TVET CERTIFICATE IV in CROP PRODUCTION

CRPIO401

BASIC KNOWLEDGE OF ORGANIC AND INORGANIC CHEMISTRY

Demonstrate basic knowledge of organic and inorganic chemistry

Competence

Learning hours

40

Credits: 4

Sector: Agriculture

Sub-sector: Crop production

Module Note Issue date: September, 2020

Purpose statement

This module describes the skills and knowledge required to apply the knowledge of chemistry provided for level four. The module will allow the participant to describe, demonstrate and identify different components of chemistry as basis of other related modules

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	1.2 Proper calculation of atomic mass.	
	1.3 Appropriate description of chemical elements and molecules	
2. Apply the basic knowledge of inorganic chemistry	2.1 Proper simple chemical equilibrium	24
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	2.3 Proper description of solution	
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3.Apply/understand/acquire / the basic knowledge on organic chemistry	3.1 Proper classification of Organic compounds (Functional group).	21
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	3.3 Appropriate description of organic products carbohydrates Proteins, Vitamins and lipids".	

Total Number of Pages: 76

LO 1.1 – . DESCRIBE ATOM PARTICLES

Introduction

The word **Chemistry** refers to **the study of the composition, structure and properties of substances under different conditions.**

All the stuff you see around you is made of atoms and you are also made of atoms.
Examples:
Breathing O_2 binds to Fe in Hemoglobin in red blood cells and thousands of other chemical processes going on in our cells and bodies
We get energy from food we eat
sugar + oxygen → carbon dioxide + water + energy
We get energy for industry by combustion
butane + oxygen →carbon dioxide + water +energy

Chemical reactions involve rearrangement of atoms but atoms are not created or destroyed - This idea of mass not changing is also called "Conservation of Mass"

In 1808 Dalton published *A New System of Chemical Philosophy,* in which he presented his theory of atoms:
Dalton's Atomic Theory
1. Each element is made up of tiny particles called atoms.
2. The atoms of a given element are identical; the atoms of different elements are different in some fundamental way or ways.
3. Chemical compounds are formed when atoms of different elements combine with each other. A given compound always has the same relative numbers and types of atoms.
4. Chemical reactions involve reorganization of the atoms—changes in the way they are bound together. The atoms themselves are not changed in a chemical reaction.

- Description and structure of atoms

Definition of an atom

Discussion corner!

1. How did you define matter in o-level?
2. (a) What is the smallest thing that you know that occupies volume, has weight and cannot be divided any further?
(b)Why do you think it is the smallest?
3. As the secretary of a class consisting of thirty-two students, you have a single orange; all the class members need a share of the fruit:
(a) How will you share it with everyone equally?
(b)Is there a point at which the orange will be too small that you can no longer divide it?
(c) How small were the pieces?

I have discovered that…

When you sub-divide an orange, it reaches a point at which further sub-division cannot take place. Although the portion is small, it still occupies volume and has weight. Hence it is considered to be ***matter****. Just like the orange, substances can be sub-divided to the point where further division cannot take place. At this point, we call the particles atoms. Atoms are the smallest basic units of matter.*

All objects from our everyday life that we touch or see are composed of small particles called **atoms**. Atoms are very tiny and invisible. An atom ***is the smallest particle into which an element can be divided without losing the chemical properties of the element.***
Atoms are usually very small such that they cannot even be seen with the aid of a microscope.

Main components of an atom

Activity
1. Regardless of some exceptions, all atoms are composed of the same components. True or False? If this statement is true why do different atoms have different chemical properties?
2. Using your knowledge about atom, what is the role each particle plays in an atom?

The components of the atom are known as sub-atomic particles.

- ✓ **Nucleus**
  The nucleus of an atom is positively charged due to the presence of protons which hare positively charged. Neutrons carry no charge.
  -P**rotons** are sub-atomic particles of an atom
  - **neutrons** are sub-atomic particles of an atom
  Protons and neutrons are found in a central location within an atom. This location is called the **nucleus**.
- ✓ **Electrons**
  Electrons are negatively charged and keep on moving round the nucleus experiencing a force of attraction from the nucleus.
  Electrons occupy special positions known as **levels**.
  Electrons are found outside the nucleus but within the atom.

A sub-atomic particle consists of characteristic charge, mass and its symbol.
Table Properties of sub-atomic particles

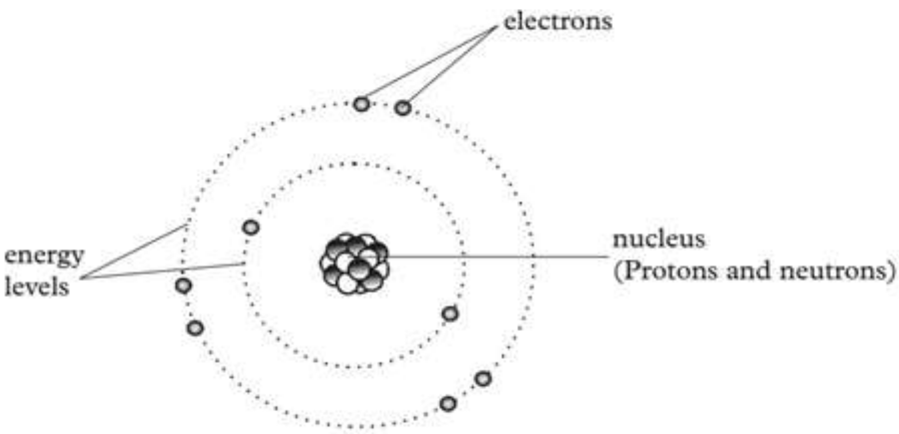
Particle	Absolute charge(Coulomb)	Relative charge	Mass (kg)	Relative masse
Neutron(n)	0	0	1.675×10^{-27}	1.0087 amu*
Proton(p or p^+)	$+1.6\times10^{-19}$	+1	1.673×10^{-27}	1.0073 amu
Electron(e-)	-1.6×10^{-19}	-1	9.11×10^{-31}	000054858 amu.

(*amu: atomic mass unity, 1 amu=1.67×10^{-27}kg)

The mass of an electron is very small compared with the mass of either a proton o ra neutron.
The charge on a proton is equal in magnitude, but opposite in sign, to the charge on an electron.

The number of protons is equal to the number of electrons. For this reason, the atom is considered **electrically neutral**.

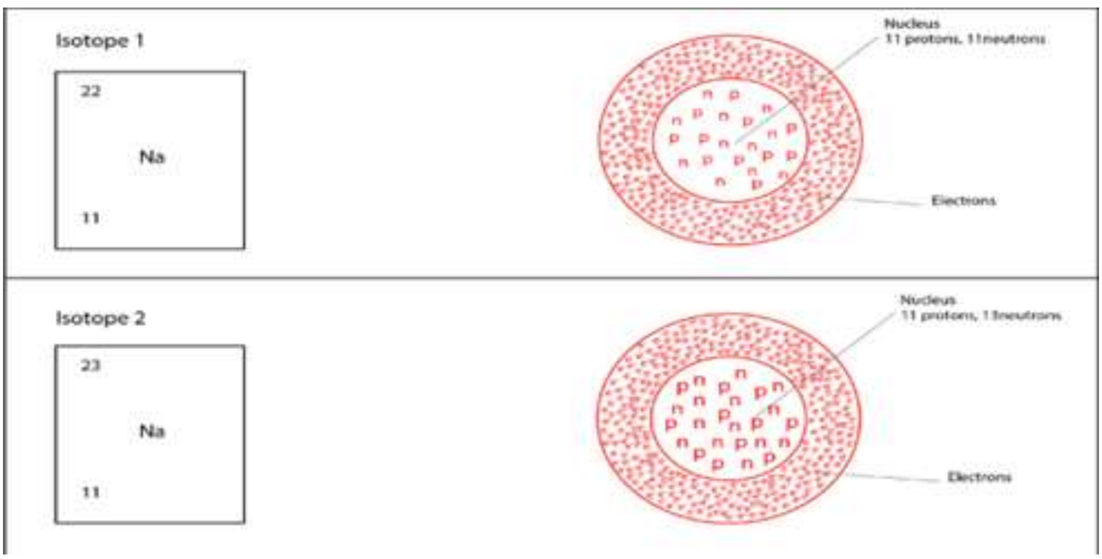
Fig . Atomic structure showing nucleus and electrons



Checking up
1. In an experiment, it was found that the total charge on an oil drop was 5.93×10^{-18} C. How many negative charges does the drop contain?
2. All atoms of the elements contain three fundamental particles. True or false? Give an example to support your answer.
3. Compare the atom constituents
a. in terms of their relative masses
b. in terms of their relative charges
4. Using the periodic table as a guide, specify the number of protons and electrons in a neutral atom of each of these elements.
a. carbon (C) b. calcium (Ca) c. chlorine (Cl) d. chromium (Cr)

LO 1.2 – CALCULATE THE ATOMIC MASS.

Activity
The diagram below shows a representation of sodium isotopes. Observe it and answer to the questions below



1. Compare the two sodium isotopes in the figures above.
2. From your observation, how do you define the isotopes of an element?
3. How is the mass number, A, determined?
4. What information is provided by the atomic number, Z?
5. What is the relationship between the number of protons and the number of electrons in an atom?
6. Where are the electrons, protons, and neutrons located in an atom?
7. Why is the mass of an atom concentrated in the center?
8. Sodium-24 and sodium-23 react similarly with other substances. Explain the statement
9. Say which one(s) of the following statements is(are) correct and which one(s) is(are) wrong: (i) isotopes differ in their number of electrons, (ii) isotopes differ in their mass numbers, (iii) isotopes differ in their number of protons, (iv) isotopes differ by their number of neutrons, (v) all the statements are wrong

- Mass of an atom

In a given atom, the number of protons, also called "atomic number" are equal to the number of electrons because the atom is electrically neutral. The sum of the number of protons and neutrons in an atom gives the mass number of that atom.

Mass number = number of protons + number of neutrons
= atomic number + neutron number

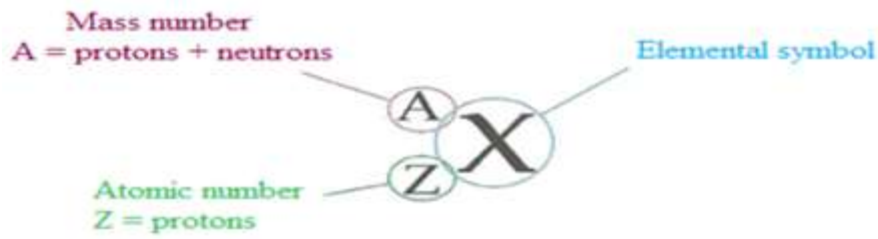
Relative atomic mass, symbolized as R.A.M or Ar, is defined as the mass of one atom of an element relative to 1/12 of the mass of an atom of carbon-12, which has a mass of 12.00 atomic mass units. The relative atomic mass, also known as the atomic weight or average atomic weight, is the average of the atomic masses of all of the element's isotopes.

- Relationship between atomic mass, number of Neutron and number of protons.

The number of protons in the nucleus of the atom is always referred to as the **atomic number**.
The symbol for the atomic number is (**Z**).
When the total number of protons and neutrons are added we get the **mass number**. The symbol for mass number is (**A**).
We can use the numbers on any atom to work out the number of protons, neutrons and electrons.
Using the relationship:
A= Z+N
Where **N** is the number of neutrons in an atom, we can work out the number of neutrons as follows.
Example
The mass number of sodium is 23 and the number of protons in its atom is 11.
Calculate the number of neutrons in an atom of sodium.
If **A= Z+N** then **N=A-Z**

Therefore, **N**=23-11=12, the number of neutrons is hence 12.
The difference in number of neutrons leads to difference in mass number.

The **atomic number** denotes the number of protons in an atom's nucleus. The mass number denotes the total number of protons and neutrons. Protons and neutrons are often called nucleons. By convention, the atomic number is usually written to the left subscript of the elemental notation, and the mass number to the left super script of the elemental notation as represented by the example below, where X represents any elemental symbol.



- Description of isotopes

Activity 1. Argon has three naturally occurring isotopes: argon-36, argon-38, and argon-40. Based on argon's reported relative atomic mass from the periodic table, which isotope do you think is the most abundant in nature? Explain. 2. Calculate the average atomic mass of an element with two naturally occurring isotopes: 85X (72.15%, 84.9118 amu) and 87X (27.85%, 86.9092 amu). Identify this element? 3. Boron has two naturally occurring isotopes. Find the percent abundances of 10B and 11B given the isotopic mass of 10B = 10.0129 amu and the isotopic mass of 11B = 11.0093 amu

Some atoms of the same of element have the same atomic number, but different mass numbers. This means a different number of neutrons. Such atoms are called isotopes of the element.
Isotopes are atoms of the same element with different masses; they are atoms containing the same number of protons but different numbers of neutrons

Relative isotopic mass is like relative atomic mass in that it deals with individual isotopes. The difference is that we are dealing with *different forms* of the same element but with different masses.
Thus, the different isotopic masses of the same elements and the percentage abundance of each isotope of an element must be known in order to accurately calculate the relative atomic mass of an element.
Notice: Remember that mass number is not the same as the relative atomic mass or isotopic mass! The mass number is the number of protons + neutrons; while relative atomic mass (or isotopic mass) is the mass if you were to somehow weigh it on a balance.

Let A1, A2, A3,…, An be an abundance of n isotopes of the same chemical element with atomic mass M1, M2, M3,…, Mn respectively, the relative atomic mass(R.A.M) is given by the following equation:

$$RAM = \frac{A_1M_1 + A_2M_2 + A_3M_3 + \cdots + A_nM_n}{A_1 + A_2 + A_3 + \cdots + A_n}$$

Example 1: Oxygen contains three isotopes 16O, 17O, and 18O. Their respective relative abundances are 99.76%, 0.04%, and 0.20%. Calculate the relative atomic mass of oxygen.

Solution:

Relative isotopic mass of 16O is 16 and its relative abundance is 99.76%;

Relative isotopic mass of 17O is 17, abundance 0.04%;

Relative isotopic mass of 18O is 18, abundance 0.20%.

$$\boxed{\frac{(99.76\times16)+(0.04\times17)+(0.20\times18)}{99.76+0.04+0.20} = 16.0044}$$

By applying the same formula, the relative abundance of the isotopes may be calculated knowing the relative atomic mass of the element and the atomic masses of the respective isotopes.

Example 2: Chlorine contains two isotopes 35Cl and 37Cl, what is the relative abundance of each isotope in a sample of chlorine if its relative atomic mass is 35.5?

Solution:

$$35.5 = \left(\frac{A\times35}{100}\right) + \left(\frac{100-A}{100}\right)\times37$$

Note that if the abundance of the isotope of atomic mass 35 is A%, the abundance of the isotope of mass 37 will be (100 – A) %.

0.35A + (100 – A) x 0.37 = 35.5

0.35A– 0.37A =35.5 – 37

– 0.02A = – 1.5

A= 1.5/0.02 = 75

Therefore, the abundance of the isotope of relative atomic mass 35.5 is 75% while that for the isotope 17Cl is 100 – 75 = 25%.

LO 1.3 – DESCRIBE CHEMICAL ELEMENT AND MOLECULES

- Description of chemical element

Chemical Element

Discussion corner!

1. (a) Why are people given names at birth?
(b)And what do people consider when naming?
2. Why do people have different names?
3. When you go to shop, how do you know that a certain product is actually manufactured by a particular company?
4. How can I know that a product I am about to buy is of quality and safe for consumption?

I have discovered that…

When babies are born, they are given names, which are used to identify them. Some babies are named after other people. In other cases, seasons are considered. Having specific names for every person enables effective communication as it prevents confusion while referring to people. A given symbol on a product enables a consumer to know the manufacturer. A quality symbol on a product also enables us to ascertain that a product has been certified by the standards board.

An **element *is a type of matter composed of atoms that all have the same atomic number. Atoms of the same element or different elements can combine. When they do so, they form molecules.***

In Chemistry, elements are given certain names considering either where they were discovered or who discovered them. It is from these names that symbols are obtained. The symbols of elements are different and specific to each element. This creates orderliness when organizing information about different elements.

The system of writing symbols uses letters taken from the name of the element.

This could be the English or Latin name of the element.

The symbol of an element may consist of one or two letters. The first letter of a chemical symbol must always be a capital letter. The letters should not be joined.

These symbols are an **international code**. This means that all over the world, they are written in the same way no matter how people spell the name of the element in their language. The symbol of an element thus remains the same in all languages.

Table of First 20 elements with their symbols

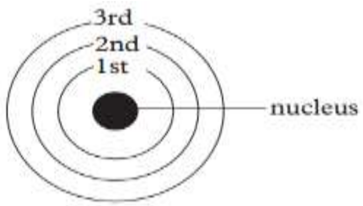
	Element	Symbol
1	Hydrogen	H
2	Helium	He
3	Lithium	Li
4	Beryllium	Be
5	Boron	B
6	Carbon	C
7	Nitrogen	N
8	Oxygen	O
9	Fluorine	F
10	Neon	Ne
11	Sodium	Na
12	Magnesium	Mg
13	Aluminium	Al

14	Silicon	Si
15	Phosphorus	P
16	Sulphur	S
17	Chlorine	Cl
18	Argon	Ar
19	Potassium	K
20	Calcium	Ca

- ✓ **Electronic configuration**

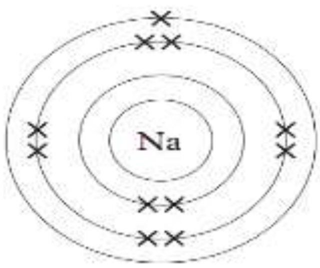
Energy level is a possible location around an atom where electrons are found. They are represented as n =1,2,3 etc We use this formula to know the maximum number of electro that a given energy level may hold **$2n^2$.** as shown in the following figure. There are 7 possible energy levels in atoms.

Fig . Energy Levels



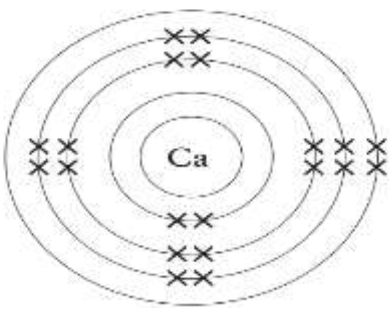
The first energy level (**labelled 1**) can hold up to only two electrons.
The second energy level (**labelled 2**) can hold a maximum of eight electrons. This energy level is filled after the first energy level and before the third level.
The third energy level (**labelled 3**) can hold a maximum of 18 electrons, however when 8 electrons are in the third level there is a degree of stability; and other electrons are added to the fourth energy level.
Electrons must first fill the first energy level before they start occupying the second energy level.

Fig .Shell model for sodium (Na)



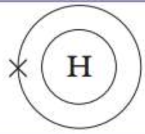
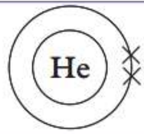
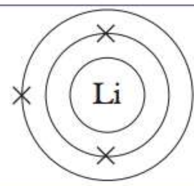
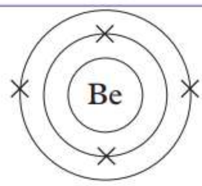
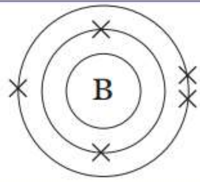
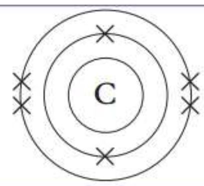
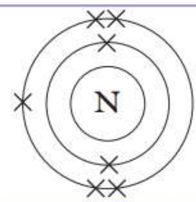
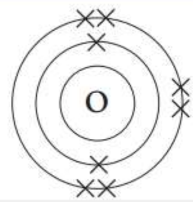
Number of protons=11, number of electrons=11

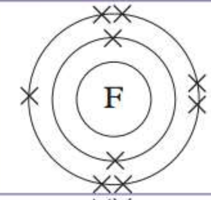
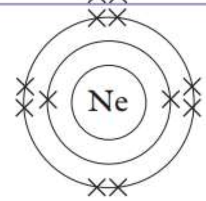
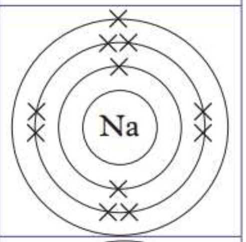
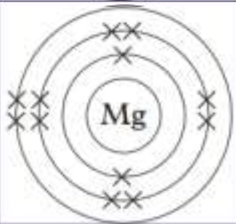
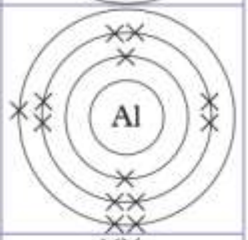
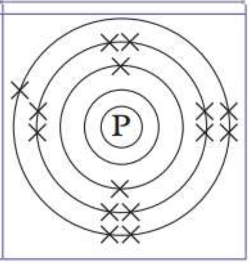
Fig . Shell model for calcium (Ca)

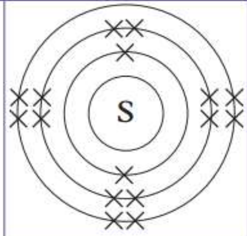
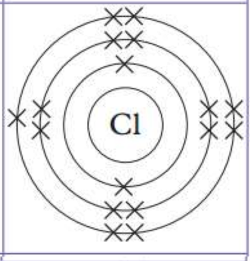
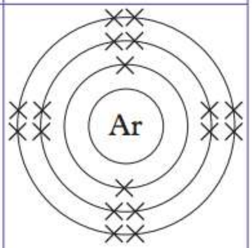
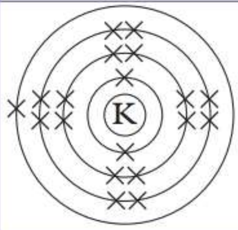
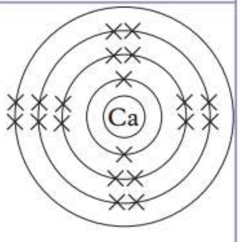


Number of protons = 20, number of electrons = 20

Table of Atomic number, mass numbers, sub-atomic particles and electronicconfiguration of the first 20 elements in the Periodic Table

Element and symbols	Atomic number (z)	Mass number (A)	Sub-atomic particles			Electronic structure	Lewi's electronic arrangement
			p	n	e^-		
Hydrogen (H)	1	1	1	0	1	1	
Helium (He)	2	4	2	2	2	2	
Lithium (Li)	3	7	3	4	3	2.1	
Beryllium (Be)	4	9	4	5	4	2.2	
Boron (B)	5	11	5	6	5	2.3	
Carbon (C)	6	12	6	6	6	2.4	
Nitrogen (N)	7	14	7	7	7	2.5	
Oxygen (O)	8	16	8	8	8	2.6	

Element and symbols	Atomic number (z)	Mass number (A)	Sub-atomic particles			Electronic structure	Lewis electronic arrangement
			p	n	e^-		
Fluorine (F)	9	19	9	10	9	2.7	
Neon (Ne)	10	20	10	10	10	2.8	
Sodium (Na)	11	23	11	12	11	2.8.1	
Magnesium (Mg)	12	24	12	12	12	2.8.2	
Aluminium (Al)	13	27	13	14	13	2.8.3	
Silicon (Si)	14	28	14	14	14	2.8.4	
Phosphorus (P)	15	31	15	16	15	2.8.5	

Element and symbols	Atomic number (z)	Mass number (A)	Sub-atomic particles			Electronic structure	Lewis electronic arrangement
			p	n	e⁻		
Sulphur (S)	16	32	16	16	16	2.8.6	
Chlorine (Cl)	17	35	17	18	17	2.8.7	
Argon (Ar)	18	40	18	22	18	2.8.8	
Potassium (K)	19	39	18	20	19	2.8.8.1	
Calcium (Ca)	20	40	20	20	20	2.8.8.2	

Self evaluation test

1. Given the following substances; common salt, water, nails, sand and kerosene. Classify them into elements and compounds
2. A student is sent by the teacher to the laboratory to collect samples of cobalt and copper from the laboratory storage room. On reaching the shelves, she finds that all the metals are labelled using their symbols. How can she recognise the two metals before collection?
3. While at home, you come across lead acid battery containing this symbol Pb.
   (a) What is the meaning of this symbol?
   (b)Why should this be a matter of concern to you?
   (c) Suggest the most appropriate ways of disposing of used lead acid batteries.

✓ **Symbols**

Activity

1. (a) Why are people given names at birth?
   (b)And what do people consider when naming?
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Table of 10 first elements with their symbols

	Element	Symbol
1	Hydrogen	H
2	Helium	He
3	Lithium	Li
4	Beryllium	Be
5	Boron	B
6	Carbon	C
7	Nitrogen	N
8	Oxygen	O
9	Fluorine	F
10	Neon	Ne

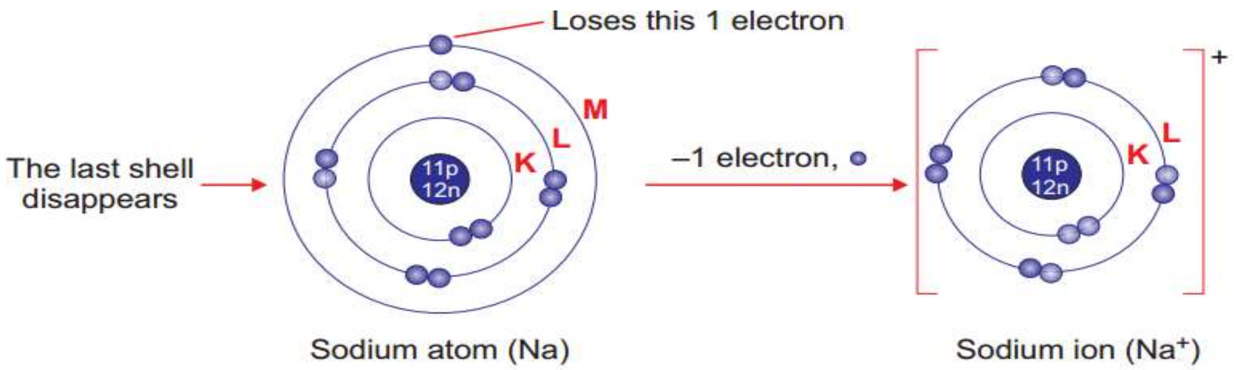
- ✓ **Metal and their properties**

Metals

An ion is formed when an atom loses or gains one or more electrons. The atom may be of a metal or a non-metal.
A metal readily loses its outermost electron or electrons to form a positive ion or **cation**.
The number of positive charges carried on a cation is equal to the number of electron(s) lost by the metal atom.

Metal ions carry positive charges because the number of positively charged protons in the nucleus becomes greater than the number of negatively charged electrons surrounding it. For example, in a sodium atom there are 11 protons in the nucleus and 11 electrons surrounding it. Loss of one electron to form a sodium ion means that there are 11 protons but only 10 electrons. There is a net charge of 1+. This charge is written as a superscript at the right of the symbol of the element

Figure Formation of a sodium ion.



The majority of known elements (about80%) are metals. All the metals are solid, except mercury which is a liquid metal at **r**oom temperature and pressure. Metals are elements which conduct electricity and heat. Metals are also shiny, hard and produce ringing sound when struck. Metals are widely used in our daily life for a large number of purposes. The common metals we use are iron, copper, aluminium, tin, zinc, gold, etc.. The electric fan, machines, bicycle, cars, aeroplane, cooking utensils are all made of metals or mixture of metals.

Physical properties of metals

ACTIVITY . Illustrating Physical Properties of Metals
• Collect some metallic objects. • Note physical properties of collected objects (Get more information from internet, if available). • Make a report on the findings.

Physical Properties of Metals are:

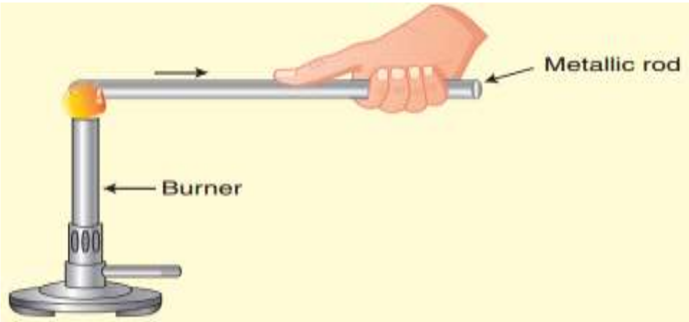
1. Metals are good conductors of heat and electricity. This means that metals allow heat and electricity to pass through them easily. Silver metal is the best conductor of heat. Copper metal is a better conductor of heat than aluminium metal.
2. Metals are lustrous (or shiny). This means that metals have a shiny appearance.

3. Metals are usually **strong**. For example, iron metal (in the form of steel) is very strong when freshly cut and is used in the construction of bridges, buildings and vehicles. Some metals are not strong. For example, sodium and potassium.
4. Metals are **ductile**. This means that metals can be drawn (or stretched) into thin wires.
5. Gold and silver are among the best ductile metals.
6. Metals are **malleable**. This means that metals can be hammered into thin sheets.

ACTIVITY : Illustrating Thermal Conductivity in Metals

Caution: Be careful while heating the objects.

- Take a steel spoon, a brass key, aluminium or copper wire (10 cm), and iron rod.
- Light the burner.
- Hold one end of iron rod in your hand.
- Keep the other end of iron rod to the flame of burner for 3 to 4 minutes as shown in figure.



Conduction of heat in metals

What do you feel?

- Repeat the activity with other metallic objects.
- State your observation in each case. Does the metal wire melt?

Deduction

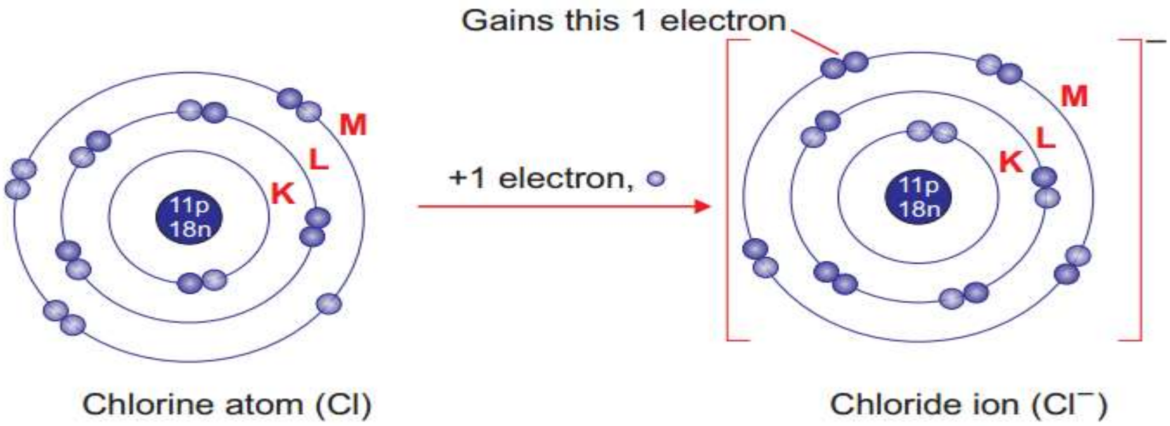
The activity tells that metals are good conductors of heat and have high melting points. The best conductors of heat are silver and copper. Lead and mercury are comparatively poor conductors of heat.

✓ **Non metal and their properties**

Non-metals readily gain one or more electrons into their outermost shell to form a negative ion or **anion**. The number of negative charges an anion carries is equal to the number of electron(s) gained by the non-metal atom.

Non-metal ions carry negative charges because the number of negatively charged electrons surrounding the nucleus becomes greater than the number of positively charged protons in it. For example, in a chlorine atom there are 17 protons in the nucleus and 17 electrons surrounding it. Gain of one electron to form a chloride ion means that there are 18 electrons and only 17 protons. There is a net charge of –1. This charge is written as a superscript at the right of the symbol of the element

Figure : Formation of a chloride ion

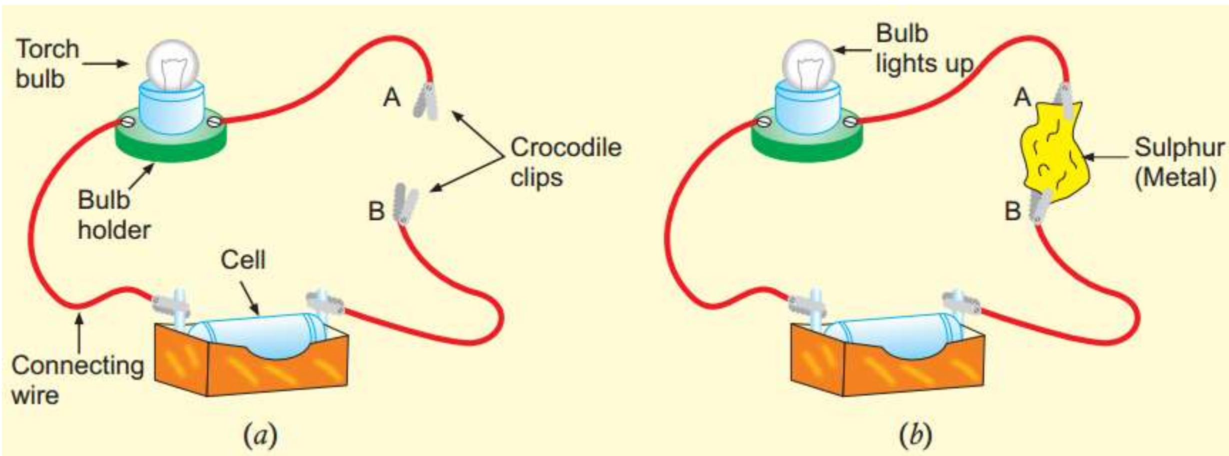


There are only 22 non-metals. Out of these, 10 non-metals are solid, 1 non-metal is liquid and the remaining 11 are gases. Thus, all non metals are solids and gases, except bromine which is a liquid non-metal. Non-metals are elements which do not conduct heat and electricity. The only exception is graphite. Non-metals are brittle and have dull appearance. They are soft, not strong and not shiny.

Physical properties of non-metals

ACTIVITY : Showing that a Non-metal does not Conduct Electricity

We take a dry cell , a torch bulb fitted in a holder and some connecting wires (copper wires) with crocodile clips, and connect them as shown in Figure (a) to make an electric circuit.
There is a gap between the ends of crocodile clips *A* and *B* so no current flows in the open circuit shown in figure (a). Let us now insert
(i) a piece of sulphur (which is a non-metal) between the crocodile clips *A* and *B* as shown in figure (b)
(ii) a carbon rod between the crocodile clips *A* and *B* as shown in figure (b).



Explanation
In both cases, we see that the bulbs do not light up at all. This means that both sulphurand carbon do not allow electric current to pass through them and no current flows in thecircuit. This activity shows that non-metals do not conduct electricity

Physical Properties of Non-Metals are:

1. **Non-metals are not lustrous:** Non-metals do not have a shining surface. The only non-metals having a shining surface is iodine.
2. **Non-metals are neither hard nor strong:** Most of the solid non-metals are soft, they can be broken easily. For example, sulphur and phosphorus.
   Only one non-metal carbon in the form of diamond is very hard. Diamond is the hardest substance known on earth.
3. **Non-metals are neither malleable nor ductile:** Non-metals are brittle which means that they break into pieces when hammered or stretched.
   Therefore, non-metals cannot be hammered with a hammer to form thin sheets. They cannot be stretched to form wires.
   The property of breaking easily is called **brittleness**. Brittleness is the characteristic property of non-metals.
   Note: Brittleness is not applicable to liquid and gaseous non-metals.
4. **Non-metals do not conduct heat and electricity:** Non-metals do not conduct heat and electricity because they have no free electrons which are necessary to conduct heat and electricity. However, there is one exception. Carbon in the form of graphite is a good conductor of electricity. Therefore, graphite is used for making electrodes.
5. **Non-metals have low melting and boiling points:** Non-metals have comparatively low melting and boiling points. Only one non-metal diamond (allotropic form of carbon) has high melting point. The melting point of diamond is 3500°C.
6. **Non-metals have low density:** The density of non-metals is low, that is, they are light substances.

Non-metals are non-sonorous: Non-metals do not produce ringing sound when hit with an object

- ✓ **Groups and periods**

Activity

1. The table below shows the names, symbols and atomic numbers of the first twenty elements. Complete the table by writing the electronic configuration, group and period to which each element belongs
2. Write a report and present it to the class.
   In your report:
   (a) Classify elements with same number of electrons in the outermost energy level.
   (b)Classify elements with same number of energy levels.

Element	Symbol	Atomic number	Electronic configuration	Group	Period
Hydrogen	H	1			
Helium	He	2			
Lithium	Li	3			
Beryllium	Be	4			
Boron	B	5			
Carbon	C	6			
Nitrogen	N	7			
Oxygen	O	8			
Fluorine	F	9			
Neon	Ne	10			
Sodium	Na	11			
Magnesium	Mg	12			
Aluminium	Al	13			
Silicon	Si	14			
Phosphorus	P	15			
Sulphur	S	16			
Chlorine	Cl	17			
Argon	Ar	18			
Potassium	K	19			
Calcium	Ca	20			

The modern Periodic Table

All modern forms of the Periodic Table are based on Mendeleev's original idea.

The principal characteristics of the modern Periodic Table are:

(i) The elements are arranged in a strict order of their **atomic numbers**.

The **Periodic law** states that **"properties of elements are a periodic function of their atomic numbers."**

(ii) The elements are arranged in vertical columns called **groups**. There are eighteen groups. Properties of elements conform to some pattern. Therefore, knowing the properties of one element in a given group makes it possible to predict the properties of other elements in the same group.

(iii) Similarly, gradation of properties from one group of elements to the next group conforms to some pattern.

(iv) Some groups of elements are given special names as follows:

Group 1: - **Alkali metals**

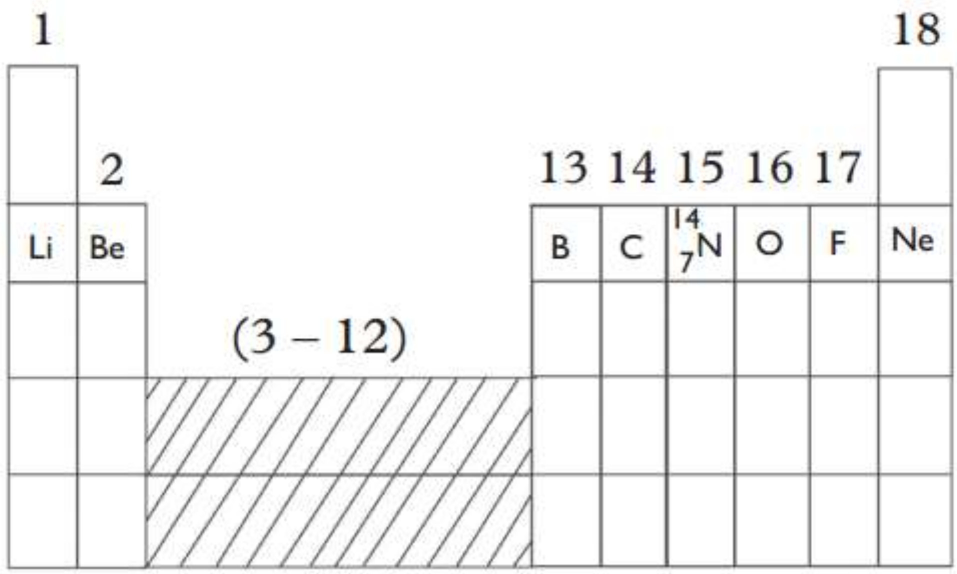
Group 2- **Alkaline Earth metals**

Group 17- **Halogens**

Group 18 or O or **Noble gases**

(v) The elements are also arranged in horizontal rows called **periods**. There are7 periods. In each period, there is transition from metallic to non-metallic characteristics.

(vi) The following grid shows a section of the Periodic Table.



(vii) For each element of the Periodic Table, there is a symbol, name and atomic number.

From the periodic table, it is clear that the:
(i) Elements in the same group have the same number of electrons in the outermost energy level. For example all group I elements have one electron in their outermost energy levels. All group two elements have two electrons in their outermost energy levels. Thus the group number to which an element belongs is equivalent to the number of electrons in its outermost energy level.
(ii)Elements in the same period have the same number of energy levels. If we want to know the period of a given element, we count the number of energy levels it has. If it has 3 energy levels, it is in the third period.

Note!
Hydrogen is placed at the top of group 1. This is done because:
1. It has one valence electron like group 1
2. It forms an ion by losing one electron.
Helium is placed in the eighth group because it has a fully filled outermost energy level. Its properties also match those of the other elements in the group.

Self evaluation test

1. (a) Explain how the elements are arranged in the modern Periodic Table.
(b)"The properties of elements in a group conform to some pattern". Explain the meaning of this statement.
(c) How many periods and how many groups does the modern Periodic table have?
2. (a) How does the nature of the elements change across a period?
(b)Give the special names of group1, group 2, group 17 and group 18 elements.
3. Why do you think helium is placed in group 18 and not 2?

- Bonding

The atoms combine with one another to achieve the inert gas electron arrangement and become more stable. So, when atoms combine to form compounds, they do so in such a way that each atom gets 8 electrons in its outermost shell or 2 electrons in the outermost n shell.

An atom can achieve the inert gas electron arrangement in three ways:
•*By losing one or more electrons* (to another atom).
Atoms with 1, 2 or 3 electrons in the outermost shell lose electrons to achieve stability.
•*By gaining one or more electrons* (from another atom).
Atoms with five, six or seven electrons in the outermost shell gain three, two or one electron respectively to achieve stability.
•*By sharing one or more electrons* (with another atom).
Atoms with four to seven electrons in outermost shell may achieve stability by sharing them with each other.

There are three types of chemical bonds:

- Ionic bond

- covalent bond

- Metallic bond

Ionic bonding

The compounds which are made up of ions are known as ionic compounds. In an ionic compound, the positively charged ions (cations) and negatively charged ions (anions) are held together by the strong electrostatic forces of attraction. The forces which hold the ions together in an ionic compound are known as **ionic bonds** or **electrovalent bonds**. Since an ionic bond consists of an equal number of positive ions and negative ions, the overall charge on an ionic compound is zero. For example, sodium chloride (NaCl) is an ionic compound which is made up of equal number of positively charged sodium ions (Na^+) and negatively charged chloride ions ($Cl^–$).

Ionic compounds are made up of a metal and a non-metal (except ammonium chloride which is an ionic compound made up of only non-metals). So, whenever a bond involves a metal and a non-metal, we call it ionic bond.

EXERCISES

1. Give two examples of ionic compounds. write their chemical formulae.
2. The overall charge of an ionic compound is zero.(true of false).
3. Name the ions present in calcium nitrate.
4. Ionic compounds are made up of ……… and a ……….
5. Give an example of an ionic compound made up of only non-metals.

Properties of ionic compounds

You may have observed the following general properties for ionic compounds:
• Ionic compounds are usually crystalline solids.
• Ionic compounds have high melting and high boiling points.

• Ionic compounds are usually soluble in water but insoluble in organic solvents
like petrol and kerosene.
• Ionic compounds conduct electricity when dissolved in water or when melted. When we dissolve the ionic solid in water or melt it, the crystal structure is broken down to form ions. These ions help in conducting electricity.

Covalent bonding

The chemical bond formed by sharing of electrons between two atoms is known as a **covalent bond**. The compounds containing covalent bonds are known as covalent compounds. A covalent bond is formed when both the reacting atoms need electrons to achieve the inert gas electron arrangement. Now, the non-metals have usually 5, 6 or 7 electrons in the outermost shells of their atoms. So, all the non-metal atoms need electrons to achieve the inert gas structure. They get these electrons by mutual sharing. Thus, whenever a non-metal combines with another non-metal, covalent bond is formed.

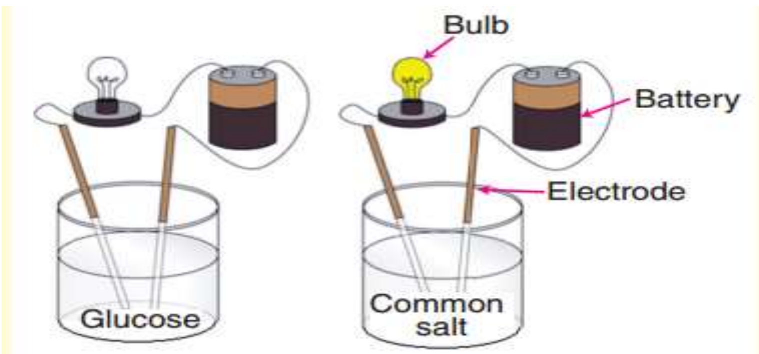
Properties of covalent compounds

You have observed the following properties of covalent compounds:
• Covalent compounds are usually liquids, gases or solids. For example, alcohol, benzene, water and cooking oil are liquids. Methane, ethane and chlorine are gases. Glucose, urea, and wax are solid covalent compounds.
• Covalent compounds have usually low melting points and low boiling points.
• Covalent compounds are usually insoluble in water, but they are soluble in organic solvents. Some of the covalent compounds like glucose, sugar and urea are soluble in water.

• Covalent compounds do not conduct electricity because they do not contain ions

Activity . detecting an ionic bond or covalent bond

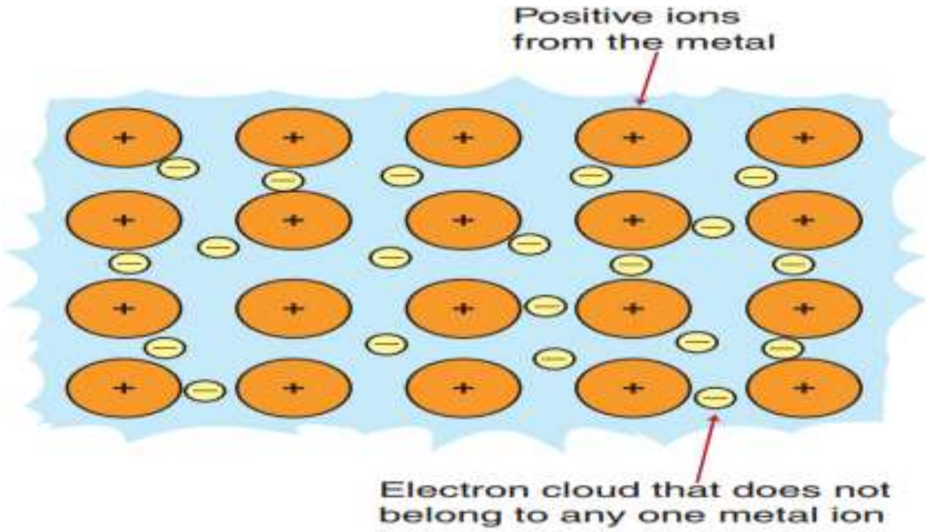
- Take the sample such as common salt (NaCl) provided.
- Try to dissolve it in water.
- If it dissolves, chances are it is likely to be an ionic compound. But, you already know some covalent compounds like glucose, urea and sugar are soluble in water.
- Now, perform electrical conductivity test.
- If the NaCl sample dissolves in water, arrange a circuit with two electrodes and a bulb.
- Figure out whether the bulb glows or not. According to your observation conclude the bond present in the sample.
- Make a report on the properties of ionic and covalent compounds.



Metallic bonding

The force which binds various metal atoms together is called **metallic bond**. The metallic bond is neither a covalent bond nor an ionic bond because these bonds are not able to explain properties of metals.
For example, metals are very good conductors of electricity but in solid state. Both ionic and covalent compounds cannot do so with the exception of graphite.

Formation of metallic bond.



- Description of molecules

  A Molecule

An **element** is a type of matter composed of atoms that all have the same atomic number. When two or more elements combine a compound is formed.
A **compound** is therefore defined as a pure substance made up of two or more elements chemically combined. Consider, Hydrogen peroxide (H2O2) is a compound composed of two elements, hydrogen and oxygen, in a fixed ratio (2:2)

A **molecule** is the smallest particle of a substance that retains the chemical and physical properties of the substance and is composed of two or more atoms.

Examples of some molecule and their chemical formula

1. Magnesium chloride **MgCl2**
2. Sodium chloride **NaCl**
3. Aluminum oxide **Al2O3**
4. Calcium nitrate **Ca(NO3)2**

✓ **Classification of molecules**

Molecules can be classified according to the chemical element which form them as: acid, base or alkali, salt or oxides.

Acid

An acid is a substance which when dissolved in water dissociates to give hydrogen ion(s) (H+) as the only positively charged ions.

Examples of acids and the ions they produce on dissociation include:

$$HCl\ (aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

$$HNO_3\ (aq) \longrightarrow H^+(aq) + NO_3^-(aq)$$

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

The hydrogen ion (H+) gives acids their characteristic properties.

Acids can either be commercial or natural.

Note: Hydrochloric acid is also found naturally in our stomachs.

Gases like carbon dioxide, hydrogen chloride and chlorine also show acidic properties when dissolved in water.

A base

A base is a substance which when dissolved in water dissociates to give hydroxide ions (OH-) as the only negatively charged ions.

Examples of bases include NaOH, KOH and Ca(OH)2. On dissociation these bases give the following ions:

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

$$KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$$

$$Ca\ (OH)_2(aq) \longrightarrow Ca^{2+}(aq) + 2OH^-(aq)$$

The hydroxide ion (OH–) gives bases their characteristic properties. A base that dissolves in water is called an **alkali**.

Salts

A salt is a compound formed when the hydrogen ions of an acid are wholly or partially replaced by a metal or ammonium ion. The process involves a **neutralization** reaction in which a base neutralizes an acid.

When all the hydrogen ions of an acid are replaced by a metal or ammonium radical, we get a **normal salt**. All chloride and nitrate salts are normal salts. When the hydrogen ions of an acid are partially replaced, we obtain an **acid salt**.

The number of hydrogen atoms in each molecule of an acid, replaceable directly or indirectly by a metal or ammonium radical, is called the **basicity** of that acid. Both hydrochloric and nitric acids contain only one replaceable hydrogen atom hence are said to be **monobasic**. Monobasic acids form one series of salts i.e normal salts.

Sulphuric and carbonic acids have two replaceable hydrogen atoms per molecule of the acid. They are **dibasic**. These acids form two series of salts.

(i) Where all hydrogen atoms of the acid are replaced i.e, normal salt.

Consider the following illustrations.

$$Zn(s) + H_2SO_4(aq) \longrightarrow \underset{\text{(normal salt)}}{ZnSO_4(aq)} + H_2(g)$$

Note that zinc sulphate has no hydrogen atom which can be replaced. It is therefor edescribed as a **normal salt**.

(ii)Where only a part of the hydrogen atoms of the acid are replaced i.e, acid salt.

Consider the following illustrations.

$$H_2SO_4(aq) + NaOH(aq) \longrightarrow \underset{\text{(acid salt)}}{NaHSO_4(aq)} + H_2O(l)$$

Note that sodium hydrogen sulphate salt still has one hydrogen atom, which can be replaced by a metal or ammonium ion. This is why it is known as an acid salt.

Names of salts are derived from the metal or ammonium ion from which they are formed and the parent acid. When naming salts the name starts with the metal or ammonium ion in the salt, followed by the respective acid radical.

Table Naming of salts from different acids

Acid and its formula	Class of salt	Type of salt	Examples
Hydrochloric acid (HCl)	Chlorides	Normal	Sodium Chloride Calcium Chloride
Nitric acid (HNO_3)	Nitrates	Normal	Potassium nitrate Lead(II) nitrate
Sulphuric acid (H_2SO_4)	Sulphates	Normal	Magnesium sulphate Lithium Sulphate
	Hydrogen sulphates	Acid	Sodium hydrogen sulphate
Carbonic acid (H_2CO_3)	Carbonates	Normal	Zinc carbonate Calcium carbonate
	Hydrogen carbonates	Acid	Calcium hydrogen Carbonate

Oxides

An oxide is a compound formed when an element react with oxygen.

Oxides can

be prepared by the following methods:

- Direct combination of an element with oxygen.
- Thermal decomposition of hydroxides, carbonates and nitrates.

Metals react with oxygen to produce metal oxide. Metal oxides are generally *basic* in nature.

$$\text{Metal} + \text{Oxygen} \longrightarrow \text{Metal Oxide}$$

Non-metals react with oxygen to produce non-metal oxide. Non-metal oxides are generally *acidic* in nature.

Non-metal + Oxygen $\longrightarrow$ Non-metal oxide

Classification of oxides

Depending upon the nature and the properties exhibited by these oxides, they are classified into:

- Acidic oxides
- Basic oxides
- Neutral oxides
- Amphoteric oxide

Acidic Oxides

Acidic oxides are oxides of non-metals. They turn blue litmus solution to red. Most acidic oxides are soluble in water. They react with water to produce an acid. All acidic oxides react with alkali to give salt and water.

Basic Oxides

Basic oxides are oxides of metals. They turn red litmus solution to blue. Most metal oxides are insoluble in water but some of these dissolve in water to form alkalis. They react with water to produce a base. All basic oxides react with acid to form salt and water

Neutral Oxide

Some non-metals react with oxygen to form neutral oxides. Neutral oxides do not show acidic nor basic characteristics. For example, carbon monoxide, nitric oxide, and nitrous oxide. These compounds are also called neutral compounds.

Amphoteric Oxides

Some metals react with oxygen to produce amphoteric oxides. Amphoteric oxides exhibit both acidic and basic characteristics.

These oxides react with acids as well as bases to form salt and water. Example of amphoteric oxides are aluminium oxide and zinc oxide.

exercises

1. ______ is a gaseous form of water.
2. Metals react with water to form metal oxide and hydrogen gas. (True or False)
3. Metals react with steam to form metal hydroxide and hydrogen gas. (True or False)
4. Complete and balance the following equations:
   (*i*) $Na(s) + H_2O(l) \longrightarrow ?$
   (*ii*) $Mg(s) + HCl(aq) \longrightarrow ?$
   (*iii*) $Cu(s) + H_2SO_4(aq) \longrightarrow ?$
   (*iv*) $Al(s) + Cl_2(g) \longrightarrow ?$
   (*v*) $K(s) + O_2(g) \longrightarrow ?$
5. Gold and Silver ______ react with dilute acids.
6. Hydrogen gas is not evolved when a metal reacts with ______ .
7. All metal chlorides are ionic in nature. (True or False)
8. Non-metals do not displace hydrogen gas from acids. (True or False)
9. Non-metals react with ______ to form covalent chlorides.
10. Non-metals react with oxygen to form (*a*) acidic oxides (*b*) neutral oxides (*c*) both (*a*) and (*b*) (*d*) none of these
11. Complete the following equations:
    (*i*) $H_2\ (g) + O_2\ (g) \longrightarrow ?$
    (*ii*) $C\ (s) + Cl_2\ (g) \longrightarrow ?$
12. Non-metal chlorides do not conduct electricity. (True or False)

- Calculation of molar mass and mole number

1. Molar mass

The mass of 1 mole of a substance is equal to its relative atomic or molecular mass in grams. The atomic mass of an element gives us the mass of one atom of that element in atomic mass units (u). To get the mass of 1 mole of atom of that element, we have to take the same numerical value but change the units from ‘u’ to ‘g’. The mass of atoms in grams is also known as gram atomic mass. For example, atomic mass of hydrogen = 1 u. So, gram atomic mass of hydrogen = 1 g.

1 u hydrogen has only 1 atom of hydrogen, 1 g hydrogen has 1 mole atoms, that is, 6.022×1023 atoms of hydrogen.

Molecular Mass (M)

The average mass of one molecule of a compound in atomic mass unit is called **molecular mass** (M). Hence,

Molecular mass (M) = M*r* × 1 u = M*r* u

The molecular mass (M) has the unit of mass, i.e., **g**, **kg** or **u**. Note that the magnitudes of molecular mass (M) and relative molecular mass (M*r*) are equal. They differ only in their units.

Example

Calculate the molecular mass of glucose ($C_6H_{12}O_6$).

Solution

Molecular mass of glucose

= (6 × atomic mass of C)+ (12 × atomic mass of H) + (6 × atomic mass of O)

= 6 × 12 u + 12 × 1 u + 6 × 16 u

= 72 u + 12 u + 96 u = **180 u**

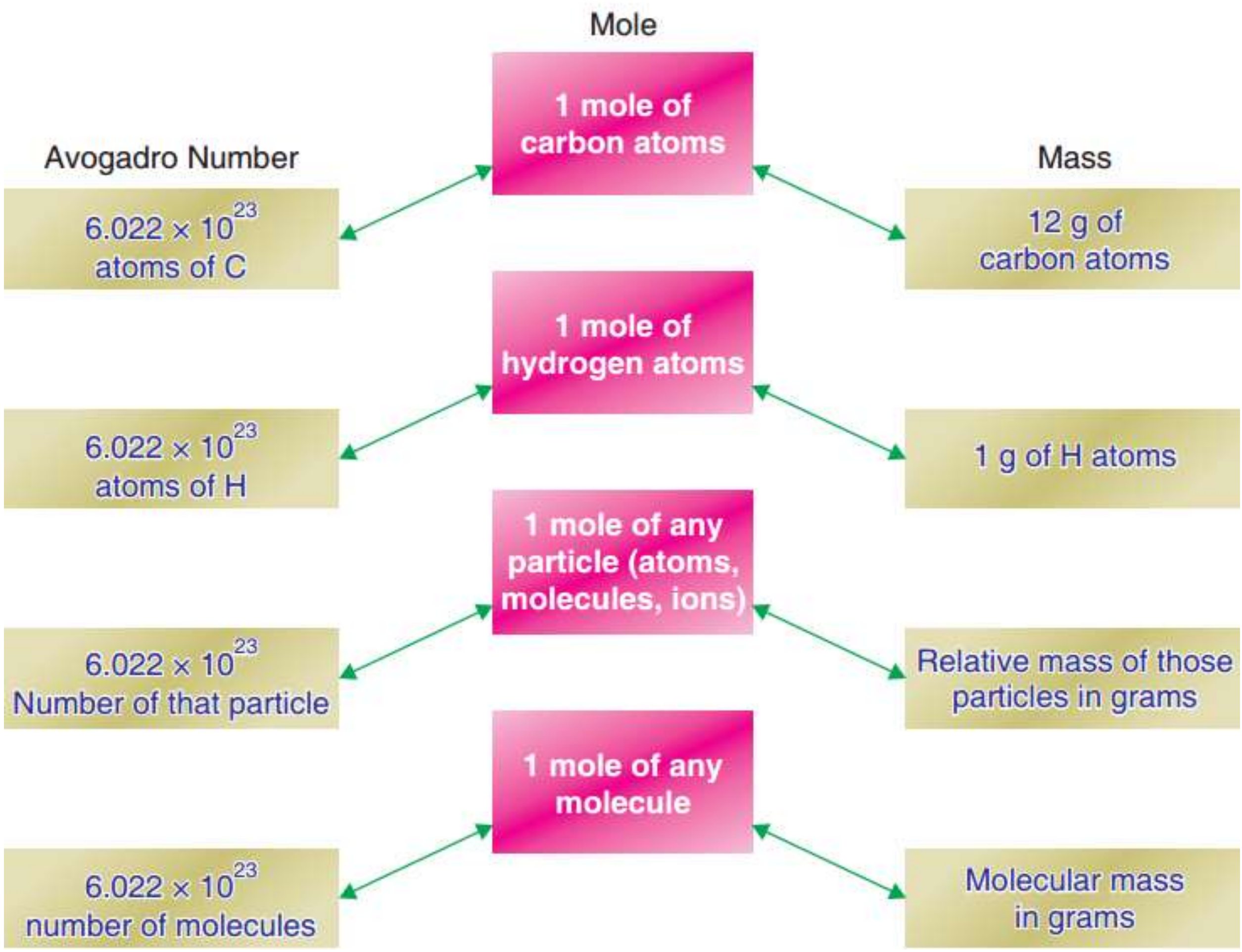
2. **Mole number**

A mole is defined as the amount of matter that contains as many objects (atoms, molecules, ions or whatever objects we are considering)as the number of atoms in exactly 12 g of pure ^{12}C (carbon-12 isotope).

A mole is a group of6.022×10^{23} particles (atoms, molecules or ions) of a substance. The SI symbol of mole is mol.

From numerous experiments, scientists have determined the number of atoms in 12 g of 12C (carbon-12) is found to be 6.022×1023. This number is called **Avogadro's number**, in honor of an Italian scientist Amedeo Avogadro. In this book, we will use 6.02×10^{23} or 6.022×10^{23} for Avogadro's number. Avogadro's number is represented by NA.

Figure 1.3.4.1: Relationship between Avogadro number, mole and mass.



Exercises

1. Avogadro number is the number of atoms in 1 gram-atom of an element.(True or False)
2. A group of 6.022 × 1023 particles (atoms, ions, or molecules) of a substance is called a______ of that substance.
3. One mole of CO2 contains _______ atoms of carbon.
4. The mass of one Avogadro number of nitrogen atoms is equal to:
(*a*) 14 amu (*b*) 14 g (*c*) 28 g (*d*) 14 kg
5. Mole is a link between the _________ of atoms (or molecules) and the _____ of atoms(or molecules).

Calculation of the number of mole

(*i*) The number of moles

$$= \frac{\text{Given no. of particles}}{\text{Avogadro number}}$$

(*ii*) The number of moles

$$= \frac{\text{mass of element in grams}}{\text{molar mass of element}}$$

(*iii*) The number of moles

$$= \frac{\text{mass of molecule in grams}}{\text{molar mass of the molecule}}$$

Example 1

Calculate the number of moles of 12.044×10^{23}helium atoms.

Solution

$$\text{No. of moles} = \frac{\text{No. of particles}}{\text{Avogadro number}}$$
$$= \frac{12.044 \times 10^{23}}{6.022 \times 10^{23}} = \mathbf{2\ mol}$$

Example 2

Calculate the number of particles in 0.1 mole ofcarbon atoms.

Solution

No. of particles = No. of moles× Avogadro number

$= 0.1 \times 6.022 \times 10^{23}$

$= \mathbf{6.022 \times 10^{22}}$ **moles**

Example 3

Calculate the number of moles of 3.011×10^{23}hydrogen atoms.

Solution

Number of moles

$$= \frac{\text{No. of particles}}{\text{Avogadro number}} \qquad = \frac{3.011 \times 100^{23}}{6.022 \times 10^{23}} = \mathbf{0.5\ mole}$$

Example 4

(a) Distinguish between N2 and 2N.

Calculate the mass of particles in the following:

(b) 0.5 mole of nitrogen atom

(c) 0.5 mole of nitrogen molecule

(Given: Atomic mass of nitrogen atom = 14 g

Molecular mass of nitrogen molecule = 28 g)

Solution

(*a*) N2 stands for 1 mole of nitrogen molecules and 2N stands for 2 moles of nitrogen atoms.

(*b*) Mass of nitrogen atoms = Atomic mass of nitrogen atom × number of moles

Mass of nitrogen atoms = ?

Atomic mass of nitrogen = 14 g

Molecular mass of N2 = 28 g

No. of mole = 0.5

= 14 × 0.5 = 7 grams

The mass of 0.5 mole of nitrogen atom is 7 grams.

(*c*) Mass of nitrogen molecules (N2)

= Molecular mass of molecules × no. of moles

= 28 × 0.5 = **14 grams**

Example 5

Calculate the no. of moles in

(i) 0.478 g of magnesium

(ii) 12 g oxygen gas

Given: Atomic mass of Mg = 24

Atomic mass of O = 16.

Solution

$$(i)\ \text{No. of moles} = \frac{\text{Mass of Mg in grams}}{\text{Atomic mass}}$$

$$= \frac{0.478}{24} = \mathbf{0.019\ mole}$$

(*ii*) In Oxygen gas, there are two atoms of Oxygen (O2)

Molecular mass of O2 = 16 × 2 = 32 g

$$\text{No. of moles in 12 g } O_2 = \frac{12}{32} = \mathbf{0.375\ mole}$$

Example 6

Calculate the mass of 6.022 × 1023 nitrogen molecules.

Solution

The mass of nitrogen molecules (N2) = 14 × 2 = 28 g
1 mole of N2 molecules = 6.022 × 10^2 molecules of nitrogen.
1 mole of N2 molecule weighs = 28 g
Therefore, 6.022 × 10^{23} molecules will weigh 28 grams

Learning Unit 2 – APPLY BASIC KNOWLEDGE OF INORGANIC CHEMISTRY

LO 2.1 – ACQUIRE THE KNOWLEDGE OF CHEMICAL EQUILIBRIUM

- Difference between indices and coefficient number

According to the law of conservation of matter, the total mass of the reactants in a chemical reaction is equal to the total mass of the products. Atoms are neither created nor can they be destroyed. They are only rearranged to form new substances.
This means that, chemical equations must always be balanced. In order to balance an equation, whole numbers called **coefficients** are placed in front of the whole symbol/formula of substances in the equation. While **indices** is the number of atoms of an element present in a molecule is written after the symbol of this element.

- Difference between reactants and products

Chemical equations

During chemical reactions, the starting substances **(reactants**) are changed into new substances (**products**). The changes that take place during chemical reactions are normally represented using chemical equations. Therefore, a chemical equation is a short hand way of representing a reaction. A chemical reaction can be represented by a word equation or by use of chemical symbols.

For example: A + B $\longrightarrow$ C + D where A and B are reactants and C and D are products.

The arrow means "to give".

A chemical equation is however complete if written with state symbols. It is only meaningful when it is balanced.

The equations are balanced in order to comply with the **law of conservation of matter.** The law states that ***matter cannot be created nor destroyed but can be changed from one state to another.***

This means that we must have the same number and types of atoms after a chemical change as were present before the chemical reaction took place.

Thus a chemical equation is a shorthand way of representing a chemical reaction.

More over, it is a convenient and efficient way to communicate chemical information.

It is used as an international code for describing chemical change

Example: The word equation for burning magnesium in oxygen therefore would be:

Magnesium + oxygen $\longrightarrow$ Magnesium oxide
(Reactants) (Product)

This word equation indicates that when magnesium and oxygen react, magnesium oxide is formed.

Activity

To observe what happens when a mixture of iron flings and sulphur powder are heated together

Materials and reagents

- Iron flings (grey in colour)
- Sulphur powder (yellow in colour)
- Test tube and test tube holder

- Bunsen burner

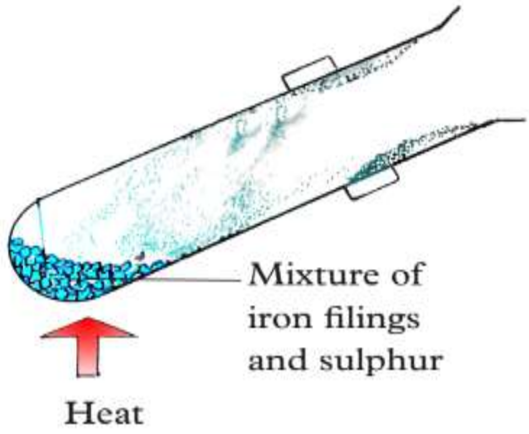
5. Spatula

Procedure

(This is a demonstration to be done by the teacher).

1. Put a spatula full of iron flings in the test tube.
2. Add another spatula full of sulphur powder and shake to get a homogenous mixture.
3. Using the test tube holder, heat the mixture until no further change occurs.
4. Record your observations.

Heating sulphur powder and iron flings



The easily observable features (or changes) which take place as a result of chemical reactions are known as **characteristics of chemical reactions**.

The important characteristics of chemical reactions are:

(*i*) Evolution of a gas

(*ii*) Formation of a precipitate

(*iii*) Change in colour

(*iv*) Change in temperature

(*v*) Change in state

Any one of these general characteristics can tell us whether a chemical reaction has taken place or not. For example, if on mixing two substances, any of the above characteristics occurs, then we can say that a chemical reaction has taken place.

- Types of reaction

There are various types of chemical reactions which exist depending on the classification. For our purpose we will examine a few:

- ✓ **Reversible reaction**

Chemistry normally distinguish between " **reversible**" reactions and those which **"go to complete or irreversible".**

Reversible reaction is a reaction capable of going in either direction of equation.

Example:

a. $HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$

b. $H_2PO_4^-(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + HPO_4^{2-}(aq)$

c. $NH_4^+(aq) + CN^-(aq) \rightleftharpoons HCN(aq) + NH_3(aq)$

d. $H_2SO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$

the sign $\rightleftharpoons$ means the reaction can proceed in either direction.

- ✓ **Irreversible**

A go to completion or irreversible reaction are those that go to completion all the reactants are turned into products. That is the reaction goes in one direction.

Examples: 1.

$$\underset{\text{Carbon monoxide}}{2CO(g)} + \underset{\text{Oxygen}}{O_2(g)} \xrightarrow{\text{Combination}} \underset{\text{Carbon dioxide}}{2CO_2(g)}$$

2.

$$\underset{\text{Hydrogen}}{H_2\,(g)} + \underset{\text{Chlorine}}{Cl_2\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Hydrogen chloride}}{2HCl\,(g)}$$

3.

$$\underset{\text{Potassium chloride}}{2KClO_3} \xrightarrow[\text{Decomposition}]{\text{Heat}} \underset{\text{Potassium chloride}}{2KCl(s)} + \underset{\text{Oxygen}}{3O_2(g)}$$

- Methods of equilibrium

To balance a chemical reaction means that the nature of the substances and their proportions are taken into account during the chemical reaction.

Activity

1. Put 1 kg mass on the left side of a weighing balance. Observe what happens.
2. Put some sand in a bag and place it on the right of the weighing balance. Observe what happens.

3. Now, add sand into the bag until the beam balances.
4. What can you conclude about the weights on both sides?
5. Repeat the procedure, but this time, remove the sand on the right hand side. What happens?
6. In pairs. Give reasons for the findings in the experiments above. How does this compare with balancing chemical equations? Discuss these with your friends in other pairs.
7. Now, look at this equation:

$$H_2SO_4 + NaOH \longrightarrow Na_2SO_4 + H_2O$$

Study questions

(i) How many sodium atoms are there on the right hand side of the equation? What about the left hand side?
(ii) How many hydrogen atoms are there on the right hand side? What about the left hand side?
(iii) Compare the number of sulphur and hydrogen atoms on both sides of the equation as well.
(iv) Are the atoms on the right hand and left hand sides balanced?
(v) How can the imbalance be corrected?

Law of conservation of matter and chemical equations

A correctly written chemical equation representing a chemical reaction obeys the law of conservation of matter. ***"The mass of all the reactants must be equal to the mass of all the products"***. Also during a chemical reaction “atoms are neither destroyed nor created, but they are all transformed ”. this is the law of Lavoisier, French chemist,(1743-1794).

Examples: 1. $Na(s) + Cl_2(g) \longrightarrow NaCl(s)$

1. *$H_2\,(g) + S(s) + O_{2(g)} \longrightarrow H_2SO_{4\,(aq)}$*
   This equations are unbalanced

Rules of balancing chemical equations

1. Write a word equation to show the reactants and products.
2. Write the correct symbol and formula under each reactant and each product.
3. Write the physical state of each substance after its symbol or formula.
4. Count the number of atoms of each element on the left-hand side of the equation (reactants) and on the right hand side (products). If the numbers of atoms of each element on each side of the equation are the same, then the equation is balanced.
5. If the numbers are not the same, then the chemical formulae must be multiplied by the lowest appropriate whole number in order to balance.
Balance by placing coefficients (whole numbers) in front of the formulae.
(**Note:** Do not change the actual formula).
6. The process of going through an equation from the left hand side to the right hand side, while ensuring that the number of each kind of atom on the reactants side equals the number of the same kind of atom on the products side is what we call balancing an equation.
7. Check the entire equation to ensure that all the atoms are balanced.
8. Make sure that the coefficients are in their lowest possible ratio.

Method to balance equation

To balance chemical equation we can use simple method:

- **Algebraic equilibrium (method of the Dumb coeficients)**
- **Randomly equilibrium (Tatonnement) by inspection**

✓ **Randomly equilibrium (Tatonnement) by inspection**

When guessing (putting coefficients on both side and verifying equality of the atoms) one can easily balance certain equations.

Example: $H_2 + Cl_2 \longrightarrow HCl$ unbalanced equation

To balance by guessing the equation in the reactants we have 2H and in the product we have only 1H. to have two H atoms in the products, we place 2 before the formula HCl and then the equation becames.

$$H_2 + Cl_2 \longrightarrow 2\,HCl \text{ balanced equation}$$

✓ **Algebraic equilibrium (method of the Dumb coefficients)**

Dumb coefficients are attributed and we proceed for the resolution of mathematical equations.

Example:1. $a\,MnO_2 + bHCl \longrightarrow c\,MnCl_2 + d\,H_2O + e\,Cl_2$

In order to have conservation of the number of atoms, one will for

Manganese (Mn) have : $a=c$: equation (1)

Oxygen (O): $2a= d$: equation (2)

Hydrogen (H): $b= 2d$: equation (3)

Chlorine (Cl): $b= 2c+ 2e$: equation (4)

We finally have 4 simple equation with five unknown coefficients. To solve this equation system, we attribute to one unknown coefficient a given value. For example let us suppose that $a = 1$

Thus the equation (1) give $a = c =1$
The equations become
Equation (2) gives $d = 2a = 2$
Equation (3) gives $b = 2d = 2 * 2 = 4$
Equation (4) Gives $b = 2c + 2e$ or $2e = b - 2c$
$2e = 4 - 2c$ as $b = 4$ then
$2e = 4 - 2 = 2$ as $c = 1$
$e = 1$

finally the chemical equation is written as :

$$1\ MnO_2 + 4\ HCl \longrightarrow 1MnCl_2 + 2\ H_2O + 1\ Cl_2$$

The coefficient one is not written and the equation becomes:

$$MnO_2 + 4\ HCl \longrightarrow MnCl_2 + 2\ H_2O + Cl_2$$

Example 2. $a\ Cr_2O_3 + b\ Na_2CO_3 + c\ KNO_3 \longrightarrow x\ Na_2CrO_4 + y\ CO_2 + z\ KNO_2$

Cr: $2a = x$ (1)

O: $3a + 3b + 3c = 4x + 2y + 2z$ (2)

Na : $2b = 2x$ (3)

C : $b = y$ (4)

K: $c = z$ (5)

N : $c = z$ (6)

Let us say that $a=1$, then $x= b=y=2$

Replace the known value in (2)

$3a + 3b + 3c = 4x + 2y + 2z$

$3 + 6 + 3c = 8 + 4 + 2z$

$3c + 9 = 2z + 12$ as $c = z$

$3c - 2c = 12-9$

$c = 3$ $z = 3$

if one replaces the letters by their values found , the equation becomes:

$$Cr_2O_3 + 2\ Na_2CO_3 + 3KNO_3 \longrightarrow 2\ Na_2CrO_4 + 2CO_2 + 3KNO_2$$

The coefficient 1 is not written.

We have check atom conservation and it is important to write the physical state of reacting substances.

(s): for solid state

(g): for the gaseous state

(l):for liquid state

(aq): for dissolved in water or aqueous

LO 2.2 – . DIFFERENTIATE ACIDS BASES AND AMPHOTERE

- Properties of acids and bases

a) Properties of acids

Activity Experiment to determine properties of acids

Apparatus and reagents

Dilute hydrochloric acid, litmus paper (blue and red), oranges, vinegar, lemon, paper, concentrated sulphuric acid, test tubes, magnesium ribbon, sour milk and nitric acid.

Procedure

1. Taste these substances and record your observations in a table.

Substance	Taste (bitter, sour)
Sour milk	
Oranges	
Vinegar	
Lemon	

2. Drop a piece of magnesium ribbon into a test tube containing dilute hydrochloric acid. Write down what you observe.
3. Place some concentrated sulphuric acid in a test tube. Drop a piece of paper in the acid. Record your observations.
4. Drop blue and red litmus papers in dilute hydrochloric acid contained in a test tube. What do you observe?
5. Add phenolphthalein and methyl orange indicators to nitric acid in different test tubes. What do you observe?

The following are properties of acids:

1. Acids have **sour** taste.
2. Acids turn blue litmus paper red.
3. A piece of paper placed in concentrated sulphuric acid gets charred. This is because acids are **corrosive**.

4. When magnesium ribbon is dropped in dilute hydrochloric acid, bubbles of a colorless gas are seen. This shows that dilute acids react with metals to produce hydrogen gas as one of the products.

b) Properties of alkali

ACTIVITY ***TO DETERMINE PROPERTIES OF BASES***

Apparatus and reagents

Test tubes, droppers, sodium hydroxide solution, dilute hydrochloric acid litmus papers, phenolphthalein indicator, methyl orange indicator.

Procedure

1. Add pieces of litmus paper to sodium hydroxide solution. Record your observations.
2. Add two drops of phenolphthalein to 2 cm3 of dilute sodium hydroxide solution. Record your observation. Repeat this experiment using methyl orange in place of phenolphthalein. Record your observations.
3. Add equal volumes of sodium hydroxide and hydrochloric acid in the same test tube. Test the resulting solution with red and blue litmus papers. What do you observe?

The following are properties of alkali:
1. Alkalis have a bitter taste.
2. Alkalis have a soapy feel.
3. Alkalis turn red litmus paper blue, methyl orange indicator yellow, and phenolphthalein indicator pink.

- PH of solution (Acidic, Neutral and alkaline)

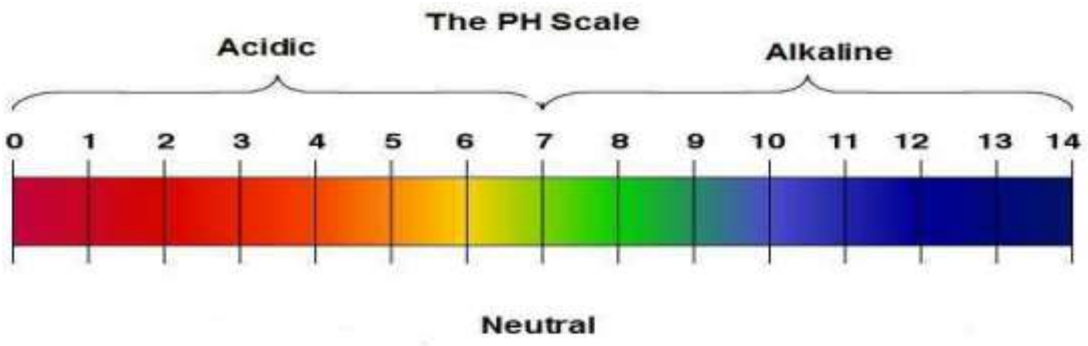
After cultivation, appropriate measures are taken, such as addition of lime so as to lower acid levels or addition of gypsum to increase acidity of the soil. The pH of a solution is a measure of the acidity or alkalinity of the solution.

The pH values for acids range from zero to just below seven. Substances such as rain water and lemon juice are considered acidic and have pH values, which range between 4 and 7.
They are said to be **weak acids**. Solutions of hydrochloric acid and sulphuric acid have pH values, which range between 0 and 4. These solutions are said to be **strong acids**. As the pH values decrease from 7 to 0, the strength of acids increases.
A pH value of 7 implies the solution is neither acidic nor basic and it is hence said to be **neutral**. Distilled water is neutral hence has a pH of 7. The pH values of bases range between 7 and 14. Ammonia solution and calcium hydroxide solution have pH values between 7 and 10 and are said to be **weak bases**. An example of a naturally occurring weak base is a wood ash. Sodium hydroxide and potassium hydroxide solutions have pH values above 10 and are said to be **strong bases**. As the pH values increase from 7 to14, the strength of the bases also increases.

The pH scale

The pH scale measures how acidic or basic a substance is. It has numbers ranging from 0 to 14. A pH of 7 shows that a solution is neutral while a pH below 7 shows that a solution is acidic. A pH higher than 7 indicates that a solution is basic.

Example of a standard pH colour chart



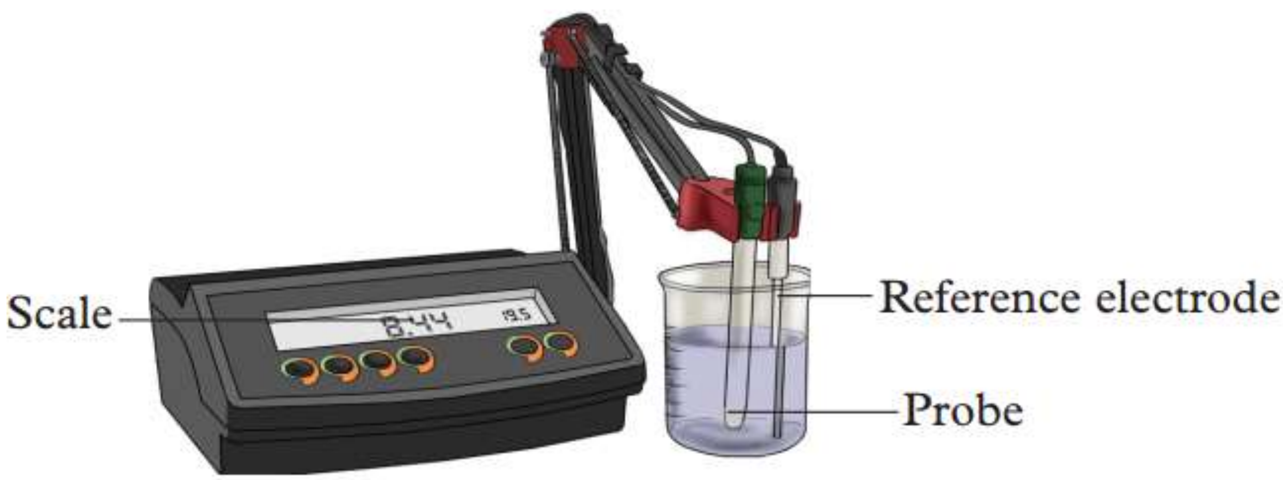
Using a pH-meter

Activity

Your teacher will bring a pH meter in class for you to observe.

1. Observe the pH meter carefully. (Be careful not to break the apparatus)
2. Compare it to the chart given by your teacher.
3. Draw and label the various parts.
4. Use the pH meter to determine the pH of the solutions provided by the teacher.
5. Use the readings to group the solutions as strong or weak acids and bases

A pH meter



A pH meter is used to make rapid and accurate measurements of pH of various solutions. A pH meter is made up of two electrodes one of which is the **reference electrode** connected to a multi-voltmeter and the second one called the **probe.**

Both are dipped into the solution being tested. The result is directly converted into pH and shown on the screen. The electrodes have to be rinsed in distilled water before being dipped into another solution of unknown pH for testing.

pH meters are usually used in hospitals to determine pH of blood and urine samples for diagnostic purposes.

- Some examples acids and bases

Examples of acids are:

(i) Commercial acids such as

- ✓ sulphuric acid
- ✓ hydrochloric acid
- ✓ nitric acid.

(ii) Natural acids such as

- ✓ Acetic acid (Vinegar)
- ✓ lactic acid (Milk)a
- ✓ citric acid(from citrus fruits and vegetables,)
- ✓ Ascorbic acid (vitamin C, as from certain fruits)

Examples of soluble bases (alkali) are

- ✓ Sodium hydroxide
- ✓ magnesium hydroxide
- ✓ potassium hydroxide among others.

LO 2.3 – DESCRIBE SOLUTION

- Definition of Solution

A solution is a **solution** is a special type of homogeneous mixture composed of two or more substances. In such a mixture, a solute is a substance dissolved in another substance, known as a solvent.

Solute + Solvent = Solution

- Properties of solute and solvents

a. **Properties of solvents**

A solvent (from the Latin solvō, "loosen, untie, solve") is a substance that dissolves a solute, resulting in a solution. A solvent is usually a liquid but can also be a solid, a gas, or a supercritical fluid. The quantity of solute that can dissolve in a specific volume of solvent varies with **temperature**.

Types of Solvents: The chemical classification of a solvent is based on its chemical structure.

- ***Hydrocarbon solvents*** are classified into three sub-groups based on the type of “carbon skeleton” of their molecules, giving us the aliphatic, aromatic and paraffinic solvents families. Paint thinner is a common example of a hydrocarbon solvent.
- ***Oxygenated solvents*** are produced through chemical reactions from olefins (derived from oil or natural gas), giving us the following sub-groups: alcohols, ketones, esters, ethers, glycol ethers and glycol ether esters. The human body naturally produces ketones when it burns fat.
- ***Halogenated solvents*** are solvents that contain a halogen such as chlorine, bromine or iodine. Many people recognize perchloroethylene as an example – a highly effective solvent used in dry cleaning.

Solvents can be divided as Polar and Non-Polar.

Polar solvents have high dielectric constant and have one or more electronegative atoms like N, H or O. Alcohols, ketones, carboxylic acids, and amides are the common examples of the functional group present in polar solvents. Polar solvents are made of polar molecules and can dissolve polar compounds only.

The Polar solvent is further divided as polar protic solvents and polar aprotic solvents. Water and methanol are polar protic molecules as they are capable of forming the hydrogen bond with the solutes. On the other hand, acetone is said as polar aprotic solvent as they are incapable of forming the hydrogen bond with the solute, but create dipole-dipole interactions with the ionic solutes.

Non-polar solvents contain bonds with similar electronegative atoms like C and H. These are made up non-polar molecules and can dissolve non-polar compounds or solutes.

Characteristics of the Solvent

- Solvent has the low boiling point and gets easily evaporate.
- Solvent exists as liquid only but can be solid or gaseous as well.
- The commonly used solvents contain the carbon element and hence called as organic solvents, while others are called as inorganic solvents.
- Solvents have characteristic color and odor.
- Acetone, alcohol, gasoline, benzene, and xylene are the commonly used organic solvents and are of great importance in chemical industries.
- Solvents are also used in regulating the temperature in a solution, either to absorb the heat generated during some chemical reaction or to enhance the speed of the reaction with the solute.

b. Properties of solute

Characteristics of the Solute

- Solute has higher **boiling points** than solvent.
- These can be solid, liquid or gas.
- By increasing the surface area of the particles of the solute, the **solubility** will increase. ...
- In case of gaseous solutes, the **solubility** is affected by the pressure, besides the volume and **temperature**.

Given below are the key differences between solute and the solvent:

1. **Solute** can be defined as the substance that gets dissolved by the solvent in a solution, while the substance that dissolves the solute is called as the **solvent**. Therefore the solute is present in the lesser amount than the solvent.
2. Solute can be found in solid, liquid or gaseous state, while the solvent is mainly found in the liquid state, but can be solid or in the gaseousstate as well.
3. The **boiling point** is higher of the solute than solvent. The properties of both solute and solvent are interdependent of each other.

- Properties of solution

Solutions are likely to have properties similar to those of their major component—usually the solvent. However, some solution properties differ significantly from those of the solvent. Here, we will focus on liquid solutions that have a solid solute, but many of the effects we will discuss in this section are applicable to all solutions.

Colligative Properties

Solutes affect some properties of solutions that depend only on the concentration of the dissolved particles. These properties are called colligative properties. Four important colligative properties that we will examine here are vapor pressure depression, boiling point elevation, freezing point depression, and osmotic pressure.

Vapor Pressure Depression

All liquids evaporate. In fact, given enough volume, a liquid will turn completely into a vapor. If enough volume is not present, a liquid will evaporate only to the point where the rate of evaporation equals the rate of vapor condensing back into a liquid. The pressure of the vapor at this point is called the vapor pressure of the liquid.

Boiling Point and Freezing Point Effects

A related property of solutions is that their boiling points are higher than the boiling point of the pure solvent. Because the presence of solute particles decreases the vapor pressure of the liquid solvent, a higher temperature is needed to reach the boiling point. This phenomenon is called boiling point elevation. For every mole of particles dissolved in a liter of water, the boiling point of water increases by about 0.5°C.

The presence of solute particles has the opposite effect on the freezing point of a solution. When a solution freezes, only the solvent particles come together to form a solid phase, and the presence of solute particles interferes with that process. Therefore, for the liquid solvent to freeze, more energy must be removed from the solution, which lowers the temperature. Thus, solutions have lower freezing points than pure solvents do. This phenomenon is called freezing point depression. For every mole of particles in a liter of water, the freezing point decreases by about 1.9°C.

Osmotic Pressure

The last colligative property of solutions we will consider is a very important one for biological systems. It involves osmosis, the process by which solvent molecules can pass through certain membranes but solute particles cannot. When two solutions of different concentration are present on either side of these membranes (called *semipermeable membranes*), there is a tendency for solvent molecules to move from the more dilute solution to the more concentrated solution until the concentrations of the two solutions are equal. This tendency is called osmotic pressure. External pressure can be exerted on a solution to counter the flow of solvent; the pressure required to halt the osmosis of a solvent is equal to the osmotic pressure of the solution.

Osmolarity (osmol) is a way of reporting the total number of particles in a solution to determine osmotic pressure. It is defined as the molarity of a solute times the number of particles a formula unit of the solute makes when it dissolves (represented by *i*):

$$\text{osmol} = M \times i$$

If more than one solute is present in a solution, the individual osmolarities are additive to get the total osmolarity of the solution. Solutions that have the same osmolarity have the same osmotic pressure. If solutions of differing osmolarities are present on opposite sides of a semipermeable membrane, solvent will transfer from the lower-osmolarity solution to the higher-osmolarity solution.

Types of solutions

A **saturated solution** contains the maximum amount of a solute that will dissolve in a given solvent at a specific temperature.
An **unsaturated solution** contains less solute than the solvent has the capacity to dissolve at a specific temperature.
A **supersaturated solution** contains more solute than is present in a saturated solution at a specific temperature.

- Some examples of solute, solvents and solution.

Some examples of solvents are water, ethanol, toluene, chloroform, acetone, milk, etc.

Examples of solutes include, sugar, salt, oxygen, etc.

There are numerous examples of solutions. For example milk (solvent) and sugar (solute) makes sweet milk. Air is made up of gases (oxygen, nitrogen, carbon dioxide, etc.). Alloys are made up of metals, for example, bronze is made up of copper and tin. River water contains water (solvent) and dissolved oxygen (solute).

- Solution preparation

Activity
To prepare 1mol dm–3 sodium chloride solution (1M NaCl solution).

Apparatus and chemicals

- Weighing balance
- 1dm3 volumetric flask
- Stirrer/glass rod
- Wash bottle
- Distilled water
- Beaker
- Funnel
- Sodium chloride

Procedure

1. Weigh 58.5 g of sodium chloride crystals.
2. Place 400 cm3 of distilled water in a beaker.
3. Add a little salt into the water, stirring continuously until all the salt dissolves.

4. Using a filter funnel transfer the salt solution into a 1 dm3 volumetric flask

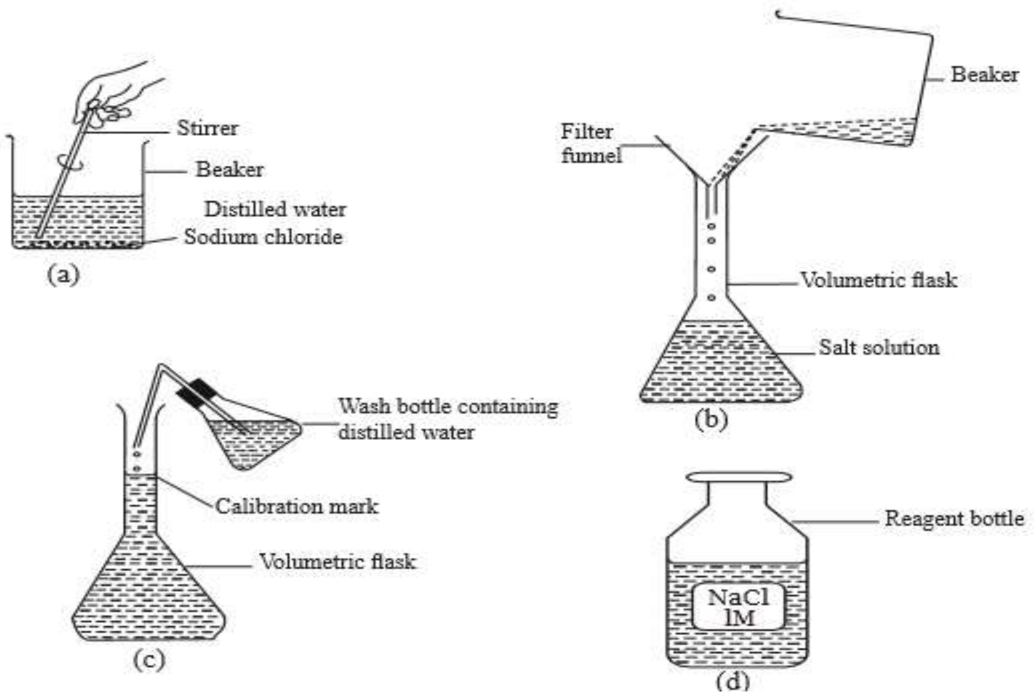
5. Rinse the beaker with distilled water carefully and add the washing into the flask.

6. Add more distilled water and shake the flask well. Add more water until the solution level is just below the calibration mark. Using a wash bottle add more distilled water drop by drop until the bottom of the meniscus is at the same level with calibration mark.

7. Use a stopper to cover the flask and invert it several times to make sure the solution mixes thoroughly.

8. Transfer the solution into a reagent bottle and label it 1M NaCl,

How to make 1M solution of sodium chloride



Study questions

1. Did all the salt dissolve in the water?

2. Write the formula of sodiumchloride.

3. Using the relative atomic masses, find the molar mass of sodium chloride.

The preparation of the solution requires a reagent that is so to say the quantity to be weighed for mass if the reagent is in solid state ; The volume to be pipetted using pipette if the solute is a liquid and then after dissolve it in water, So the solution can be prepared by two methods such as **dissolution method and dilution method**.

In the preparation of solution, glasses, volumetric flask, pipette, glass rod, measuring cylinder, analytical balance, spatula, beakers, magnetic stirrer and other laboratory devices are used.

The term concentration refers to the amount of solute dissolved in a specific volume of solvent. In a given amount of water, the more solute dissolved, the more concentrated the solution. If the solution contains a little solute, it is described as a dilute solution. Similarly, if a large volume of water is added to a small amount of solute, then the solution is dilute.

LO 2.4- CALCULATION OF CONCENTRATION (MOLALITY (m), NORMALITY, MOLARITY (M), PERCENTAGE).

- Definition of (Molality (m), Normality, Molarity (M), Percentage).

Many chemical reactions are performed with reagents in solution. It is necessary to know the amount of solute dissolved in a given amount of solution. If a large volume of water is added to a small amount of solute, the solution is said to be dilute. Similarly if the solution contains a larger amount of solute, the solution becomes more said to be concentrated.

Concentration of solutions

The term concentration refers to the amount of solute dissolved in a specific volume of solvent. In a given amount of water, the more solute dissolved, the more concentrated the solution. If the solution contains a little solute, it is described as a dilute solution.

Similarly, if the solution contains large amount of solute, then the solution becomes more concentrated.

How concentrations are expressed?

In consumer and industrial world, the most common method of expressing the concentration is based on the quantity of solute in a fixed quantity of solution. The quantities referred to here can be expressed in weight (w), volume (v) or both (that is the weight of a solute in a given volume of solution). In order to distinguish among these possibilities the abbreviations (w/w, v/v and w/v) are used. There are a number of ways to express the relative amounts of solute and solvent in a solution.

✓ **Percent composition (by mass)**

We can consider percent by mass (or weight percent, as it is sometimes called) in two ways:

• The parts of solute per 100 parts of solution.

• The fraction of a solute in a solution multiplied by 100. You need two pieces of information to calculate the percent by mass of a solute in a solution:

• The mass of the solute in the solution.

• The mass of the solution. Percent by mass is the mass of the solute divided by the mass of the solution (mass of solute plus mass of solvent), multiplied by 100.

✓ **Molarity (M)**

Molarity (M) tells us the number of moles of a solute in exactly one litre of a solution.

Note: Molarity is spelled with an "r" and is represented by a capital M.

We need two pieces of information to calculate the molarity of a solute in a solution:

- The moles of solute present in the solution.
- The volume of solution (in litres) containing the solute.

Molarity is probably the most commonly used unit of concentration. It is the number of moles of solute per litre of solution (not necessarily the same as the volume of solvent!).

✓ **Molality(m)**

Molality (*m*), or molal concentration, is the amount of a substance dissolved in a certain mass of solvent. It is defined as the moles of a solute per kilograms of a solvent.

Molarity vs molality

	Molarity (M)	**Molality (m)**
Measure of	Concentration	Concentration
Definition	The moles of a solute per liters of a solution	The moles of a solute per kilograms of a solvent
Units	*M*	*m*
Equation	*M* = moles solute / liters solution	*m* = moles solute / kg solvent
Ratio of moles to:	Volume (in liters)	Mass (in kilograms)

✓ **Normality**

Normality is a measure of concentration equal to the gram equivalent weight per liter of solution. Gram equivalent weight is the measure of the reactive capacity of a molecule. The solute's role in the reaction determines the solution's **normality**. **Normality** is also known as the equivalent concentration of a solution.

- Formula used for calculating: Molality (m), Normality , Molarity (M), Percentage.

✓ **Percent composition (by mass)**

Use the following equation to calculate percent by mass:

$$\text{Percent mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

Worked example

Determine the percent composition by mass of a 100 g salt solution which contains 20 g salt.

Solution

$$\frac{20\text{g}}{100\text{g solution}} \times 100\% = 20\%$$

✓ **Molarity (M)**

To calculate molarity we use the equation:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in litres}}$$

Worked example

What is the molarity of a solution made when water is added to 11g calcium chloride(CaCl2) to make 100 ml of solution? (Ca=20, Cl=35)

Solution

RMM of CaCl2 = 40 + (35 × 2)=110

$$\text{Moles} = \frac{\text{mass of } CaCl_2}{\text{RMM } CaCl_2} \text{ given} = \frac{11}{110} = 0.10 \text{ moles}$$

$$110 \text{ g} \longrightarrow 1 \text{ mole}$$

$$11 \text{ g} \longrightarrow x \text{ mole}$$

$$x = \frac{11 \times 1}{110} = 0.1 \text{ mole}$$

$$100 \text{ ml} \longrightarrow 0.1 \text{ mole}$$

$$1000 \text{ ml} \longrightarrow x$$

$$x = \frac{1000 \times 0.1}{100} = 1 \text{ M}$$

$$\therefore \text{molarity} = 1 \text{ M}$$

✓ **Molality(m)**

Molality formula and units

The units of morality are *m* or mol/kg.

Molality equation

C= moles solute / kilograms solvent

✓ **Normality**

Normality Formula

- Normality = Number of gram equivalents × [volume of solution in litres]$^{-1}$
- Number of gram equivalents = weight of solute × [Equivalent weight of solute]$^{-1}$
- N = Weight of Solute (gram) × [Equivalent weight × Volume (L)]
- N = Molarity × Molar mass × [Equivalent mass]$^{-1}$
- N = Molarity × Basicity = Molarity × Acidity

Normality is often denoted by the letter N. Some of the other units of normality are also expressed as eq L^{-1} or meq L^{-1}. The latter is often used in medical reporting.

How to Calculate Normality?

There are certain tips that students can follow to calculate normality.

1. The first tip that students can follow is to gather information about the equivalent weight of the reacting substance or the solute. Look up the textbook or reference books to learn about the molecular weight and the valence.
2. The second step involves calculating the no. of gram equivalent of solute.
3. Students should remember that the volume is to be calculated in litres.
4. Finally, normality is calculated using the formula and replacing the values.

Calculation of Normality in Titration

Titration is the process of gradual addition of a solution of a known concentration and volume with another solution of unknown concentration until the reaction approaches its neutralization. To find the normality of the acid and base titration:

$N_1 V_1 = N_2 V_2$

Where,

- N_1 = Normality of the Acidic solution
- V_1 = Volume of the Acidic solution
- N_2 = Normality of the basic solution
- V_3 = Volume of the basic solution

Normality Equations

The equation of normality that helps to estimate the volume of a solution required to prepare a solution of different normality is given by,

Initial Normality (N_1) × Initial Volume (V_1) = Normality of the Final Solution (N_2) × Final Volume (V_2)

Relation Between Normality and Molarity

The formula of molarity is given as:

⇒ Molarity (M) = No. of moles of solute × [volume of the solution in litres]$^{-1}$

⇒ Normality = [Molarity × Molar mass] × [Equivalent mass]$^{-1}$

⇒ Normality = Molarity × Basicity

⇒ Normality = Molarity × Acidity

⇒ N = M × number of equivalents

Examples

Question 1. In the following reaction calculate and find the normality when it is 1.0 M H_3PO_4

H_3AsO_4 + 2NaOH → Na_2HAsO_4 + $2H_2O$

Solution:

If we look at the given reaction we can identify that only two of the H^+ ions of H_3AsO_4 react with NaOH to form the product. Therefore, the two ions are 2 equivalents. In order to find the normality, we will apply the given formula.

N = Molarity (M) × number of equivalents

N = 1.0 × 2 (replacing the values)

Therefore, normality of the solution = 2.0.

Question 2. Calculate the normality of 0.321 g sodium carbonate when it is mixed in a 250 mL solution.

Solution:

First, you have to know or write down the formula for sodium carbonate. Once you do this you can identify that there are two sodium ions for each carbonate ion. Now solving the problem will be easy.

N of 0.321 g sodium carbonate

N = Na_2CO_3 × (1 mol/105.99 g) × (2 eq/1 mol)

N = 0.1886 eq/0.2500 L

N = 0.0755 N

Differences Between Normality and Molarity

Normality	Molarity
Also known as equivalent concentration.	Known as molar concentration.
It is defined as the number of gram equivalent per litre of solution.	It is defined as the number of moles per litre of solution.
It is used in measuring the gram equivalent in relation to the total volume of the solution.	It is used in measuring the ratio between the number of moles in the total volume of the solution.
The units of normality are N or eq L^{-1}	The unit of molarity is M or Moles L^{-1}

- Application of (Molality (m), Normality, Molarity (M), Percentage). Solution Dilution

✓ **Uses of Normality**

Normality is used mostly in three common situations:

- In determining the concentrations in acid-base chemistry. For instance, normality is used to indicate hydronium ions (H^3O^+) or hydroxide ions (OH^-) concentrations in a solution.

- Normality is used in precipitation reactions to measure the number of ions which are likely to precipitate in a specific reaction.
- It is used in redox reactions to determine the number of electrons that a reducing or an oxidizing agent can donate or accept.

Limitations in Using Normality

Many chemists use normality in acid-base chemistry to avoid the mole ratios in the calculations or simply to get more accurate results. While normality is used commonly in precipitation and redox reactions there are some limitations to it. These limitations are as follows:

- It is not a proper unit of concentration in situations apart from the ones that are mentioned above. It is an ambiguous measure and molarity or molality are better options for units.
- Normality requires a defined equivalence factor.
- It is not a specified value for a particular chemical solution. The value can significantly change depending on the chemical reaction. To elucidate further, one solution can actually contain different normalities for different reactions.

Examples.

question 1. What will the concentration of citric acid be if 25.00 ml of the citric acid solution is titrated with 28.12 mL of 0.1718 N KOH?

Solution:

$N_a \times V_a = N_b \times V_b$

Na × (25.00 mL) = (0.1718N) (28.12 mL)

Therefore, the concentration of citric acid = 0.1932 N.

Question 2. Find the normality of the base if 31.87 mL of the base is used in the standardization of 0.4258 g of KHP (eq. wt = 204.23)?

Solution:

0.4258 g KHP × (1 eq/204.23g) × (1 eq base/1eq acid):

= 2.085×10^{-3} eq base/0.03187 L = 0.6542 N

Normality of the base is = 0.6542 N.

Question 3. Calculate the normality of acid if 21.18 mL is used to titrate 0.1369 g Na2CO3?

Solution:

0.1369 g Na_2CO_3 × (1 mol/105.99 g) × (2 eq/1 mol) × (1 eq acid/1 eq base):

= 2.583×10^{-3} eq acid/0.02118 L = 0.1212 N

Normality of the acid = 0.1212 N.

⇒ **Try this:**

Question4 : What is the concentration of aluminium in a 3.0 M solution of aluminium sulfate?

Answer: 6.0 M Al_3^+

- ✓ **use of Molality**

Molality is the favoured concentration transmission approach because the solution's mass of solute and solvent does not change.

Concentrations expressed in **molality** are used when studying properties of solutions related to vapor pressure and temperature changes. **Molality** is used because its value does not change with changes in temperature. The volume of a solution, on the other hand, is slightly dependent upon temperature

Examples

Question: Calculate the molality of a solution where 0.5 grams of toluene (C_7H_8) is dissolved in 225 grams of

Benzene (C_6H_6) .

Calculate the moles of given solute.

Toluene – Molecular weight = $C_7H_8 = 7 \times 12 \times +1 \times 8 = 92 moles/gram$

Using the formula:

Moles of Toluene = $\frac{Mass\ in\ grams}{Molecular\ weight}$ = 0.054 mole.

So, the mole of toluene is 0.054 mole.

Now calculate the kilogram of solvent.

$\frac{225\ grams\ of\ Benzene}{1000}$ = 0.225 kilogram

As the final step, calculate the molality using the formula.

Molality (m) = $\frac{Moles\ of\ Toluene}{Mass\ of\ Benzene\ in\ grams}$

Molality (m) = $\frac{0.054 moles}{0.225 kg}$

Molality = 0.24 m

- ✓ **use of Molarity**

Molarity (M) indicates the number of moles of solute per liter of solution (moles/Liter) and is one of the most common units used to measure the concentration of a solution. Molarity can be used to calculate the volume of solvent or the amount of solute.

The relative amount of oxygen in a planet's atmosphere determines its ability to sustain aerobic life. The relative amounts of iron, carbon, nickel, and other elements in steel (a mixture known as an "alloy") determine its physical strength and resistance to corrosion. The relative amount of the active ingredient in a medicine determines its effectiveness in achieving the desired pharmacological effect. The relative amount of sugar in a beverage determines its sweetness

- ✓ **application of percentage**

Percentages are often used for calculations involving money in daily life. The prices of most quantities increase (or decrease) by a percentage over time. Profit, loss, discount, commission and rates of interest can all be expressed as percentages

Examples

1. Mother's weight is 25 % more than that of daughter. What percent is daughter's weight less than mother's weight?

Solution:

Let daughter's weight be 100 kg.

Then mother's weight = (100 + 25) kg = 125 kg

If mother's weight is 125 kg, then daughter's weight is 100 kg.

If mother's weight is 1 kg, then daughter's weight is 100/125 kg

If mother's weight is 100 kg, then daughter's weight = (100/125 × 100) kg

Therefore, daughter's weight is 20 % less than that of mother.

2. Kelly requires 36 % to pass. She gets 196 marks and falls short by 20 marks. Find the maximum numbers she could have got.

Solution:

Tanya had scored 20 marks more, she could have scored 36%.

Now, 20 marks more than 196 marks.

Therefore, 196 + 20 = 216

Let maximum marks be x, then 36% of m = 216

⇒ 36/100 × m = 216

⇒ m = 216 × 100/36 ⇒ m = 600

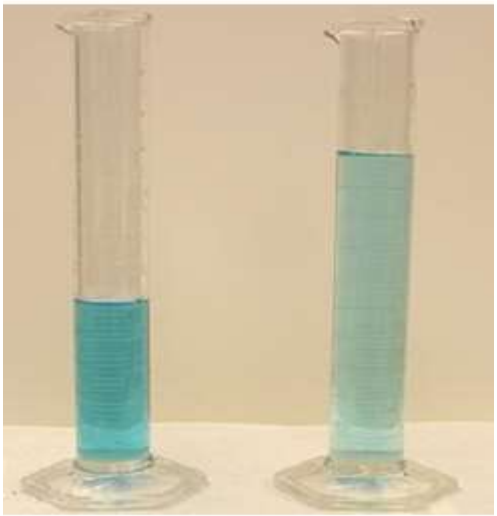
Thus, the maximum marks she could have got were 600.

- Dilution of Solutions

Dilution is the process whereby the concentration of a solution is lessened by the addition of solvent. For example, we might say that a glass of iced tea becomes increasingly diluted as the ice melts. The water from the melting ice increases the volume of the solvent (water) and the overall volume of the solution (iced tea), thereby reducing the relative concentrations of the solutes that give the beverage its taste.

Dilution is also a common means of preparing solutions of a desired concentration. By adding solvent to a measured portion of a more concentrated *stock solution*, we can achieve a particular concentration. For example, commercial pesticides are typically sold as solutions in which the active ingredients are far more concentrated than is appropriate for their application. Before they can be used on crops, the pesticides must be diluted. This is also a very common practice for the preparation of a number of common laboratory reagents

Figure below *Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent.*



A simple mathematical relationship can be used to relate the volumes and concentrations of a solution before and after the dilution process. According to the definition of molarity, the molar amount of solute in a solution is equal to the product of the solution's molarity and its volume in liters:

n=ML

Expressions like these may be written for a solution before and after it is diluted:

n1=M1L1 n2=M2L2

where the subscripts "1" and "2" refer to the solution before and after the dilution, respectively. Since the dilution process *does not change the amount of solute in the solution,* $n_1 = n_2$. Thus, these two equations may be set equal to one another: M1L1=M2L2

This relation is commonly referred to as the dilution equation. Although we derived this equation using molarity as the unit of concentration and liters as the unit of volume, other units of concentration and volume may be used, so long as the units properly cancel per the factor-label method. Reflecting this versatility, the dilution equation is often written in the more general form:

C1V1=C2V2   where CC and VV are concentration and volume, respectively.

Examples 1. Determining the Concentration of a Diluted Solution

If 0.850 L of a 5.00-*M* solution of copper nitrate, $Cu(NO_3)_2$, is diluted to a volume of 1.80 L by the addition of water, what is the molarity of the diluted solution?

Solution

We are given the volume and concentration of a stock solution, V_1 and C_1, and the volume of the resultant diluted solution, V_2. We need to find the concentration of the diluted solution, C_2. We thus rearrange the dilution equation in order to isolate C_2:

C1V1=C2V2

$$C_2 = \frac{C_1 V_1}{V_2}$$

Since the stock solution is being diluted by more than two-fold (volume is increased from 0.85 L to 1.80 L), we would expect the diluted solution's concentration to be less than one-half 5 *M*. We will compare this ballpark estimate to the calculated result to check for any gross errors in computation (for example, such as an improper substitution of the given quantities). Substituting the given values for the terms on the right side of this equation yields:

$$C_2 = \frac{0.850\ \mathrm{L} \times 5.00\ \dfrac{\mathrm{mol}}{\mathrm{L}}}{1.80\ \mathrm{L}} = 2.36\ M$$

This result compares well to our ballpark estimate (it's a bit less than one-half the stock concentration, 5 *M*).

Learning Unit 3 – APPLY BASIC KNOWLEDGE OF ORGANIC CHEMISTRY

LO 3.1 CLASSIFY ORGANIC COMPOUNDS

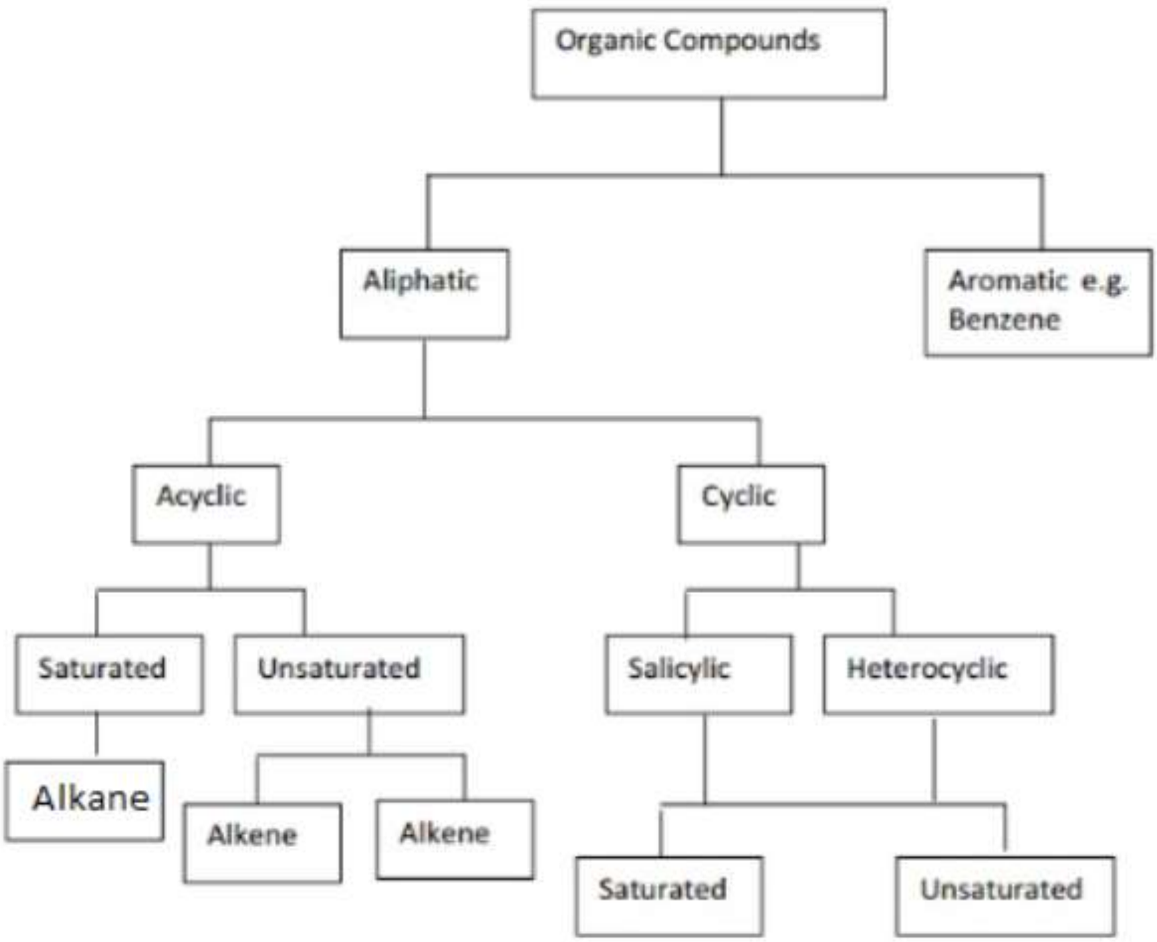
Introduction

Organic chemistry is defined as the study of the compounds mainly composed by carbon and hydrogen atoms, and sometimes oxygen, nitrogen, phosphorus, sulphur and halogens atoms. The study of the rest of the elements and their compounds falls under the group of inorganic chemistry. However, there are some exceptions such as carbonates, cyanides, carbides, carbon oxides, carbonic acid, carbon disulphide which are considered as inorganic compounds. Since various organic compounds contained carbon associated with hydrogen, they are considered as derived from hydrocarbons. Thus, a more precise definition of organic chemistry is: “the study of hydrocarbons and the compounds which could be thought of as their derivatives’’.

The organic and inorganic compounds can be differentiated based on some of their properties as summarised in the following table.

Organic compounds	Inorganic compounds
Form covalent bond	Most form ionic bond
Classified as alcohols, aldehydes, carboxylic acids, etc. with characteristic properties	Exist as acids, bases and salts
Lower melting and boiling points	Higher melting and boiling points
Insoluble in water but soluble in organic solvents such as n-hexane, ethanol, acetone…	Less soluble in water and insoluble in organic solvents
Highly volatile and inflammable	Not volatile and not inflammable
Reactions are generally slow	Reactions are generally fast
Exhibit the phenomenon of isomerism	Very few isomers

<u>Classifcation of organic compounds</u>



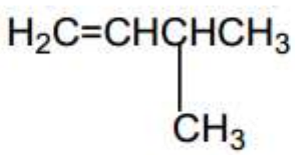
1. **Aliphatic compounds**

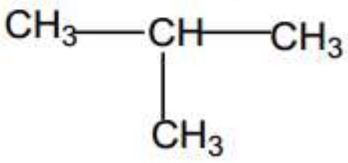
Aliphatic compounds are organic compounds in which the carbon atoms are arranged in a straight or branched chain.

Examples

1) CH3-CH2-CH2-CH3

2) CH3-CH=CH-CH3

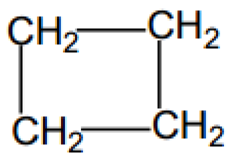
3) 

4) 

2. Alicyclic compounds

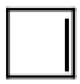
Alicyclic compounds are organic compounds that contain one or more carbon rings that may be saturated or unsaturated.

Example: 1) cyclobutane



2) Cyclopentane

3) Cyclobutene



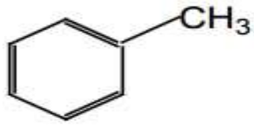
3. Aromatic compounds

Aromatic compounds are compounds that contain a closed ring that consists of alternating single and double bonds with delocalised pi electrons.

Example:

1. Benzene (C6H6)

2.Toluene (C6H5-CH3)



Note: Heterocyclic compounds: Are also classified as cyclic compounds which include one or two atoms other than carbon (O, N, S) in the ring. Thus furan, thiophene and pyridine are heterocyclic compounds.

- Functional group

In **organic chemistry**, **functional groups** are specific substituents or moieties within molecules that may be responsible for the characteristic **chemical** reactions of those molecules. The same **functional group** will undergo the same or similar **chemical l**reaction(s) regardless of the size of the molecule it is a part of.

✓ **Alkanes**

Alkanes are the simplest class of organic compounds. They are made of carbon and hydrogen atoms only and contain two types of bonds, carbon-hydrogen (C-H) and carbon-carbon (C-C) single covalent bonds. They do not have functional groups. Alkanes form a homologous series with the general formula C_nH_{2n+2} where n is the number of carbon atoms in the molecule. The first member of the family has the molecular formula CH_4 (n=1) and is commonly known as methane and the second member with molecular formula is C_2H_6 (n=2) is called ethane.

These compounds are also known as saturated hydrocarbons. This name is more descriptive than the term “alkane’’ because both their composition (carbon and hydrogen) and the fact that the four single covalent bonds of each carbon in their molecules are fully satisfied or ‘’saturated’’.

The name alkane is the generic name for this class of compounds in the IUPAC system of nomenclature. These hydrocarbons are relatively unreactive under ordinary laboratory conditions, but they can be forced to undergo reactions by drastic treatment. It is for this reason that they were named paraffins (Latin parum affinis = little activity).

✓ **Alkene**

Activity

1. Describe the formation of a carbon-carbon double bond. What is the hybridisation state of a carbon doubly bonded?
2. What is the shape of the molecule around the double bond? Explain.

Alkenes are a homologous series of hydrocarbons which contain a carbon-carbon double bond. Since their skeleton can add more hydrogen atoms, they are referred as unsaturated hydrocarbons.
The general formula of alkenes is C_nH_{2n}.

✓ **Alkynes**

Activity

1. Explain the formation of a carbon-carbon triple bond.
2. What is the hybridisation state of a carbon atom triply bonded and what is the shape of the structure around it.
3. Differentiate between the following compounds
4. $HC\ CCH_2CH_2CH_2CH_3$ and $CH_3CH_2C\ CCH_2CH_3$

A triple bond consists of one sigma bond and two pi bonds. Each carbon of the triple bond uses two sp orbital to form sigma bonds with other atoms. The unhybridised 2p orbitals which are perpendicular to the axes of the two sp orbitals overlap sideways
to form pi bonds.
According to the VSEPR model, the molecular geometry in alkynes include bond angle of 180o around each carbon triply bonded.Thus, the shape around the triple bond is linear.

There are two types of alkynes: terminal alkynes and non-terminal (internal) alkynes
A terminal alkyne has a triple bond at the end of the chain e.g.: : R-C ≡ *C* - *H*
A non-terminal alkyne has a triple bond in the middle of the chain: *R* - *C* ≡ *C* - *R*'
Examples
HC ≡CHCH2CH2CH3, a terminal alkyne
CH3C≡ CCH2CH3, a non-terminal alkyne

- ✓ **Alcohol**

Activity

1. Look at the following compounds and classify them in their homologous series.
a. CH3CH2CH2CH2CH3
b. CH3CH=CHCH2CH3
c. CH3CHOHCH2CH3
d. CH3CH2-O-CH3
e. CH3CH2CH2CH2CH2OH
f. (CH3)2CHCHOHCH3
2. By doing your own research, distinguish the rules used to name alcohol compounds.

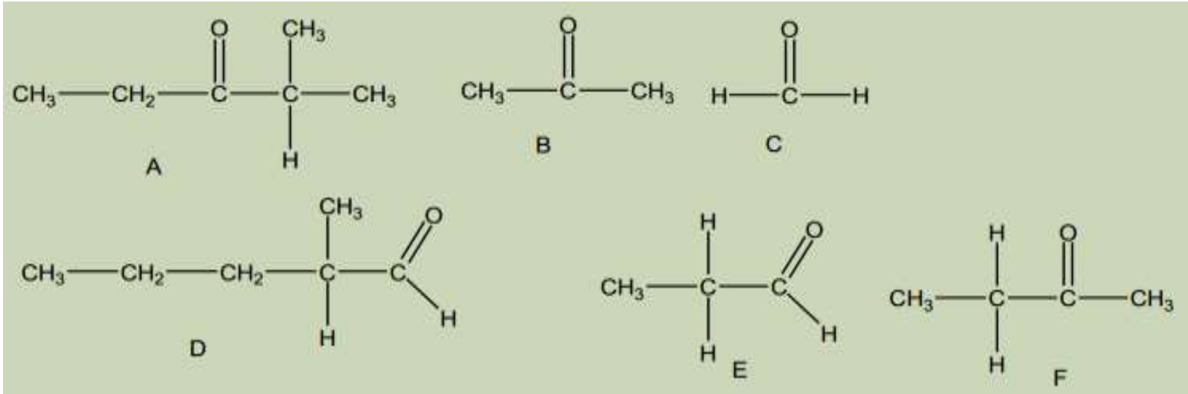
Alcohols are organic compounds that are derivatives of hydrocarbons where one or more hydrogen atoms of hydrocarbon is or are replaced by hydroxyl (-OH) group.
They are represented by the general formula: **CnH2n+1OH** or **ROH** where R is a radical: alkyl group made by a chain of carbon atoms.
Alcohols are called monohydric if only one hydroxyl group is present (eg: CH3CH2-OH). Dihydric alcohols are those with two hydroxyl group (diol: vicinal and gem), trihydric (triols) and polyhydric are those with many – C-OH groups. The functional group attached is –OH group to any atom of carbon.

- ✓ **carbonyl compounds: aldehydes and ketones**

Carbonyl compounds are compounds that contain carbon-oxygen double bond (C=O). Carbonyl compounds are classifed into two general categories based on the kinds of chemistry they undergo. In one category there are aldehydes and ketones; in the other category there are carboxylic acids and their derivatives. This unit looks on category of aldehydes and ketones.

Activity

Observe the following molecules and answer to the questions.



1. Categorize the above molecules

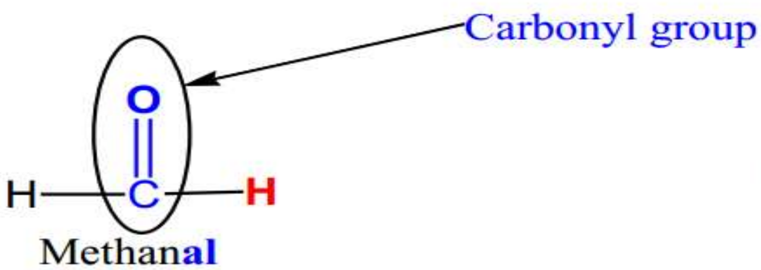
2. What criteria have you used to categorize?

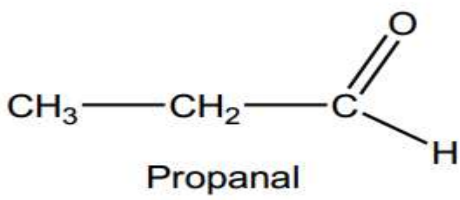
3 Name those categories

4. Name individual molecules

- ✓ **Aldehydes**

For aldehydes, the carbonyl group is attached to hydrogen atom and alkyl group as shown in the molecule of propanal below. Methanal is the smallest aldehyde, it has two hydrogen atoms attached to carbonyl group

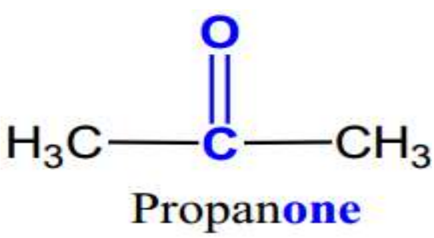




If you are going to write this in a condensed form, you write aldehyde as –CHO, **don't write it** as -COH, because that looks like an alcohol functional group.

- ✓ **Ketones**

Ketone has two alkyl groups attached to the carbonyl group. Examples:



Important: ketones don't have a hydrogen atom attached to the carbonyl group

- Nomenclature of organic compound

✓ **alkanes**

IUPAC Rules for the nomenclature of alkanes

a. Find and name the longest continuous carbon chain.

b. Identify and name groups attached to this chain.

c. Number the chain consecutively, starting at the end nearest a substituent group.

d. Designate the location of each substituent group by an appropriate number and name.

e. Assemble the name, listing groups in alphabetical order.

The saturated hydrocarbon form homologous series (series in which members have similar chemical properties and each differs from the preceding by a methylene group –CH2-).

The first four members are known by their common names, from C5 and above the Roman prefixes indicating the number of carbon atoms is written followed by the ending "ane" of the alkanes.

Note: Alkyl groups are obtained when one hydrogen atom is removed from alkanes; therefore their names are deduced from the corresponding alkanes by replacing "ane" ending with "yl" desinence

Examples of naming alkanes

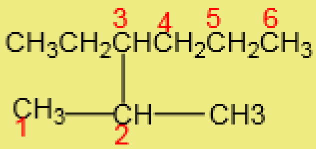
n	Name of residue	R-H	Alkane	alkyl R	Residue	Abbreviation
1	Meth	ane	CH_4	-yl	CH_3-	Me
2	Eth	ane	CH_3CH_3	-yl	CH_3CH_2-	Et
3	Prop	ane	$CH_3CH_2CH_3$	-yl	$CH_3CH_2CH_2$-	Pr
4	But	ane	$CH_3(CH_2)_2CH_3$	-yl	$CH_3(CH_2)_2CH_2$-	Bu
5	Pent	ane	$CH_3(CH_2)_3CH_3$	-yl	$CH_3(CH_2)_3CH_2$-	Pe
6	Hex	ane	$CH_3(CH_2)_4CH_3$	yl	$CH_3(CH_2)_4CH_2$-	Hex
7	Hept	ane	$CH_3(CH_2)_5CH_3$	yl	$CH_3(CH_2)_5CH_2$-	Hep
8	Oct	ane	$CH_3(CH_2)_6CH_3$	yl	$CH_3(CH_2)_6CH_2$-	Oct
9	Non	ane	$CH_3(CH_2)_7CH_3$	yl	$CH_3(CH_2)_7CH_2$-	Non
10	Dec	ane	$CH_3(CH_2)_8CH_3$	yl	$CH_3(CH_2)_8CH_2$-	Dec
11	Undec	ane	$CH_3(CH_2)_9CH_3$	yl	$CH_3(CH_2)_9CH_2$-	-
12	Dodec	ane	$CH_3(CH_2)_{10}CH_3$	yl	$CH_3(CH_2)_{10}CH_2$-	-
13	Tridec	ane	$CH_3(CH_2)_{11}CH_3$	yl	$CH_3(CH_2)_{11}CH_2$-	-
14	Tetradec	ane	$CH_3(CH_2)_{12}CH_3$	yl	$CH_3(CH_2)_{12}CH_2$-	-
15	Pentadec	ane	$CH_3(CH_2)_{13}CH_3$	yl	$CH_3(CH_2)_{13}CH_2$-	-
16	Hexadec	ane	$CH_3(CH_2)_{14}CH_3$	yl	$CH_3(CH_2)_{14}CH_2$-	-
17	Heptadec	ane	$CH_3(CH_2)_{15}CH_3$	yl	$CH_3(CH_2)_{15}CH_2$-	-
18	Octadec	ane	$CH_3(CH_2)_{16}CH_3$	yl	$CH_3(CH_2)_{16}CH_2$-	-
19	Nonadec	ane	$CH_3(CH_2)_{17}CH_3$	yl	$CH_3(CH_2)_{17}CH_2$-	-
20	icos	ane	$CH_3(CH_2)_{18}CH_3$	yl	$CH_3(CH_2)_{18}CH_2$-	-

Note: n is the number of carbon atoms

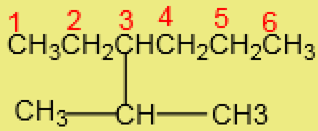
Prefixes di, tri, tetra, sec, tert, are not considered when alphabetizing.

f. In case of chains of the same length, the priority is given for part where many branched of alkyl groups appear.

Example



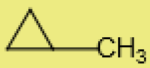
3-ethyl-2-methylhexane
(two substituents)



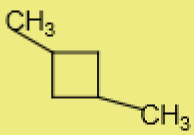
not
3-isopropylhexane
(one substituent)

g. For cyclanes or cycloalkanes, the prefix "cyclo" is recommended, followed by the name of the alkanes of the same carbon number. But in case of ramified cyclanes, the priority is for the ring.

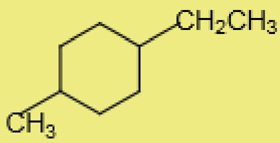
Examples:



methylclopropane



1,3-dimethylcyclobutane



1-ethyl-4-methylcyclohexane

✓ **Nomenclature of alkenes**

IUPAC names of alkenes are based on the longest continuous chain of carbon atoms that contains the double bond.

The name given to the chain is obtained from the name of the corresponding alkane by changing the suffix from –**ane** to –**ene**.

If the double bond is equidistant from each end, number the first substituent that has the lowest number.

If there is more than one double bond in an alkene, all of the bonds should be numbered in the name of the molecule, even terminal double bonds. The numbers should go from lowest to highest, and be separated from one another by a comma.

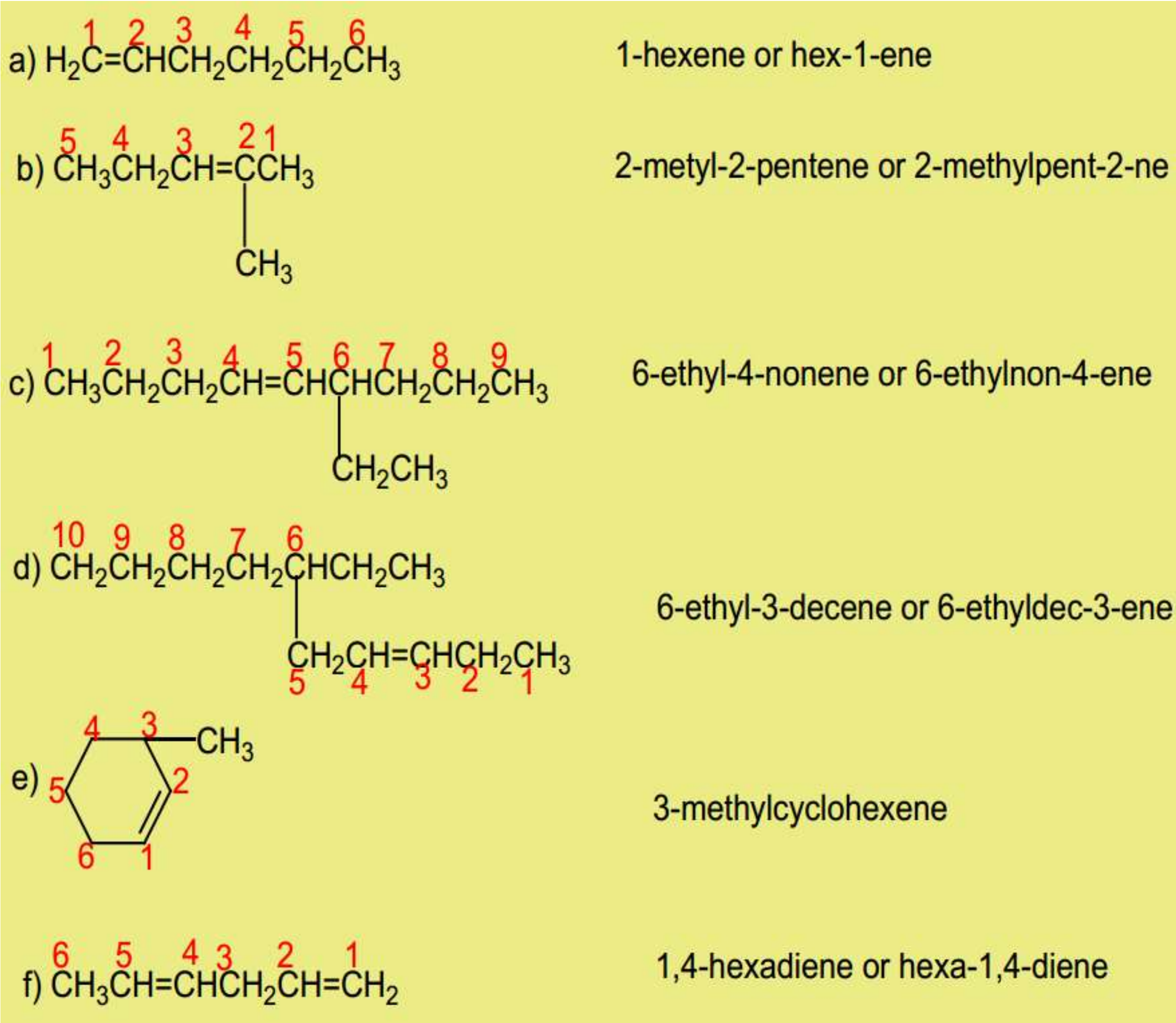
The chain is always numbered from the end that gives the smallest number for the location of the double bond.

In naming cycloalkenes, the carbon atoms of the double bond are numbered 1 and 2 in the direction that gives the smallest numbers for the location of the substituents.

If a compound contains two or more double bonds, its location is identified by a preffix number. The ending is modified to show the number of double bonds:

- **a diene** for two double bonds,
- **a triene** for two three bonds
- **a tetraene** for four double bonds

Examples



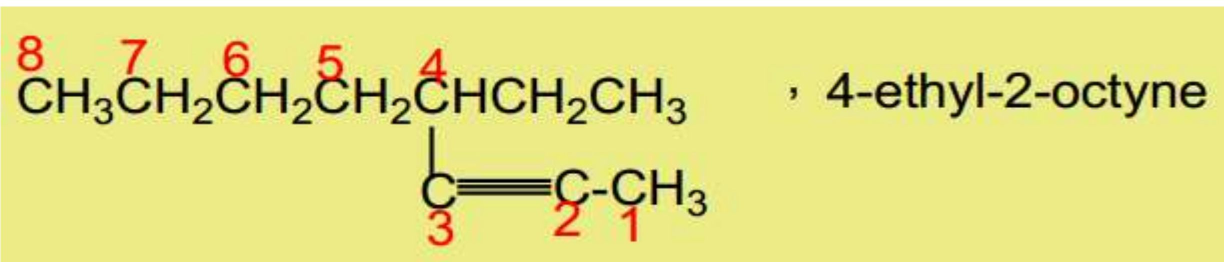
- ✓ **Nomenclature of alkynes**

Alkynes are named by identifying the longest continuous chain containing the triple bond and changing the ending **–ane** from the corresponding alkane to **–yne**.

Examples

H-C≡ C-H, ethyne

HC ≡C-CH2CH2CH3, pent-1-yne



Alkynes with four or more carbon atoms have structural isomers.

$HC\equiv CCH_2CH_2CH_3$ and $CH_3C\equiv CCH_2CH_{3w}$

- ✓ **Nomenclature of alcohol**

According to IUPAC system, alcohols are named by replacing the fnal ‘’**e**’’ of the parent hydrocarbon with‘’**ol**’’, then specify the position of **-OH** group before ending by **ol**

Examples:

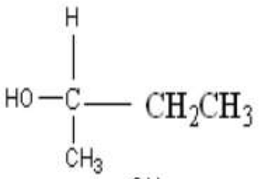
CH_3OH: Methanol

CH_3-CH_2-OH: ethanol

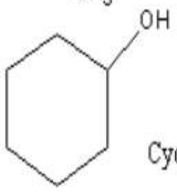
CH_3-CH_2-CH_2-OH: Propan-1-ol

CH_3-CHOH-CH_3: Propan-2-ol

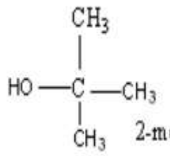
CH_3-CH_2-CH_2-CH_2OH: Butan-1-ol



butan-2-ol



Cyclohexanol

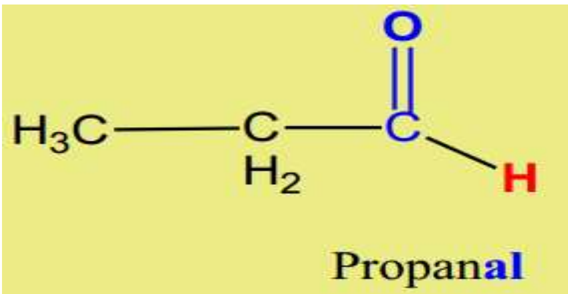


2-methyl propan-2-ol or tertiobutanol

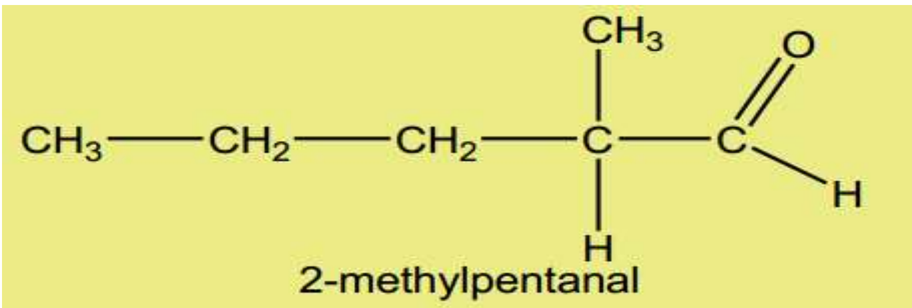
- ✓ **Nomenclature of Aldehydes**

The systematic name of an aldehyde is obtained by replacing the terminal **"e"** from the name of the parent hydrocarbon with **"al."**In numbering the carbon chain of an aldehyde, the carbonyl carbon is numbered one.

1stExample: Write formula of propanal.
Propanal has a chain of 3 carbons without carbon-carbon double bonds.
The suffix ‘al’ indicates the presence of the -CHO group. The carbon in carbonyl functional group is counted as first of the chain.



2nd Example: Write formula of 2-methylpentanal.
This molecule has 5 carbons in the longest chain, including the one in the -CHO group. There aren’t any carbon-carbon double bonds. A methyl group is attached to the number 2 carbon



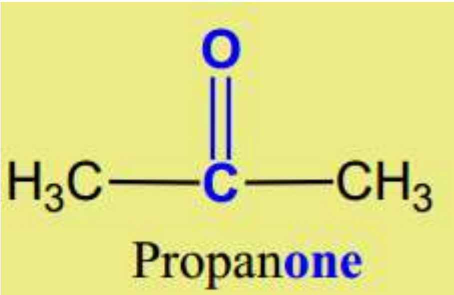
Important: in aldehydes, the carbon in the -CHO group is always counted as first carbon.

✓ **Nomenclature of Ketones**

The systematic name of a ketone is obtained by removing the terminal **"e"** from the name of the parent hydrocarbon and adding **"one."** The chain is numbered in the direction that gives the **carbonyl carbon the smallest number**. Ketone contains a carbon-oxygen double bond just like aldehyde, but for ketone carbonyl group is bonded to two alkyl groups.

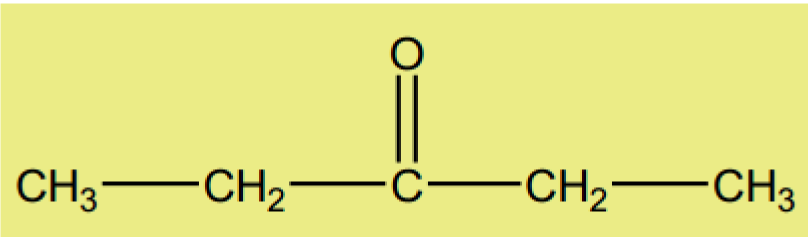
1st Example*: *Write formula of propanone.

Propanone has a chain of 3 carbons. The **sufx"one"** indicates the presence of the >C=O in the middle of the carbon chain. The carbon-oxygen double bond has to be in the middle of the carbon chain, for this molecule, the carbonyl group is on carbon 2.



2nd Example: Write formula of pentan-3-one.

In pentanone, the carbonyl group could be in the middle of the chain or next to the end - giving either pentan-3-one or pentan-2-one.
The position of the carbon-oxygen double bond has to be indicated because there is more than one possibility. This molecule has its carbon-oxygen double bond at carbon 3. If it was on the second carbon, it would be pentan-2-one.



- <u>Function of organic compound</u>

The organic chemistry is a subject that plays an important role in modern life. In general, there is no art, science or industry where knowledge of organic chemistry is not applied.
Examples where organic chemistry is applied:

1)Application in daily life.

In our day-to-day life, we find many substances or materials that are commonly used and the later are made of organic compounds.

- Food: starch, fats, proteins, vegetables,...
- Clothes: cotton, wool, nylon, dacron, ...
- Fuels: petrol, diesel oil, and kerosene
- Dyes of all kinds
- Cosmetics (body lotion,…)
- Soaps and detergents
- Medicine: cortisone, sulphonamide, penicillin,…
- Drugs: morphine, cocaine,...
- Stationery: pencils, paper, writing ink,…
- Insecticides,rodenticides,ovicides …

2)Applications in industry

The knowledge of organic chemistry is required in many industries such as manufacture of food, pharmacy, manufacture of dyes and explosives, alcohol industry, soil fertilisers, petroleum industry, etc.

3)Study of life processes

Organic chemistry in other words is the chemistry of life. For example the vitamins, enzymes, proteins and hormones are important organic compounds produced in our body to ensure its proper development.

LO 3.2. POLYMERISE ORGANIC SIMPLE MOLECULES

- Properties of organic simple molecules (alcohol, Ethylene, acetic acid and ester)

✓ **alcohol**

Properties of alcohols

Physical properties:

Solubility in water

Alcohols are soluble in water. This is due to the hydroxyl group in the alcohol which is able to form hydrogen bons

Boiling point

The boiling point of alcohols also increase as the length of hydrocarbon chain increases.
The reason why alcohols have a higher boiling point than alkanes is because the intermolecular forces of alcohols are hydrogen bonds, unlike alkanes with van der Waals forces as their intermolecular forces.

Viscosity

Viscosity is the property of a fluid that resists the force tending to cause the fluid to flow.

The viscosity of alcohols increase as the size of the molecules increases. This is because the strength of the intermolecular forces increases, holding the molecules more firmly in place.

Polarity

Amide > Acid > Alcohol > Ketone ~ Aldehyde > Amine > Ester > Ether > Alkane
Amide is the most polar while alkane is the least.
Alcohol is ranked third in terms of polarity due to its hydrogen bonding capabilities and presence of one

oxygen atom in an alcohol molecule. Carboxylic acids are more polar than alcohols because there are two oxygen atoms present in a carboxylic acid molecule.

Flammability

The flammability of alcohols decrease as the size and mass of the molecules increases. Combustion breaks the covalent bonds of the molecules, so as the size and mass of the molecules increases, there are more covalent bonds to break in order to burn that alcohol. Hence, more energy is required to break the bonds, therefore the flammability of alcohols decrease as size and mass of molecules increases.

Chemical properties:

Combustion

Alcohols burns in oxygen to produce carbon dioxide and water. Alcohols burn cleanly and easily, and does not produce soot. It becomes increasingly more difficult to burn alcohols as the molecules get bigger.

The general molecular equation for the reaction is:

$$C_nH_{2n+1}OH + (1.5n)O_2 \rightarrow (n+1)H_2O + nCO_2$$

e.g. combustion of ethanol:

$$C_2H_5OH\ (l) + 3\ O_2\ (g) \rightarrow 2\ CO_2\ (g) + 3\ H_2O\ (g);\ (\Delta H_c = -1371\ kJ/mol)$$

Dehydration

- alcohol to alkene

Dehydration of alcohols is done by heating with concentrated sulfuric acid, which acts as the dehydrating agent, at 180°C. This reaction uses alcohols to produce corresponding alkenes and water as byproduct.

Oxidation - alcohol to carboxylic acid

Alcohols can be oxidised into carboxylic acids.

e.g. oxidation of ethanol:

$$C_2H_5OH + [O] \rightarrow CH_3COOH + H_2O$$

Ethanol, if left exposed to air, can oxidise and become ethanoic acid. An example is wine turning sour as the alcohol content, which is ethanol, is oxidised by atmospheric oxygen.

Esterification

Alcohols can be reacted with carboxylic acid to form esters. More of this will be explained under *Formation of esters*

✓ **ethylene**

Properties of Ethene (ethylene)

- colourless gas at room **temperature** and pressure it **Melting point** = -169°C. **Boiling point** = -104°C.
- slightly sweet smell.
- flammable.

- non-polar molecule soluble in non-polar solvents. ...
- reactive: the active site is the double bond For example, ethene readily undergoes addition reactions.

✓ **acetic acid**

Properties of Acetic Acid

Even though ethanoic acid is considered to be a weak acid, in its concentrated form, it possesses strong corrosive powers and can even attack the human skin if exposed to it. Some general properties of acetic acid are listed below.

- Ethanoic acid appears to be a colourless liquid and has a pungent smell.
- At STP, the melting and boiling points of ethanoic acid are 289K and 391K respectively.
- The molar mass of acetic acid is 60.052 g/mol and its density in the liquid form is 1.049 $g.cm^{-3}$.
- The carboxyl functional group in ethanoic acid can cause ionization of the compound, given by the reaction: $\mathbf{CH_3COOH \rightleftharpoons CH_3COO^- + H^+}$
- The release of the proton, described by the equilibrium reaction above, is the root cause of the acidic quality of acetic acid.
- The acid dissociation constant (pK_a) of ethanoic acid in a solution of water is 4.76.
- The conjugate base of acetic acid is acetate, given by CH_3COO^-.
- The pH of an ethanoic acid solution of 1.0M concentration is 2.4, which implies that it does not dissociate completely.
- In its liquid form, acetic acid is a polar, protic solvent, with a dielectric constant of 6.2.

Uses: Acetic acid is largely used in the food industry as vinegar, and as an acidity regulator.

The metabolism of carbohydrates and fats in many animals is centred around the binding of acetic acid to coenzyme A. Generally, this compound is produced via the reaction between methanol and carbon monoxide (carbonylation of methanol).

✓ **ester**

The physical properties of fats and oils

Solubility in water

Fats and oils are not water soluble. The chain lengths are so great that far too many hydrogen bonds between water molecules must be broken; this is not an energetically profitable arrangement.

Melting point

Melting points determine whether the substance is a fat (a solid at room temperature) or an oil (a liquid at room temperature). Fats normally contain saturated chains. These allow more effective van der Waals dispersion forces between the molecules: more energy is required to separate the chains, increasing the melting point.

A greater number of double bonds, or <u>degree of unsaturation</u>, in the molecules results in a lower melting point, because the van der Waals forces are less effective. The efficacy of van derWaals forces depends on the ability of molecules to pack closely together. The presence of carbon-carbon double bonds in the chains disrupts otherwise tidy packing.

- <u>Equilibrium of organic simple molecules</u>

1. Defnition of monomer, polymer and polymerization

- ✓ **Monomer**

  The term monomer comes from mono "one" and meros "part",which expresses a single unit or a small molecular subunit that can be chemically bind to another identical or different molecule to form larger molecule (polymer).The monomer is repeated in the polymer chain and it is the basic unit which makes up the polymer.

  For instance in the large compound formed by nA → -An- where A is a monomer and the polymer is given by the repeated monomers in the chain; i.e. –A-A-A-A-A-A-. The larger molecules such as carbohydrates, lipids, nucleic acids and proteins are found in living systems, like our own bodies.

  nCH2= CH2 → [- CH2 — CH2 —]n, ethylene (CH2 = CH2) is a monomer

- ✓ **Polymer**

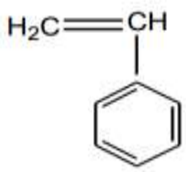
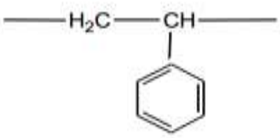
A polymer is a large molecule (macromolecule or giant molecule) composed of smaller molecules (monomers) linked together by intermolecular covalent bonds.

Polymers have a high molecular weight in the range of 103 to 107. A polymer can be represented as (–An–) or (– A – A – A – A – A-……) which is a polymer of the monomer A.

Example: nCH2 = CH2 → [- CH2 — CH2 —]n,

Polyethylene, [- CH2 — CH2 —]n is a polymer while ethene,CH2 = CH2 is a monomer

Some examples of polymers and monomers

Monomer		repeating unit/ monomeric unit	Polymer
$H_2C{=}CH_2$	Ethylene	$—H_2C—CH_2—$	Polyethylene
$H_2C{=}CH(Cl)$	vinyl chloride	$—H_2C—CH(Cl)—$	poly(vinyl chloride)
	Styrene		Polystyrene

- Polymerisation of organic molecules

Polymerization is the process in which monomer units are linked by chemical reaction to form long chains (polymers).
For example, a gaseous compound, Butadiene, with a molecular weight of 54 g/mole combines nearly 4000 times by polymerization and gives a polymer, known as polybutadiene.
Butadiene + butadiene + butadiene + … + butadiene → Polybutadiene
Note: 1.The degree of polymerization (n) is defned as the number of monomeric units in a macromolecule or polymer or oligomer (a polymer consisting of few number of monomers units) molecule.
2. A polymer formed by identical monomers is called **homopolymrer** while a polymer formed by different monomers is a **copolymer**

Checking up
1. Draw two repeat units of the polymer formed from each of the following
monomers or pair of monomers.
a. CH2=CH-OCOCH3
b. CH2=CCl-CH=CH2
c. CH2=CHCl and CH2=CH-CH3
2. Draw the monomers from which the following polymers are made:
a. ---CH2-CH(CN)-CH2-CH(CN)-CH2—
b. ---CH2-C(CN)=CH-CH(CN)-CH2-C(CN)=CH-CH(CN)-CH2---

1. Types of polymerization

There two types of polymerization reaction: Addition polymerization and condensation polymerization.

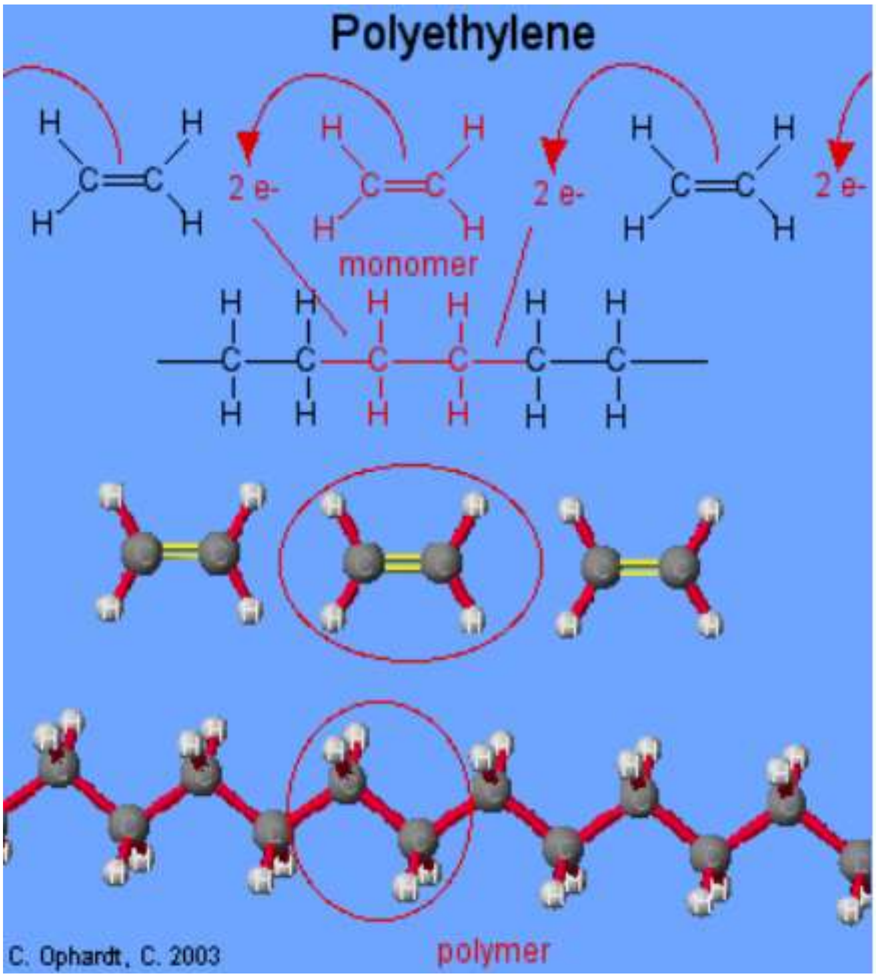
1.1. Addition polymerization

Addition polymerization is a process where monomers are linked together to form a
polymer, without the loss of atoms from the molecules. When the monomer molecules add up to form the polymer, the process is called "***Addition polymerization***".
This involves the combination together of monomer units to give new product (polymer) having the same empirical formula to the monomer but having relative molecular weight.The monomer units are usually unsaturated compounds.
Some examples of addition polymers are polyethylene, polypropylene, polystyrene, polyvinyl acetate, polyvinyl chloride (PVC), rubber, polytetrafluoroethylene (Teflon), etc.

Example 1: Formation of polyethene or polyethylene
Polyethene is formed by addition of CH2= CH2 monomer molecules linked together as follows

$$n\,CH_2 = CH_2 \rightarrow [-\,CH_2 - CH_2 -]_n$$

Ethene Polyethene



1.2. Condensation polymerization

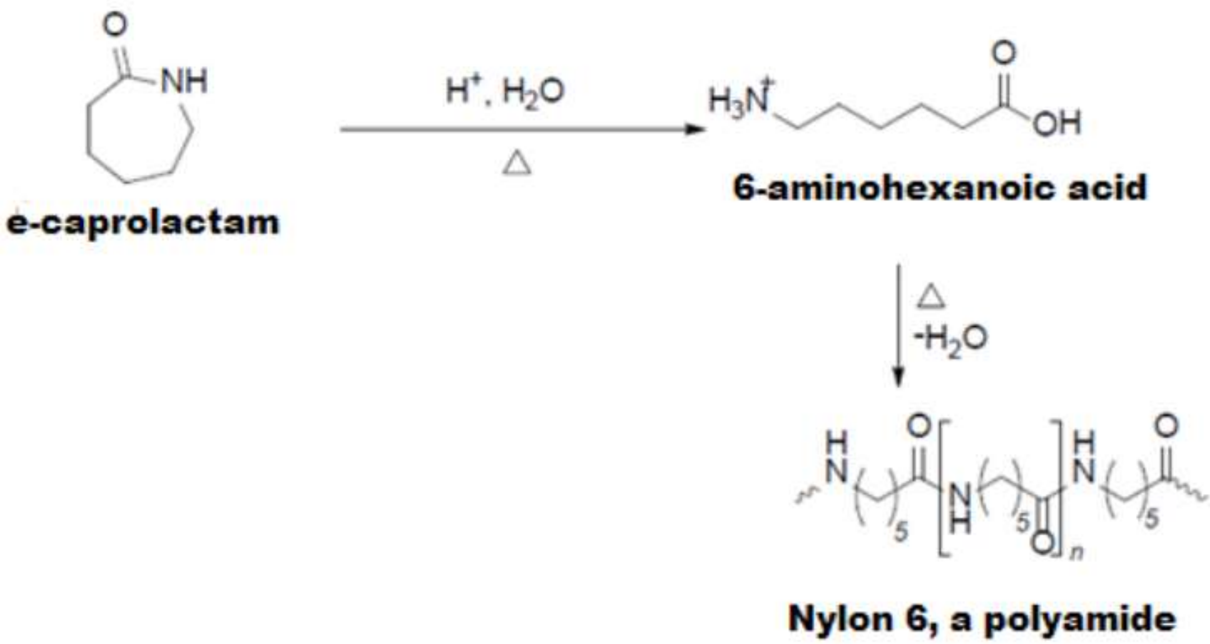
Condensation polymerization is a process where two or more monomers chemically combine to form a polymer with elimination of simple molecules like water, ammonia, hydrogen chloride, alcohol, etc.
In this type of reaction, each monomer generally contains two functional groups for the condensation reaction to take place.
There are two main types: **polyamides** which are formed between a diol with a dicarboxylic acid and **polyesters** which are formed when a dicarboxylic acid and a diamine react to form nylons.The condensation polymers include Nylons, Terylene, Kevlar polymer, proteins, cellulose, Bakelite, Dacron, etc.

Example 1: formation of polyamides

This type of polymers is formed by the result of generation of amide bonds in the polymerization reaction.

i. Formation of Nylon

Nylon 6 and nylon- 6, 6 are important examples for this type of polymers nylon 6 is synthesized from e-caprolactam, which on heating decomposes into 6-aminohexanoic acid that polymerizes into nylon 6. Here the number 6 represents the number of carbon atoms present in the monomer unit.



Nylon -6,6 is produced by the condensation reaction between two monomer units adipic acid and 1,6-hexanediamine in the presence of heat. This is formed from a sixcarbon diacid and a six carbon diamine as shown below.

$$nH_2N{-}(CH_2)_6{-}NH_2 + nHOOC{-}(CH_2)_4{-}COOH \xrightarrow{-2nH_2O} -(NH{-}(CH_2)_6{-}NH{-}CO{-}(CH_2)_4{-}CO)_n-$$

1, 6-Hexanediamine   Adipic acid   Nylon -6,6

Comparison between addition and condensation polymerization

Addition polymerization	Condensation polymerization
1. Involves unsaturated monomer like ethylene, vinyl chloride, styrene etc.	1. Involves substances with at least 2 different reacting functional groups like ethylene glycol (2-COOH groups).
2. Fast addition of monomers.	2. Step-wise slow addition
3. Initiator is necessary to catalyze the polymerization	3. Catalyst is not necessary
4. No elimination products	4. Small molecule like H_2O, HCl, or CH_3OH is often eliminated.
5. Polymers made are, for example, polyethene, polypropylene, polybutadiene, polyvinylchloride.	5. Polymers made are, for example, terylene, nylon, formaldehyde-resins, silicones

LO 3.3. DESCRIBE ORGANIC PRODUCTS

• Different organic products: carbohydrates Proteins, Vitamins and lipids”.

- ✓ **What are carbohydrates?**

Carbohydrates are found in a wide array of both healthy and unhealthy foods—bread, beans, milk, popcorn, potatoes, cookies, spaghetti, soft drinks, corn, and cherry pie. They also come in a variety of forms. The most common and abundant forms are sugars, fibers, and starches.

Foods high in carbohydrates are an important part of a healthy diet. Carbohydrates provide the body with glucose, which is converted to energy used to support bodily functions and physical activity.

- ✓ **Proteins**

Protein is a macronutrient that is essential to building muscle mass. It is commonly found in animal products, though is also present in other sources, such as nuts and legumes. There are three macronutrients: **protein**, fats and carbohydrates. Macronutrients provide calories, or energy

- ✓ **Vitamins**

A ***vitamin*** is an organic molecule that is an essential micronutrient which an organism needs in small quantities for the proper functioning of its metabolism.

Water-soluble vitamins

Water-soluble vitamins travel freely through the body, and excess amounts usually are excreted by the kidneys. The body needs water-soluble vitamins in frequent, small doses. These vitamins are not as likely as fat-soluble vitamins to reach toxic levels. But niacin, vitamin B6, folate, choline, and vitamin C have upper consumption limits. Vitamin B6 at high levels over a long period of time has been shown to cause irreversible nerve damage.

A balanced diet usually provides enough of these vitamins. People older than 50 and some vegetarians may need to use supplements to get enough B12.

Water-soluble vitamins

Nutrient	Function	Sources
Thiamine (vitamin B1)	Part of an enzyme needed for energy metabolism; important to nerve function	Found in all nutritious foods in moderate amounts: pork, whole-grain or enriched breads and cereals, legumes, nuts and seeds
Riboflavin (vitamin B2)	Part of an enzyme needed for energy metabolism; important for normal vision and skin health	Milk and milk products; leafy green vegetables; whole-grain, enriched breads and cereals
Niacin (vitamin B3)	Part of an enzyme needed for energy metabolism; important for nervous system, digestive system, and skin health	Meat, poultry, fish, whole-grain or enriched breads and cereals, vegetables (especially mushrooms, asparagus, and leafy green vegetables), peanut butter
Pantothenic acid	Part of an enzyme needed for energy metabolism	Widespread in foods
Biotin	Part of an enzyme needed for energy metabolism	Widespread in foods; also produced in intestinal tract by bacteria
Pyridoxine (vitamin B6)	Part of an enzyme needed for protein metabolism; helps make red blood cells	Meat, fish, poultry, vegetables, fruits
Pyridoxine (vitamin B6)	Part of an enzyme needed for protein metabolism; helps make red blood cells	Meat, fish, poultry, vegetables, fruits
Folic acid	Part of an enzyme needed for making DNA and new cells, especially red blood cells	Leafy green vegetables and legumes, seeds, orange juice, and liver; now added to most refined grains
Cobalamin (vitamin B12)	Part of an enzyme needed for making new cells; important to nerve function	Meat, poultry, fish, seafood, eggs, milk and milk products; not found in plant foods
Ascorbic acid (vitamin C)	Antioxidant; part of an enzyme needed for protein metabolism; important for immune system health; aids in iron absorption	Found only in fruits and vegetables, especially citrus fruits, vegetables in the cabbage family, cantaloupe, strawberries, peppers, tomatoes, potatoes, lettuce, papayas, mangoes, kiwifruit

Fat-soluble vitamins

Fat-soluble vitamins are stored in the body's cells and are not excreted as easily as water-soluble vitamins. They do not need to be consumed as often as water-soluble vitamins, although adequate amounts are needed. If you take too much of a fat-soluble vitamin, it could become toxic.

A balanced diet usually provides enough fat-soluble vitamins. You may find it more difficult to get enough vitamin D from food alone and may consider taking a vitamin D supplement or a multivitamin with vitamin D in it.

Fat-soluble vitamins

Nutrient	Function	Sources
Vitamin A (and its precursor*, beta-carotene) *A precursor is converted by the body to the vitamin.	Needed for vision, healthy skin and mucous membranes, bone and tooth growth, immune system health	Vitamin A from animal sources (retinol): fortified milk, cheese, cream, butter, fortified margarine, eggs, liver Beta-carotene (from plant sources): Leafy, dark green vegetables; dark orange fruits (apricots, cantaloupe) and vegetables (carrots, winter squash, sweet potatoes, pumpkin)
Vitamin D	Needed for proper absorption of calcium; stored in bones	Egg yolks, liver, fatty fish, fortified milk, fortified margarine. When exposed to sunlight, the skin can make vitamin D.
Vitamin E	Antioxidant; protects cell walls	Polyunsaturated plant oils (soybean, corn, cottonseed, safflower); leafy green vegetables; wheat germ; whole-grain products; liver; egg yolks; nuts and seeds
Vitamin K	Needed for proper blood clotting	Leafy green vegetables such as kale, collard greens, and spinach; green vegetables such as broccoli, Brussels sprouts, and asparagus; also produced in intestinal tract by bacteria

- ✓ **Lipids**

Lipids are molecules that contain hydrocarbons and make up the building blocks of the structure and function of living cells. Examples of **lipids** include fats, oils, waxes, certain vitamins (such as A, D, E and K), hormones and most of the cell membrane that is not made up of protein.

Lipid, any of a diverse group of organic compounds including fats, oils, hormones, and certain components of membranes that are grouped together because they do not interact appreciably with water. One type of lipid, the triglycerides, is sequestered as fat in adipose cells, which serve as the energy-storage depot for organisms and also provide thermal insulation. Some lipids such as steroid hormones serve as chemical messengers between cells, tissues, and organs, and others communicate signals between biochemical systems within a single cell. The membranes of cells and organelles (structures within cells) are microscopically thin structures formed from two layers of phospholipid molecules. Membranes function to separate individual cells from their environments and to compartmentalize the cell interior into structures that carry out special functions. So important is this compartmentalizing function that membranes, and the lipids that form them, must have been essential to the origin of life itself.

- Importance of carbohydrates Proteins, Vitamins and lipids".

✓ **Carbohydrates**

They serve several key functions in your body. They provide you with energy for daily tasks and are the primary fuel source for your brain's high energy demands. Fiber is a special type of carbone that helps promote good digestive health and may lower your risk of heart disease and diabetes

✓ **Proteins**

The Power of **Protein**

- Your body uses **protein** to build and repair tissues.
- You also use **protein** to make enzymes, hormones, and other body chemicals.
- **Protein** is an **important** building block of bones, muscles, cartilage, skin, and blood

✓ **Vitamins**

Vitamins and minerals are considered essential **nutrients**

- because acting in concert,
- they perform hundreds of roles in the body.
- They help shore up bones, heal wounds, and bolster your immune system.
- They also convert food into energy, and repair cellular damage.

✓ **Lipids**

Lipids are essential for all life on Earth.

- They play many **important** roles in maintaining the health of an organism.
- Arguably the most **important** function **lipids** perform is as the building blocks of cellular membranes.
- Other functions include energy storage, insulation, cellular communication and protection.

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