WEEK ONE/ UNIT ONE: Atomic structure (Sub-Topic: Sub-Atomic particles) Course Learning Outcome:

(a) describe the structure of the atom in terms of the location of protons, neutrons and electrons in an atom
(b) define terms associated with the numbers of protons , electrons and neutrons
(c) describe the Bohr's model of the atom
(d) explain the occurrence of isotopes

(NTS 2b, 2c, 2e p. 13, 3h, 3j, p. 14).

Teaching Activity

You are to review previous knowledge on the particulate nature of matter and Dalton atomic theory. Review some experimental evidence that suggest that the atom is made of sub-particles. Learn the names and properties of the three subatomic particles and terms associated with their numbers each atom (atomic number and mass number), Use the suggested video and animation to get better understanding of the location of the sub-atomic particles emphasizing the atomic nucleus. Pay particular attention to fact that atoms of the same elements can have different mass numbers leading to the concept of Isotopy. Also search online for other e-books and learning materials on the structure of the atom. You note that any model of the atom is a mental picture based some experimental evidence of what is in reality.

Practice questions.

1. Which sub-atomic particles are responsible for the existence of isotopes and where are they located in an atom?
2. Describe the pictorial view of the atom according to the Bohr model

Read the following:

1. Abbey T. K., & Essia J. W. *GAST Core Science for senior secondary school*.
2. Galyoun I, Eghan J. M., Darkwa K. B. Owusu-SekyereJ. R. Zugle, (2008). *Black Star Series Integrated Science for Senior High Schools,* pp 38-40 Useful links

https://www.youtube.com/watch?v=hgCui4Qpswk

https://www.youtube.com/watch?v=EMDrb2LqL7E&list=RDCMUCiX8pAYWBppIbtUZTfGnRJw&start_radio=1&t=24

https://www.youtube.com/watch?v=EMDrb2LqL7E&list=RDCMUCiX8pAYWBppIbtUZTfGnRJw&index=1

WEEK TWO/ UNIT TWO: Structure of the Atom (Electronic Configuration) Course Learning Outcome:

(a) Number/label Bohr's electronic shells (orbits) in order of increasing energy

(b) allocate electrons into the various allowed energy levels/ shells of the Bohr's Model of the atom
(c) write out the electronic configuration in terms of the number of electrons in each shell (NTS 2c, 2e, 2f. p. 13, 3h, 3j, p. 14).

Teaching Activity

You will be required to review the terms atomic number and mass number as well as the Bohr's electronic shells. Carefully note the order of magnitude of the energies of these shell. **Pay particular attention to the maximum number of electrons allowed in each shell.** Practice filling the shells using the first twenty elements of the periodic table.

Practice questions.

1. An atom has atomic number of thirteen (13) and a mass number of twenty four (24). Write down its electronic configuration using the Bobr's model
2. What is the maximum number of electrons that are allowed in N shell of the Bohr's atom?

Read the following:

1. Abbey T. K., & Essia J. W. *GAST Core Science for senior secondary school*.
2. Ga;youn I, Eghan J. M., Darkwa K. B. Owusu-SekyereJ. R. Zugle, (2008). *Black Star Series Integrated Science for Senior High Schools,* pp 38-40

Useful links

https://www.youtube.com/watch?v=EMDrb2LqL7E

https://www.youtube.com/watch?v=hgCui4Qpswk

https://www.youtube.com/watch?v=EMDrb2LqL7E&list=RDCMUCiX8pAYWBppI
btUZTfGnRJw&start_radio=1&t=24

WEEK THREE/ UNIT THREE: Chemical Bonding (Ionic Bonding) Course

Learning Outcomes:

(a) explain the concept of chemical bonding as force of attraction (b) describe the processes involved in ionic bond formation.
(c) State properties of ionic compounds (NTS 2c, 2e, 2f. p. 13, 3h, 3j, p. 14).

Teaching Activity

You are to review the formation of ions (cations and anions) from neutral atoms. You are to use the electronic configurations of the noble gases to explain why they do not normally combine with other atoms. You are to explain an ionic bond as attraction between opposite charges.
Illustrate ionic bond formation using the Bohr's shells and the Lewis symbols. Use video and animation to get better understanding. Try to identify elements on the periodic table that can easily for ionic bonds/compounds and give the properties of these compounds.

Practice questions.

1. Use the Lewis symbols to show the formation of ionic bond between magnesium and nitrogen
2. Why is aluminum oxide solid and has a very high melting point? **Read the following:**

1. Abbey T. K., & Essia J. W. *GAST Core Science for senior secondary school*.
2. Galyoun I, Eghan J. M., Darkwa K. B. Owusu-SekyereJ. R. Zugle, (2008). *Black Star Series Integrated Science for Senior High Schools*, pp 42-48

Useful links

https://www.slideshare.net/FJHScience/chemical-bondsppt
https://www.youtube.com/watch?v=OTgpN62ou24
https://www.youtube.com/watch?v=pa3Id5hxmuY
https://www.youtube.com/watch?v=eNsVaUCzvLA

WEEK FOUR/ UNIT FOUR: Chemical Bonding (Covalent Bonding) Course

Learning Outcome:

(a) explain the concept of covalent bonding as force of attraction (b) describe the processes involved in covalent bond formation.
(c) state properties of covalent compounds (NTS 2c, 2e, 2f. p. 13, 3h, 3j, p. 14).

Teaching Activity

You are to explain how covalent bonds involving same atoms and different atoms are formed base on the octet and duplet rules. You are to use diagrams (Bohr's electronic shells and Lewis electron dots) to show the formation of covalent compounds. Use video and animation to get better understanding. Try to identify elements on the periodic table that can easily form covalent bonds/compounds and give the properties of these compounds.

Practice questions.

1. Use a period table to determine which of the following elements will form covalent bond/s with one another.
   Sodium, chlorine nitrogen, magnesium, hydrogen, oxygen

2. Use Lewis electron dots as well as the Bohr electronic shells to show the covalent compound, ammonia is formed.

Specific relevant materials

1. Galyoun I, Eghan J. M., Darkwa K. B. Owusu-SekyereJ. R. Zugle, (2008). *Black Star Series Integrated Science for Senior High Schools*, pp 42-48
2. P. W. Atkins and J. A. Beran, (1990). *General Chemistry 2*rd ed. W. H. Freeman & Co., New York, USA, Chapter 9. Useful links

- https://www.slideshare.net/FJHScience/chemical-bondsppt
- https://www.youtube.com/watch?v=OTgpN62ou24
- https://www.youtube.com/watch?v=pa3Id5hxmuY
- https://www.youtube.com/watch?v=eNsVaUCzvLA

WEEK FIVE/ UNIT FIVE

TOPIC: The Mole and formula mass Course

Learning Outcome:

(a) explain what is meant by the term 'amount of substance'.
(b) define the mole as a unit of measurement of amount of substance
(c) write the relationship between the mole (n), number of entities (N) and Avogadro's constant (L)

(d) solve simple calculations, solve calculations involving amount concentration and formula mass of a compound
(NTS 2b, 2c, 2e p. 13, 3h, 3j, p. 14).

Teaching Activity

You are to review fundamental quantities and their corresponding units of measurements. You are also to revise how to determine the formula mass of a compound. **Try to relate the mole as representing a specified number of entities just as everydayterms such as dozen (12), ream (500) do. You are to identify the Avogadro's number as representation of the mole.** Try to look at the sub and multiple quantities of the mole and relate the mole to other quantities or measurements (mass and concentration) in chemistry.

Practice questions

1. What is the specific numerical figure that the Avogadro's number represent?
2. What is meant by amount of substance?

Specific relevant materials

1. Ameyibor, K., &Wiredu M. B. (1991).*GAST chemistry for senior secondary school*. London: Macmillan Education Limited.
2. Galyoun I, Eghan J. M., Darkwa K. B. Owusu-SekyereJ. R. Zugle, (2008). *Black Star Series Integrated Science for Senior High Schools*, pp 49-50

Useful links

- https://www.youtube.com/watch?v=04AOsVQI9Bo
- https://www.youtube.com/watch?v=wI56mHUDJgQ

WEEK SIX/ UNIT SIX

TOPIC: Chemical Formulae and Equations Course

Learning Outcome:

(a) write the correct chemical formulae of common compounds
(b) give IUPAC names to simple inorganic compounds
(c) translate simple chemical reactions into equations

(d) balance simple chemical equations with the state of the substances attached to them. (NTS 2c, 2e, 2f. p. 13, 3h, 3j, p. 14).

Teaching Activity

You are to review valency or combing power of atoms and the law of chemical combinations. You have to note that chemical compounds are named according a system (IUPAC) of laid down rules which you have to learn. Distinguish trivial names from the IUPAC names. Try use the rules to name inorganic compounds including those with radical cations and anions as well as oxoacids. You begin the study of chemical reactions with simple processes such as burning, rusting and then identify chemical identities of the reactants and products

Practice questions

1. When hydrogen gas burns in oxygen (air) water is formed as the only product. Write a balanced chemical equation of the reaction indicating the states of the substances.
2. Write down the chemical formulae of the following compounds
   (a) Copper(I) oxide
   (b) Iron(III) oxide
3. Write down the IUPAC name for the following compounds
   (a) Sodium tetraoxosulphate (VI)
   (b) Iron(II) trioxonitrate(V)

Specific relevant materials

1. P. W. Atkins and J. A. Beran, (1990). *General Chemistry 2rd* ed. W. H. Freeman & Co., New York, USA, Chapter 2.
2. Galyoun I, Eghan J. M., Darkwa K. B. Owusu-SekyereJ. R. Zugle, (2008). *Black Star Series Integrated Science for Senior High Schools,* pp 49-50

   Useful links

   - https://www.youtube.com/watch?v=pa3Id5hxmuY
   - https://www.youtube.com/watch?v=eNsVaUCzvLA

WEEK SEVEN/ UNIT SEVEN

TOPIC: Pure and Impure Substances (Mixtures) Course

Learning Outcome:

(a) identify pure and impure substances
(b) identify some common mixtures
(c) describe methods of purification of impure substances
(d) describe steps involved in separating mixtures with three or more components.

(e) state the importance of purification of impure substances. (NTS 2b, 2c, 2e p. 13, 3h, 3j, p. 14).

Teaching Activity

You are to define a pure and impure (mixtures) substances and try to identify various mixtures in the environment including homes. You introduce purification methods by using experiences in everyday life involving hand picking, sieving, decanting etc., and state the difference in the properties of the components that is making it possible for the separation to occur. You can now try explaining the underlying principles of some laboratory separation techniques involving both organic and inorganic substances. Finally try to explain you would separate three or more component mixture. **Practice questions**

1. Why is brine a mixture
2. Outline steps as to how you would separate a mixture of sand, sugar and salt
3. State and explain the differences in the property of components of mixtures that make the following separation process possible
   (a) Filtration
   (b) Paper chromatography
   (c) Decantation
   (d) Column chromatography
   (e) Solvent extraction

Specific relevant materials

- Solomons, G. T. W., and Fryhle, C. B., (2007). *Organic Chemistry*, 9th ed., USA: John Wiley & Sons, Inc., Chapters 1-3

WEEK EIGHT/ UNIT EIGHT

TOPIC: Acid, Base and Salt

The Content Learning Outcomes (CLOs) for this topic are for student-teachers to be able to define acids and bases from the perspectives of Arrhenius, Bronsted Lowry and Lewis with examples, state the sources of acids bases and salts, describe the physical properties of acids, bases and salts, identify acid-base indicator, describe pH and indicators. (NTS 2b, 2c, 2e p. 13, 3h, 3j, p. 14). You have to try to classify compounds into various groups. Then try to identify those with similar effect with dyes especially litmus paper. Try to identify some common household/food items that are acidic or basic. You are to look at the definitions that various scientists to give to acids and basic. Note carefully the limitations of some these definitions. Obtain a chart should the index of acidity (pH) of some food items and other substance. Find out the importance of pH values in life and some industrial processes.

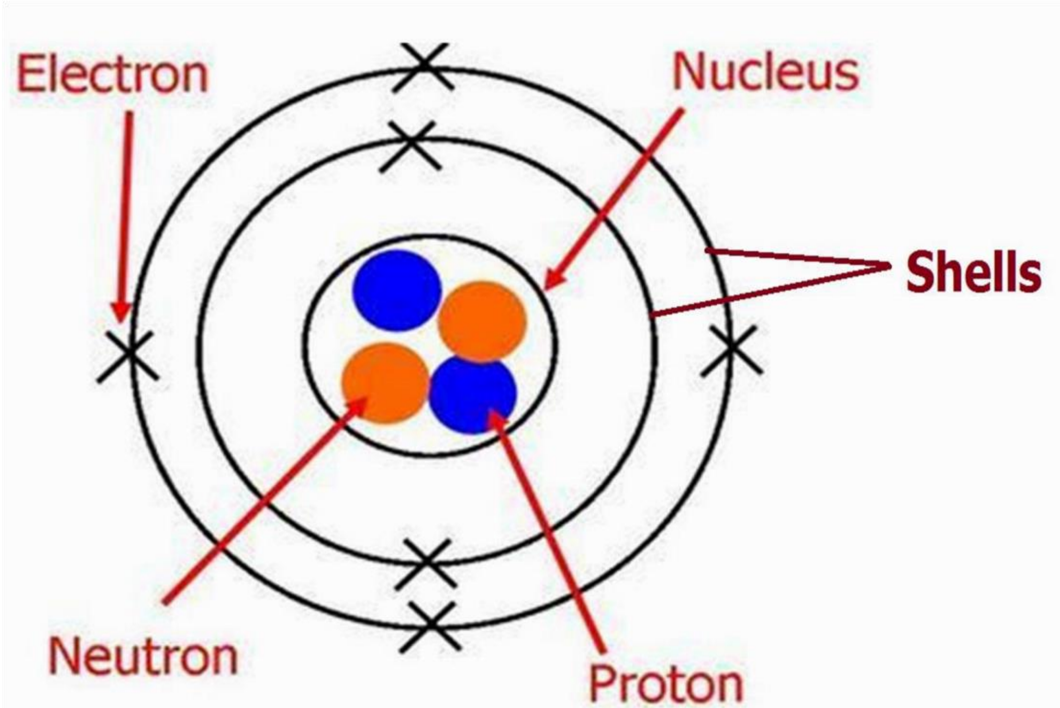
Specific relevant materials

- Petrucci, R. A., Hardwood, W. S., Herring, F. G., and Madura, J. D., (2007). *General Chemistry: Principles and Modern Applications*, 9th ed., Prentice Hall, New Jersey, USA, Chapter 15.
- P. W. Atkins and J. A. Beran, (1990). *General Chemistry 2*rd ed. W. H. Freeman & Co., New York, USA, Chapter 14 and 15.
- Ameyibor, K., &Wiredu M. B. (1991).*GAST chemistry for senior secondary school*. London: Macmillan Education Limited.

UNIT 1: STRUCTURE OF AN ATOM

An atom is the smallest particle of an element that can take part in chemical reactions. Atoms are made up of a nucleus of positively charged protons and neutral neutrons, surrounded by negatively charge electrons in the shell. An atom is electrically neutral because it has equal number of protons and electrons

The electrons, according to Bohr, move in circular orbits or shells around the nucleus



The number of electrons in atom is described as its atomic number. You should note that the number of electrons in an atom is always equal to the number of its protons in the nucleus. The total number of neutrons and protons in the nucleus of an atom is called its mass number

The atomic number of atoms of the same element is the same. However, their mass number could be different. This is because the number of neutrons in such atoms could be different. Then the atoms are said to be isotopes.

Dalton’s Atomic Theory [John Dalton (1808)]

The main points of Dalton’s atomic theory are:

1. Everything is composed of atoms, which are the indivisible building blocks of matter and cannot be destroyed.
2. All atoms of an element are identical.
3. The atoms of different elements vary in size and mass.
4. Compounds are produced through different whole-number combinations of atoms.
5. A chemical reaction results in the rearrangement of atoms in the reactant and product compounds.

An atom is electrically neutral because it has equal positive charged protons and negative charged electrons **where the negative charged electrons will cancel the positive charged protons** to form **neutral atom.**

SUB-ATOMIC PARTICLES

Particles	Relative change	Relative Mass	Location
Proton	+1	1	Inside the nucleus
Electron	-1	1/1840 (negligible)	On the shells or orbitals
Neutron	0	1	Inside the nucleus

- Proton has a positive charge with mass number 1 found inside the nucleus of an atom
- An electron has a negative charge with 1/1840 mass number and found on the shells of an atom
- Neutron has no charge but has mass number of 1 and it is found inside the nucleus of an atom.

NOTE: Electron has 1/1840 mass number, meaning; when we take 1840 protons or neutrons, we will get 1 electron out of the 1840 protons or neutrons.
This means electrons are much smaller than protons and neutrons that is why we sometimes refer to the relative mass of an electron as **negligible.**

ATOMIC NUMBER (Z)

- Atomic number of an atom is the number of protons found in the nucleus of an atom.
- Atomic number is represent by the symbol **Z**.
- When atom is in its neutral form, the atomic number is the same as electron number and the same as the proton number which is written as **Z = p = e**
- Atomic number (Z) is written as $_{6}C$ where **6** is the **atomic number or proton number or the electronic number** of the Carbon atom.

MASS NUMBER (A) / NUCLEON NUMBER

- Mass number is the **total number** of protons (p) and neutrons (n) found in the nucleus of an atom.
- Mass number is represented by the symbol **A**
- Mass number (A) is computed as A = p + n where p is the proton number and n is the neutron number found in the nucleus of an atom.
- Mass number (A) is written as ^{12}C where **12** is the mass number of the Carbon atom.

Solved Examples

1. Find the mass number (A) of Lithium atom with 3 protons and 4 neutrons.

   Solution

   Mass number (A) = number of proton (p) + number of neutron (n)
   Mass number (A) = 3 + 4

Mass number (A) = 7

2. An atom with mass number (A) of 24 and neutron (n) number of 12. Find
   i. the proton number
   ii. the type of atom
   iii. the number of electrons

<u>Solution</u>

i. Mass number (A) = proton number (p) + neutron number (n)
   Given mass number = 24,
   Neutron number = 12
   Proton number (p) = ?

   $A = p + n$

   $24 = p + 12$

   Change of subject

   $24 - 12 = p$

   $12 = p$

Therefore proton number (p) = 12

ii. As far as the proton number is 12 then it is Magnesium (Mg) atom.

iii. The atom has 12 electrons. Remember proton number (p) is the same as electron number (e) when the atom is at its neutral state.

3. Find the neutron number and electron number of Chlorine atom with 14 mass number (A) and 6 proton number (p).
   <u>Solution</u>

   Given mass number (Z) = 14
   Proton number (p) = 6
   Neutron number (n) = ?
   Electron number (e) = ?
   Mass number (Z) = proton (p) + neutron number (n)
   $14 = 6 + n$
   Change of subject
   $14 - 6 = n$
   $8 = n$ Therefore the neutron number = **8 neutrons**

If proton number is 12 then electron number = **8 electrons**

<u>Try</u>

1. The atom number of an element Y is 17 and its mass number is 37, calculate the number of i) protons ii) electrons and iii) neutron

2. Given two chlorine atoms with mass numbers 35, 37 with proton number (p) 17 for both atoms. Calculate the; neutron number for both atoms

ELECTRONIC CONFIGURATION

MEANING OF ELECTRONIC CONFIGURATION

- Electronic configuration is the representation of the arrangement of electrons distributed among the orbital shells and subshells.
- Electronic configuration shows how electrons are arranged on the shells of an atom.

BOHR'S ATOMIC MODEL

- In 1913, Neils Bohr proposed his quantized shell model of the atom to explain how electrons are positioned in an atom.
- The maximum number of electrons which each shell must occupy is calculated by $\mathbf{2n^2}$ where n is the number of shells or orbitals.
- The shells are assigned with letters K, L, M, L.
- K-Shell take a maximum number of 2 electrons
- L-Shell takes a maximum number of 8 electrons
- M-Shell can take 18 electrons but due to scientific reasons we will be using **8 electrons as maximum for M-Shell.**
- N-Shell takes a maximum of **32** electrons

) ARRANGEMENT OF ELECTRONS IN THE ATOM AND THE PERIODIC TABLE

Bohr put forward a theory of electron ~~position~~ positioning which is still generally accepted for chemical purposes.

Bohr suggested the existence of certain circular orbits (or shells) at definite distances from the nucleus in which the e's may rotate. The orbits are associated with definite energy content of the e', increasing outwards from the nucleus. An electron in a given orbit does not radiate energy, however if the e's absorbs energy, ~~the~~ it may jump to one of the outer orbits of greater energy. If it later falls back to an inner orbital of lower energy energy will be radiated as light of definite colour (or frequency).

Bohr's work and other subsequent works lead to the following conclusions.

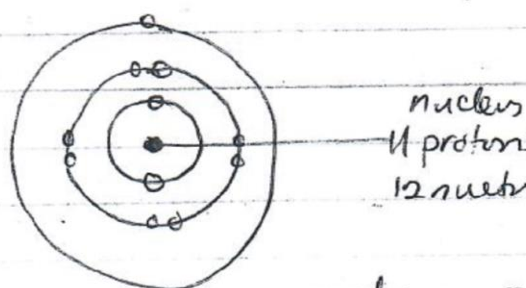
1, Several gps of e's may occur in an atom and each

grp is known as an electron shell. Shells are numbered 1, 2, 3 4, 5, 6 7 etc
or K L M N O P Q etc
outwards from the nucleus. All e's in a given shell have approximately equal energy. This energy increases in successive shells outwards from the nucleus

2. The ~~Maximum~~ possible number of e's in a shell numbered n is $2n^2$ i.e in successive shells 2, 8, 18, 32 -- electrons.

3. In the outermost shell of an atom, the maximum number of electrons possible is 8.

A diagramatic representation of a typical atom is as shown below



electrons revolving in orbits 2, 8, 1

NOTE:

- The orbit closer the nucleus has minimum energy and the orbits far from the nucleus has maximum energy.
- This means, the movement of electrons which are closer to the nucleus (the first two electrons) moves round the nucleus with minimum speeed or energy whilst the movement of electrons which are far from the nucleus (the other eight electrons on the next shells) move round the nucleus with maximum speed or energy.

TABLE SHOWING ARRANGEMENT OF ELECTRONS IN SHELL
OF ATOMS OF SOME ELEMENTS

ELEMENT	PROTONS (№)	ELECTRONS IN SHELL K	L	M	N
hydrogen	1	1			
helium	2	2			
lithium	3	2	1		
beryllium	4	2	2		
boron	5	2	3		
carbon	6	2	4		
Nitrogen	7	2	5		
oxygen	8	2	6		
flourine	9	2	7		
neon	10	2	8		
Sodium	11	2	8	1	
magnesium	12	2	8	2	
aluminium	13	2	8	3	
~~Phosphorus~~ Silicon	14	2	8	4	
phosphorus	15	2	8	5	
Sulphur	16	2	8	6	
chlorine	17	2	8	7	
argon	18	2	8	8	
Potassium	19	2	8	8	1
calcium	20	2	8	8	2

CHEMICAL BONDING
(INTERATOMIC BONDS)

Most elements exist naturally in a combined state with other elements. That is they exist as compounds. Gases such as argon and neon are however generally chemically inert and have great diffulty in forming compounds with other elements. That is they are Self-Satisfied. These gases neon and argon have eight electrons in their outermost electron shell and are said to have an octet of electrons. In the simpler noble gas, helium, the duplet (two electrons in outermost shell) is equally stable and functions like the octet.

The tendency of other elements is to try to attain this noble gas structure of a stable outer octet (or duplet) of electrons and their chemical behaviour is a reflection of this's tendancy.

① THE OCTET RULE

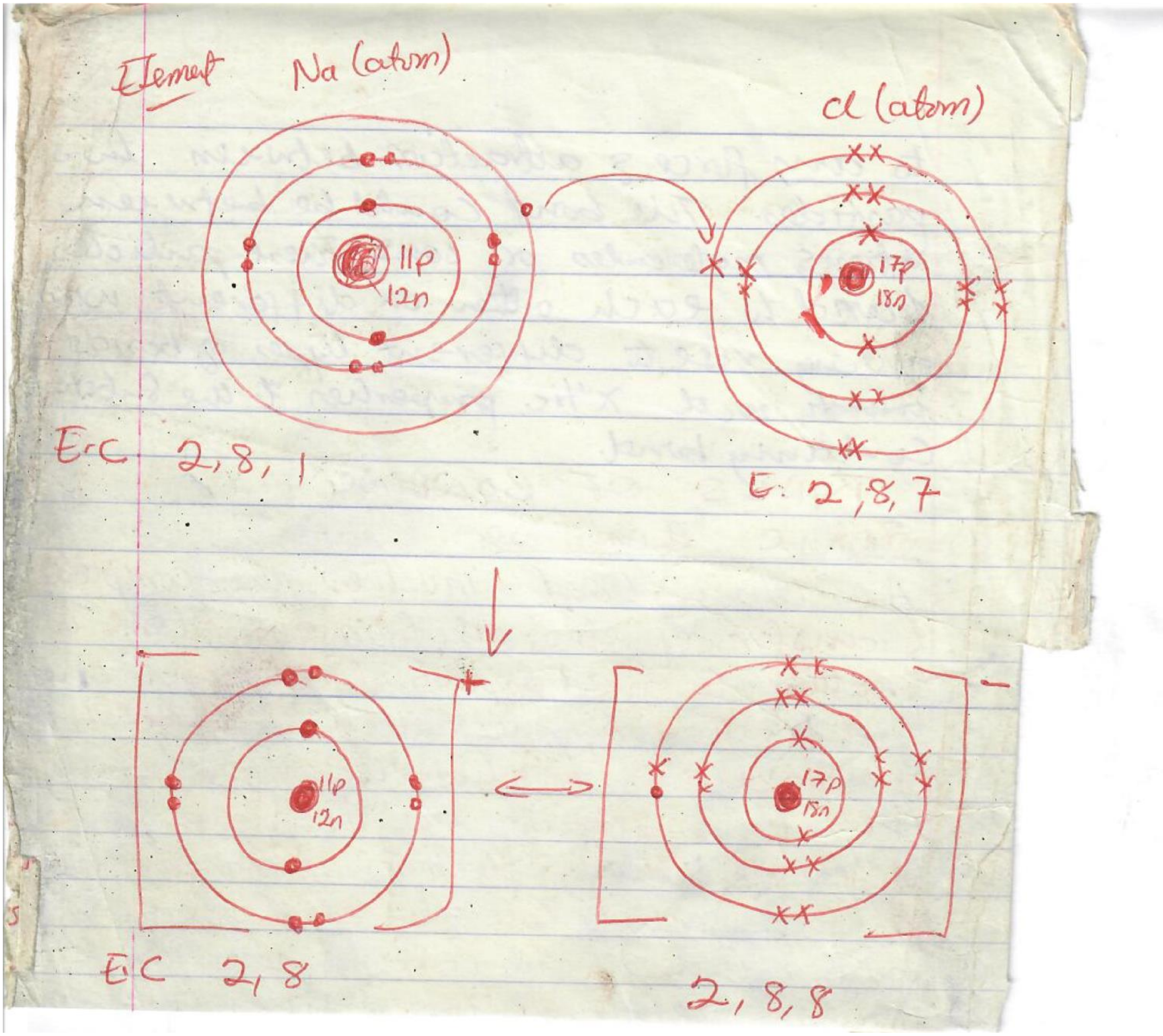
This is the process of achieving a stable electronic configuration of 2 or 8 electrons in the outer shell.

A chemical bond, or simply a bond refers to any force of attraction between two particles. The bond could be between atoms, molecules or ions. These particles bond to each other in different ways giving rise to different types of bonds ~~which~~ with x'tic properties to the substance containing bond.

TYPES OF BONDING

(1) IONIC BONDING

An ionic bond involves the complete transfer of electrons from one atom to another to give oppositely charged ions held together by electrostatic ~~attract~~ force of attraction. In this type of bonding, the atoms attain the noble gas structure by either losing electrons or gaining electrons. The atoms that lose the electrons form positive ions called cations and are usually metals or Those atoms that gain electrons form negative ions called anions. These atoms are usually the nonmetals. An ionic bond bond between sodium (a metal or an electropositive element) and chlorine (an electronegative element or nonmetal) [illegible] illustrated below



NOTE:

- Ionic bonding occurs between chemical combination of metallic and non metallic atoms. Some metallic atoms are Sodium (Na), Aluminium (Al), Calcium (Ca), Zinc (Zn), Magnesium (Mg), Gold (Au), Lead (Pb) etc.
- Ome non-metallic atoms are oxygen (O), sulphur (S), Nitrogen (N), hydrogen (H), chlorine (Cl), carbon (C), boron (B), etc.
- Some elements have complete maximum number of electrons occupied on its shells and they are called *nobel or inert gasses*; example, Helium (He), Neon (Ne), Argon (Ar), Krypton (kr) etc. (Rare gasses are used to fill balloons and filled fluorescent bulbs)

- Sodium atom can lose its electron to another atom which needs an electron. When this happens, we say the sodium atom (Na) is now called **sodium ion (Na^+).**
- The ionic bond will be $Na^+ + Cl^-$ (this means sodium atom is **losing** its 1 electron on the third shell in order to **transfer** it to chlorine atom who needs or is **gaining** 1 electron from sodium in order to be stable)
- Ionic bonding is the **chemical combination** of two or more **ions** (an atom which has lose or gain an electron to be stable).
- When there is an ionic bonding, it forms what we call **ionic compound.** Example**;**

  i) $Na^+ + Cl^- \longrightarrow$ **NaCl**

  ii) $Mg^{2+} + O^{2-} \longrightarrow MgO$

  <u>Explanation</u>

  Magnesium (Mg) is a metallic atom with electronic configuration of **2,8,2 (meaning it can lose the 2 to be stable)**

  Oxygen (O) is a non-metallic atom with electronic configuration **2,6** (**meaning it needs 2 electrons to complete the octet rule and to be stable**) so Magnesium ion can easily bond with oxygen ion to form *Magnesium oxide* which is an ionic compound.

COVALENT BONDING

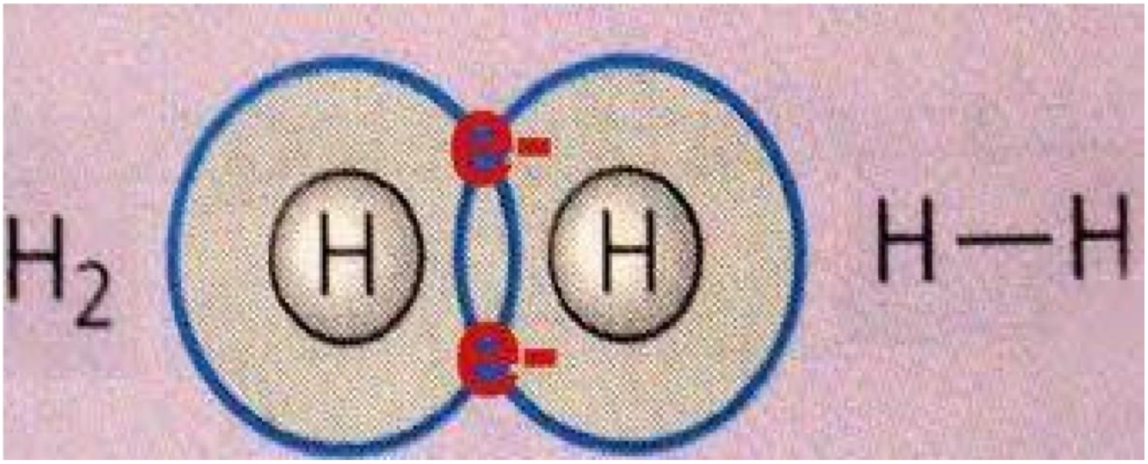
A **covalent bond** is a chemical bond formed by the sharing between atoms of their outermost electrons such that each atom acquires a noble gas electron configuration. The shared valence electrons between two non-metal atoms is called a *covalent bond.*

A covalent compound, usually a molecule, is formed when two or more non-metal atoms *bond* by sharing valence electrons.

A covalent bond can be single, double and triple depending on the number of electrons shared between atoms.

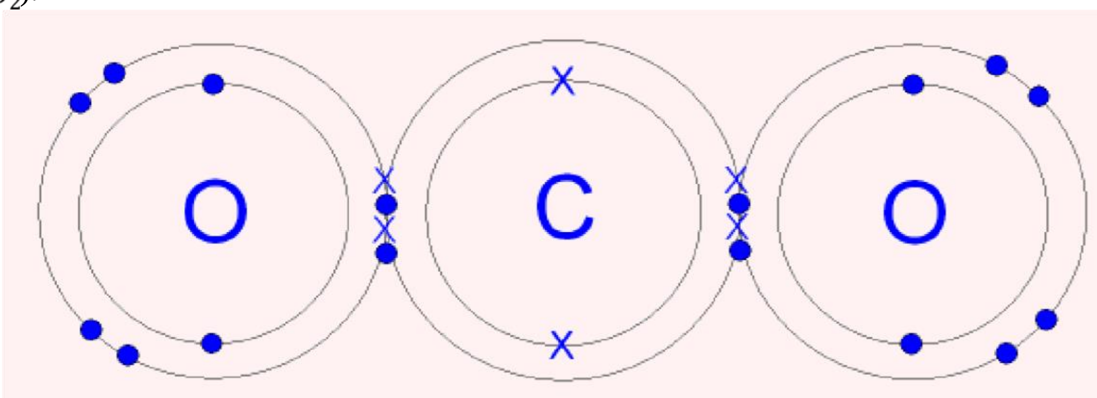
Single covalent bond

It is the sharing of two electrons between two atoms (one electron from each atom). Eg. NH_3, F_2, Cl_2, Br_2, etc. An illustration of a single covalent bond between two hydrogen atoms to form molecular hydrogen is shown below.



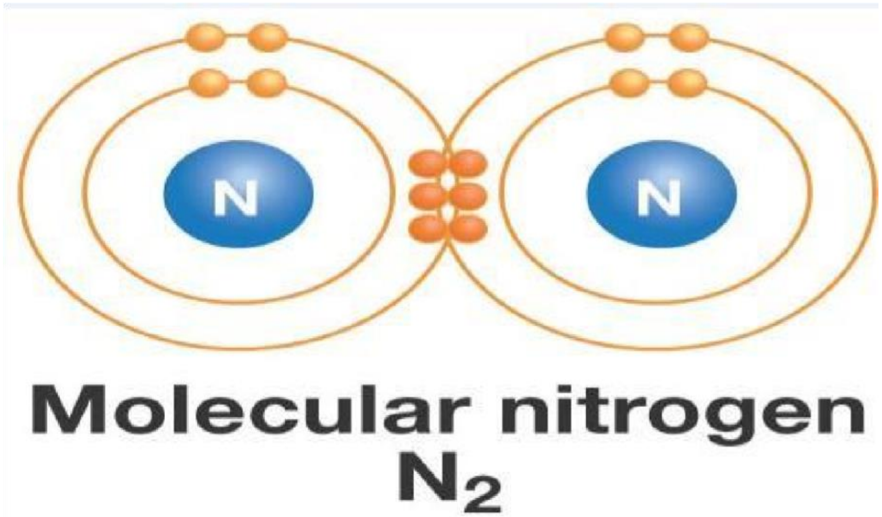
Double Covalent bond

Is the sharing of four electrons between two atoms (two electrons from each atom). E.g.: Carbon dioxide (CO_2).

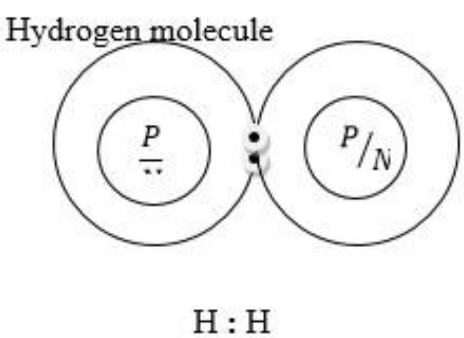


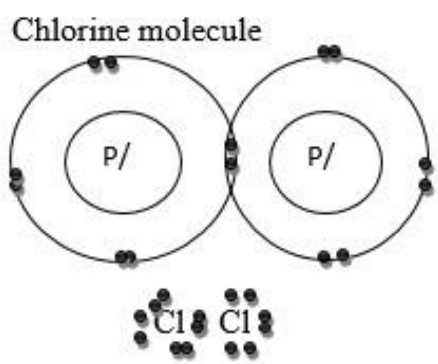
Triple Covalent bond

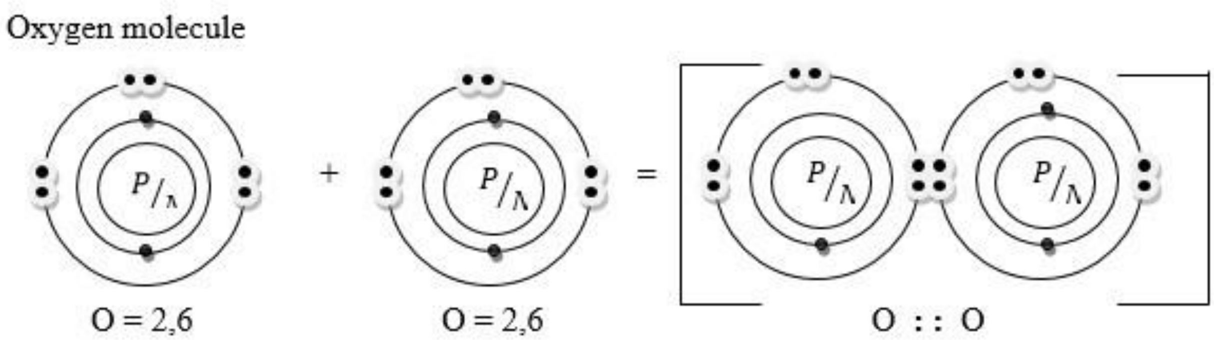
This is the sharing between two atoms of six electrons (three electrons from each atom). E.g. Nitrogen gas, N_2



Other examples of covalent bond formation is shown below with the corresponding electron dot representations.



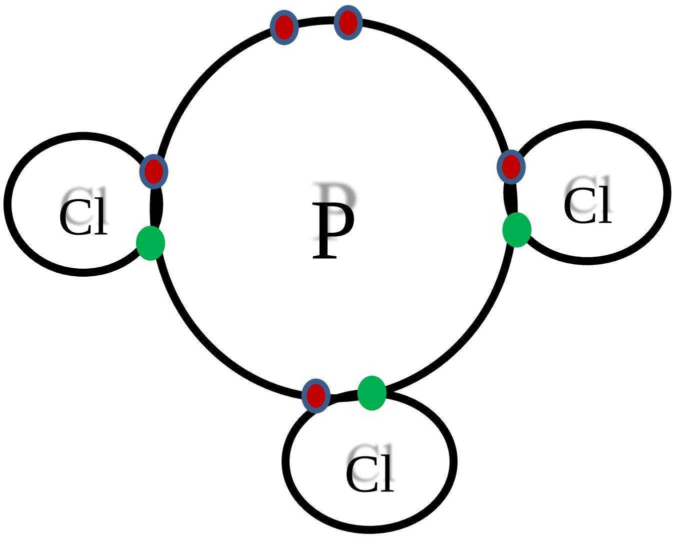




Work Examples

1. Using only the valence electrons, draw diagrams to show the covalent bonding in the following compound PCl_3

Solution



LEWIS DOT STRUCTURES

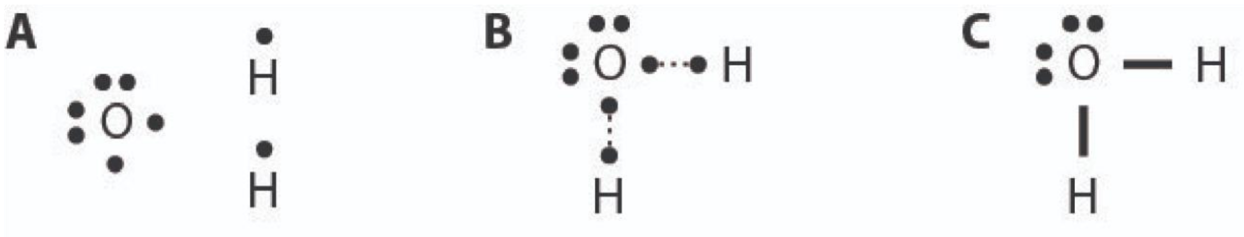
Lewis dot structures are one way to represent how atoms form covalent bonds. A table of Lewis dot symbols of nonmetal elements that form covalent bonds is shown below.

Element	H	C	N	O	F
Symbol	H	C	N	O	F

Important things to note about electron representation of atoms and molecules

1. Dots are placed around the symbol of the element to represent the number of valence electrons in the element.
2. There can be up to eight dots, for eight valence electrons.
3. The first four electrons are placed as single electrons, then the remaining four are paired.
4. The number of bonds that each element is able to form is usually equal to the number of unpaired electrons.
5. In order to form a covalent bond, each element has to share one unpaired electron.

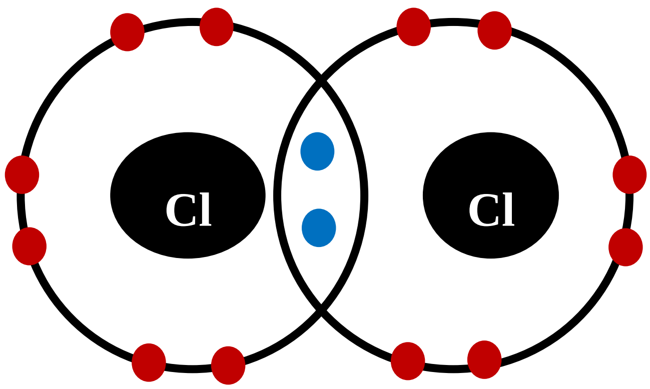
For example water (H_2O)



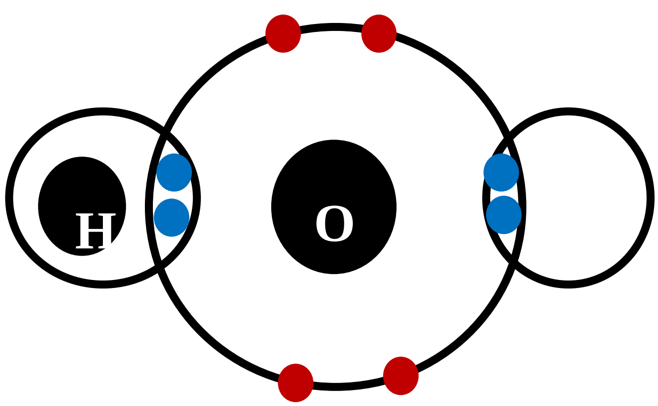
LONE PAIRS OF ELECTRONS

Pairs of outermost shell electrons that are not used in a covalent bond formation are known as *lone pairs* of electrons

Consider the chlorine molecule, Cl_2, each chlorine atom contains three lone pairs of electrons



For a molecule of water,H_2O, there are two lone pairs of electrons on the oxygen atom.



Properties of covalent compounds

1. They have low melting and boiling points. This explain why most covalent compounds exist as gases (e.g carbon dioxide, hydrogen, hydrogen sulphide), liquids (water, kerosene) or solids with low melting points (e.g. naphthalene, shear butter)

2. Covalent compounds are insoluble in water, e.g. kerosene or oil cannot dissolve in water
3. They do not conduct electricity and are often used as insulators
4. They are soluble in organic solvents like petrol, kerosene and diesel.

Ionic Verses Covalent Compound

Ionic Compounds	**Covalent**
They are made up of charged particles (ions)	They are made up of molecules
They have strong chemical bonds between ions	They have strong chemical bonds inside the molecule
Are usually water soluble	Are usually water soluble
Have high melting and boiling points	Have low melting and boiling points
Conduct electricity when molten or aqueous	Do not conduct electricity
Are crystalline solids	Are often gases or volatile liquids

THE MOLE AND FORMULA MASS

What is the moles?

In everyday life, we often use terms to represent a specific numerical number of items. For instance, the term *dozen* refers to a collection of 12 items, and a *score* is also a collection of 20 items, a *realm* refers to 500 items.
The *mole* could be said to be such a term representing a collection of 6.02×10^{23} particles called the Avogadro's number.
Scientifically, the mole is the amount of substance which contains the same number of elementary particles as the number of atoms (6.02×10^{23}) contained in 12g of carbon - 12 isotope.

Some basic facts about the mole

1. The **mole** is the unit of measurement of an amount of substance.
2. One mole of any substance contains **$6.02\text{x}10^{23}$** particles of that substance.
3. One mole of every substance is represented by its **chemical formula**.

Avogadro Constant (L = 6.02×10^{23} particles/mol)

The Avogadro's constant is defined as the number of particles in 1 mole of a substance.

Relationship between the mole (n), the number of particles (N) and the Avogadro's constant (L) is given by:

$$\boxed{N = n \times L}$$

WORKED EXAMPLES

1. How many atoms are in 0.2mole of calcium? ($L = 6.02 \times 10^{23} mol^{-1}$)

Solution

Number of moles of calcium, (n) = 0.2moles
Avogadro's Constant, (L) = 6.02×10^{23}
Number of atoms of calcium = n x L

$= 0.2 \times 6.02 \times 10^{23}$
$= \mathbf{1.204 \times 10^{23}} \boldsymbol{atoms}$

2. How many atoms are there in 0.4moles of hydrogen molecules? [L= 6.02x1023] **Solution** moles (n) of H_2 = 0.4

N (H_2) = 0.4 x 6.02x1023 = 2.408x1023 molecules
But the question demands hydrogen atoms
Therefore, N (H) = 2 x 2.408x1023 = **4.8x1023 atoms**
Note that, H_2 = 2 x H. i.e. 1 molecule of hydrogen = 2 atoms of hydrogen

Relative Molecular Mass (M_r)

The relative molecular mass of a molecule of an element or compound is the number of times the mass of one molecule of the element or the compound is as heavy as one – twelfth (1/12) of the mass of one atom of carbon – 12.

The relative molecular mass is similar to the relative atomic mass in principles. The main difference is that M_r deals with molecules whereas A_r deals with atoms.

Mathematically, M_r can be expressed as

$$M_r = \frac{average\ mass\ of\ 1\ molecule\ of\ a\ substance}{1/12\ of\ the\ mass\ of\ 1\ atom\ of\ carbon\ 12}$$

Worked Examples

Determine the relative molecular masses of the following:
a. Sodium chloride (NaCl) **b.** Sodium Hydroxide (NaOH) **c**. Sulphric acid (H_2SO_4) [Na=23, H=1, O=16, Cl=35.5, S= 32]
Solution

a. Sodium Chloride = NaCl

One atom of sodium = 23
One atom of chlorine = 35.5

Mr. of NaCl = 23+35.5 = 58.5

b. Sodium Hydroxide = NaOH

One atom of sodium = 23
One atom of oxygen = 16
One atom of hydrogen = 1

Mr of NaOH = 23+16+1 = 40

c. Sulphric acid = H_2SO_4

2 atoms of hydrogen = 2x1
1 atom of sulphur = 32
4 atoms of oxygen = 4 x16
Mr of the H_2SO_4 =2 + 32 + 64 =98

Molar mass of a substance is the mass of 1 mole of that substance.

The unit of measure of molar mass is the g/mol.Molar mass of a substance is determined by summing the atomic masses of the individual atoms that make up the substance.
Numerically, molar mass = relative molecular mass.
Molar mass is related to the amount of substance as follows
$n = \frac{m}{M}$,
Where n = amount of substance (mole), m = mass of the substance in grams, M = molar mass of the substance

Worked Examples

1. Calculate the molar mass of H_2SO_4
[H=1, O=16, S=32]
Solution

a. Note: 1 mole of H_2SO_4 contains 2 atoms of H, 1 atom of S and 4 atoms O.
∴Molar mass of H_2SO_4 = (1×2)+32+(16×4)
=2 +32+64

= **98g/mol**

2. Calculate the molar mass of Calcium hydroxide

[H=1, O=16, Ca=40]
Solution

Chemical formula for calcium hydroxide is $Ca(OH)_2$
Molar mass of $Ca(OH)_2$=40+2(16+1)=**74g/mol**

3. Calculate the molar mass of Sugar ($C_{12}H_{22}O_{11}$)

[Relative atomic masses: H=1, O=16, C=12]

Solution

$M(C_{12}H_{22}O_{11})=(12\times12)+(1\times22)+(16\times11)=$ **342 g/mol**

4. Calculate the molar mass of 0.025moles of carbon (IV) oxide gas which has a mass of 1.1g

Solution

Mass of CO_2 = 1.1g, Amount of substance (n) = 0.025moles

$$M = \frac{m}{n}$$

$$= \frac{1.1g}{0.025} = \mathbf{44.0}\boldsymbol{g/mol}$$

1. Calculate the molar mass of the following substances:

a. Sodium chloride

b. Ammonium sulphate

c. $CuSO_4$

[Na = 23, Cl = 35.5, S = 32, O = 16, N = 14, Cu = 64, H = 1]

2. Calculate the number of molecules in 6.4g of sulphur (IV) oxide gas. (S= 32, O= 16, L= 6.02×10^{23})

Number of Moles in a given Mass of a Substance

Calculate the number of moles in the following mass of substances.

a. 5.0g of NaOH

b. 13.05g of CH_3CH_2OH

c. 1.2g of HNO_3 [H=1, C=12, N=14, O=16, Na=23]

Solution

a. Mass of NaOH, m = 5.0g

Molar mass of NaOH, M=23+16+1 = 40g/mol

Number of moles, $n = \frac{m}{M}$

$$= \frac{5.0g}{40g/mol} = \mathbf{0.125}\boldsymbol{mol}$$

b. $\boldsymbol{n} = \mathbf{0.284}\boldsymbol{mol}$

$\boldsymbol{n} = \mathbf{0.019}\boldsymbol{mol}$

Work out the steps

Mass of a substance in a given number of moles

♣ From $n = \frac{m}{M}$, $\Rightarrow m = n \times M$

Worked Examples

Find the mass of 0.5mol of ethanol molecules (C_2H_5OH). [H=1, C=12, O=16]
Solution

Number of moles, n = 0.5mol
Molar mass, $M = (12 \times 2) + (1 \times 5) + 16 + 1 = 46g/mol$
Mass, $m = n \times M = .5mol \times 46g/mol = \mathbf{23g}$

CONCENTRATION OF SOLUTION

In everyday life we express concentration effects in the form of taste (how sweet or bitter something is) or in the intensity of colour. For instance, four cubes of sugar in tea in a normal tea cup will taste sweeter than the one with only one cube of sugar. Similarly, ten grams of a blue dye in a gallon of solvent will appear deeper than when only two grams of the dyes is placed in a gallon of solvent.

In chemistry we normally express concentration as the amount of substance dissolved in $1dm^3$ of solution.
It is given by the symbol C and has a unit $moldm^{-3}$ or M.
Mathematically: concentration,

$$c = \frac{quantity\ of\ solute(n)}{volume\ of\ solution(v)}$$

Types of Concentration

Mass Concentration: It is the amount of substance in grams dissolved in $1dm^3$ of solution.

- $= \frac{\qquad (\qquad)}{}$
- $= -$
- The S.I. unit is g/dm^3

Molar Concentration

This is defined as the quantity of solute in moles dissolved in $1dm^3$ of solution.

- ❑ $molar\ concentration(c) = \frac{quantity\ of\ solute(in\ moles)}{volume\ of\ solution(v)}$

$$\rightarrow c = \frac{n}{v}$$

It follows that:

- ❑ $molar\ concentration(c) = \frac{mass\ concentration(\rho)}{molar\ mass(M)} \rightarrow c = \frac{\rho}{M}$
- ❑ The S.I. unit is mol/dm^3.

Worked Examples

1. When 0.2 mol of sodium trioxocarbonate (IV) is dissolved in water to make $0.5dm^3$ of solution. Calculate the concentration of the solution

   Solution

   Number of moles, n = 0.2 mol, volume of solution, $v = 0.5dm^3$
   Molar concentration, $c = \frac{n}{v} = \frac{0.2\ mol}{0.5\ dm^3}$
   $= \mathbf{0.4\ mol/dm^3}$

2. Calculate the molar concentration if 15g of sodium hydroxide (NaOH) is dissolved in $2dm^3$ of solution.

   [Na = 23, O =16, H= 1]
   Solution

   Concentration, $c = \frac{n}{v} = \frac{m}{v \times M} = \frac{15}{2 \times 40}$
   $= \mathbf{0.1875 mol/dm^3}$

3. Calculate the concentration in grams per dm^3 when

   a. 20g of NaOH dissolved in $5dm^3$ of solution

   b. 0.4mol NaOH dissolved in $800cm^3$ of solution

   (Na = 23, O= 16, H=1)

   Solution

   a. Mass conc., $\rho = \frac{m}{v} = \frac{20g}{5dm^3} = \mathbf{4g/dm^3}$

b. Mass conc., $\rho = \frac{m}{v}$, but m = n $\times M$

$$\rightarrow \rho = \frac{n \times M}{v} = \frac{0.4 \times 40}{0.8} = \boldsymbol{20gdm^{-3}}$$

4. What volume of solution would have a concentration of 0.01mol/dm³ when 5.0g of $Ca(OH)_2$ is completely dissolved in it? $\left[\boldsymbol{Ca = 40, O = 16, H = 1}\right]$

Solution

Molar conc., c = 0.01mol/dm³, mass = 5.0g, v = ?
Molar mass of $Ca(OH)_2$= 40+2(16+1) = 74g/mol
From $c = \frac{n}{v}$ and $n = \frac{m}{M}$ $\Rightarrow c = \frac{m}{v \times M}$

$$\therefore v = \frac{m}{c \times M} = \frac{5}{0.01 \times 74} = 6.757dm3 \approx \boldsymbol{6757cm3}$$

5. Calculate the mass concentration of a 0.75M H_2SO_4.

(H=1, S=32, O = 16)

Solution molar conc. (C) =

0.75M

$M(H_2SO_4)$ = 2 (1) + 32 + 4 (16) = 98gmol^{-1}

Then from;
Mass Conc., $\rho = C \times M$
$= 0.75 \times 98 =$

= **73.5gdm^{-3}**

PREPARATION OF SOLUTION

Terminologies

Solution: A homogeneous mixture of two or substances. A solution is formed when a solute dissolves in a solvent.
A **solute** is a substance that is able to dissolve in a solvent.
A **solvent** is any liquid in which a solute dissolves. E.g.: water, kerosene.

Types of Solution

Dilute solution is the solution in which small amount of solute is dissolved in a large volume of solvent.
Concentrated Solution is a solution in which a large amount of solute is dissolved in a given volume of solution. The bigger the amount of solute in a solution, the more concentrated the solution is.
Saturated Solution is a solution which can't dissolve any solute at a particular temperature.
Unsaturated Solution this is a solution which can dissolve more solute when added to it.
Standard Solution is a solution whose concentration is exactly known.

Preparation of Standard Solution

In the preparation of a standard solution, the following care must be taken:

1. Don't wash your hands with an unknown liquid, since some liquids (such as acids) are corrosive

2. Don't drink an unknown liquid in the laboratory, they can be toxic **Basic**

   Apparatus used

Weighing scale /chemical balance, Spatula, Volumetric flask Beaker, Wash bottle, Stirring rod

Steps involved in the preparation of Standard Solution

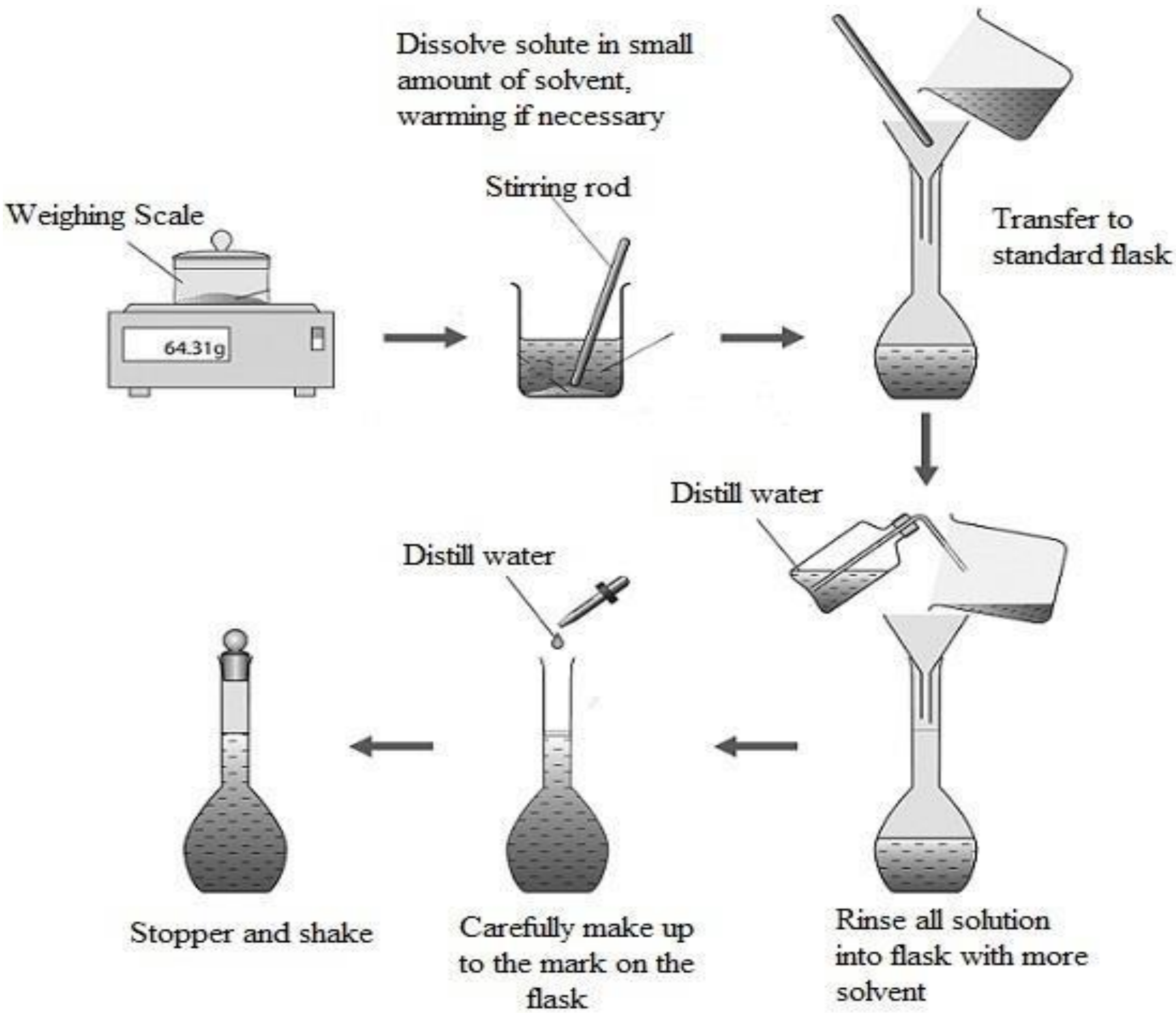
1. Calculate the molar mass (M) of the solute (substance) by using the chemical formula.
2. Calculate the mass (m) in grams of the solute using the relation: $\boldsymbol{m = C \times M \times V}$, where m = mass C = concentration in mol/dm^3 and V = volume in dm^3
3. Weigh the calculated mass of the solute using the chemical balance

4. Dissolved the weighed solute in about 250cm^3 of distilled water in a clean dry beaker.

5. Transfer the resulting solution into a 1000cm^3 or 1dm^3 volumetric flasks.

6. Add more distilled water to the flask until the 1000cm^3 mark is reached. Cover the volumetric flask and shake well to dissolve the solute particles.

These steps are summarised in the flow chart.



CHEMICAL FORMULA AND EQUATIONS

Introduction

In chemical sciences, name of elements and compounds are most often presented in some standard abbreviated forms called chemical symbols and chemical formulae respectively.

Chemical symbols of all elements are usually presented in a special tabular form called **Periodic Table**. The periodic table allows a continuous addition of the chemical symbols of elements yet to be discovered. The special features of the periodic table are outline below.

The periodic table

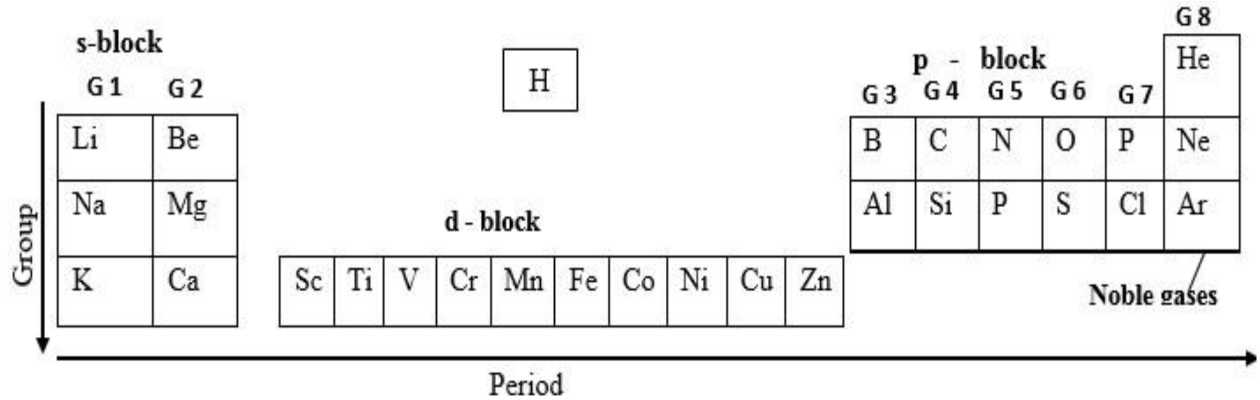
The periodic table shows the arrangement of elements in horizontal rows ('periods') and vertical columns ('group') according to the number of protons in the nucleus or atomic number.

Period is a set of elements which have the same number of shells and whose atomic numbers increase by one unit from one atom to the next. E.g. K and Ca are in the same period.

Group is a set of elements with the same number of valence electron(s) whose atomic number increase. E.g. Li, Na and K are in the same group.

Another set of classification on the Periodic Table is by blocks. The periodic table is blocked according to the electronic configuration of the elements and increasing atomic number. For instance, elements in the s-block have their valence electrons in the s-orbital. On the other hand, elements with their valence electrons in the d-orbital occupy the d-block and are called transition elements.

The basic features of the Periodic table is shown below.



Chemical Symbol

Chemical elements are represented symbolically by one or two letters normally derived from their names. A very few elements have symbols which appear to have no relationship with their names. In those cases, the origin of the symbol used is given. A chemical symbol represents one atom of an element. For example, H means one atom of hydrogen.

Chemical Formula

A **chemical formula** is a way of presenting information about the chemical proportions of atoms that constitute a particular chemical compound or molecule, using chemical element symbols, numbers, and sometimes also other symbols, such as parentheses, dashes, brackets, commas and *plus* (+) and *minus* (−) signs.

For example, is represented by the formula NH_3. This represents the molecule of ammonia. It contains an atom of Nitrogen and three atoms of Hydrogen.

Also, Potassium chloride has formula KCl. It represents the formula unit of potassium chloride

The below shows the names and chemical formulae of some everyday substances.

Compound	**Elements Present**	**Formula**
Water	**Hydrogen; Oxygen**	H_2O
Sodium chloride	**Sodium; Chlorine**	**NaCl**
Sugar (glucose)	**Carbon; Hydrogen; Oxygen**	$C_6H_{12}O_6$
Sand	**Silicon; Oxygen**	SiO_2
Ammonia	**Nitrogen; Hydrogen**	NH_3
Methane	**Carbon; Hydrogen**	CH_4
Ethanol	**Carbon; Hydrogen; Oxygen**	CH_3CH_2OH
Chalk	**Calcium; Carbon; Oxygen**	$CaCO_3$

Steps in Writing Chemical Formulae of Binary Compounds

1. Write the symbols of the elements involved.
2. Indicate their valencies below them.
3. Interchange their valencies and write them as subscripts.
4. Reduce the numbers to simplest ratios.

Example 1

Magnesium   Oxygen

Mg   O

2   2

Mg_2O_2

MgO

Example 2

Aluminium   Oxygen

O   Al

3   2

Al_2O_3

Other examples involving sodium chloride and water are shown below.

1. Sodium   Chlorine

Na   Cl

1   1

$Na_1Cl_1 = NaCl$

2. Hydrogen   Oxygen

H   O

1   2

$H_2O_1 = H_2O$

Steps in Writing Chemical Formulae of Non-Binary Compounds

1. Write the symbols of the metal and the radical involved.

2. Indicate their valencies below them.

3. Interchange their valencies and write them as subscript.

Example 1

Sodium Carbonate

Na CO_3

1 2

Na_2CO_3

Example 2

Aluminium Sulphate

Al SO_4

3 2

$Al_2(SO_4)_3$

Chemical Reaction

The interaction between elements and compounds to produce new species is called a **chemical reaction.** In a chemical change, also called a chemical reaction, one or more **reactants** (starting materials) are converted into one or more **products.** Chemical reactions occur all around us. They fuel and keep alive the cells of living tissues; they occur when we light a match, cook dinner, start a car listen to a portable radio, or watch television. There are several types of chemical reactions which are normally represented by a chemical equation. Some types of chemical reactions are discussed below.

Chemical Equations

A chemical equation is an equation which expresses in symbols the various elements or compounds reacting to form new substances. Chemical equation is the statement, in formulae, that expresses the identities and quantities of the substances involved in a chemical or physical change. It is a short hand notation for chemical reactions. When chemical reactions happen, atoms are not created or destroyed. This means that when you write a chemical equation, there must be the same number of each type of atom present in the reactant and the products.

Illustration of chemical equation

$$A + B \longrightarrow C + D$$

Reactants Products

Chemical equations give us information about:

1. Reactants and products
2. The physical states of reactants
3. The feasibility of the reaction
4. The solvent and the experimental conditions

Word Equations

A word equation should state the reactants (starting materials), products (ending materials), and direction of the reaction in a form that could be used to write a chemical equation.

Examples

Hydrogen gas + Oxygen gas → Water vapour (word equation)

$H_2 (g) + O_2(g) \rightarrow 2\ H_2O(g)$ (chemical reaction)

Copper + Oxygen → Copper(II) oxide (word equation)

$Cu + O_2 \rightarrow 2CuO$ (chemical reaction)

Balancing Chemical Equations

A balanced equation is an equation in which the total number of atoms of each element on both sides of the equation is equal. A chemical equation can be balanced either by inspection or by equation. The inspection involves looking through the equation and applying the necessary number of moles needed to balance the atoms of the various elements in the equation. Consider the reaction below:

H_2SO_4 + NaOH → Na_2SO_4 + H_2O (unbalanced equation)

If you look through the above equation, you will find that Na is 1 on the left hand side but 2 on the right hand side.

Therefore, we need to put 2 before NaOH to get 2NaOH.

This affects the number of oxygen and hydrogen,

H_2SO_4 + 2NaOH → Na_2SO_4 + H_2O.

Now the number of hydrogen atoms on the left side is 4 whiles there are only 2 hydrogen atoms on the right side. This requires that we apply 2 to H_2O to obtain $2H_2O$. The oxygen is also balanced. The new balanced equation becomes:

- ❑ H_2SO_4 + 2NaOH → Na_2SO_4 + $2H_2O$

Steps in Balancing Chemical Equations

1. Write out the names of the compounds first and replace them by their chemical formulae or symbols.
2. Balance any atom, except oxygen and hydrogen, by changing the coefficient of the compound and not the composition.
3. Finally, balance oxygen and hydrogen atoms.

Example 1

$$Lithium + Oxygen \rightarrow Lithium\ oxide$$

$$Li \quad + \quad O_2 \quad \rightarrow Li_2O \text{ (unbalanced equation)}$$

Li = 1, O = 2 Li = 2, O = 1

Li=1×2=2×2=4, O = 2 Li = 2×2 = 4, O = 1 ×2 = 2

$$4Li \quad + \quad O_2 \quad \rightarrow \quad 2Li_2O \text{ (balanced equation)}$$

Now, Li = 4, O = 2 Li = 4, O = 2

Example 2

$2HCl_{(aq)} + Na_2CO_{3(s)} \qquad 2NaCl_{(s)} + H_2O_{(l)} + CO_{2(g)}$

NB: Use the symbols (g = gas; l = liquid; s = solid; aq = in water solution) to denote the state substance in the equation.

The above equation is balanced because the total number of atoms of each element on the left hand side is equal to the total number of atoms of the same element on the right hand side. You may check for the number of atoms of each kind on both sides of the equation.

Types of Chemical Reactions

These are some general classifications of chemical reactions.

1. Combination/addition/synthesis reactions
2. Thermal decomposition reaction
3. Double decomposition reaction
4. Displacement/replacement reactions
5. Reversible reaction
6. Neutralization reaction
7. Oxidation-Reduction/Redox reaction

Combination/Addition/Synthesis Reactions

A combination reaction (synthesis reaction), is a reaction in which two or more substances combine to form a single new substance.

The general form of a combination reaction is:
A+B→AB
One combination reaction is two elements combining to form a compound.

Examples

1. Solid sodium metal reacts with chlorine gas to product solid sodium chloride.

$2Na_{(s)}+Cl_{2(g)} \rightarrow 2NaCl$

2. Reaction between iron and sulphur to form iron (II) sulphide

$$\boldsymbol{Fe + S \rightarrow FeS}$$

3. Reaction between calcium oxide and water to form calcium hydroxide

$$\boldsymbol{CaO + H_2O \rightarrow Ca(OH)_2}$$

4. Reaction between carbon dioxide and water to form carbonic acid

$$\boldsymbol{CO_2 + H_2O \rightarrow H_2CO_3}$$

Thermal Decomposition reaction

A decomposition reaction is a reaction in which a compound breaks down into two or more simpler substances. The general formula of this reaction is:

$AB \rightarrow A+B$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas.

$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

A reaction is also considered to be a decomposition reaction even when one or more of the products is still a compound. For instance, a metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Also, a metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water.

$2NaOH(s) \rightarrow Na_2O(s) + H_2O(g)$

Single Displacement/Replacement Reaction

This is a reaction in which one element replaces a similar element in a compound. The general formula of a single-displacement (also called single-replacement or *metathesis reaction*) reaction

is: $A + BC \rightarrow AC + B$

Example

$Fe(s) + CuSO_{4\,(aq)} \rightarrow FeSO_{4\,(aq)} + Cu_{(S)}$

Here, both iron and copper have the same valence.
One metal cation takes the place of the other bonding to the sulphate anion.
One metal cation takes the place of the other bonding to the sulphate anion.
Many metals react easily with acids, and, when they do so, one of the products of the reaction is hydrogen gas.

Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen i.e.

$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

Double-Displacement Reactions

A double-displacement reaction is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds. The general form of a doublereplacement (also called double-replacement) reaction is:

$AB + CD \rightarrow AD + CB$

In order for a reaction to occur, one of the products is usually a solid precipitate, a gas, or a molecular compound such as water. A precipitate forms in a double-replacement reaction when the cations from one of the reactants combine with the anions from the other reactant to form an insoluble ionic compound. When aqueous solutions of potassium iodide and lead (II) nitrate are mixed, the following reaction occurs.

$2KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2KNO_{3(aq)} + PbI_{2(s)}$

Combustion Reactions

A combustion reaction is a reaction in which a substance reacts with oxygen gas, releasing energy in the form of light and heat. Combustion reactions must involve O_2 as one reactant. The combustion of hydrogen gas produces water vapor (see figure below).

$2H_2(g)+O_2(g)+2H_2O(g)$

Many combustion reactions occur with a hydrocarbon, a compound made up solely of carbon and hydrogen. The products of the combustion of hydrocarbons are always carbon dioxide and water. Many hydrocarbons are used as fuel because their combustion releases very large amount of heat energy. Propane (C_3H_8) is a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$C_3H_8\ (g) + 5O_2\ (g) \rightarrow 3CO_2\ (g) + 4H_2O\ (g)$

Oxidation and Reduction Reaction

An oxidation reaction is the one in which oxygen is added to a substance. It is also a reaction in which hydrogen is removed from a substance. Reduction is the addition of hydrogen to a substance or the removal of oxygen from a substance.

Consider the reaction:

$CuO+H_2 \rightarrow Cu+H_2O.$

The copper (II) oxide (CuO) is losing oxygen. It is reduced. The hydrogen is gaining oxygen. It is being oxidized.

If a substance loses oxygen during a reaction, it is reduced. If a substance gains oxygen during a reaction, it is oxidized.

Reduction and oxidation always take place together in a reaction. So the reaction is called a Redox reaction.

In the above reaction, the copper (II) oxide is described as the oxidizing agent and the hydrogen is the reducing agent.

An oxidizing agent is a substance which transfers oxygen to another substance or removes hydrogen from that substance.

A reducing agent is a substance which transfers hydrogen to another substance or removes oxygen from that substance.

Neutralization Reaction

It is the reaction between an acid and a base to produce **salt and water only**. It is called neutralization because in the reaction, both the acid and the base properties are lost.

Examples:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

$$2HNO_3 + CaO \rightarrow Ca(NO_3)_2 + H_2O$$

$$H_2SO_4 + 2KOH \rightarrow K_2SO_4 + H_2O$$

Reversible and Irreversible Reaction

Certain reactions take place only in the forward reaction, they continue until one reactant is completely used up.

E.g. burning of wood or $2HCl + Zn \rightarrow ZnCl_2 + H_2$.

These reactions are known as **irreversible reactions** and they are denoted by the symbol $\rightarrow$

Some other reactions can however occur in both backward and forward reactions. In these reactions none of the reactant is completely used up so that the reaction does not go in either direction depending on the conditions.

To show that a reaction is reversible, the symbol $\rightleftharpoons$ is used in the equation. A **reversible reaction** is a reaction where the reactants and products react together to give the reactants back.

PURE AND IMPURE SUBSTANCES

Pure Substances

A pure substance is a substance that is made up of just one chemical element or compound. A pure substance contains only one kind of particle throughout. Only a few pure substances can be found in nature. Almost all pure substances we use have been made pure by humans. We take the raw material that contains them and separate out the substance we want. E.g. table sugar, salt, aluminum foil.

Properties of Pure Substance

1. It has uniform composition and properties in all its parts.
2. It has certain characteristics such as specific temperatures at which it melts and boils.
3. The constituents cannot be separated by physical means.

Mixtures

A mixture is a physical combination of two or more substances which do not react chemically. A mixture is a physical combination of two or more substances that can be separated or reversed by physical means. E.g. Air, soil, sea water, smoke, soft drinks, milk, blood, alloy etc. The ease with which you can separate a mixture depends on the degree to which the components have been mixed.

Points to Note:

1. The substances that form a mixture are held together by physical forces.
2. Components of a mixture can be separated.
3. Components of a mixture do not have a constant composition.
4. Mixture may be solids, liquids and gases or a combination of any of the three states of matter.

Examples of mixtures

Mixtures	Compounds
Air (mixture of gases)	N_2 (78%), O_2 (21%), water vapour (varies), CO_2 (0.036%), Inert gases (0.09%).
Petroleum (mixture of liquids)	Refinery gas, gasoline (petrol), kerosene (or paraffin oil), diesel oil, lubricating oil, bitumen (or asphalt).
Brine (salt solution)	Salt and water
Brass, alloy (mixture of metals)	Copper and zinc

Types of Mixtures

Basically, there are three types of mixtures depending on their states. These include;

1. Solid mixtures (e.g. Alloy)
2. Liquid mixtures (e.g. Blood, soft drink)
3. Gaseous mixtures (e.g. Air)

Classification of Mixtures

Different types of mixtures can be obtained by combining any two of the states of matter. These include;

1. **Solid – Solid mixture:** It is a mixture of two solid substances. e.g. sugar and sand, Sulphur and iron fillings, Bronze (Copper + Tin), Steel (Iron + Carbon), Brass (Copper + Zinc), Duralumin (Aliminium + Copper), etc.

   Features: One may be soluble and the other insoluble, both may be soluble or insoluble, one may be insoluble and the other sublime.

2. **Liquid – liquid mixture**: It is a mixture of two different liquid substances. e.g. oil and water, kerosene and oil, water and kerosene/petrol, Alcoholic drinks (alcohol + water + flavourings), Cough mixtures (honey + fruit juices + some medicine), Crude oil, etc.

   Feature: They may be either miscible or immiscible.

3. **Solid – liquid mixture:** It is a mixture of solid and liquid substances. e.g. sugar and water, Gari and water, Milk and sugar, Chalk particles and water, etc.

   Features: The solid may be soluble or insoluble in the liquid. The solid may form suspension or colloid in the liquid.

4. **Solid – Gas mixture**: It is a mixture of solid and gaseous substances. e.g. dust (fine sand particles mixed with air), harmattan wind, smoke (soot + air).

5. **Liquid – Gas mixture**: It is a mixture of liquid and gaseous substances. e.g. mist or fog (water droplets mixed with air), carbon dioxide in soft drinks.

6. **Gas – Gas mixture**: It is a mixture of gaseous substances. e.g. air.

Homogeneous Mixture

A homogeneous mixture is any mixture that is uniform in composition throughout the mixture. The following are some everyday examples of homogeneous mixture.

1. Natural water itself is an example of homogeneous mixture. It often contains dissolved minerals and gases, but these are dissolved throughout the water.
2. Blood plasma is a homogeneous mixture.
3. A cup of coffee and sugar is considered homogeneous since the sugar dissolves completely throughout.
4. The air we breathe is a homogeneous mixture of oxygen, nitrogen, carbon dioxide along with other components in smaller quantities.

Heterogeneous Mixture

A mixture is said to be heterogeneous when we can see two or more phases separated by boundaries. The phases might be a solid and a liquid or two liquids or other combinations of solid, liquid and gases. The components of heterogeneous mixtures can most often be separated by simple mechanical processes such as filtration or decantation.
Examples are: water and kerosene, water and palm oil, gari in excess amount of water etc.

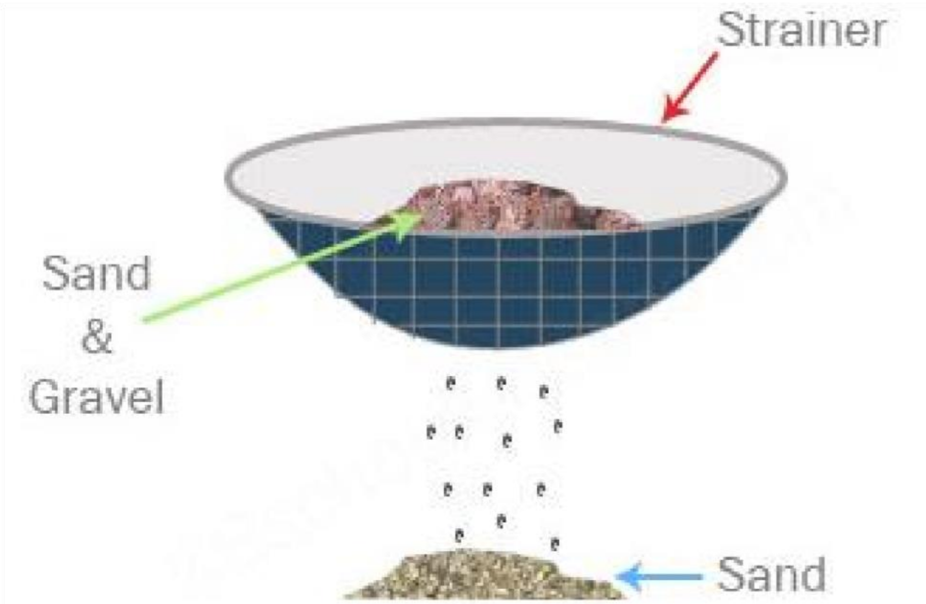
Emulsion : It is a mixture of two or more liquids in which one is present as droplets of microscopic size distributed throughout the other. Emulsions are formed from the component liquids either spontaneously or more often by mechanical means such as agitation, provided the liquids have no mutual solubility. Example is milk (a dispersion of fat droplets in aqueous solution).

Suspension: This is a solid and liquid mixture, in which the solid particles are insoluble hence remain dispersed visibly in the liquid medium. Example: chalk powder in water, charcoal powder in water, etc.

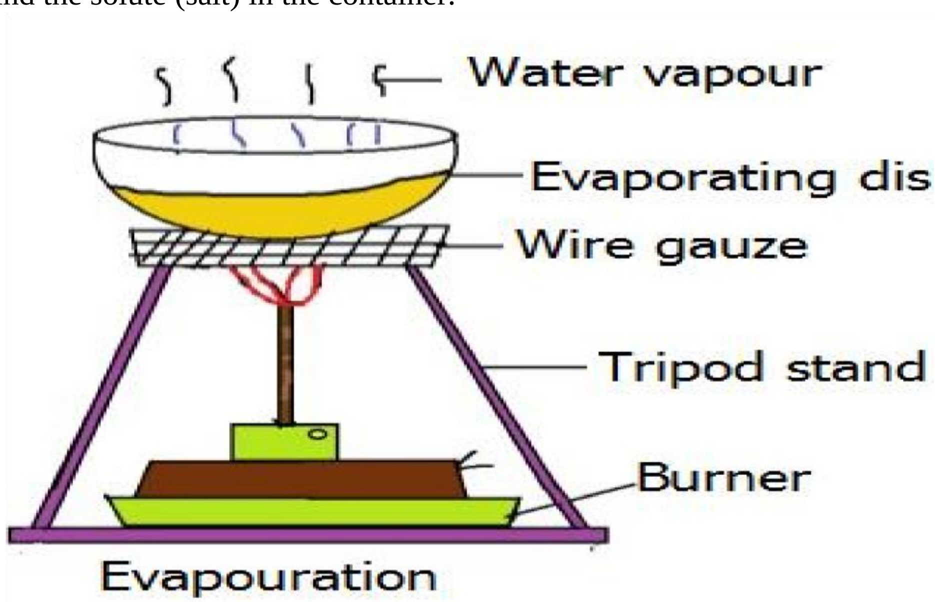
Methods of Separating Mixtures

The components of mixtures retain their individual properties. The methods of purification make use of physical properties of the chemicals (substances) involved. These include the boiling point and solubility in different solvent. These characteristics of mixture can be employed to separate mixtures. The following are methods that can be used to separate mixtures: Sieving, filtration, evaporation, distillation, chromatography, fractional distillation, crystallization, Sublimation, Magnetization and use of separation funnel.

Sieving and filtration are physical mechanisms of particle removal, where a particle is denied access through a pore or passageway that is smaller than the particle itself. The smaller particles pass through the mesh of the sieve and the larger particles are collected on the mesh of the sieve.
For example, the larger particles of powdered kokonte or gari can be separated from the smaller particles by sieving. The process is illustrated below

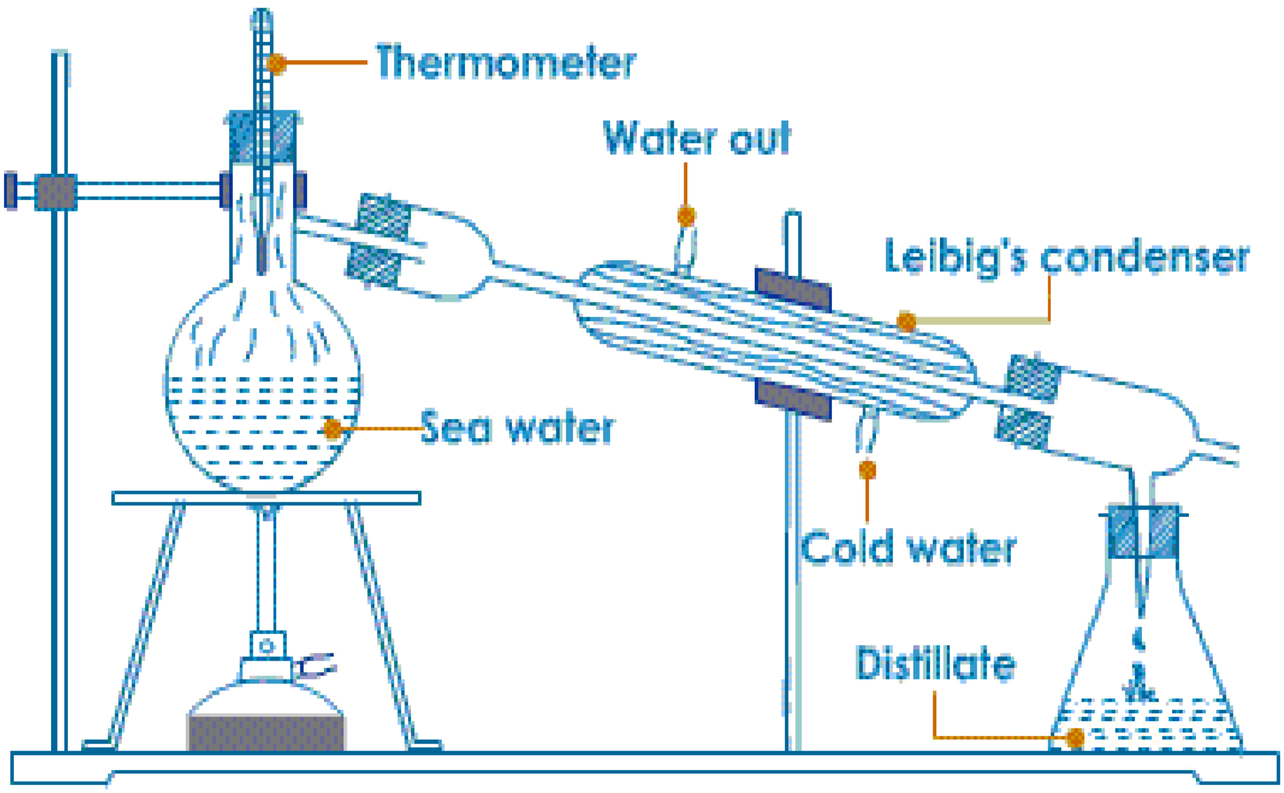


Evaporation: Is a method of separating solid solute from a solvent by exposure to heat or the direct effect of the sun rays.
E.g. separating salt and water in a salt solution. In this case, water changes from liquid phase to gas leaving behind the solute (salt) in the container.



Simple distillation: This is a method used to separate two miscible liquid of different boiling points. The distillation process involves boiling (evaporating) the liquid and cooling (condensing) the vapour back into the pure liquid.

Examples: ethanol from water is made possible because ethanol boils at a much lower temperature (78 ^oC) than water (100 ^oC) and is vapourised before the boiling point of water is reached.



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Fractional Distillation: Is a method of separating two or more miscible liquids with different boiling points. This method is normally used to separate a mixture of liquids with different boiling point which are close or multicomponent mixture of liquids such as crude oil. The Tema Oil Refinery operates on the principle of fractional distillation to give the kerosene, petrol, diesel, etc

The **distillation setup** is similar to that of the simple distillation except a fractionating column is added which responsible for the separation of the various components. Examples of mixtures that can be separated by fractional distillation are; petroleum (crude oil) and liquid air

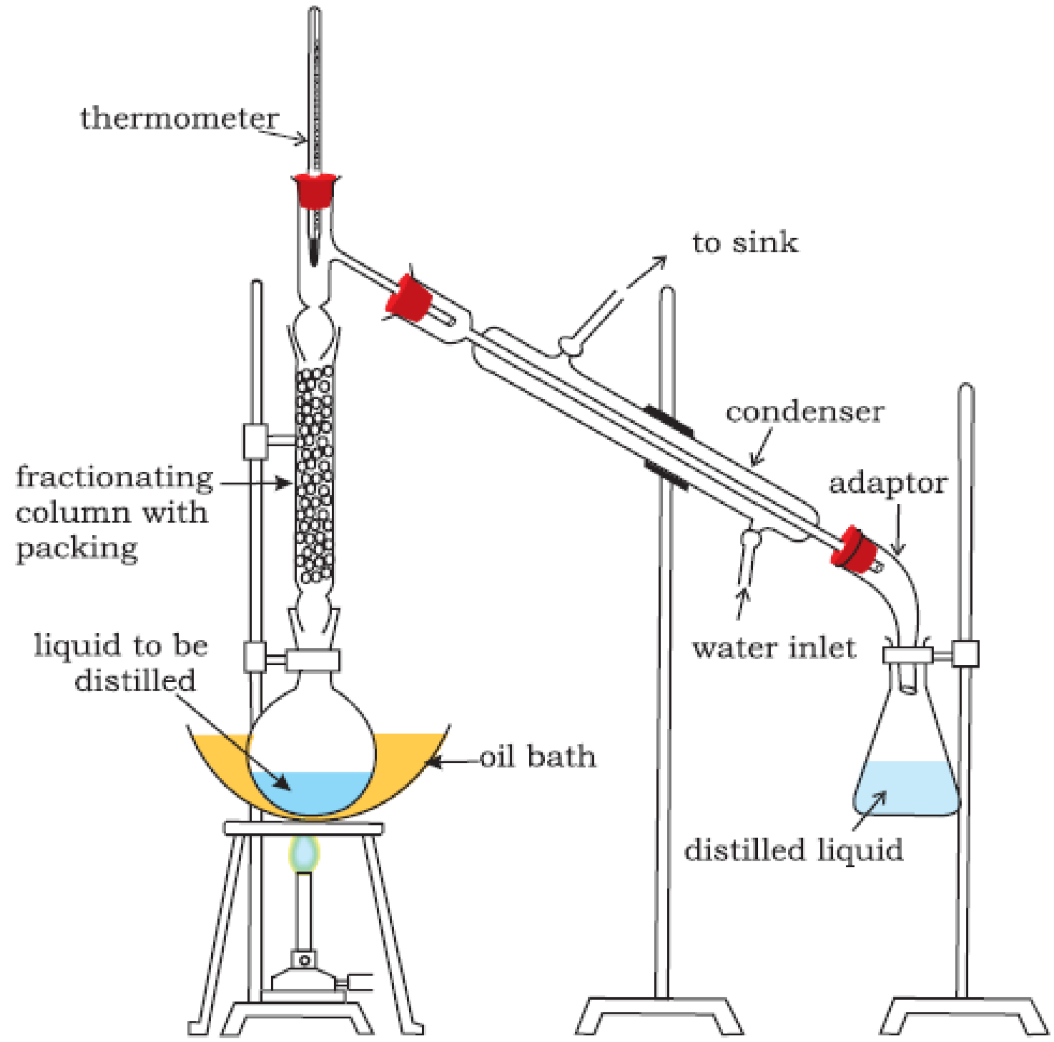
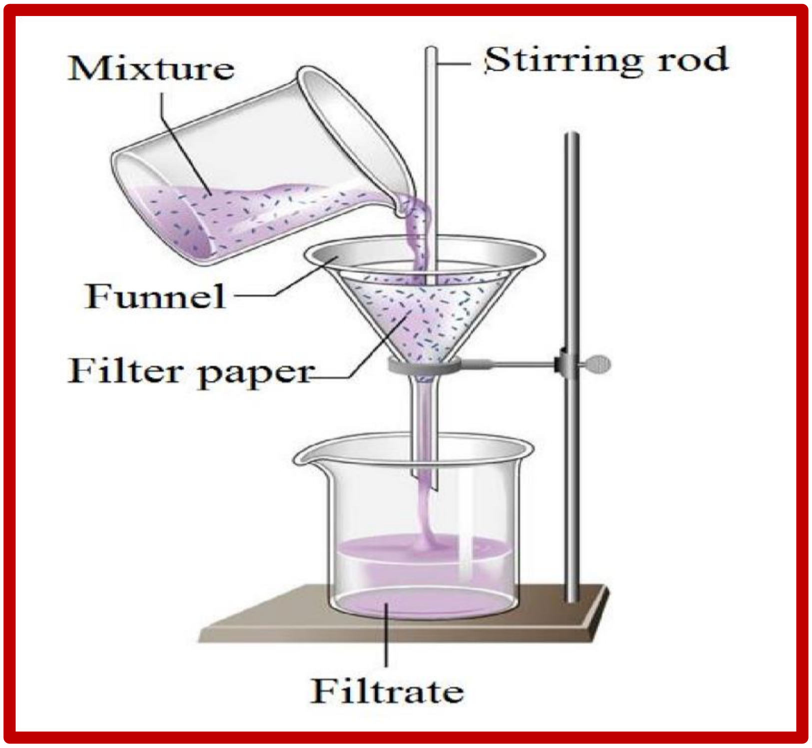
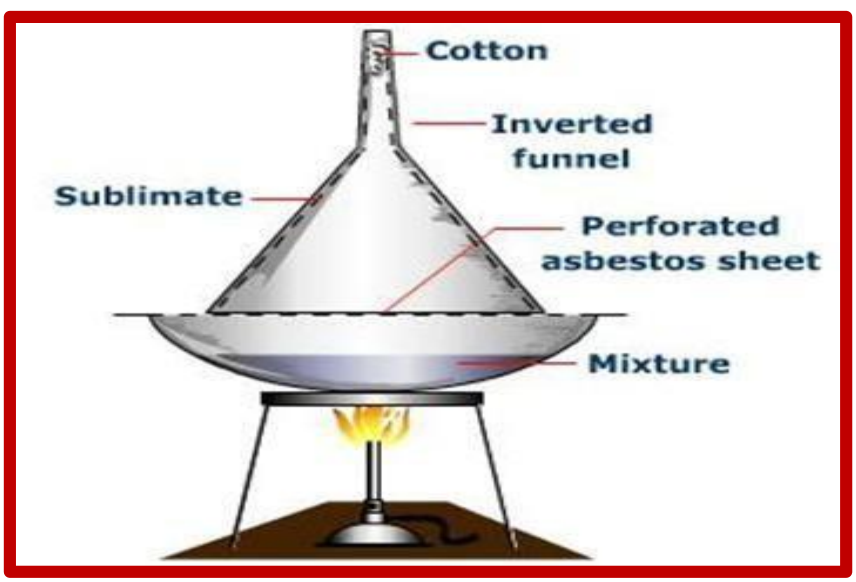


Fig.12.6 Fractional distillation. The vapours of lower boiling fraction reach the top of the column first followed by vapours of higher boiling fractions.

Filtration is a separating method used to separate insoluble substances from a solution by passing the mixture through a porous material such as filter paper. Examples: Separating fine sand particles from water, Water and powdered Chalk.

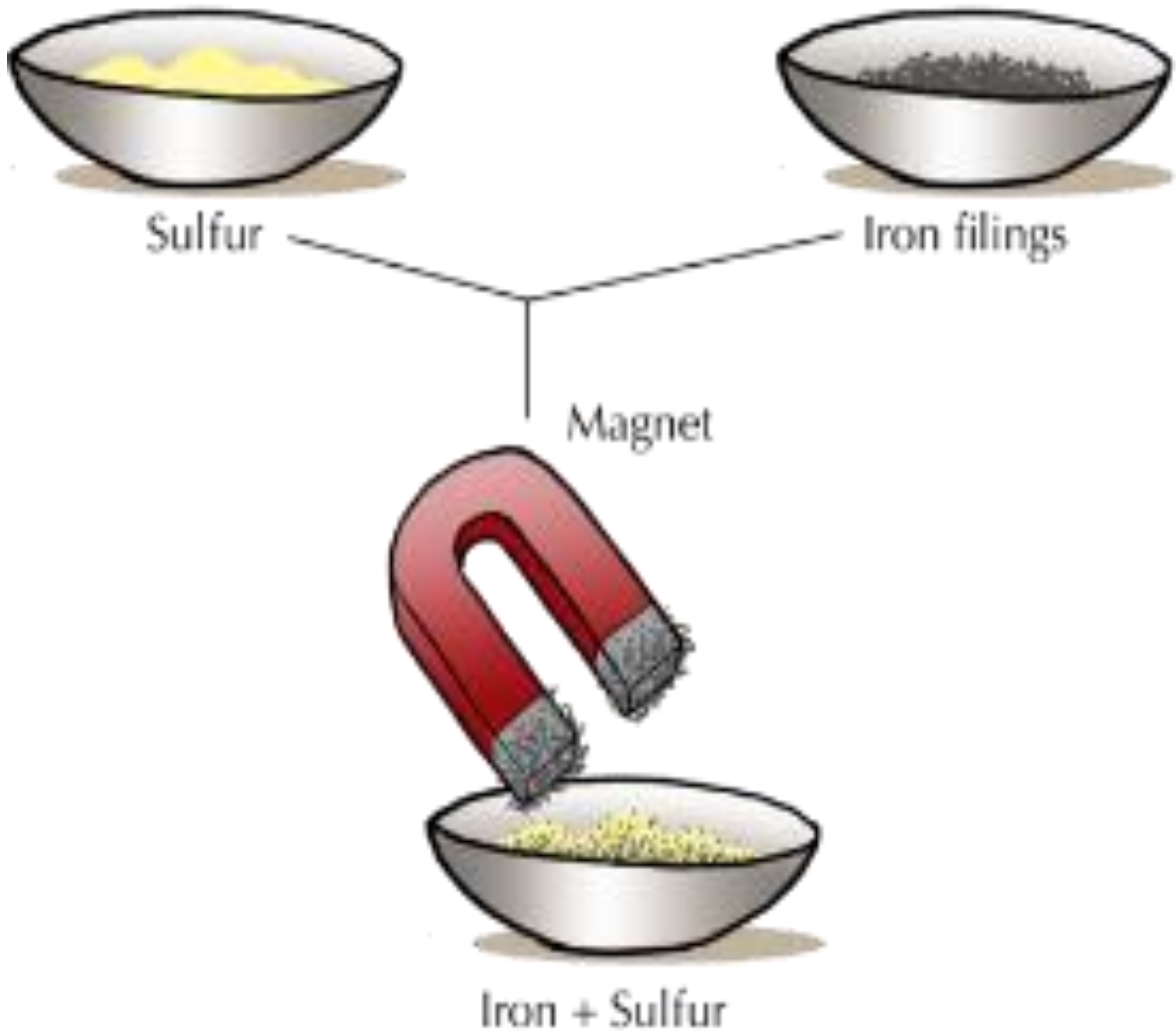


Sublimation: is the direct change of state from a solid to a gas (vapour) on heating or from gas to solid on cooling. This method is used to separate a solid which sublimes from solid which does not sublime. Examples of substances that sublime are: iodine, solid carbon dioxide (dry ice), ammonium chloride, camphor, naphthalene ball. For example if a mixture of ammonium chloride and sodium chloride is heated, the ammonium chloride turns directly to vapour but the sodium chloride remains unchanged.



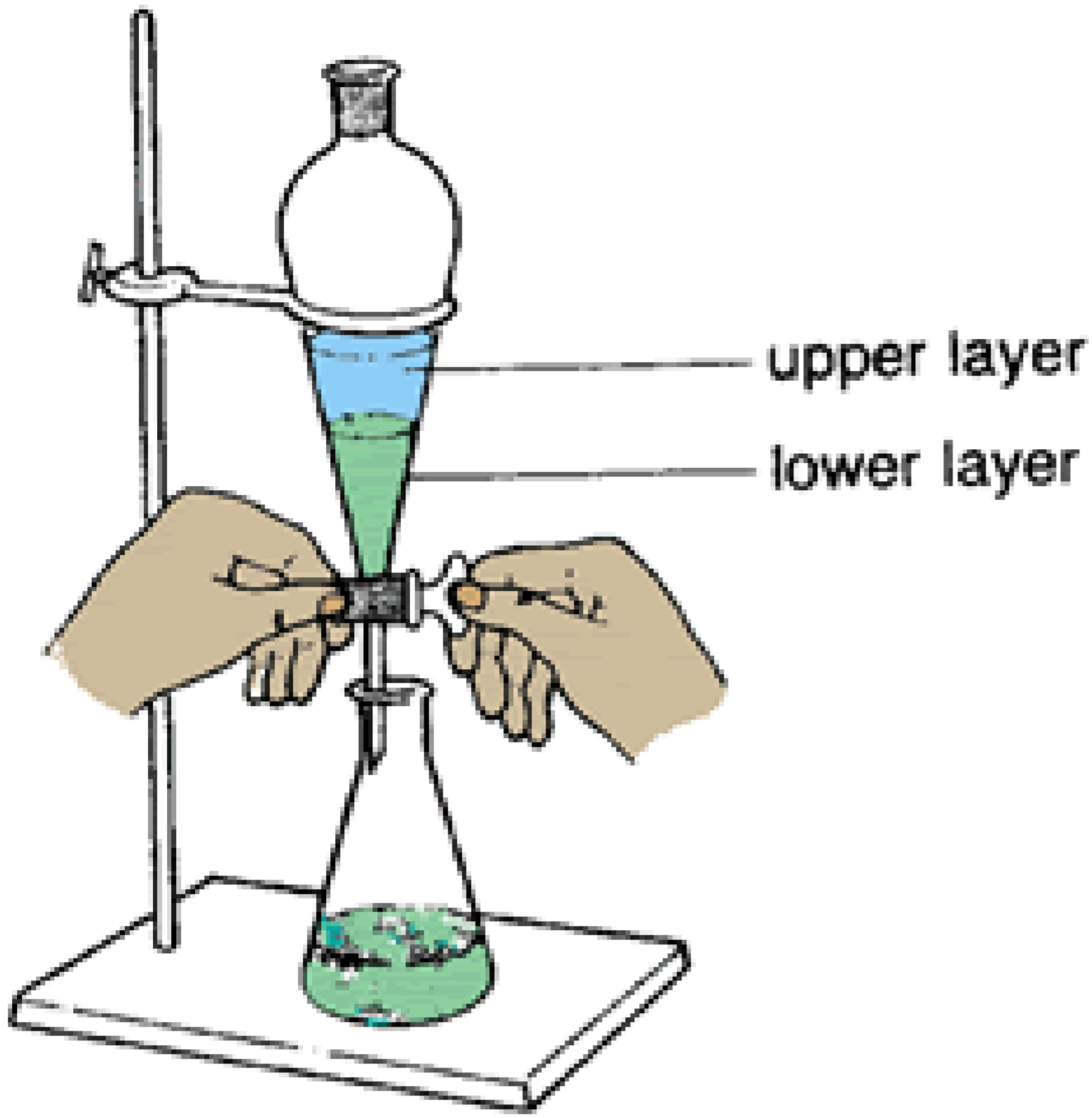
Magnetization

This method is used to separate a mixture of a non-magnetic substance and a magnetic substance. For example, a mixture of iron fillings and sulphur.



Use of Separating Funnel

A separating funnel is a funnel with a tap used to separates a mixture of immiscible liquids (such as oil and water). Two liquids are said to be immiscible when they form two layers in a container after being put together and left to stand a while. The lighter liquid (e.g oil) collects above the heavier liquid (water). When the tap is opened the water is run out, but the tap is closed before the oil reaches the bottom.

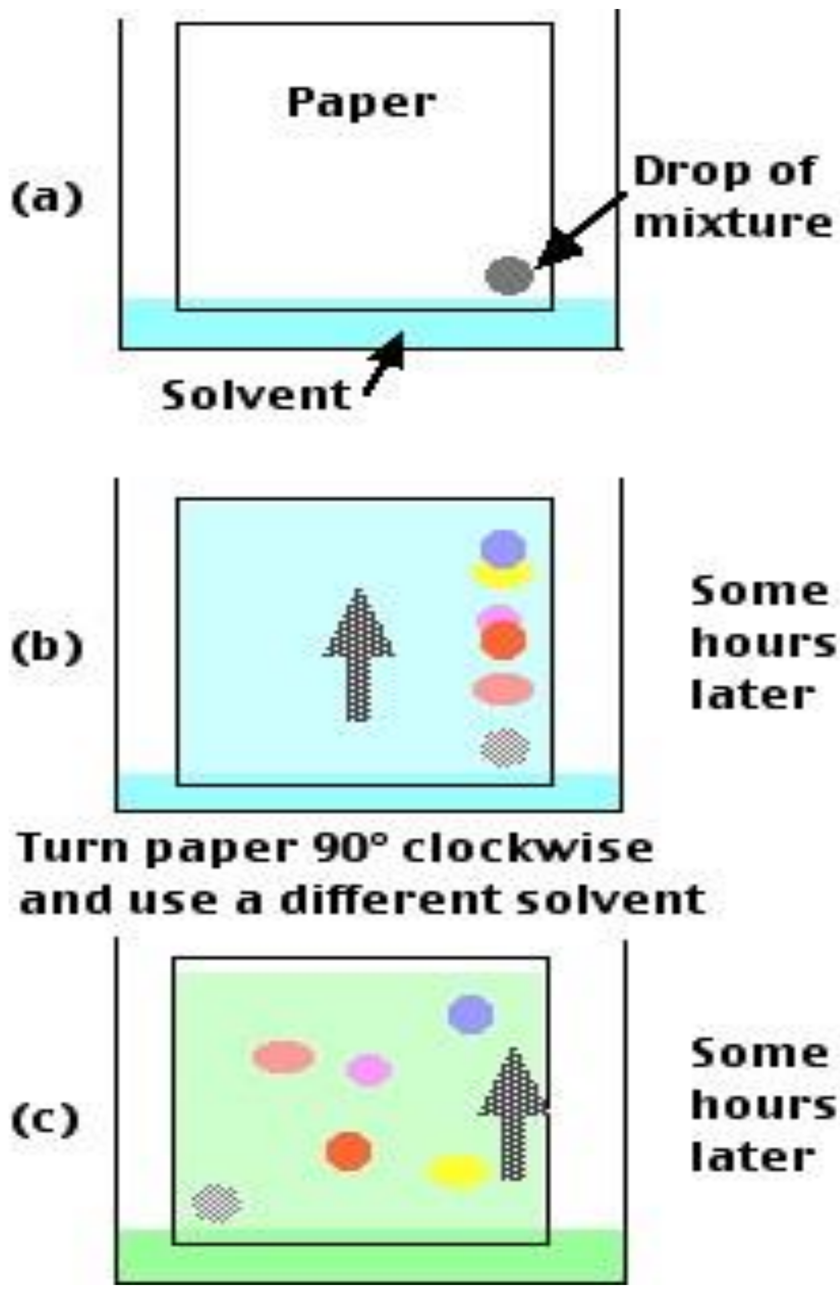


Chromatography

It is a method of separating a mixture by taken advantages of different rates of movement in a solvent over an absorbent material. Chromatography is usually used in separating organic compounds in a liquid mixture.

Steps involved

1. Spot a drop of black ink at the centre of a paper.
2. Suspend the tip of the filter paper by means of sliced cork in a boiling tube such that its end is dipping into some ethanol solvent in the tube.
3. Now set apart the apparatus for some few hours.



Centrifuging (use of centrifuge)

A centrifuge is an apparatus for the separation of substances by rotating them in a tube in a horizontal circle at high speed. It can be used to separate fine insoluble particles in a liquid suspension or denser liquid from less ones. Centrifuge is used to separate blood cells from blood plasma and cream from milk.

ACIDS, BASES AND SALTS

DEFINITION OF ACIDS AND BASES

Introduction

In concept of acids and bases has evolved over time with various definitions given to them at any particular time. These definitions mostly could not cover the entire concept and has to be replaced. We discuss some of these definitions in the next few sections.

Arrhenius Acid

An acid is a compound, which dissolves in water to produce hydrogen ions (H^+) or hydroxonium ions (H_3O^+).

Example 1: $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

In the example 1 above HCl in aqueous solution produced hydrogen ion (H^+). Therefore, it is an Arrhenius' acid.

Example 2:

$$HCl\,(l) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-\,(aq)$$

$$H_2SO_4\,(l) + 2H_2O\,(l) \rightarrow 2H_3O^+\,(aq) + SO_4^{2-}(aq)$$

In the example 2 above, HCl and H_2SO_4 are both Arrhenius acids because each produces hydroxonium ion in aqueous solution.

Arrhenius Base

A base is any substances which produces hydroxyl (OH^-) ions in aqueous solution.

Examples: $NaOH_{(aq)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$

$Mg(OH)_2 \rightarrow Mg^{2+} + OH^{-}$

NH3(l) + H2O(l) → NH4 (aq) + OH-

In the examples above: NaOH, $Mg(OH)_2$, and NH_3 are Arrhenius Bases because each produces hydroxyl (OH^-) ions in aqueous solution.

Bronsted- Lowry Acid

An acid is proton donor or an acid is any substance that donates proton.

Examples: $\mathbf{HCl} + H_2O \rightleftharpoons H_3O^+ + Cl^-$

$$\mathbf{H_sSO_4} + SO_4^{2-} \leftrightarrows HSO_4^- + HSO_4^-$$

In the examples above HCl and H_2SO_4 are both Bronsted-Lowry's acids because each donates a proton (H^+) to the H_2O and SO_4^{2-} respectfully.

Bronsted- Lowry Base

A Base is any substance that accept a proton

Examples:$CH_3COOH_{(l)}$ + $\mathbf{H_2O_{(l)}}$$\rightleftharpoons \mathrm{CH3COO - (aq) \ + \ H3O + (aq)}$

$$\mathrm{H2SO4(aq)} \ + \ \mathbf{SO42 - (aq)} \leftrightarrows \mathrm{HSO4 - (aq) \ + \ HSO4 - (aq)}$$

H_2O and SO_4^{2-} are Bronsted-Lowry bases because both accept a proton from respective acids.

Lewis Acid

An acid is any substance that accepts electron pair to form a coordinate covalent bond.

Example: NH3 (l) + **H +**(aq) → NH4+(aq)

In the example above, H^+ is the Lewis acid since it accepts electrons from the NH_3 to form NH_4^+

Other examples of Lewis acids are:

1. all cations, especially transition metals: Mn^{2+}, Fe^{3+}, $Zn^{2.}$
2. Molecules with empty orbitals e.g. orbitals: BeF_2, BCl_3, $AlBr_3$

Lewis Base

Lewis Bases is any substance that donates electron pair to form coordinate covalent bond.

Example: $H_2O_{(l)} + H^{+}_{(aq)} \rightarrow H_3O^{+}_{(aq)}$

In the example above, H_2O is a Lewis base because it donates electron pair to H^+ form H_3O^+

Other examples are:

H_2O, SO_2, NH_3, CH_3NH_2, Cl^-, S^{2-}, Br^-, O^{2-} etc.

Physical Properties of Acids

1. It is colourless
2. It has a sharp/sour taste
3. It is very corrosive
4. Acid turns blue litmus paper to red
5. It conducts electricity when molten/aqueous
6. Smooth in between fingers

Chemical Properties of Acids

1. Acid reacts with alkalis (bases) to form salt and water

   Example: $HCl + NaOH \rightarrow NaCl + H_2O$

   Acid base Salt water

2. Acid reacts with trioxocarbonate (iv) (CO_3^{2-}) to produce carbon dioxide gas (CO_2).

   Example: $2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$

3. Acid reacts with reactive metals to produce hydrogen gas.

   Example: $2\ HCl_{(aq)} + Ca_{(s)} \rightarrow CaCl_{2(aq)} + H_{2\ (g)}$

4. Acids react with metals based on the reactivity series of metals.

Neutralization Reaction

It is a reaction between an **acid** and a **base** to form **salt and water only**.

Acid + Base → Salt + water

Neutralization reaction is usually accompanied by heat. Strong acids or bases usually give out more heat than weak acids and the bases in the reaction.

Examples of Neutralization reaction

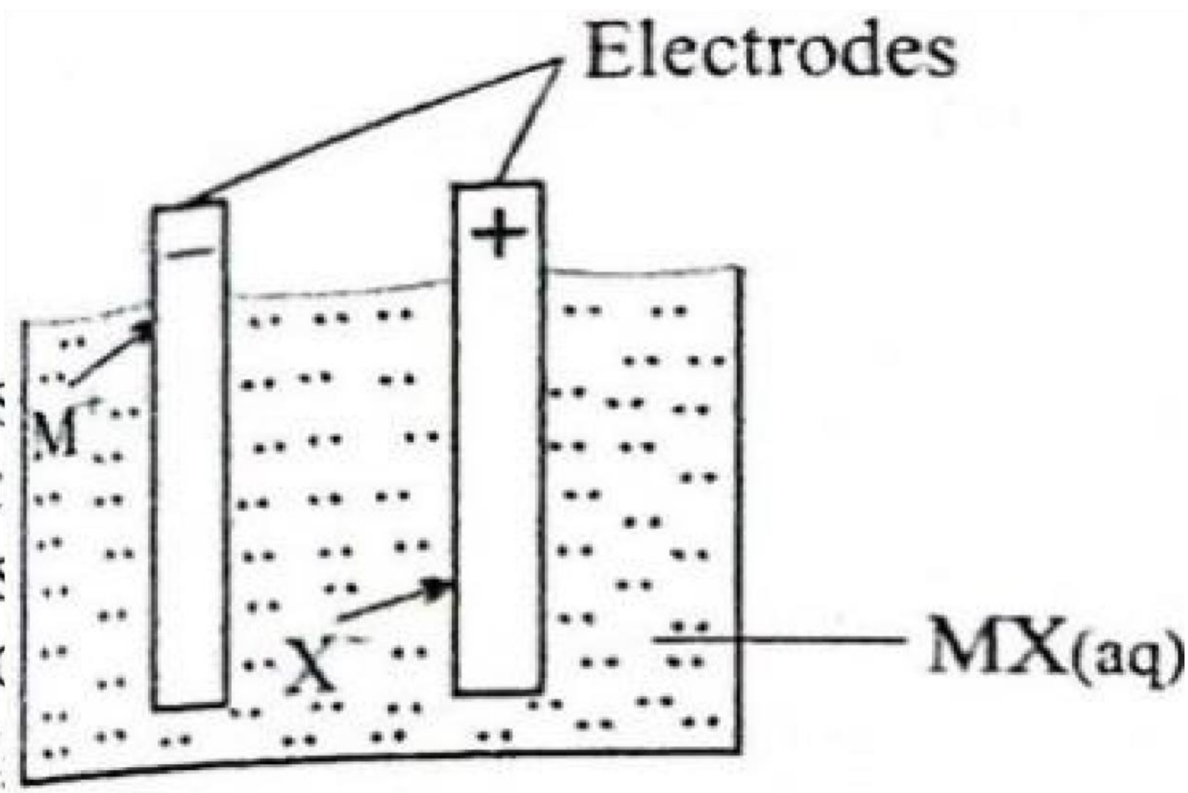
$$HCl(aq) + NaOH(aq) \rightarrow NaCl\ (aq) + H2O(l)$$

H_2SO_4 (aq) +2NaOH $\rightarrow$ $NaSO_4$ (aq) + 2H_2O (l)

H^+ (aq) + OH^-(aq) $\rightarrow$ H_2O (l)

Acids, Bases and Salts as Electrolytes

An electrolyte is a substance which conducts electricity either in aqueous solution or in molten form and undergoes a chemical change in the process. Ions are responsible for the conduction of electricity in aqueous solutions or molten forms.



The

H_2S

HC

CuS

NaC

CH

$$NH_3(l) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

electrolytes therefore ionize completely or partially and the ions produced migrate to the appropriate poles of a potential difference put across the solution. The negative ions migrate to the positive pole and the positive ions migrate to the negative pole. In this way, current flows through the system. The poles are called **electrodes**. The ions gain or lose electrons at electrodes by a process called **electrolysis**.

The Strength of an Acid

It is a measure of the ability of the acid to ionize in solution to conduct electricity. Strong acids conduct electricity strongly because they ionize in solution to produce greater number of ions.

Weak acids conduct electricity weakly because they ionize in solution to produce few ions.

Classification of Acids

Weak Acid:

It is a type of acid which partially dissociates or ionizes in aqueous solution to produce few ions.

OR

It is a type of acid which does not ionize or dissociates completely in aqueous solution.

Some of the molecules in weak acid remain unionized in the aqueous solution. Weak acids are usually organic acids

Examples of weak acid includes: ethanoic acid, methanoic acid, citric acid trioxocarbonate (IV) acid (H_2CO_3), nitrous acid (HNO_2) etc.

Most weak acids exists largely as molecules and their ionization is reversible

- $CH_3COOH_{(aq)} \leftrightharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$

- $CH_3CH_2COOH_{(aq)} \leftrightharpoons CH_3CH_2COO^-_{(aq)} + H^+_{(aq)}$

In water they dissociates as follows:

- $HNO_{2\,(aq)} \leftrightharpoons H^+_{(aq)} + NO_2^-{}_{(aq)}$

- $H_2CO_{3\,(aq)} \leftrightharpoons 2H^+ + CO_3^{2-}{}_{(aq)}$

Strong Acids:

It is a type of acid which completely dissociates or ionizes in solution to produce greater number of ions. OR

It is a type of acid which completely ionizes or dissociates in aqueous solution.
Strong Acids are usually inorganic (mineral) acids. Strong acids are also referred to as mineral acid because they are obtained from mineral sources. Strong acid contains high concentration of hydrogen ions.

Examples of strong acid include:

Hydrochloric acid (HCl), Trioxonitrate (V) acid (HNO_3), Tetraoxosulphate (VI) acid (H_2SO_4) and Tetraoxophosphate (V) acid (H_3PO_4).

In water, they (strong acids) dissociates as follows:

$$HCl\ (aq) \rightarrow H+\ (aq) + Cl-(aq)$$

$$HNO3\ (aq) \rightarrow H+\ (aq) + NO3-(aq)$$

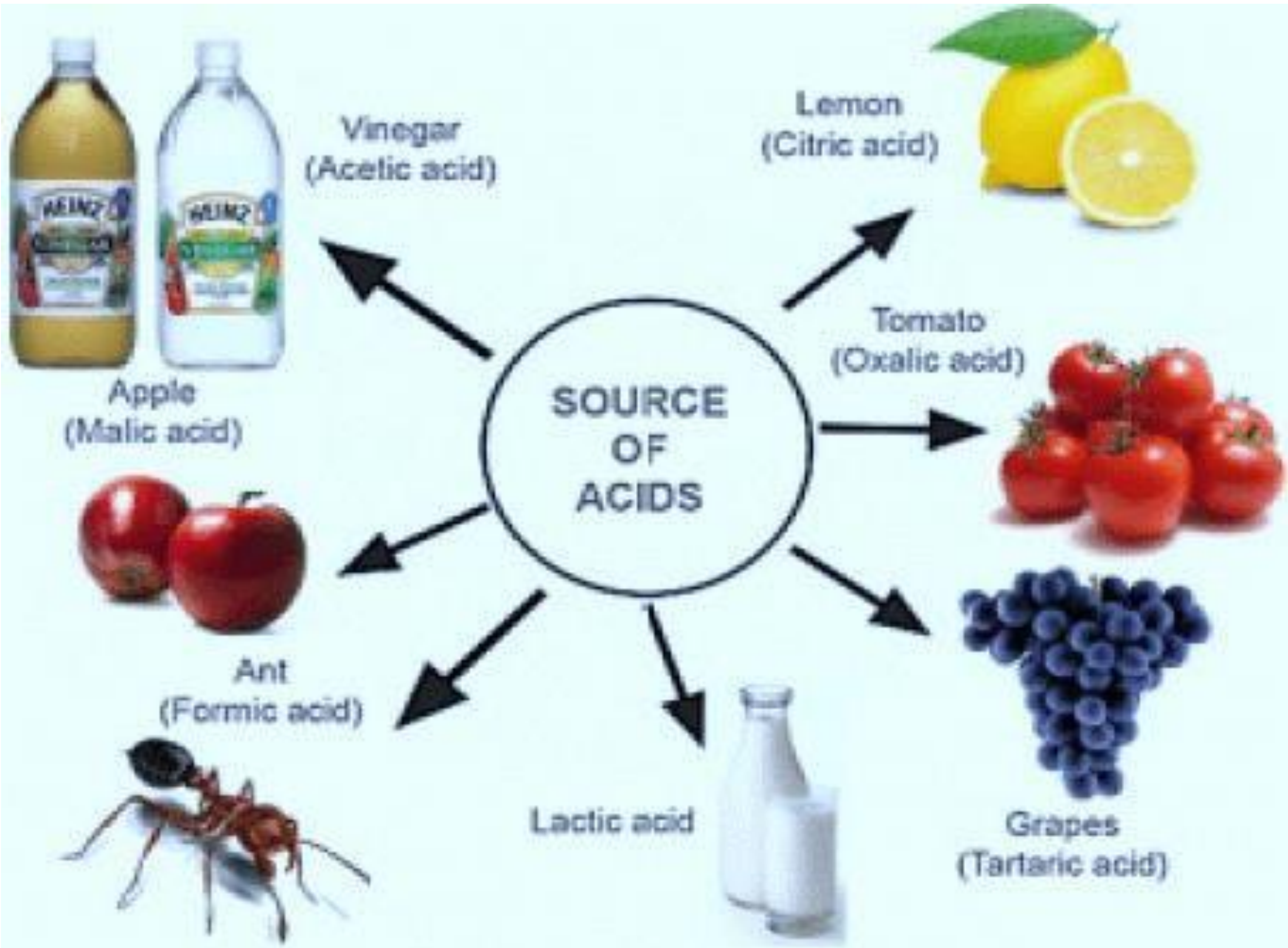
$$H_2SO_4\ (aq) \rightarrow 2H^+\ (aq) + SO_4^{2}\ (aq)$$

$$H_3PO_4\ (aq) \rightarrow\ 3H^+{}_{(}aq) + PO_4^{3-}{}_{(}aq)$$

Organic and Inorganic Acids

Organic Acids: Natural or organic acids are obtained in nature or from living things. They are usually weak acids. The table below shows some organic acids and their sources.

ACIDS	**SOURCES**
Methanoic acid	Ants, bees
Ethanoic acid	Vinegar, beer
(Acetic acid)	Palm wine
Citric acid	Citrus fruits
Ascobic acid **(Vitamin C)**	Green vegetables, Vitamin C tablets
Lactic acid	Sour milk



Acids and their sources

Inorganic Acids (Mineral Acid)

These are acids that originated from minerals sources. Inorganic acid are strong acid (except H_3PO_4, H_2CO_3 which are weak acid). Some examples are given in the table below.

FORMULA	COMMON NAME	IUPAC NAME
HCl	Hydrochloric Acid	Hydrochloric acid
H_2SO_4	Sulphuric acid	Tetraoxosulphate(VI) acid
HNO_3	Nitric acid	Trioxonitrate (V) acid
H_2CO_3	Carbonic acid	Trioxocabonate (IV)acid
H_3PO_4	Phosphoric Acid	Tetraoxophosphate (V)acid

The Basicity of an Acid

Basicity: It is the number of replaceable hydrogen ions in an acid that can be replaced by a metal. Generally: Acids with basicity of one are referred to as **monobasic acid.** Acids with basicity of two are called **dibasic acids**. Acids with basicity of three are called **tribasic acids.**

Acid	**Basicity**
HCl	1/(monobasic)
HNO_3	1/(monobasic)
CH_3COOH	1/(monobasic)
H_2SO_4	2/(dibasic)
H_2CO_3	2/(dibasic)
H_3PO_4	3/(tribasic)

Sources of Some Bases

The table below shows some common bases and their sources

Bases	**Sources**
Potassium hydroxide, KOH	Ashes of plants. eg. plantain peels
Ammonia, NH_3	Decomposed organic matter / Urine
Amine, RNH_2	Fish
Calcium Oxide, CaO	Heating limestone, $CaCO_3$
Iron(III) oxide, Fe_2O_3	Iron ore

Physical Properties of Bases

1. Some bases have bitter taste. e.g. NaOH, KOH
2. They are slippery in between fingers
3. Turn red litmus blue
4. Turn phenolphthalein indicator pink

5. Strong bases are good conductors of electricity

Chemical Properties of Bases

1. React with acid to form salt and water

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

2. They react with ammonium salt to give ammonia gas

$NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow NH_{3(g)} + NaCl_{(aq)} + H_2O_{(l)}$

The Strength of Bases

Strong bases are referred to as alkalis.Strong bases are alkalis that ionize completely in aqueous solution.

Examples of strong bases include: potassium hydroxide (KOH) and sodium hydroxide (NaOH)
They have typical pH of 13-14

Equations to illustrate the process of producing ions

$$NaOH_{(aq)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$
$$KOH(aq) \rightarrow K^{+}(aq) + OH^{-}(aq)$$

Weak bases: They are bases which are partially ionized in aqueous solution.
Examples of weak alkali / base include:

- calcium hydroxide $Ca(OH)_2$,
- magnesium hydroxide $Mg(OH)_2$,
- Ammonia NH_3,
- Ammonium hydroxide $NH_4(OH)$,
- Sodium hydrogen trioxocarbonate (IV) ($NaHCO_3$), urea.

MEANING OF SALT

A **salt** is defined as a substance formed when all or part of the replaceable hydrogen atoms in an acid is replaced by a metal or ammonium ion in a reaction. A **salt** is an ionic compound whose cation comes from a base and anion comes from an acid. A **salt** is a compound consisting of positive metallic ions derived from a base and negative ions derived from an acid

Examples of salts:

- Sodium chloride (NaCl)
- Potassium chloride (KCl)
- Zinc trioxonitrate (V) $Zn(NO_3)_2$
- Copper (II) tetraoxosulphate (VI) $CuSO_4$
- Iron (II) sulphide FeS
- Sodium hydrogen trioxocarbonate (IV) ($NaHCO_3$)
-
- Sodium nitrate or Sodium trioxonitrate(V) ($NaNO_3$)
- Calcium tetraoxosulphate (VI) or Calcium sulphate $CaSO_4$

Sources of Salts

Salts	Salts
Sodium chloride (NaCl)	Sea water, Rocks salts
Sodium trioxonitrate(V) ($NaNO_3$)	Saltpetre
Iron (II) sulphide FeS	Iron pyrites
Calcium trioxocarbonate (IV) $CaCO_3$	Chalk, Limestone, Marble
Potassium chloride (KCl)	Brine, Rock salts

Calcium tetraoxocarbonate (VI) $CaSO_4$	Gypsum

Types of Salts

Normal salt: it is a type of salt formed when **all replaceable** hydrogen ions of an acid are replaced by a metal. **Equations**: $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Examples of normal salts:

- Sodium chloride (NaCl)
- Potassium chloride KCl
- Potassium sulphate or potassium tetraoxosulphate (VI) K_2SO_4
- Calcuim carbonate or calcium trioxocarbonate (VI) $CaCO_3$
- Soduim sulphate or sodium tetraoxosulphate (VI) Na_2SO_4
- Sodium nitrate or sodium trioxonitrate(V) ($NaNO_3$)

Acid salt: it is a type of salt formed when only part of the replaceable hydrogen ions of an acid are replaced by a metal or ammonium ion.

Examples include:

- Sodium hydrogen trioxocarbonate (IV) ($NaHCO_3$)
- Sodium hydrogen tetraoxocarbonate (V) (NaH_2CO_4)
- Sodium hydrogen tetraoxosulphate (VI) ($NaHSO_4$)
- Potassium hydrogen tetraoxosulphate (VI) ($KHSO_4$)

Basic salt: it is a type of salt formed when insufficient acid is available to replace all the hydroxide or oxide ions of a base.

Examples of basic salts include:

- Zinc chloride hydroxide Zn(OH)Cl

- Magnesium chloride hydroxide $Mg(OH)Cl$
- Bismuth (II) chloride oxide $BiOCl$

Double salt: it is the type of salt formed when an equimolar of two salts containing two different cations in different oxidation states are mixed together. Double salts in which the cations are in +1 and +3 oxidation states are referred to as **alum**

Examples of double salt:

- Potassium Aluminium bis tetraoxosulphate (VI) dodecahydrate ($KAl(SO_4)_2.$ $12H_2O$)
- Iron (III) ammonium bistetraoxosulphate (VI) decahydrate ($Fe(NH_4)(SO_4)_2$ $.10H_2O$) also
- Iron (II) ammonium hexahydrate $Fe(NH_4). 6H_2O$

Complex salt: it is the type of salt which contains a complex cations or anions as a stable species.

Examples of complex salt:

- Tetraamine Copper (II) tetraoxosulphate (VI) $[Cu(NH_3)_4]^{2+} [SO_4]$
- Potassium hexacynoferrate (III) $K_3 [Fe(CN)_6]$
- Tetraaqua zinc (II) tetraoxosulphate (VI) $Zn(H_2O)_4SO_4$
- Tetraamine nickelate (II) trioxocarbonate (IV) $Ni(NH_3)_4CO_3$

Preparation of Salts

There are several methods of preparing salts some of these include:

- Neutralization e.g. $NaCl$, $NaSO_4$, KNO_3.
- The action of acid on metals e.g. $MgSO_4$, $FeSO_4$
- The action of acid on metal carbonates e.g. $CaCl_2$, $ZnCl_2$

- The reaction of acids with metal oxides and hydroxides.
- Double decomposition e.g. $PbSO_4$, $CaSO_4$, AgCl, $CaCO_3$.
- Direct combination e.g. FeS, $ZnCl_2$, AlCl, $CaCO_3$.

Uses of Salts

Salt	**Uses**
Sodium chloride (NaCl)	- Used as food preservative. - Used to manufacture sodium hydroxide (NaOH).
Ammonium chloride (NH_4Cl)	- Used as flavouring agents. - Used as a systemic and acidifying agent.
Copper (II) tetraoxosulphate (VI) $CuSO_4$	- Used to kill fungi. - Used to manufacture pure copper.
Potassium nitrate (KNO_3)	- Used to manufacture gunpowder.
Magnesium tetraoxosulphate (VI) ($MgSO_4$)	- Used as medicine for stomach upset
Sodium hydrogen trioxocarbonate (IV) ($NaHCO_3$)	- Used in preparation of baking powder
Calcium tetraoxosulphate (VI) $CaSO_4$	- Used for plaster of Paris (POP)

Ammonium tetraoxosulphate (VI) $(NH_4)_2SO_4$		Used to manufacture fertilizer.
Sodium trioxocarbonate (IV) (Na_2CO_3)		Used to manufacture cement, mortar and concrete.
Lead(II) trioxocarbonate (IV) ($PbCO_3$)		Used to manufacture paint.
Sodium octadecanoate (sodium stearate)		Used as soap.

pH of a Solution

The pH of a solution is an is an indication or a measure of the concentration of hydrogen ions $[H^+]$ or hydroxonium ion $[H_3O^+]$ the solution contains.

pH is defined as negative logarithm of hydrogen $[H^+]$ or hydroxonium ion $[H_3O^+]$ concentration.

Mathematically: $\mathbf{pH = -\log_{10}[H^+] \text{ or } -\log_{10}[H_3O^+]}$

When a solution has more hydrogen ions concentration H^+ than its hydroxyl ion concentration OH^-, then the solution is **acidic solution**.

When a solution has hydroxyl ion OH^-concentration than hydrogen ion H^+ concentration, then the solution is **alkali solution.**

Note: $\mathbf{pH + pOH = 14}$

Worked example

1. Calculate the pH of aqueous solution of 3.64×10^{-8} M hydrogen ion concentration.

Answer pH= -log $[H^+]$, $[H^+] = 3.64 \times 10^{-8}$ pH= - log $[3.64 \times 10^{-8}]$ = **7.44**

2. What is hydrogen ion concentration of a solution whose pOH is 11.5?

Answer pH + pOH = 14, but pOH = 11.5

pH + 11.5 = 14 →pH= 14-11.5 = **2.5**

3. Calculate the pH of aqueous solution 0.0003M hydrogen ions.

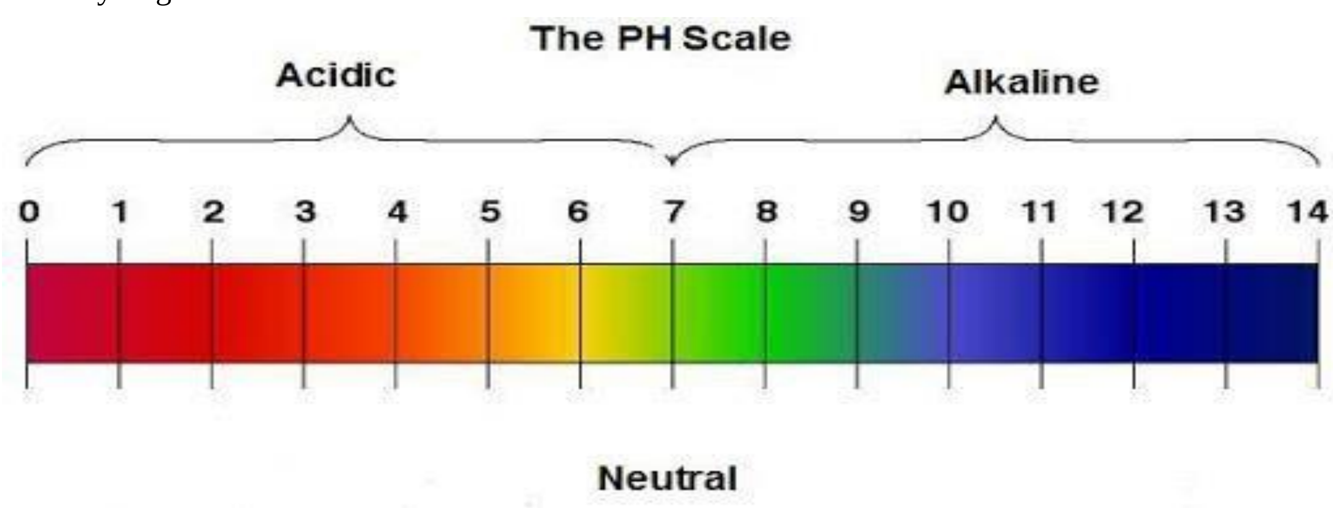
Answer

$[H^+] = 0.0003M = 3x10^{-4}M$, pH = -log $[H^+]$, pH = -

log $3x10^{-4}$ = **3.523**

pH Scale

The pH scale is an index used to determine the acidity or alkalinity of a particular solution. The scale consists of numbers ranging from 0 to 14. These numbers correspond to acidity/alkalinity or hydrogen ion concentration of the solution.



Increasing Increasing

Significant of pH Scale

- It enables us to determine whether a substance is acidic, alkaline or neutral as well as the degree of acidity or alkalinity of the substance.

- In soap industry it is important to know that the final product is neutral and hence the final soap solution must have a pH of 7.

- In the pharmaceutical and medical fields, various drugs are prepared at given pHs which must be determined.
- In agriculture, soils which are too acidic are harmful to crops and information is obtained by determining the pH of the soil solution. Neutralization of excess acid is carried out in practice by adding some slaked lime ($Ca(OH)_2$).

Acid-Base Indicators

Acid base indicators are weak complex organic acids or bases which ionize slightly in water.

Examples of acid-base indicators are;

- Litmus
- Methyl orange
- Methyl red,
- Phenolphthalein etc.

Each of these indicators produces a particular colour in a given solution. The colour variation is due to differences in the concentration of H^+ and OH^- in the solution. Usually, indicators are used to find out whether the solution is acidic, basic or neutral.

Acid – Base Indicator Table

Indicator	**Colour in Acidic solution**	**Colour in alkaline solution**	**Colour in neutral solution**
Litmus	Red	Blue	Purple
Methyl orange	Red/ Pink	Yellow	Orange
Phenolphthalein	Colourless	Pink	Colourless
Universal indicator	Red	Purple	Green

Universal Indicators

A universal Indicator is a mixture of acid-base indicators which shows series of colour changes at different pH values. Universal Indicator can be used to estimate the pH of a solution. It is available in solution or paper form accompanied by standard colour chart and matching pH values as shown above.

pH Values of Some Common Materials

MATERIALS	pH
Battery acid	0.5
Gastric juice	1.0 – 3.0
Lemon juice	2.2 – 2.4
Vinegar	2.4 – 3.4
Tomato juice	4.0 – 4.4
Carbonated beverages	4.0 – 5.0
Black coffee	5.0 – 5.1
Urine	5.5 – 7.5

Rain (unpolluted)	6.2
Milk	6.3 – 6.6
Saliva	6.5 – 7.5
Pure water Blood	7.0 7.35 – 7.45
Bile	6.8 – 7.0
Pancreatic fluid	7.8 – 8.0
Sea water	8.0 – 9.0
Soap	8.0 – 10.0
Milk of magnesia	10.5
Household ammonia	11.7
Lye (1.0M NaOH)	14.0

CHEMISTRY OF CARBON COMPOUNDS II

Introduction

This unit discusses the concept of isomerism among compounds of carbon. It then describe the chemistry of some specific classes of carbon carbons namely alkanes, alkenes, alkynes and alkanols.

Isomerism

Isomerism is the occurrence of two or more compounds with the same molecular formula but different structural formula

Example: Molecular formula C_4H_{10}, the isomers are:

Isomers: A. $CH_3CH_2CH_2CH_3$ Butane

B. CH_3CHCH_3 2-methylpropane
$\quad\quad\quad|$
$\quad\quad\; CH_3$

Structures A and B above are isomers that have the same molecular formula, C_4H_{10} but different structural formula.

For more information on isomerism, **click on the link below;** https://youtu.be/uW7z2TS5KqU

Types of isomerism

There are two main types of isomers, they are;

- Structural isomerism
- Stereoisomerism

Structural isomerism is defined as the occurrence of compounds of the same molecular formula but different structural formula.
There are three main sub-division of structural isomerism, they are;

1. **Chain isomers**: - this arises because of the possibility of branching of carbon chain.

Example: Molecular formula: C_5H_{12}

Isomers: Example: Molecular formula: C_5H_{12} $CH_3CH_2CH_2CH_2CH_3$,

Pentane

$CH_3CH_2(CH_3)CHCH_3$, 2-methylbutane

$CH_3C(CH_3)_2CH_3$, 2,2-dimethylpropane

2. **Positional isomers**: this arises because of the possibility of changing positions of functional groups on carbon chain. ➢ Example: Molecular formula: C_3H_8O ➢ Isomers:

❖ $CH_3CH_2CH_2OH$ → 1-propanol

❖ $CH_3C(OH)HCH_3$ → 2-propanol

3. **Functional group isomers:** this arises due to the possibility of interchanging certain pairs of functional groups.

Example; Molecular formula: $C_4H_{10}O$

Isomers:

$CH_3CH_2CH_2CH_2OH$ → 1-butanol

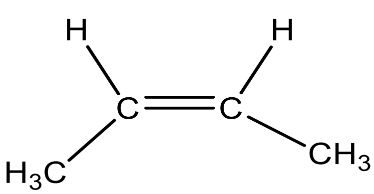
$CH_3CH_2\text{-}O\text{-}CH_2CH_3$ → Diethyl ether or ethoxyethane/ Diethyl oxide

Stereoisomerism: This is the occurrence of two or more molecules of the same structural formula but different arrangement of atoms in space.

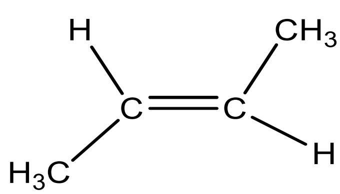
Two main types of stereoisomerism exist. These are;

1. **Geometric stereoisomerism**: This is the occurrence of compounds of the same structural formula but different arrangement of atoms in space due to hindrance to rotation.

Example: but-2-ene has two Geometric stereoisomers.



Cis-2-butene



Trans-2-butene

NB: Trans-isomers are stable than cis-isomers.

Alkanes

Alkanes are saturated hydrocarbons containing carbon-carbon single bonds. They form homologous series with general molecular formula CnH_{2n+2}. All carbon atoms of alkanes undergo SP^3 hybridization. The shape of alkane molecule is tetrahedral. The carbon-carbon bonds are sp^3-sp^3 sigma bond overlap

Sources of alkanes

- Methane is sourced from natural gas (formed by biological decomposition of marine organisms) and destructive distillation of wood (burning of wood under limited supply of wood)
- other alkanes are produced from crude oil (or petroleum).

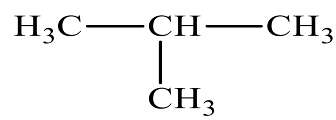
Isomerism of alkanes

Alkanes exhibits chain isomerism. This is due to the possibility of branching from straight chain

Example C_4H_{10} Isomers:

$CH_3CH_2CH_2CH_3$

n-butane



2-methylpropane

Physical Properties of Alkanes

Alkanes exist in all the three states of matter under normal conditions.

- (C_1-C_4) are gases. (C_5-C_{17}) are liquids and (17 upwards) are waxy solids.
- They are less dense than water and also insoluble in water.
- They have very low melting and boiling point but increases as the molecular mass increases along the homologous series.

Chemical properties of alkanes

Alkanes burns completely in excess oxygen to produce carbon dioxide, water and energy.

- Example: $CH_3CH_3 + O_2$ $CO_2 + H_2O$ + Energy

- However, in a limited supply of oxygen, alkanes undergo incomplete combustion, producing poisonous carbon (II) oxide. This is also called addition reaction.

$CH_3CH_3 + O_2$ $CO_2 + CO + H_2O$ + Energy

- However, in a limited supply of oxygen, alkanes undergo incomplete combustion, producing poisonous carbon (II) oxide. This is also called addition reaction.
- This occurs in everyday devises like cars, kerosene lamps and gas cooker;
- Starting a car in an enclosed garage
- Using cooker and kerosene stove in a poorly ventilated kitchen.

Alkanes undergo substitution reaction

E.g. Halogenation reaction

$$CH_4 + Cl_2 \xrightarrow{uv} CH_3Cl + HCl$$

Alkanes undergo cracking, a reaction in which a large molecule of alkane is broken into smaller molecule.

(a) $$\underset{\text{ethane}}{C_2H_6} \longrightarrow \underset{\text{ethene}}{C_2H_4} + \underset{\text{hydrogen}}{H_2}$$

(b) $$\underset{\text{dodecane}}{C_{12}H_{26}} \longrightarrow \underset{\text{hexane}}{C_6H_{14}} + \underset{\text{ethene}}{C_2H_4} + \underset{\text{butene}}{C_4H_8}$$

Uses of Alkanes

Alkanes are used;

- As fuel for automobiles
- As source of heat for cooking
- As a stating material for many organic synthesis

Laboratory Preparation of Alkanes

i. Hydrogenation of alkene

$$\underset{\text{ethene}}{CH_2 = CH_2} + H \xrightarrow{Ni} \underset{\text{ethane}}{CH_3CH_3}$$

ii. Decarboxylation

$$CH_3CH_2COONa + NaOH \longrightarrow CH_3CH_3 + Na_2CO_3$$

iii. From alkyl halide, RX (X=Cl, Br, F)

$$2CH_3Cl + Na \longrightarrow CH_3CH_3 + NaCl$$

Alkenes

Alkenes are unsaturated hydrocarbons containing carbon-carbon double bond. They form a homologous series with general molecular formula C_nH_{2n} where $n \geq 2$. They exhibit sp^2-hybridization. The shape of alkene molecule is trigonal planar. They are more reactive due to the presence of delocalized electrons in the π-bond. The double bond consists of one strong sigma (σ) bond and one weak pi (π) bond. The bond is also shorter and stronger than single bond.

Sources of Alkenes

Alkenes are produced by:

- **Cracking of large alkanes**. Eg. Cracking of kerosene under high temperature.

  Example

$$C_{12}H_{26} \xrightarrow{673-900K} C_{10}H_{22} + CH_2 = CH_2$$

- **Conversion of alkanols (i.e. Dehydration of alkanol)**

Example:

$$CH_3CH_2OH \rightarrow CH_2 = CH_2 + H_2$$

- **Conversion of alkyl halide to alkenes** Example:

$$CH_3CH_2CH_2Br \xrightarrow{\text{alc. KOH}} CH_3CH_2 = CH_2$$

Physical Properties of Alkenes

1. They have lower boiling and melting point. These increase steadily with increasing chain length due to increasing strength of van der Waals forces.

2. They are polar, therefore insoluble in water. They dissolve in non-polar solvent like benzene and CCl_4.

3. The lower members up to butane are gases. C_5 up to C_{15} molecule are volatile liquids. Above C_{15} molecule are solids.

Chemical Properties of Alkenes

1. Alkenes are chemically more reactive than alkanes. This because alkenes are unsaturated hydrocarbons that have a double bond, C=C, between two carbon atom. Almost all of the chemical reactions of alkene occur at the double bond.

2. Alkenes can undergo:

(a) Combustion

(b) Addition reaction

(c) Polymerisation reaction

Combustion of Alkenes

Like any other hydrocarbons, alkenes burn in air or oxygen. Like alkanes, an alkene burns completely in the sufficient oxygen to produce carbon dioxide and water. Alkene burns incompletely in limited supply of oxygen to produce carbon monoxide, carbon (in the form of soot) and water.

Example

1. Complete combustion of ethene $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$
2. Incomplete combustion ethene $C_2H_4 + 2O_2 \rightarrow 2CO + 2H_2O$

Addition Reaction of Alkenes

Alkenes are unsaturated hydrocarbons, hence they undergo addition reactions. Addition reaction occurs when other atoms are added to each carbon atom of the double bond, —C=C— to form single covalent bond. Alkenes may undergo addition reaction with halogen, steam, hydrogen, halogen halide and potassium permanganate(VII).

Uses of Alkenes

1. Alkenes are used as a starting material in the synthesis of plastics, lacquers alcohols
2. They are also used as fuel
3. Ethene, the most important organic feedstock in industry for many chemical products such as polyethylene , vinyl chloride, styrene, ethanol etc

Alkynes

Alkynes are unsaturated hydrocarbons containing a carbon-carbon triple bond. They form a homologous series with general molecular formula C_nH_{2n-2} where $n \geq 2$. The shape of alkyne molecule is linear. The triple bond consist of one strong sigma (σ) bond and two weak pi (π) bond. The functional group in alkyne is carbon-carbon triple bond.

It is represented as -C≡C-

Isomerism in Alkynes

Alkynes show position and chain isomerism but not geometric because it is linear.

Molecular formula: C_4H_6

Isomers: A. $CH_3C{\equiv}CCH_3$ B. $CH{\equiv}CCH_2CH_3$

2-butyne 1-butyne

Sources of Alkyne

1. Alkynes are produced on a large scale by cracking alkane from petroleum fractions.

Example;

$$2CH_4 \xrightarrow{1800K} HC\equiv CH + 3H_2$$

2. Ethyne is produced from the action of water on calcium carbide

$$\underset{\text{Cold water}}{CaC_2 + 2H_2O} \longrightarrow \underset{\text{ethyne}}{Ca(OH)_2 + C_2H_2}$$

3. Refluxing 1, 2-dibromoalkane with alcoholic KOH. Example;

$$CH_2BrCH_2Br \xrightarrow{\text{alc. KOH}} \underset{\text{ethyne}}{HC\equiv CH} + 2HBr$$

Physical Properties of Alkynes

1. The alkynes containing up to four carbon atoms are gases, while those with more carbon atoms, depending on the size of the molecules, are either liquids or solids.

2. The alkyne homologous series show a gradual increase with molar masses
3. Symmetrical alkynes have slightly higher physical constants: boiling and melting points and densities

4. Alkynes are mainly insoluble in water but are freely soluble in organic solvents

5. Alkyne molecules containing four or more carbon atoms show chain isomerism.

   Thus, butyne (C_4H_6) has two isomers: butyne, $CH_3CH_2CH_2\equiv CH$ and

   but-2-yne $CH_3.C\equiv C.CH_3$

Chemical Properties of Alkynes

1. Alkynes are more reactive than alkenes since they are more unsaturated than the alkenes

2. They undergo a double addition reaction to give saturated compounds. For e.g.

$$\underset{\text{ethyne}}{CH\equiv CH} + H_2 \longrightarrow \underset{\text{ethene}}{CH_2=CH_2} + H_2 \longrightarrow \underset{\text{ethane}}{CH_3\text{-}CH_3}$$

Test for Unsaturated compounds

1. Unsaturated compounds rapidly decolourises a reddish brown Br_2 liquid in

   CC_4

Unsaturated compounds also decolourises an acidified purple solution of $KMnO_4$ to green which later turns brown ppt.

For more information on test for unsaturated compounds, click on the link below

https://youtu.be/E2Tqv1DnwHg

Uses of Alkyne

1. Ethyne is used as oxy-acetylene flame for cutting and welding of metals.

2. It is used as a raw material for synthesis of rubbers, PVC plastics.

3. It is used as a source of fuel for hunters' lamp.

4. It is used in the production of trichloroethane.

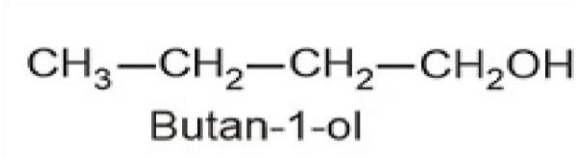
Alkanols

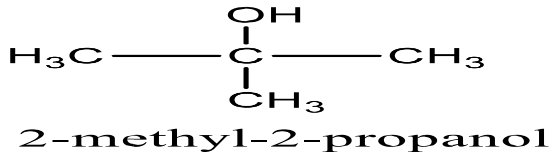
Alkanols are organic compounds which contain one or more hydroxyl group (-OH) attached to the alkyl group. Alkanols containing only one hydroxyl group (-OH) are referred to as monohydric alkanols whiles those containing more than one hydroxyl group (-OH) are known as polyhydric. Alkanols form a homologous series with general $C_nH_{2n+1}OH$. They have functional group –OH (ie. hydroxyl group). The intermolecular force of attraction in alkanols is hydrogen bonding.

Isomerism in alkanol

Alcohols exhibit following types of isomerism:

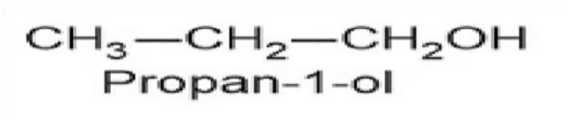
Chain isomerism: Alcohols with four or more carbon atoms exhibit this type of isomerism in which the carbon skeleton is different. Example; C_4H_9OH

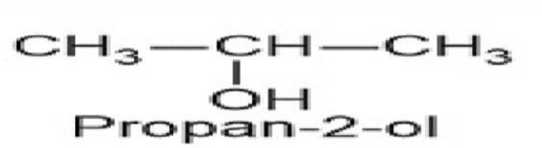




Position isomerism Alcohols with three or more carbon atoms can exhibit position isomerism. In this type of isomerism the position of the functional group i.e., the -OH group varies. In other words the carbon atoms to which the -OH group is attached is different.

Example; C_3H_7OH





Physical Properties of Alkanols

Alkanols have higher melting and boiling than corresponding alkane. This is due to strong intermolecular force of attraction (hydrogen bonding).

Lower members of alkanols are soluble in water due to intermolecular force of attraction. The lower, C_1-C_4 are liquids at normal condition.

Chemical Properties Alkanols

The functional group (-OH) in alkanols can undergo many chemical reactions.

Examples;

1. **Esterification reaction**; alkanols and alkanoic acid to form esters.

2. **Oxidation of alkanols**

- Both primary (1°) and secondary (2°) alkanol undergo oxidation except tertiary (3°) alkanols.

- Primary alkanol are oxidized to alkanals and then to alkanoic acid

Example; $\underset{\text{ethanol}}{CH_3CH_2OH} \xrightarrow{K_2Cr_2O_7/\ H^+} \underset{\text{ethanal}}{CHO} \xrightarrow{\text{Further oxidation}} \underset{\text{ethanoic acid}}{CH_3COOH}$

3. **Dehydration of Alkanols:** When alkanols are dehydrated they turn to alkenes. Example;

$$\underset{\text{ethanol}}{CH_3CH_2OH} \xrightarrow[\text{heat}]{\text{Excess } H_2SO_4} \underset{\text{ethene}}{CH_2{=}CH_2} + H_2O$$

Alkanoic Acid

They are organic compounds containing one or more carboxylic group **(-COOH).** The structured monocarboxylic acids form a homologous series with general molecular formula $\mathbf{C_nH_{2n+1}COOH}$.

Sources of Alkanoic Acid

Name	Sources
Methanoic or formic acid	Ants, bee
Ethanoic or acetic acid	Vinegar, sour wine or gin
Butanoic acid	Butter, cheese
Hexanoic acid	Coconut oil
Hexadecanoic acid/palmitic acid	Palm oil

Citric acid	Citrus fruits eg. lemon, orange
Oleic acid	Corn oil, groundnut oil
Lactic acid	Sour milk
Tartaric acid	Grapes
Ascorbic acid	Tomatoes
Malic acids	Apples

Laboratory Preparation of Alkanoic Acid
Oxidation of primary alkanols

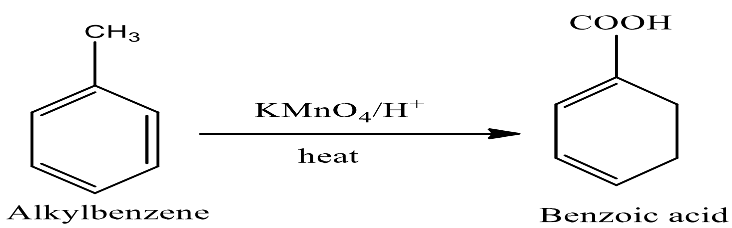
- Reagent: **$KMnO_4/H^+$ or $K_2Cr_2O_7$**

  - Condition : **heat**

- **Oxidation of Alkylbenzene**

- Reagent: $KMnO_4/H^+$ or $K_2Cr_2O_7$

- Condition: heat



- **. Acid Hydrolysis**

$$CH_3COOCH_2CH_3 \xrightarrow[\text{dil. acid}]{} CH_3COOH + CH_3CH_2OH$$

Physical Properties

1. They have sour taste

2. The first 8 straight chain members are liquids, the rest are solids with low melting point.

3. They have relatively higher boiling and melting point. This is because the molecules associate themselves through hydrogen bonding.

4. The lower members are soluble in water.

Chemical Properties of Alkanoic Acid

1. Organic acids are weaker than mineral acids.

2. Produces hydrogen gas with active metals.

3. Undergo neutralization reaction with a base.

4. Undergo esterification reaction with alkanols

Test For Alkanoic Acid

- **Reagent:** $NaHCO_3$ or Na_2CO_3

- **Observation:** Effervescence of colourless gas which turns lime water milky.

Uses of Ethanoic Acid

1. As organic solvent.

2. As a starting material for organic synthesis

3. As disinfectant

4. As a preservative

5. Treatment of skin diseases like ring worms

6. As a preservative