 1774. However, it was Lavoisier who first described the major properties of this newly discovered gas and arrived at the conclusion that the gas was essential for the combustion of substances. He erroneously thought that this gas was a major constituent of all acids and named it oxygen, which meant acid-producer.



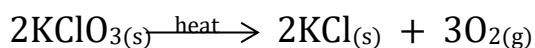
Laboratory preparation

Oxygen is usually prepared in the laboratory by

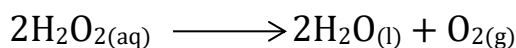
- the decomposition of potassium trioxochlorate (V) and hydrogen peroxide, and
- the oxidation of hydrogen peroxide

Other methods include the thermal decomposition of mercury (II) oxide and lead (IV) oxide, and the reaction of sodium peroxide with water.

From trioxochlorate (V): Potassium trioxochlorate (V) decomposes slowly to release all its oxygen when heated above its melting point (368°C). In the presence of manganese (IV) oxide, which acts as a catalyst, the reaction occurs at a lower temperature and at a much faster rate.



From Hydrogen peroxide: In the presence of the catalyst, manganese (IV) oxide, hydrogen peroxide readily decomposes to liberate oxygen without any heating.



INDUSTRIAL PREPARATION:

In places where a cheap source of electricity is available, oxygen may be prepared industrially by the electrolysis of water. However, in most places, it is prepared by the fractional distillation of liquid air. This preparation involves two main processes,

- The liquefaction of air, and
- The fractional distillation of the resultant liquid air.

Liquefaction of air

Air, in the gaseous form, is first passed through caustic soda to remove carbon dioxide. It is then compressed to a pressure of about 200 atm, cooled and allowed to escape rapidly through a minute aperture. The sudden expansion of the air into a region of lower pressure causes it to cool even further (since heat energy is used up in separating the molecules). The cooled air will cool the incoming stream of air. This process is repeated until the gaseous air becomes a liquid at – 200°C.

Fractional Distillation of liquid air:

The liquid air is then led to a fractionating column. On distillation, nitrogen, which has a lower boiling point of -196°C, is evolved first, leaving behind a liquid very rich in oxygen. Further heating converts the liquid oxygen to a gas at -183°C. The oxygen gas is dried, compressed and stored in steel cylinders under a pressure of about 100 atm.

Physical Properties:

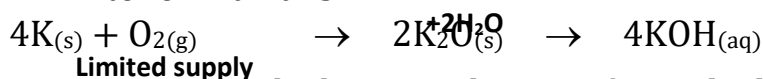
1. Oxygen is a diatomic gas which is colourless, odourless and tasteless.
2. It is neutral to moist litmus paper.
3. It is slightly soluble in water, i.e. only about 2% by volume will dissolve in water at room temperature and standard pressure. This solubility of the gas, however, is of vital importance to aquatic life.
4. Gaseous oxygen is approximately 1.1 times denser than air.
5. Gaseous oxygen liquefies at -183°C and solidifies at -225°C.

Chemical Properties:

1. Direct combination with other elements:

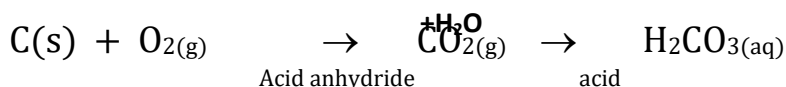
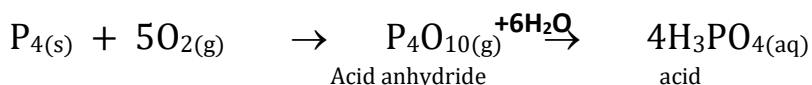
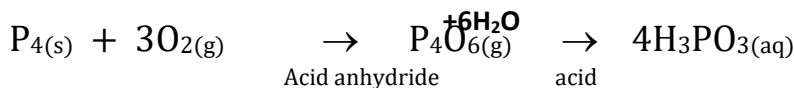
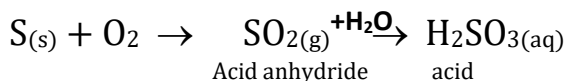
Oxygen combines readily with all other elements (except the rare gases, some halogens and some unreactive metals) to form oxides.

- a. **With metals:** most metals react with oxygen to form basic oxides. The oxides of very electro-positive metals like potassium, sodium and calcium dissolve in water to form alkalis.



However, the basic oxide is not formed when sodium is heated in a plentiful supply of oxygen. Instead, it burns with a bright yellow flame to form sodium peroxide, Na_2O_2 . Similarly, potassium burns in excess oxygen with a lilac flame to form a higher oxide, KO_2 , instead of the basic oxide, K_2O .

- b. **With non-metals:** non-metals like sulphur, carbon and phosphorus burn in oxygen to form acidic oxides. These are also known as acid anhydrides, as they dissolve in water to form acids. Acid anhydrides are oxides of non-metals which dissolve in water to form acid solution.



Reaction with compounds:

- Many compounds react with oxygen. Most hydrocarbons and compounds of carbon, hydrogen and oxygen burn in oxygen to form carbon (IV) oxide and water.
- Sulphides burn in oxygen to give sulphur (IV) oxide and another oxide.
- Ammonia does not form an oxide of nitrogen when it burns in oxygen. Instead, nitrogen and water are formed.

Test for Oxygen:

Oxygen rekindles a glowing splinter. Nitrogen (I) oxide, N_2O also does this but it can be distinguished from oxygen as given in the table below.

Differences between Oxygen and Nitrogen (I) oxide

Property	Nitrogen (I) Oxide	Oxygen
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Smell	Pleasant smell	Odourless
Solubility in water	Fairly soluble	Slightly soluble
Reaction with nitrogen (I) oxide	No reaction	Produces reddish brown fumes of nitrogen (IV) oxide
Reaction with heated copper	Residual gas N_2	No residual gas

Uses of Oxygen

- Oxygen is used as an aid to breathing where the natural supply of oxygen is insufficient, e.g. in high altitude flights, mountaineering and diving operations.
- It is used for resuscitating fainting patient, aiding breathing in some lung diseases and when anaesthetics are administered to patient.
- Oxygen is used in oxy-ethyne and oxy-hydrogen flames for cutting and welding metals.
- It is used in the steel industry for the removal of carbon, sulphur and phosphorus, i.e. in the L-D process for making steel.
- Liquid oxygen and fuels are used as propellants for space rockets.
- Oxygen is used in the manufacture of important chemical compounds like tetraoxo-sulphate (VI) acid, trioxonitrate (V) acid ethanoic acid.

EXERCISE

- Who discovered oxygen gas?
- Highlight (4) physical properties of oxygen gas.
- List (4) uses of oxygen
- Describe the discovery of oxygen gas
- How can you prepare oxygen gas in the laboratory?
- How can you identify oxygen gas in the laboratory?
- Highlight (4) chemical properties of oxygen gas.
- Explain the industrial production of oxygen
- Explain the term electrolysis
- Is oxygen a monoatomic, diatomic, tri atomic or polyatomic element

WEEK 5

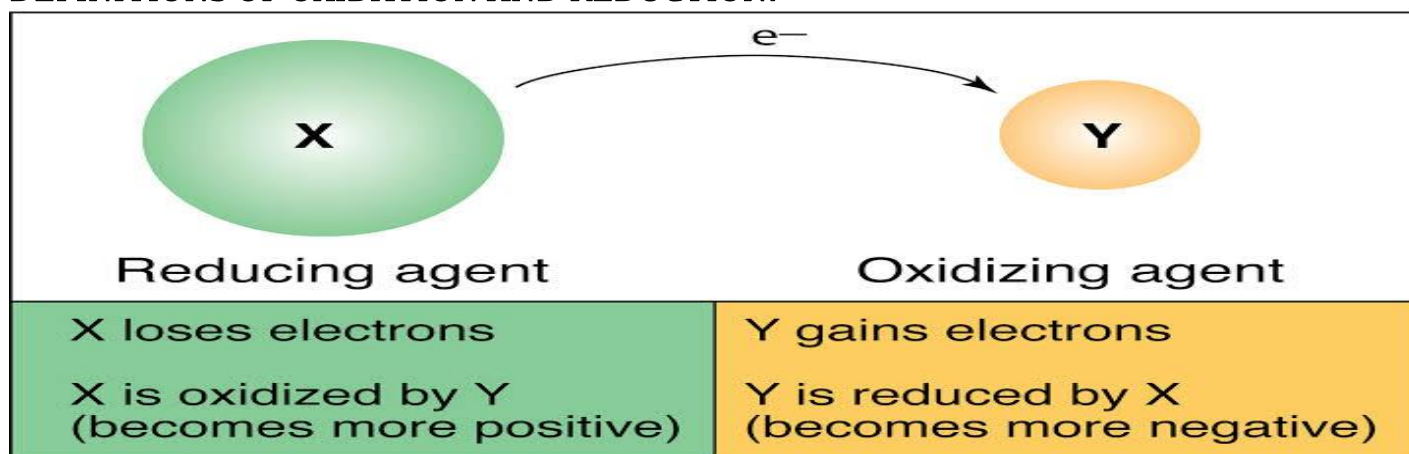
OXIDATION AND REDUCTION

OBJECTIVES: By the end of the lesson, students should be able to:

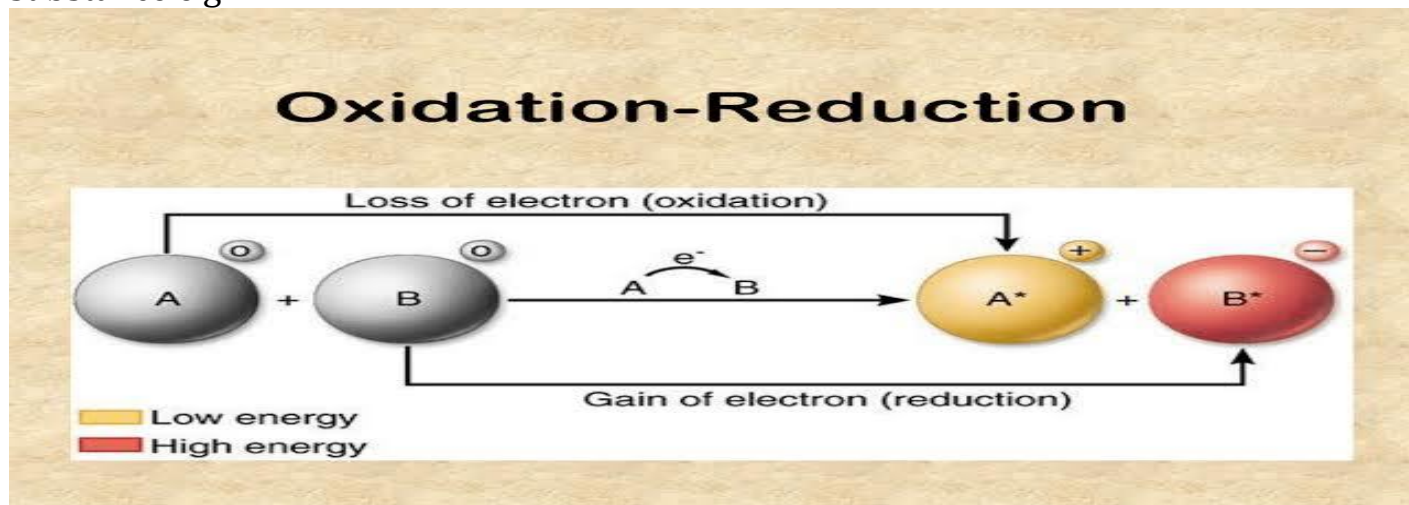
- Define oxidation and reduction.
- Identify substances that are oxidised or reduced in any given redox reaction.
- Identify the oxidising agent and the reducing agent in any given reaction.

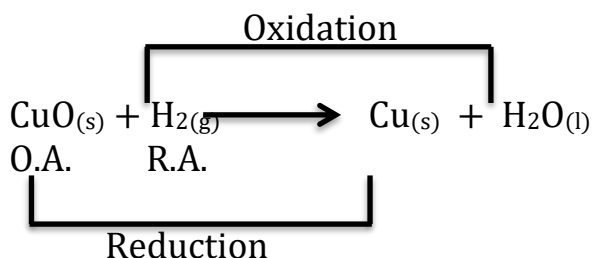
The term oxidation was originally thought of in simple terms as reactions in which oxygen is added to a substance, while reduction was described as the removal of oxygen from an oxygen-containing compound. The meanings of the terms have gradually been broadened and modified.

DEFINITIONS OF OXIDATION AND REDUCTION:

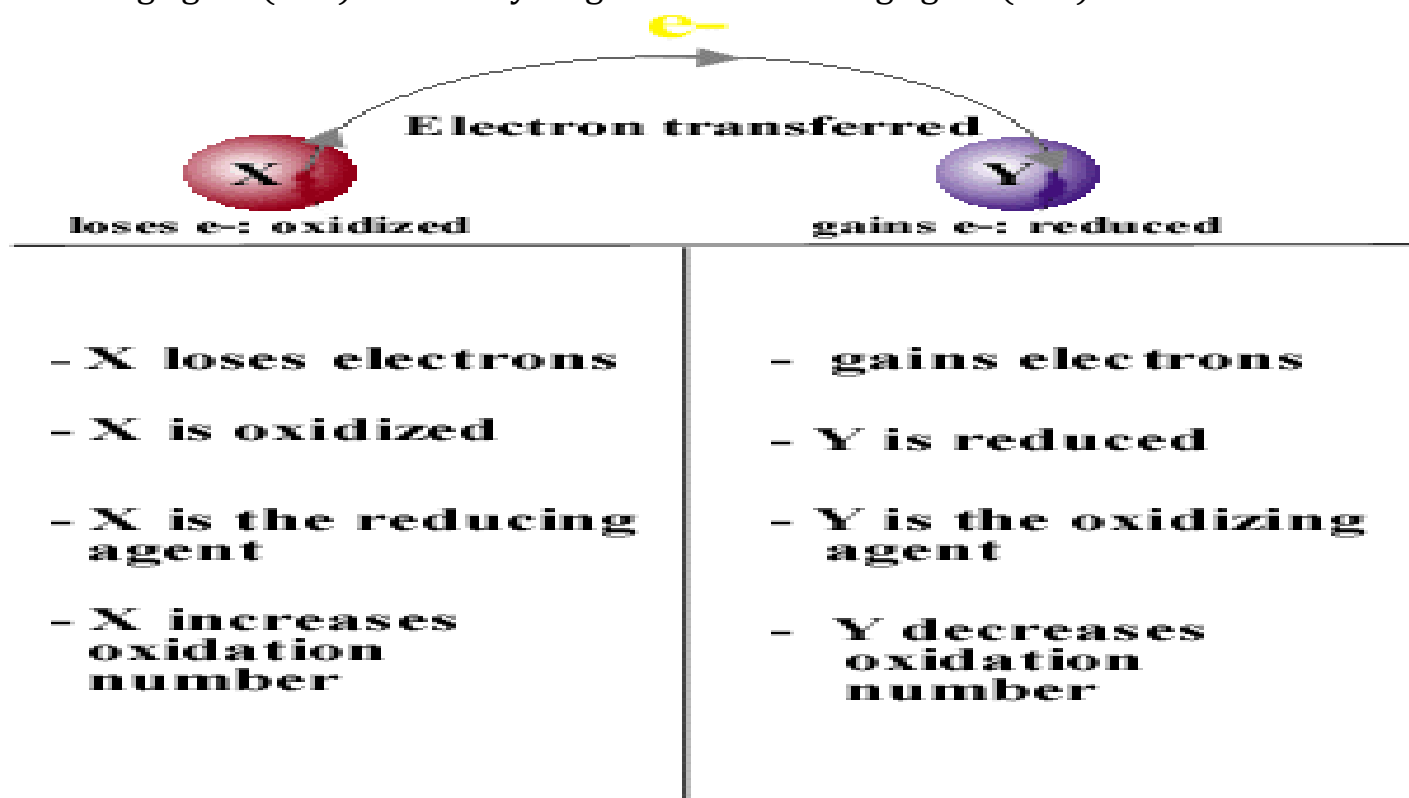


Addition of oxygen: At first, oxidation is simply a reaction in which oxygen combines with another substance, while reduction is that in which oxygen is removed from a substance e.g.

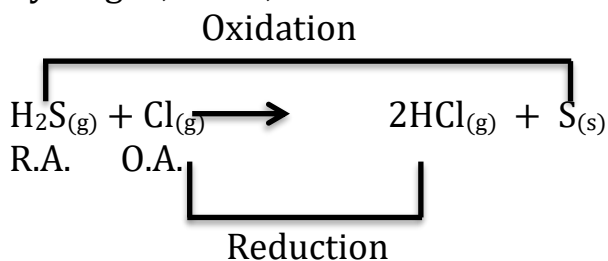




In the reaction above, the hydrogen is oxidized to water (since oxygen has been added to it) while the copper (II) oxide has been reduced to metallic copper (since oxygen has been removed from it). In such a reaction, the copper (II) oxide is described as the oxidizing agent (O.A.) and the hydrogen as the reducing agent (R.A.).



Removal of Hydrogen: The definition of oxidation is extended to mean the removal of hydrogen, while, that of reduction means the addition of hydrogen e.g.

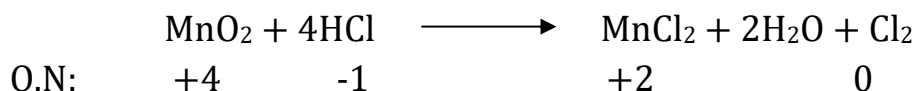


The hydrogen sulphide has been oxidized to yellow sulphur (since hydrogen has been removed from it) and the chlorine has been reduced to hydrogen-chloride (since hydrogen has been added to it). Thus, chlorine is an oxidizing agent, while, hydrogen-sulphide is a reducing agent.

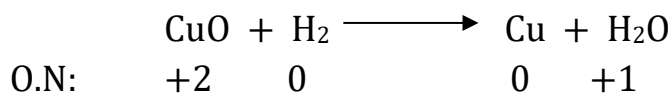
In terms of oxidation number:

Oxidation is the process in which the oxidation number of an atom increases.

Reduction is the process in which the oxidation number of an atom decreases. E.g.



Manganese (IV) oxide was reduced to manganese (II) chloride while hydrogen chloride was oxidized to chlorine gas.



Copper (II) oxide was reduced to copper while hydrogen gas was oxidized to water.

The concept of oxidation number

The oxidation number (O.N.) of an element in a molecule/ion is the electrical charge it appears to have as determined by a set of arbitrary rules. OR

O.N. is the charge an atom appears to have in a neutral molecule or in an ion. OR

O.N. is the charge an atom would have if the pair of electrons in each covalent bond were assigned to the more electronegative element.

O.N. is also referred to as **oxidation state**.

Rules for determining oxidation numbers

1. The oxidation number of each atom in an uncombined element (whether monoatomic or polyatomic) is zero. O.N. of He, Ne, Ar, Na, Mg, C, P, H₂, O₂, O₃, S₈ etc. is zero.
2. Oxidation number of one oxygen atom is -2 in its compounds **except** in peroxides (H₂O₂, Na₂O₂ etc) where it is -1, and in its binary compounds with fluorine, OF₂ where it has O.N. of +2.
3. Oxidation number of one hydrogen atom is +1 **except** in its binary compounds with metals (e.g. NaH, CaH₂ etc.) where it is -1.
4. Oxidation number in a simple ion containing one atom is equal to the charge on the ion. E.g. the O.N of Na in Na⁺ is +1; that of S in S²⁻ is -2, while that of Al in Al³⁺ is +3.
5. Group I metals such as Na and K have O.N. of +1 while group II elements such as Mg and Ca have O.N. of +2 per atom in their compounds.
6. Halogens (F, Cl, Br and I) have O.N. of -1 per atom in their binary compounds only except when bonded to a more electronegative atom. E.g. in IF, the O.N. of iodine is +1 while that of fluorine is -1.
7. In a neutral compound, the sum of the O.N. of all the atoms must be equal to zero in order to maintain electrical neutrality.

8. In a covalent molecule, the shared pair electrons are assigned to the more electronegative element. E.g. in NH_3 , the O.N. of N is -3 and each H atom is +1.
9. When two or more different atoms form an ion, the sum of their O.N. must be equal to the charge on the ion. E.g.

The charge on OH^- is -1. This is because $(-2) + (+1) = -1$

The charge on SO_4^{2-} is -2. This is because $(+6) + (-2 \times 4) = (+6) + (-8) = -2$

In terms of electron transfer:

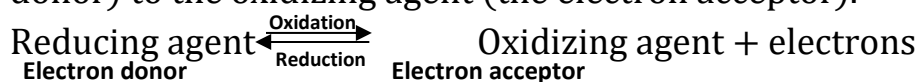
Oxidation is a process involving a loss of electron(s) while reduction is a process involving a gain of electron(s).

REMEMBER: OIL RIG means **Oxidation Is Loss**, **Reduction Is Gain**

A substance loses electron when it is oxidized and gains electron when it is reduced.

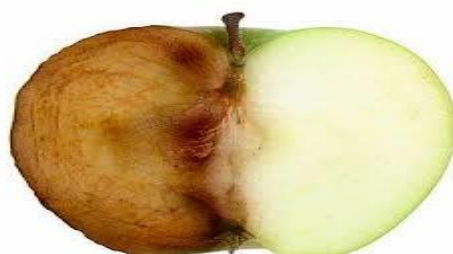
So a redox reaction is basically a transfer of electrons from the reducing agent (the electron

donor) to the oxidizing agent (the electron acceptor).



Antioxidants in action

Oxidation occurs when the apple is left exposed to air



without with

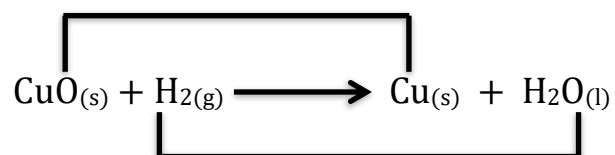
The apple is protected when dipped in orange juice containing the antioxidant vitamin C



EXERCISE:

1. Define oxidation and reduction
2. Enumerate any oxidation processes in your surroundings
3. Give examples of oxidation and reduction
4. Identify the oxidising agent and the reducing agent in any given reaction
5. Using appropriate examples, explain

- (i) oxidising agent
(ii) reducing agent
6. Give the full meaning of OIL RIG
 7. Explain oxidation and reduction in terms of electrons
 8. What is the use Vitamin C in oxidation reactions
 9. Explain oxidation and reduction in terms of oxidation number
 10. Identify oxidation and reduction process in the reaction below, and discover the oxidizing and reducing agent

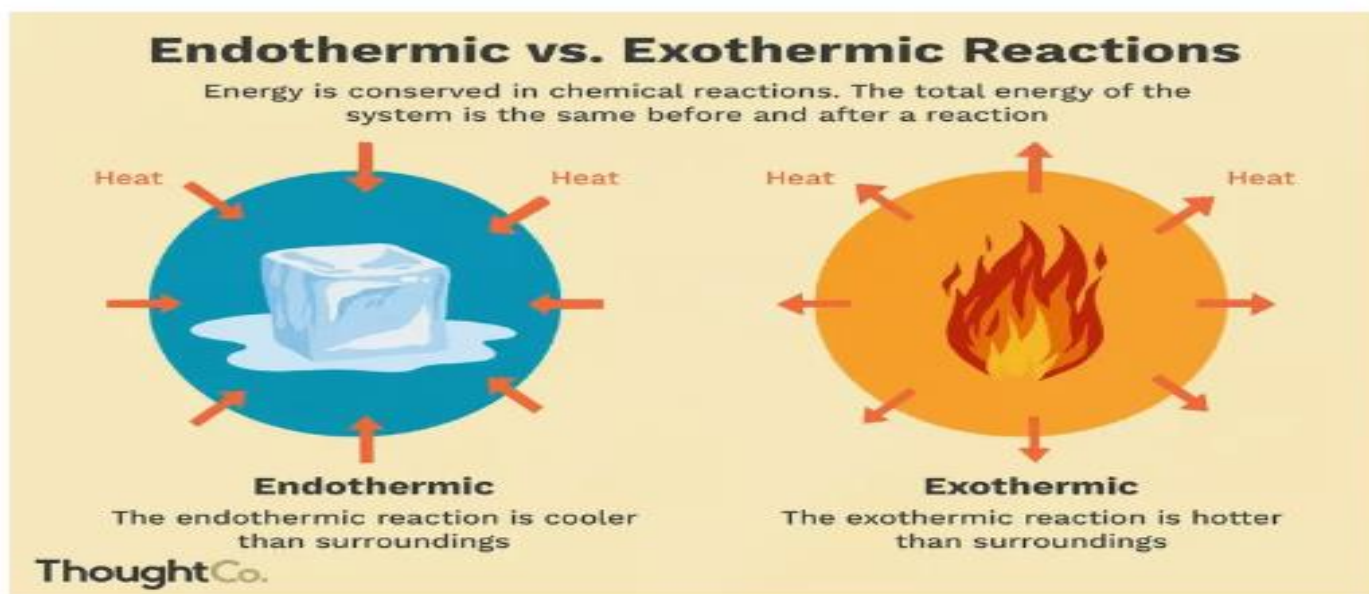


WEEK 6: ENDOTHERMIC AND EXOTHERMIC REACTION

OBJECTIVES:

Students will be able to:

1. Describe the endothermic and exothermic reactions
2. Recognize exothermic processes in the surroundings
3. Give examples of the endothermic and exothermic reactions



During a chemical reaction, heat can be evolved or absorbed depending on the relative heat content of the reactants and the products. When the total heat content of the products is less than that of the reactants, the excess energy will be given off as heat to the surrounding as the reaction occurs. This type of reaction is said to be **exothermic**.

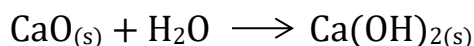
When the total heat content of the product is more than that of the reactants, heat energy must be absorbed from the surroundings for the reaction to proceed. Such a reaction is said to be **endothermic**.

An **exothermic** reaction is one during which heat is liberated to the surroundings.

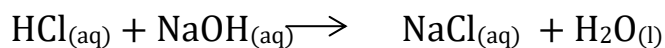
An **endothermic** reaction is one during which heat is absorbed from the surroundings.

Some examples of exothermic reactions are as follows:

1. Reaction between calcium oxide and water

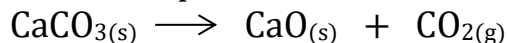


2. Reaction between an acid and a base

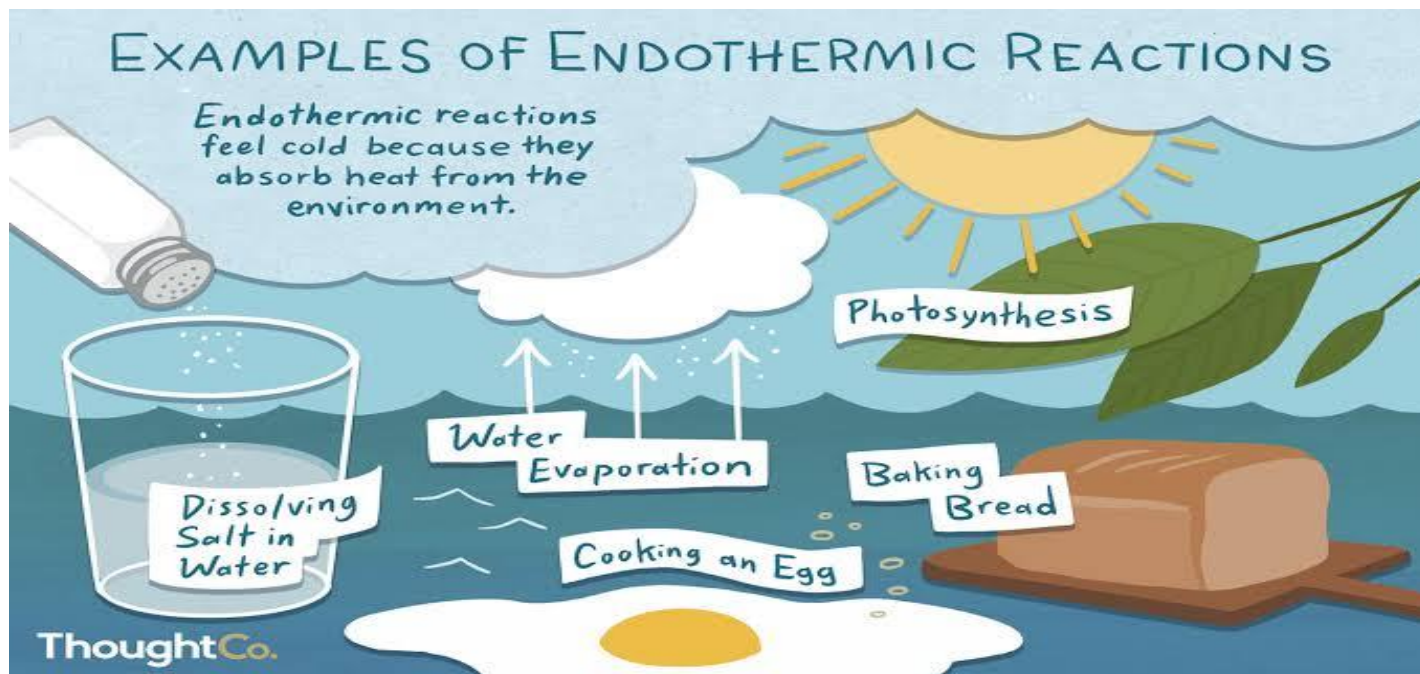


Some examples of endothermic reactions are as follows:

1. Thermal decomposition of calcium trioxocarbonate (IV)



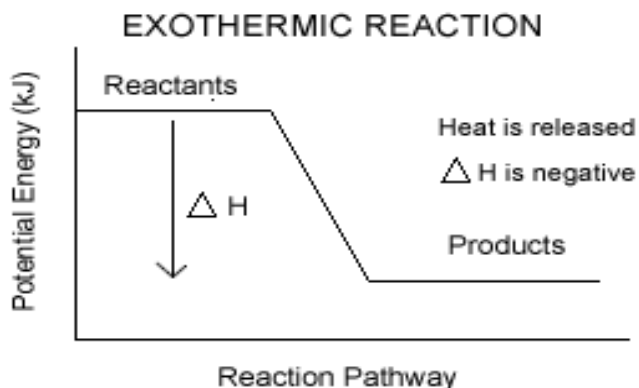
2. Thermal dissociation of ammonium chloride



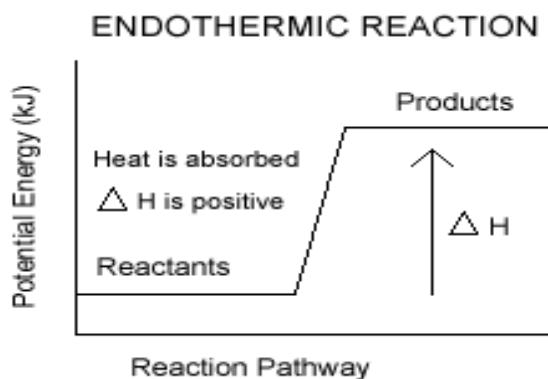
When trying to classify a process as exothermic or endothermic, watch how the temperature of the surroundings changes. An exothermic process releases heat, and causes the temperature of the immediate surroundings to rise. An endothermic process absorbs heat and cools the surroundings.

Exothermic processes	Endothermic processes
making ice cubes	melting ice cubes
formation of snow in clouds	conversion of frost to water vapor
condensation of rain from water vapor	evaporation of water

a candle flame	forming a cation from an atom in the gas phase
mixing sodium sulfite and bleach	baking bread
rusting iron	cooking an egg
burning sugar	producing sugar by photosynthesis
forming ion pairs	separating ion pairs
Combining atoms to make a molecule in the gas phase	splitting a gas molecule apart
mixing water and strong acids	mixing water and ammonium nitrate
mixing water with an anhydrous salt	making an anhydrous salt from a hydrate
crystallizing liquid salts (as in sodium acetate in chemical hand warmers)	melting solid salts
nuclear fission	reaction of barium hydroxide octahydrate crystals with dry ammonium chloride
mixing water with calcium chloride	reaction of thionyl chloride (SOCl_2) with cobalt(II) sulfate heptahydrate



Endothermic



CONCLUSION:

During a chemical reaction, heat can be evolved or absorbed depending on the relative heat content of the reactants and the products. When the total heat content of the products is less than that of the reactants, the excess energy will be given off as heat to the surrounding as the reaction occurs. This type of reaction is said to be **exothermic**.

When the total heat content of the product is more than that of the reactants, heat energy must be absorbed from the surroundings for the reaction to proceed. Such a reaction is said to be **endothermic**.

EXERCISE:

1. What do you understand by exothermic, Illustrate with equations.
2. Explain endothermic reactions? Illustrate with equations.
3. Draw a labelled diagram of energy change during:
 - i. Endothermic reaction
 - ii. Exothermic reaction
4. Recognize exothermic processes in the surroundings

5. Give examples of the endothermic and exothermic reactions
Identify the type of reactions in the following processes
6. Evaporation of water
7. Condensation of rain from water vapor
8. Melting ice cubes
9. Baking bread
10. Burning sugar

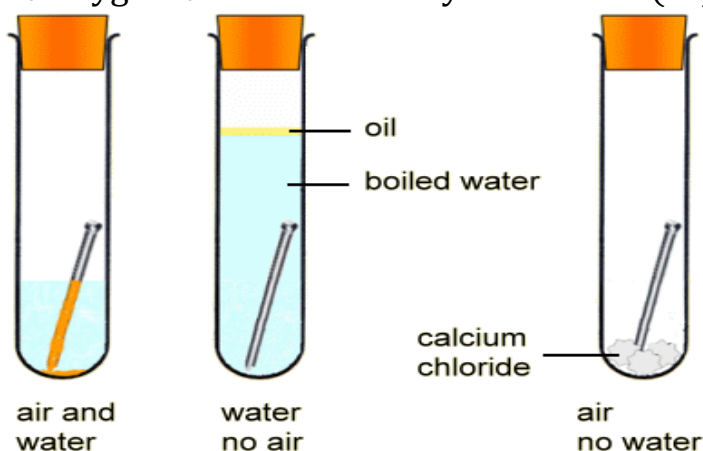
WEEK 7 CHEMICAL REACTIONS OF METALS

(RUSTING AS AN OXIDATION PROCESS)

The corrosion of iron is called rusting. The rusting of iron costs us billions of naira every year. Rust forms on the surface of iron (or most steels). Unfortunately, this rust is a soft, crumbly substance. It soon flakes off, exposing fresh iron, so that more iron can rust. Both air and water are needed for iron to rust. Rust is a form of iron (III) oxide. It has water loosely bonded to it. It is called hydrated iron (III) oxide.

We can show the reaction as:

Iron + oxygen + water \longrightarrow hydrated iron (III) oxide.

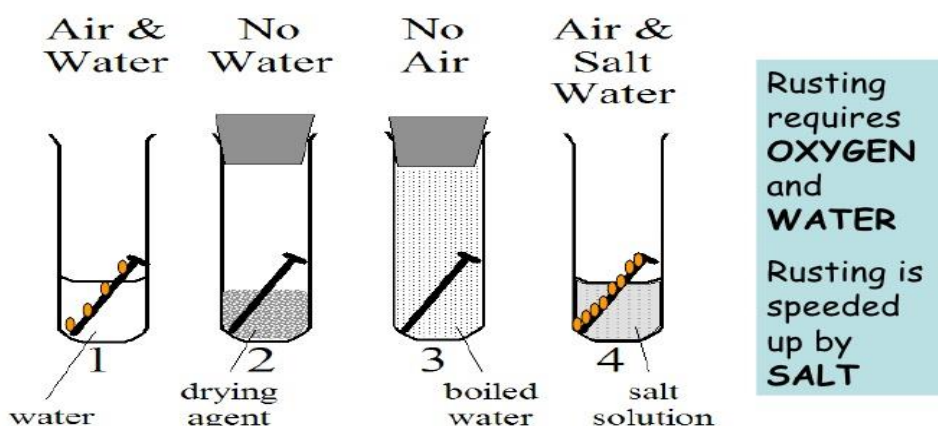


Why do metals rust?

Metal rusts because of a chemical reaction between oxygen and metal. This chemical reaction can speed up in water. Metals get rust because of oxygen. When air (which has oxygen in it) is able to get to the metal after a period of time rust starts to appear. (Rust is also called iron oxide. So, rust appears only on iron).

Chemical reaction can be speed up in water.

CHEMICALS FOR RUSTING



ECONOMIC IMPORTANCE OF RUSTING

1. Depletion of natural resources
2. Appearance as when corroded material is unpleasant to the eye
3. Loss the value of the iron
4. Changes the melting point
5. Health; for example, pollution due to escaping product from corroded equipment or due to a corrosion product itself.
6. Safety, for example, a sudden failure can cause fire, explosion, release of toxic product and construction collapse.

EXERCISE:

Define the following terms

1. Rusting
2. Combustion
3. Respiration
4. Identify how rusting, combustion, respiration is an oxidation process
5. In a tabular form, give (3) differences between burning and rusting
6. Highlight the examples of rusting, combustion, respiration
7. State the agents that aid rusting
8. What is needed for combustion to take place?
9. The chemical name for iron rust is?
10. Why is rusting process speeded up in a salt solution?

WEEK 8 CHEMICAL REACTIONS OF METALS

(FACTORS THAT INCREASE RATE OF RUSTING)

The factors without which rusting would not occur are exposure of the metal to water and oxygen. When these are both present rusting is likely to occur.

Factors which increase the rate of rusting:

1. higher temperature,
2. high water salinity,
3. higher atmospheric pressure.

Corrosion: When iron is exposed to water and oxygen, it corrodes. The corrosion products are the familiar red-brown rust. Indoors, a certain amount of moisture in the air (above about 65% relative humidity) is usually required before iron will rust. However, corrosion may occur at lower relative humidity if the object's surface is contaminated with salt, dust or other pollutants. Some iron objects have acquired an adherent rust layer from past exposure to high levels of moisture. This rust layer is often stable and will help slow corrosion of the underlying metal if the object is accidentally exposed to high moisture levels. Salt or other contaminants in the layer will stimulate further corrosion, especially at higher relative humidity. A sign of ongoing corrosion is flakes or loose powdery rust surrounding an object. Salts and oils on the skin can produce exactly the kind of corrosive environment that causes iron to react. Also exposure of iron objects to cooking materials such as salt and sauces can accelerate corrosion. If the iron object in question is considered valuable, it should not be used as it was in the past. It should be protected from excessive handling and kept in a dry location. Attics, basements and garages are not good places to store iron objects because of fluctuating environmental conditions. Iron has been used for several thousand years in many different forms.

EXERCISE:

1. What do you understand by the word 'corrosion'?
2. Highlight the factors that contribute to rusting
3. Highlight the factors that increase the amount of rusting
4. Identify the economic importance of rusting
5. Differentiate between burning and corrosion
6. Highlight the examples of rusting.
7. Explain the term 'humidity'
8. How can the humidity of a particular environment aid corrosion
9. Why does most iron rust?
10. Explain why the presence of oxygen contribute to rusting process

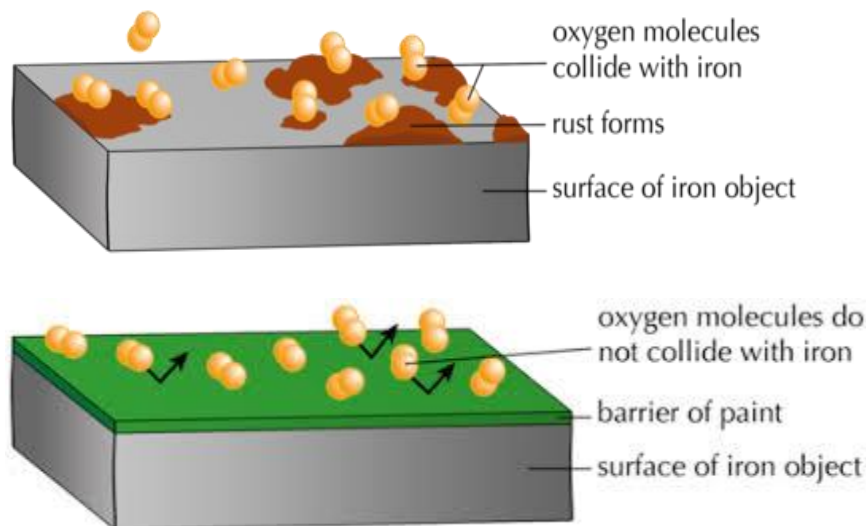
WEEK 9 &10: CHEMICAL REACTIONS OF METALS

(WAYS OF PREVENTING RUSTING)

We know that air and water are needed for iron to rust. Therefore, if we can keep these away from iron, it can't rust.

We can do this by coating the iron or steel with:

- Paint
- Oil or grease
- Plastic
- A less reactive metal, or
- A more reactive metal



Most methods for preventing rust rely on keeping the iron away from air and water. A complete barrier is placed around the iron, usually as a protective coating on its surface. The iron will rust if there is even a tiny gap in the coating. Then the rust soon spreads under the coating.

However, this does not happen if you use a more reactive metal. Even if the coating is scratched, the iron does not rust. Zinc is often used to protect iron. We say that the iron is **galvanised**. The zinc is **more reactive** than the iron. Therefore, any water or oxygen reacts with the zinc rather than the iron. This is called **sacrificial protection**. The zinc sacrifices itself to protect the iron.

Magnesium or aluminium can also be used instead of zinc to sacrificially protect iron. Sacrificial protection is used in harsh conditions, such as when the iron is in contact with sea water, which accelerates rusting. Examples are bars of zinc connected to the legs of an iron pier or a ship's hull. The method is also used to protect underground pipes which can be attached to the sacrificial metal by wires. The bars of metal need replacing eventually.

Sacrificial protection is also used where the coating is likely to be scratched, such as when emptying commercial wheelie bins. You can clearly see the coating of zinc metal on these bins and on some lamp posts.

Painting the metal prevents it from rusting by isolating it from oxygen and water. Biological corrosion is made from living organisms, and living cells and bacteria. The cells are made up into two categories, fungi and bacteria cells. Rust is the product of the reaction between iron and oxygen.

Sacrificial Anode

Definition - What does Sacrificial Anode mean?

Sacrificial anodes are easily corroded materials deliberately installed in a pipe or tank to be sacrificed to corrosion, leaving the rest of the system relatively corrosion free.

A sacrificial anode is also known as a galvanic anode.

The mechanism of the sacrificial anode protection system is very similar to the reaction mechanism of electrochemical cells. In sacrificial anodes the protected metal is placed on the cathode side and then a more reactive metal or alloy (having a larger potential difference than the protected metal) is chosen and connected to the protected metal as an anode. The redox reaction will proceed spontaneously. An oxidation reaction occurs at the anode, which means that the sacrificial metal will be consumed. At the same time, the reduction reaction occurs on the cathodic side, preventing the protected metal from erosion. Thus, corrosion on the protected metal is successfully shifted to the anode, protecting the metal.

Sacrificial anodes are normally supplied with either lead wires or cast-in straps to facilitate their connection to the structure being protected. The lead wires may be attached to the structure by welding or mechanical connections.



The materials used for sacrificial anodes are either relatively pure active metals, such as zinc or magnesium, or are magnesium or aluminum alloys that have been specifically developed for use as sacrificial anodes.

Advantages of using sacrificial anodes:

- Can be used where there is no power
- Lower initial cost
- Less supervision required
- Comparatively simple installation and additional anodes can easily be added if the initial installation proves to be inadequate

Sacrificial anodes are used to protect:

- Hulls of ships
- Water heaters
- Pipelines
- Distribution systems
- Above-ground tanks
- Underground tanks
- Refineries

The anodes in sacrificial anode-cathodic protection systems must be periodically inspected and replaced when consumed.

EXERCISE:

1. Highlight three preventive measure for rusting
Explain the following terms
2. Galvanized

3. Sacrificial anode
4. Highlight more examples of rusting.
5. Enumerate other ways of preventing rusting