General Chemistry
for
First Stage of
Astronomy department

Semester 1st

Dr. Wafaa Al-Qaysi

2020-2021

General Chemistry

Lec. 1 (Introduction)

1.1. Chemistry*:* Is the study of the composition, structure, properties, reactions of ***matter*** and the changes it undergoes. So Chemistry is largely an experimental science, and a great deal of knowledge comes from laboratory research.

1.2. Matter: is anything that occupies space and has mass**. For example:** Table, books, walls are all composed of matter.

1.2.1. States of matter
There are three states of matter:

1. Solid: Is rigid objects with definite shapes and a definite volume (for example, ice).

2. Liquid (fluid): Is less rigid than solid and its physical state characterized by an indefinite shapes and a definite volume (for example, water).

3. Gas: Is the physical state characterized by an indefinite shape and an indefinite volume. Gasses are fluid but unlike liquids, they can expand indefinitely

1.2.2. Properties of matter

There are two types of properties, ***physical and chemical***.

A physical property can be measured and observed without changing the composition or identity of a substance. For example, color, melting point, boiling point, freezing point, solubility and density. For example, we can measure melting point of ice by heating a block of ice, we can also freeze the water to form ice. Another example, when we say that helium gas is lighter than air, we are referring to a physical property.

* ***Melting point*** is the temperature at which a solid form will transition to form a liquid.

Boiling point is the temperature at which a liquid form will transition to form a gas.

A chemical property can be measured and observed with a chemical changing the composition or identity of a substance. For example, "Hydrogen gas burns in oxygen gas to form water". After the change, the original substances, hydrogen and oxygen gas, will have vanished and a chemically different substance water will have taken their place.

1.2.3. Classifications of pure matter

Pure matter consists of two types, ***elements and compounds***

1) Element is a substance that cannot be separated into simpler substances by chemical means.

At present, 118 elements have been positively identified. Chemists use alphabetical symbols to represent the names of the elements. The first letter of the symbol for an element is always capitalized, but the second letter is never capitalized. For example, Co is the symbol for the element cobalt, whereas CO is the formula for carbon monoxide, which is made up of the elements carbon and oxygen. Table 1.1 shows some of the more common elements. The symbols for some elements are derived from their Latin names for example, Au from aurum (gold), Fe from ferrum (iron), and Na from natrum (sodium)— although most of them are abbreviated forms of their English names.

Table Error! No text of specified style in document.**-1 .Some Common Elements and Their Symbols**

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bromine	Br	Iodine	I	Potassium	K
Calcium	Ca	Iron	Fe	Silicon	Si
Carbon	C	Lead	Pb	Silver	Ag
Chlorine	Cl	Magnesium	Mg	Sodium	Na
Chromium	Cr	Mercury	Hg	Sulfur	S
Cobalt	Co	Nickel	Ni	Tin	Sn
Copper	Cu	Nitrogen	N	Zinc	Zn

2) Compound is a substance composed of two or more elements chemically united in fixed proportions. Hydrogen gas, for example, burns in oxygen gas to form water, a compound whose properties are distinctly different from those of the starting materials.

1.3. Substances and Mixtures

A substance is a matter that has a definite or constant composition and distinct properties. Examples are water, silver, ethanol, table salt (sodium chloride), and carbon dioxide. Substances differ from one another in composition and can be identified by their appearance, smell, taste, and other properties.

A mixture is a combination of two or more substances in which the substances retain their distinct identities. Some examples are air, soft drinks, milk, and cement.
*Note: <u>Mixtures do not have constant composition.</u>

Mixtures are either **homogeneous or heterogeneous.**

A homogeneous mixture the composition of the mixture is the same solution such as sugar dissolves in water.

A heterogeneous mixture which the composition is not uniform such as oil to water.

1.4. Measurements

The study of chemistry depends on **measurement**. Chemists use measurements to compare the properties of different substances and to assess changes resulting from an experiment. A number of common devices enable us to make simple measurements of a substance's properties: **The meter stick measures length; the buret, the pipet, the graduated cylinder, and the volumetric flask measure volume; the balance measures mass; the thermometer measures temperature.** These instruments provide measurements of macroscopic properties, which can be determined directly. Microscopic properties, on the atomic or molecular scale, must be determined by an indirect method.

*A measured quantity is usually written as a number with an appropriate unit.

1.4.1. SI Units

The international authority on units, proposed a revised metric system called the International System of Units (abbreviated SI, from the French System International d'Unites).

Table 1.2 shows the seven SI base units. All other SI units of measurement can be derived from these base units. Like metric units, SI units are modified in decimal fashion by a series of prefixes.

TABLE Error! No text of specified style in document.**-2. SI Base Units**

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of	mole	mol
Luminous intensity	candela	cd

In Chemistry, We use metric units and SI units for length, volume, mass, temperature and time.

1.5. Experimental Quantities

1.5.1. Mass and weight

Mass describes the quantity of matter in an object. The terms *weight* and *mass*, in common usage, are often considered synonymous. They are not, in fact. ***Weight*** is the force of gravity on an object:

Weight = mass × acceleration due to gravity

The common conversion units for mass are as follows:

1 gram (g) = 10^{-3} kilogram (kg)

***Note: The mass of the apple is constant and does not depend on its location, but its weight does. For example, on the surface of the moon the apple would weigh only one-sixth what it does on Earth, because of the smaller mass of the moon**.

1.5.2. Volume

Volume is length (m) cubed, so its SI-derived unit is the cubic meter (m^3). Generally, however, chemists work with much smaller volumes, such as the cubic centimeter (cm^3) and the cubic decimeter (dm^3):

$$1cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$$

$$1dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$$

Another common, non-SI unit of volume is the liter (L). ***A liter*** is the volume occupied by one cubic decimeter. Chemists generally use L and mL for liquid volume. One liter is equal to 1000 milliliters (mL) or 1000 cubic centimeters:

$$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$$

and one milliliter is equal to one cubic centimeter:

$$1\ mL = 1\ cm^3$$

1.5.3. Density

Density is the mass of an object divided by its volume:

$$density = \frac{mass}{volume} \quad or \quad d = m/V$$

Where d, m, and V denote density, mass, and volume, respectively. Note that density is an intensive property that does not depend on the quantity of mass present.

The reason is that V increases as m does, so the ratio of the two quantities always remains the same for a given material.

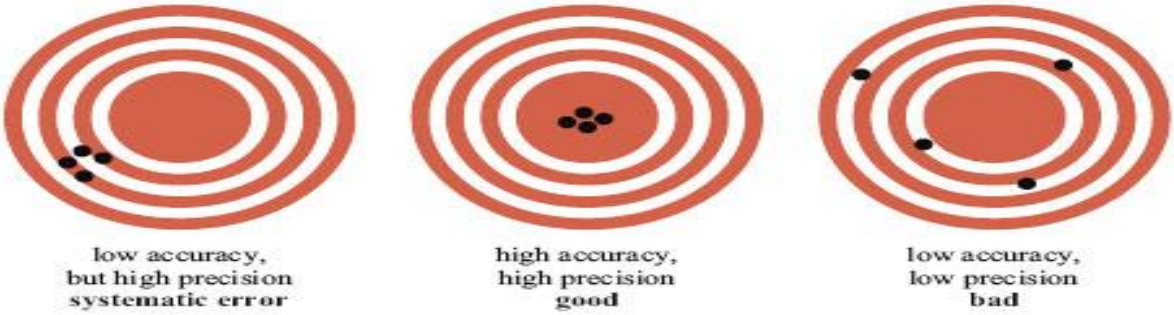
The unit for density is (kg/m^3). This unit is awkwardly large for most chemical applications. Therefore, (g/cm^3) and its equivalent, grams per milliliter (g/mL), are more commonly used for solid and liquid densities.

1.6. Accuracy and Precision

Accuracy tells us how close a measurement is to the true value of the quantity that was measured.

Precision refers to how closely two or more measurements of the same quantity agree with one another.

The difference between accuracy and precision is a subtle but important one.



Suppose, for example, that three students are asked to determine the mass of a piece of copper wire. The results are:

Student C	Student B	Student A	
2.000 g	1.272 g	1.563 g	
2.002 g	1.271 g	1.123g	
2.001 g	1.272 g	1.343	Average value

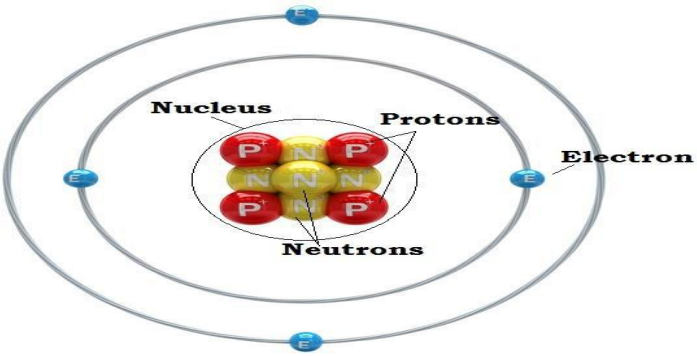
The true mass of the wire is 2.000 g. Therefore, Student B's results are a good precision but a bad accuracy Student C's results are not only a good precision, but also a good accuracy, because the average value is closest to the true value. Student A's results are bad accuracy and precision.

*Note: Highly accurate measurements are usually precise too. On the other hand, highly precise measurements do not necessarily guarantee accurate results.

Lec.2

2.1. Atomic Model

The atom is thought as a tiny solar system in which there is a central core (like the sun) with other particles traveling in circular path or orbits.



2.2. The Nucleus

The central core from the solar system model is called the nucleus. The nucleus contains two types of particles, **proton and neutron.**

1- **The Proton** is a particle that has a mass (or weight) of one amu (atomic mass unit) and a positive one (+1) electrical charge. The symbol for the proton is p^+ or H^+.

2- **The Neutron** has a mass of one amu (atomic mass unit) but has no electrical charge; that is, it is a neutral particle. The symbol for the neutron is n.

2.3. The outer structure

The particles that orbit the nucleus are called **electrons.** These particles have an electrical charge of negative one (-1) but their mass is so small that it is considered to be zero. The symbol for the electron is e^-.

2.4. Atomic Number, and Mass Number

Atomic number (Z) of an atom of an element is equal to the number of protons in the nucleus of the atom.

Note: In a neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates the number of electrons present in the atom. For example, the atomic number of nitrogen is 7; this means that each neutral nitrogen atom has 7 protons and 7 electrons.

Mass number (A): is the total number of neutrons and protons present in the nucleus of an atom of an element. Except for the most common form of hydrogen, which has **one proton and no neutrons**, all atomic nuclei contain both protons and neutrons. In general, the mass number is given by: mass number = number of protons + number of neutrons

= atomic number + number of neutrons

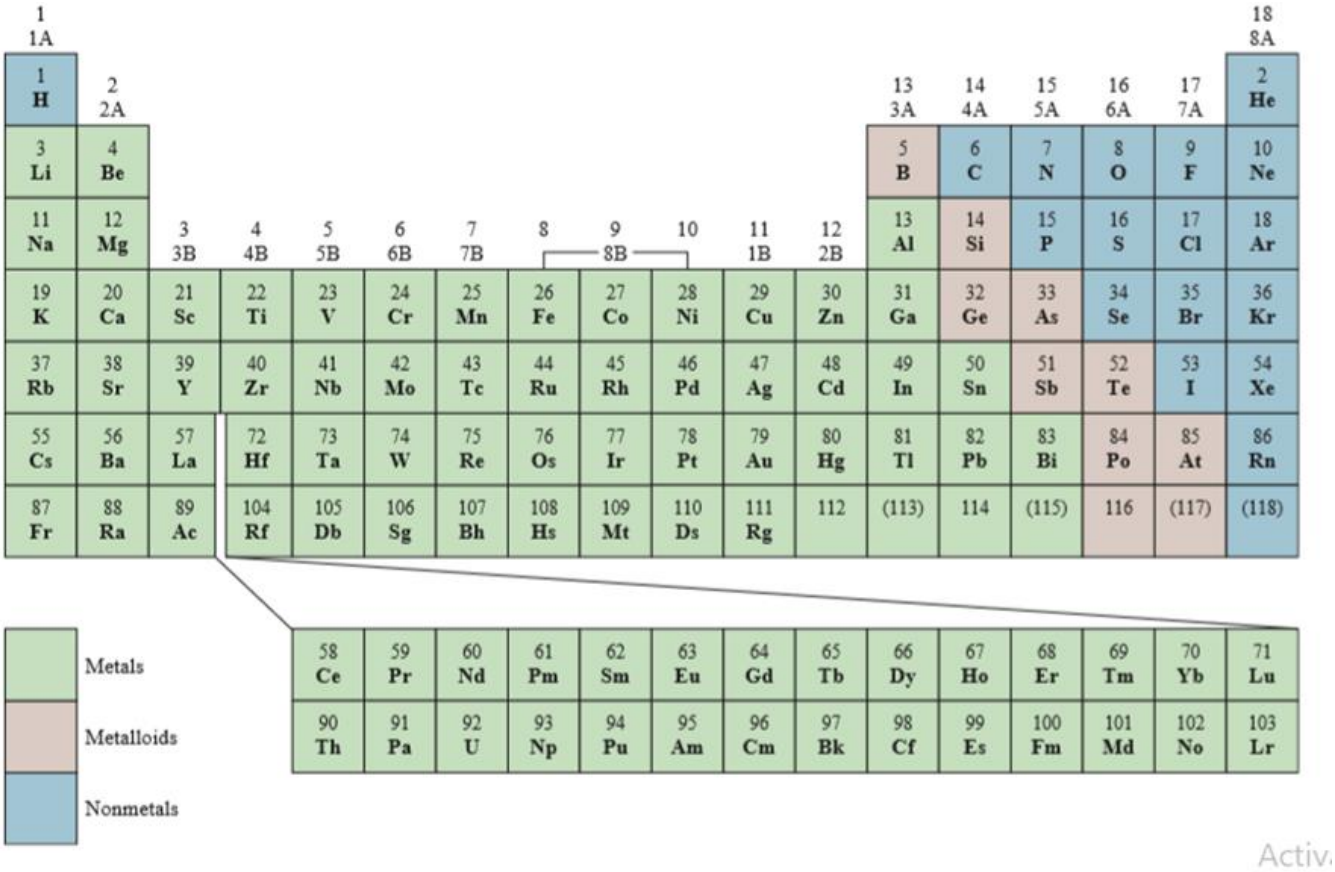
2.5. The Periodic Table

The periodic table is a chart in which elements having similar chemical and physical properties are grouped together. The periodic table is a handy tool that correlates the properties of the elements in a systematic way and helps us to make predictions about chemical behavior.
Figure 2.1 shows the modern periodic table, in which the elements are arranged by atomic number (shown above the element symbol)

1- In horizontal rows called periods
2- In vertical columns known as groups or families, So

A period is a horizontal row of element in the periodic table.

A group is a vertical column of elements in the periodic table that have similar properties.



Figer 2.1

2.5.1. Classification of groups

Some element groups have special names.

The Group 1A elements (Li, Na, K, Rb, Cs, and Fr) are called **alkali metals**.

Group 2A elements (Be, Mg, Ca, Sr, Ba, and Ra) are called **alkaline earth metals**.

Group 7A (F, Cl, Br, I, and At) are known as **halogens**.

Group 8A (He, Ne, Ar, Kr, Xe, and Rn) are called **noble gases (or rare gases)**.

2.6. Molecules and atom

2.6.1. Molecule

Is an aggregate of at least two atoms in a definite arrangement held together by chemical forces (also called chemical bonds). A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio.

A molecule is not necessarily a compound, which, by definition, is made up of two or more elements, for example, Hydrogen gas, and Water, is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom.

The hydrogen molecule, symbolized as H_2, contains only two atoms. Other elements are nitrogen (N_2) and oxygen (O_2), as well as the Group 7A elements—fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). Of course, a diatomic molecule can contain atoms of different elements. Examples are hydrogen chloride (HCl) and carbon monoxide (CO).

2.6.2. Ion

Is an atom or a group of atoms that has a net positive or negative charge.

The number of protons in the nucleus of an atom remains the same during chemical reactions, but electrons may be lost or gained.

The loss of one or more electrons is named a **cation**, with positive charge. While, an **anion** is an ion whose net charge is negative due to an increase in the number of electrons.

2.7 Chemical formula

A molecular formula shows the exact number of atoms of each element in the smallest unit of a substance. For example, H_2 is the molecular formula for hydrogen, O_2 is oxygen, O_3 is ozone, and H_2O is water.

Note that oxygen (O_2) and ozone (O_3) are allotropes of oxygen.

Empirical Formulas

The molecular formula of hydrogen peroxide, is H_2O_2. This formula indicates that each hydrogen peroxide molecule consists of two hydrogen atoms and two oxygen atoms. The ratio of hydrogen to oxygen atoms in this molecule is 2:2 or 1:1. The empirical formula of hydrogen peroxide is HO.

Calculate the empirical formula for hydrazine (N_2H_4)?

Formula of Ionic Compounds

The formulas of ionic compounds are usually contains from positive and negative charge.

For example: a solid sample of sodium chloride (NaCl) consists of equal numbers of Na^+ and Cl^- ions arranged in a three-dimensional network.

Lec.3

Homeworks1

Q1) Circle the correct answer in the following sentences

1) The SI unit for expressing the Time is

a) second b) candela c) kilogram

2) Which of the following substance is element?

a)sodium hydroxide b) Iron c) water

3) Which of the following substance is compound?

a)CO b) Co c) O_2

4) The temperature at which a solid form will transition to form a liquid is:

a) Melting point b) End point c) Boiling point

5) What is the mass of Silver…….. that occupy a volume of 15.6 cm^3 and the density is 19.3 g/cm^3.

a) 301.08 g b) 1.273g c) 0.808g d) 580.5g

5) The empirical formula of N_2H_2 is ……

a) N_2H b) NH_2 c) NH d) N_2H_2

6) Elements Mg and Ca are called……..

a) alkali metals b) alkaline earth metals c) halogens d) noble gases

7) When sugar dissolves in water called …….

a) Mole b) Mass c) Heterogeneous d) Homogeneous

Q2) Complete all sentences:

1) A substance is defied that ……., and examples are ……..

2) A heterogeneous solution is ……....such as ……

3) Chemists use ……to measure volume, while we use……to measures mass

4) Mass describes as the ……, the unit is ……., while Weight is ……..……

5) Volume is …….., and the unit is ……...

6) Density is the ……, and the unit is ……...

7) Fluid is defied that …….

8) Examples of physical property are ……

9) **Accuracy** tells us ……., while the **Precision is** **…….**

Q3) To convert this number $74.6*10^{-6}$ to scientific notation, we get:

a)0.00746 b) 0. 000746 c) 0.0000746

Q4) What is the density of a substance that has a mass of 0.6 gm and occupy a volume of 0.2 ml?

a)0.33 b) 3 c) 3.3

Q5) The mass of a sample of copper nitrate is 3.82 g. A student measures the mass and finds it to be 3.81 g, 3.82 g, 3.79 g, and 3.80 g in the first, second, third, and fourth trial, respectively. How would you describe the measurements which is **Accuracy or Precision**?

Lec.4

Analytical Chemistry (Reactions in Aqueous Solutions)

Chemistry: Is the study of the composition, structure, properties, reactions of ***matter*** and the changes it undergoes. So Chemistry is largely an experimental science, and a great deal of knowledge comes from laboratory research.

Matter: is anything that occupies space and has mass**. For example:** Table, books, walls are all composed of matter.

Branches of chemistry such as:

1- **Inorganic chemistry:** It deals with compounds that do not contain carbon.
2- **Organic chemistry:** It deals with compounds that contain carbon.
3- **Analytical chemistry:** It deals with analysis of different elements in a compound.
4- **Biochemistry:** The chemistry of living organisms of vital processes.
5- **Physical chemistry:** It deals with the relationship of chemical and physical properties of matter.

The analytical chemistry:

This kind of chemistry specific for separation and determination of the chemical substances.

Classification of Analytical chemistry:

Qualitative analysis chemistry: Reveals the chemical identity of the species in the samples (i.e. shows what elements-compounds, ions- a given substance contains).

Quantitative analysis chemistry: Establishes the relative amount of one or more of species or analytes in numerical terms.

Qualitative analysis has usually done before the quantitative analysis.

For example, if the river water is contaminated and leads to die fishes and other aquatic organism, the aim of qualitative analysis is to identify the reasons of this contamination. Is the reason, the transfer of heavy metals (such as Pb, Hg, Cd) to the river water? If the answer is positive, how much the quantity or concentration of contaminated element is? (This is Quantitative analysis job).

1.1 General Properties of Aqueous Solutions

A solution is a homogeneous mixture of two or more substances.

Homogeneous: means that the mixture has the same composition everywhere. When sugar dissolves in water.

Heterogeneous: A mixture that is not the same everywhere. When added oil to water, and (such as orange juice, which has suspended solids).

Solute: The substance present in a smaller amount (we will assume the solute is a liquid or a solid).

Solvent: The substance present in a larger amount (aqueous solutions such as water). A solution may be gaseous (such as air), solid (such as an alloy),or liquid (seawater, for example).

(mass) units: The masses of substance in unit of; gram (g), millgram (mg), microgram (µg), nanogram (ng).

1.1.1 Concentration of Solutions

The concentration of a solution is how much (mass) of solute is contained in a given volume of solvent or mass of solution. The concentration of a solution can be expressed in many different ways. Here we will consider one of the most commonly used units in chemistry.

Mole: symbol (mol.) is Avogadro's number ($6.02*10^{23}$) of particles (atoms, molecules, ions, or anything else).

$$\text{Number of moles for compounds} = \frac{mass}{Molar\ mass}$$

Atomic mass units (amu) of an element is the number of **grams** containing Avogadro's number of atoms.

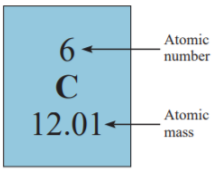
For example:

Number of moles of ion (SO_4^{2-}) = $\frac{mass}{Ionic\ mass,\ \ for\ SO4\ -2} = \frac{mass}{I\times32+(4\times16)}$

Number of moles of element (Ag^{+}) = $\frac{mass}{Atomic\ mass, for\ Ag+} = \frac{mass}{108}$

Average Atomic Mass

When you look up the atomic mass of carbon you will find that its value is not 12.00 amu but 12.01 amu.



The reason for the difference is that most naturally occurring elements (including carbon) have more than one isotope. For example, the natural abundances of carbon-12 and carbon-13 are 98.90 percent and 1.10 percent, respectively. The atomic mass of carbon-13 has been determined to be 13.00335 amu. Thus, the average atomic mass of carbon can be calculated as follows:

$$\begin{aligned}\text{average atomic mass of natural carbon} &= (0.9890)(12.00000\ \text{amu}) + (0.0110)(13.00335\ \text{amu})\\ &= 12.01\ \text{amu}\end{aligned}$$

Molar mass (M.mass) of compound is the sum of the atomic mass of all the atoms in the molecular formula of the compound.

For example:

The molecular mass of H_2O is

$$2(\text{atomic mass of H}) + \text{atomic mass of O}$$

or

$$2(1.008\ \text{amu}) + 16.00\ \text{amu} = 18.02\ \text{amu}$$

For example:

$$\text{Number of moles of } (NH_2)_2CO = \frac{mass}{(14+(2\times1))\times2+12+16}$$

Finally, note that for ionic compounds like NaCl and MgO that do not contain discrete molecular units, we use the term formula mass instead. The formula unit of NaCl consists of one Na^+ ion and one Cl^- ion. Thus, the formula mass of NaCl is the mass of one formula unit:

$$\begin{aligned}\text{formula mass of NaCl} &= 22.99\ \text{amu} + 35.45\ \text{amu}\\ &= 58.44\ \text{amu}\end{aligned}$$

and its molar mass is 58.44 g.

Molarity (M): symbol (M) is the number of moles of a substance per liter of solution. Such as "$[H^+]$" moles per liter (M), means the concentration of H^+.

$$M = \frac{number\ of\ moles\ of\ solute}{Volume\ of\ solution\ in\ liter}$$

$$M = \frac{n}{V(L)}$$

A **liter (L)** is the volume of a cube that is 10 cm on each edge.

$$M = \frac{density\ \times \%\times 1000}{Molar\ mass}$$ For Liquid

Also, $$M = \frac{(mass)}{(M.mass)} \times \frac{1000}{Vol.(mL)}$$ For solid

Notes: mass → means mass in gram unit (gm.)

vol. → means volume of solution in milliliter unit (ml.)

Lec.5

Example 1

A 500-mL solution containing 0.730 mole of $C_6H_{12}O_6$, calculate the molarity of the solution using Equation?

$$M = \text{molarity} = \frac{0.730 \text{ mol}}{0.500 \text{ L}}$$
$$= 1.46 \text{ mol/L} = 1.46\ M$$

As you can see, the unit of molarity is moles per liter, so a 500-mL solution containing 0.730 mole of $C_6H_{12}O_6$ is equivalent to 1.46 mol/L or 1.46 M.

Example 2

What is the molarity? of a solution containing (16 gm) CH_3OH in 200 ml of solution? M.mass =32.

$M = \frac{(mass)}{(M.mass)} \times \frac{1000}{Vol.(mL)}$, $M = \frac{(16\ gm)}{(32\ g\ /mol)} \times \frac{1000}{(200\ mL)}$, M = 2.5 M

Example 3

Determine the mass of Na^+ ? (22.99 g/mol) in 25.0 g Na_2CO_3 (106 g/mol).

Normality: symbol (N) is the number of milliequivalents of solute contained in 1 ml of solution, or the number of gram equivalents contained in 1 L.

$$N = \frac{number\ of\ gram-equivalent\ mass\ of\ solute}{Volume\ of\ solution\ in\ liters}$$

Number of gram-equivalent mass = $\frac{mass}{eq.mass}$

eq.mass → means equivalent mass unit = ($\frac{gm}{gm.m.eq.}$)

$$N = \frac{mass}{eq.mass} \times \frac{1000}{vol.(mL.)}$$

$$N = \frac{density \times \% \times 1000}{eq.mass}$$

3.1 General Properties of Aqueous Solutions

Molar mass (M.mass) of compound is the sum of the atomic mass of all the atoms in the molecular formula of the compound.

For example:

The molecular mass of H_2O is

= 2(atomic mass of H) + 1* (atomic mass of O)

= 2(1.008 amu) + 1* (16)

For example:

The molecular mass of $(NH_2)_2CO$ = $\left(1 * 14 + (2 \times 1)\right) \times 2 + 1 * 12 + 1 * 16$

Calculation of equivalent mass (eq.mass)

The equivalent mass of acid, is that mass of acid which contains one –gram atom of replaceable hydrogen.

Eq.wt of **acid** = $\frac{M.mass\ of\ acid}{number\ of\ replaceable(hydrogen)}$

Example: 4

eq.wt of **HCl** = $\frac{M.mass\ of\ acid}{number\ of\ replaceable(hydrogen)} = \frac{M.mass\ of\ acid}{1}$

eq.wt of $\mathbf{H_2SO_4}$ = $\frac{M.mass\ of\ acid}{number\ of\ replaceable(hydrogen)} = \frac{M.mass\ of\ acid}{2}$

eq.wt of $\mathbf{H_3PO_4}$ = $\frac{M.mass\ of\ acid}{number\ of\ replaceable(hydrogen)} = \frac{M.mass\ of\ acid}{3}$

While, the equivalent mass of base is that mass of base which contains one replaceable hydroxyl group.

Eq.wt of **base** = $\frac{M.mass\ of\ base}{number\ of\ replaceable(hydroxyl)}$

Example: 5

eq.wt of **NaOH** = $\frac{M.mass\ of\ base}{number\ of\ replaceable(hydroxyl)}$

eq.wt of $\mathbf{Ba(OH)_2}$ = $\frac{M.mass\ of\ base}{number\ of\ replaceable(hydroxyl)}$

eq.wt of **$Al(OH)_3$** = $\frac{M.mass\ of\ base}{3}$

The equivalent weight of the substance which contains or reacts with 1 gm. atm of a univalent cation M^+.

Eq.wt = $\frac{M.mass}{total\ charge\ of\ positive\ ions}$

Example: 6

eq.wt of **KCl** = $\frac{M.mass}{1}$

eq.wt of **Na_2CO_3** = $\frac{M.mass}{2}$

eq.wt of **$BaCl_2$** = $\frac{M.mass}{2}$

eq.wt of **$FeCl_3$** = $\frac{M.mass}{3}$

eq.wt of **$Ca_3(PO_4)_2$** = $\frac{M.mass}{2}$

Example: 7

How many gram – equivalent of solute are contained in 0.5 L of 0.2 N solution?

$0.2 = \frac{\text{gram – equivalent of solute}}{0.5}$

Example: 8

How many gram of solute are required to prepare 1 L of 1 N solution of NaCl, M.mass = 58.45?

Equivalence number: which one atom of the element combines with or displaces. Such as NaCl → M.mass = eq.wt while in H_2SO_4 M.mass /2

Lec.6

Percentage Composition (mass%):

The percentage of a component in a mixture or solution is usually expressed as a mass percent (mass %):

Mass percent = $\frac{mass\ of\ solute}{mass\ of\ solution\ or\ mixture} \times 100$

Example: 9

Nitric acid is solid a 70 mass % solution, which means that the reagent contains 70 g of HNO_3 per 100 g of solution.

Volume percent (vol %)

is defined as:

Volume percent = $\frac{volume\ of\ solute}{volume\ of\ solution\ or\ mixture} \times 100$

Example: 10

Ethanol is 95 vol%; this expression means 95 mL of ethanol per 100 mL of total solution. The remainder is water.

Mass- Volume percent%

(mass/v)percent = $\frac{mass\ of\ solute\ (gm)}{volume\ of\ solution\ or\ mixture\ (ml)} \times 100$

Example: 11

5 mass % aqueous silver nitrate; this expression means dissolve 5 g of silver nitrate per 100 mL of total solution. The remainder is water.

Density and specific gravity of solutions:

Density = $\frac{mass\ (g\ or\ kg)}{volume\ (ml\ or\ L)}$

Specific gravity (Sp.gr.) = $\frac{density\ of\ a\ substance}{density\ of\ water\ at\ 4°C}$

Because the density of water at 4°C is very close to 1 g/mL, specific gravity is nearly the same as density.

Parts per Million (ppm):

For very diluted solution, the concentration is more conveniently expressed is part per million (ppm), which mean grams of substance per million or billion grams of total solution or mixture.

ppm = $\frac{mass\ of\ solute\ (g)}{mass\ of\ solution} * 10^6$ mg / L **or** µg/ml

Example: 12

How much ppm in an aqueous solution containing 0.0003 % nickel?

0.0003 % nickel → means 0.0003 g nickel solute in 100 g solution.

ppm = $\frac{mass\ of\ solute\ (g)}{mass\ of\ solution} \times 10^6 = \frac{0.0003}{100} \times 10^6$

The relation between (ppm) and (molarity) is:

ppm = molarity× M.mass × 1000

The relation between (ppm) and (normality) is:

ppm = normality× eq.mass × 1000

Example: 13

An aqueous solution of $NiCl_2$ with a concentration of 500 ppm. What is the molarity and normality of this solution? At.mass of Ni =58.69, Cl = 35.5

M.mass for $NiCl_2$= 58.69+ 2×35.5

=129.69 g/mol

ppm = molarity× M.mass × 1000

500 = molarity× 129.69 × 1000 → M= 0.0038

ppm = normality× eq.mass × 1000

500 = normality×129.69/2 × 1000 → N= 0.0077

Dilution of Solutions

Dilution is the procedure for preparing a less concentrated solution from a more concentrated one.

Suppose that we want to prepare 1 L of a 0.400 M HCL solution from a solution of 1.00 M HCL?

HCL conc. = HCL dil.

$M_1*V_1=M_2*V_2$→ 1 M *V_1=0.4 M*1 L→ V_1= 0.4 L

$$\underbrace{M_iV_i}_{\text{moles of solute before dilution}} = \underbrace{M_fV_f}_{\text{moles of solute after dilution}}$$

where M_i and M_f are the initial and final concentrations of the solution in molarity and V_i and V_f are the initial and final volumes of the solution, respectively. Of course, the units of V_i and V_f must be the same (mL or L) for the calculation to work.

Lec.7

Homework

How many grams of potassium dichromate ($K_2Cr_2O_7$) are required to prepare a 250-mL (1 solution whose concentration is 2.16 *M* ?

2) What is the molarity of an 85.0-mL ethanol (C_2H_5OH) solution containing 1.77 g of ethanol?

3) In a biochemical essay, a chemist needs to add 3.81 g of glucose to a reaction mixture. Calculate the volume in milliliters of a 2.53 *M* glucose ($C_6H_{12}O_6$) solution she should use for the addition.

4) How would you prepare 500 mL of a 1.75 *N* H_2SO_4 solution, starting with an 8.61 *N* stock solution of H_2SO_4?

5) How would you prepare 250 mL of a 4 x 10^3 *ppm* NaOH solution?

6) In a titration experiment, a student finds that 25 mL of NaOH solution are needed to neutralize 0.6 g of K_2HPO_4. What is the concentration (in *ppm*) of the NaOH solution?

7) Calculate the volume required to prepare 500 mL of 3650 *ppm* HCl solution. (Sp. gr. =1.18, % = 36 %).

8) Determine the mass of Na^+ (22.99 g/mol) in 25.0 g Na_2CO_3 (106 g/mol)?

9) Suppose that we want to prepare 1 L of a 0.400 M HCL solution from a solution of 1.00 M HCL?

10) How can you prepare 500 ml of 0.1 M HCl, if the percentage is 35%, and density = 1.18 g/ml, M.mass= 36.5g/mol?

Solutions

1- How many grams of potassium dichromate (K2Cr2O7 are required to prepare a (250 mL solution whose concentration is 2.16 M?

mass? Xmass? X

Vol.= 250 ml. √

M= 2.16 M √

$$M = \frac{(mass)}{(M.mass)} \times \frac{1000}{Vol.(mL)}$$

$M.mass$ of $K_2Cr_2O_7$ = 2*39+ 2*52 +7*16 = 294.185 g/ mol.

2.16 M = $\frac{mass}{294.185 \text{ g/ mol.}} \times \frac{1000}{250\ mL}$

mass = 158.86 g

2-What is the molarity of an 85.0 mL ethanol (C2H5OH) solution containing 1.77 g? of ethanol.

mass = 1.77 g √

Vol. = 85.0 ml. √

M= ? X

M = $\frac{(mass)}{(M.mass)} \times \frac{1000}{Vol.(mL)}$

$M.mass$ of C_2H_5OH = 2*12+ 6*1 +1*16 = 46 g/ mol.

M = $\frac{1.77}{46 \text{ g/ mol.}} \times \frac{1000}{85.0\ mL}$

M = 0.452 M

3-In a biochemical essay, a chemist needs to add 3.81 g of glucose to a reaction mixture. Calculate the volume in milliliters of a 2.53 M glucose (C6H12O6 solution she should use for the addition

mass = 3.81 g √

X Vol. = ?

√ M= 2.53

M = $\frac{(mass)}{(M.mass)} \times \frac{1000}{Vol.(mL)}$

$M.mass$ of $C_6H_{12}O_6$ = 6*12+ 12*1 +6*16 = 180 g/ mol.

2.53 = $\frac{3.81}{180 \text{ g/ mol.}} \times \frac{1000}{Vol.(mL)}$

Vol. = 8.366 ml.

4- How would you prepare 500 mL of a 1.75 N H2 ,SO4 solution, starting with an 8.61 N stock solution of H2 SO4?

√ V_1= 500 ml.

√ N_1= 1.75 N

X V_2= ?

√ N_2= 8.61 N

$N_1 * V_1 = N_2 * V_2$

1.75 N * 500 ml.= 8.61 N * V_2

V_2 = 101.62 ml.

5-How would you prepare 250 mL of a 4 x 103 ppm NaOH solution

√ ppm = 4 x 10^3

√ Vol. = 250 ml.

mass = ? X

M.mass = 23*1+ 16*1+1*1 = 40 g/ mol.

ppm = M* M.mass * 1000

4 x 10^3 = M* 40 * 1000

M= 0.1 M

$$M = \frac{(mass)}{(M.mass)} \times \frac{1000}{Vol.(mL)}$$

$$0.1\ M = \frac{mass}{40\ \text{g/ mol.}} \times \frac{1000}{250\ mL}$$

mass = 1 g

That mean, we weight 1 g of NaOH and dissolve in 250 ml. of H_2Oin volumetric flask

6-In a titration experiment, a student finds that 25 mL of NaOH solution are needed to neutralize 0.6 g of K2HPO4?

What is the concentration (in ppm) of the NaOH solution .

$$M = \frac{number\ of\ moles\ of\ solute\ (NaOH)}{Volume\ of\ solution\ in\ liter} \rightarrow M = \frac{number\ of\ moles\ of\ solute}{0.025\ L.}$$

First, we need to find the number of moles of NaOH:

K_2HPO_4 being "monoprotic" means that one mole of K_2HPO_4 is one equivalent. Basically, 1 molecule of K_2HPO_4 only donates 1 H^+ ion.

We also know that NaOH is a monoprotic base, because there's only one OH^- ion in its chemical formula.

Therefore, 1 mole of K_2HPO_4 will correspond to 1 mole of NaOH in a neutralisation reaction.
In other words, 1 mole of NaOH will neutralise 1 mole of K_2HPO_4.

M.mass for K_2HPO_4 = 2*39+ 1*1+1*32+4*16 = 175 g/ mol.

Number of moles for K_2HPO_4 = $\frac{mass}{Molar\ mass}$ → Number of moles for K_2HPO_4 = $\frac{0.6\ g}{175\ \text{g/ mol}}$ → 0.003429 mol.

Because the mole ratio of K_2HPO_4 to NaOH is 1:1, 0.003429 moles of NaOH

M = $\frac{number\ of\ moles\ of\ solute\ (NaOH)}{Volume\ of\ solution\ in\ liter}$ → M = $\frac{0.003429\ \text{mol.}}{0.025\ L.}$ → M = 0.13716 M

ppm = M * M.mass * 1000 → ppm = 0.13716 M * 40 * 1000 → ppm = 5.486

7- Calculate the volume required to prepare 500 mL of 3650 ppm HCl → solution

$Vol._{dil.}$ = 500 ml. √

ppm= 3650 √

Sp. gr. =1.18 √

% = 36 % √

$Vol._{conc.}$ = ? X

$Molar\ mass$= 1*1+ 1*35.5 = 36.5 g /mol.

$M_{conc.} = \frac{density \times \% \times 1000}{Molar\ mass}$

$M_{conc.} = \frac{1.18 \times 36l100 \times 1000}{36.5}$

$M_{conc.} = 11.63$ M

ppm = $M_{dil.}$* M.mass * 1000

$M_{dil.} = 0.1$ M 3650 = $M_{dil.}$* 36.5 * 1000 →

$M_{dil.} * V_{dil.} = M_{conc.} * V_{conc.}$

0.1 M * 500 ml.= 11.63 M * $V_{conc.}$

$V_{conc.} = 4.299$ ml.

8- Determine the mass of Na g/mol) in 25.0 g Na2CO3

?(g/mol 106)

mass of Na^+ = ? X

M.mass of Na^+ = 22.99 g/mol √

mass of Na_2CO_3 = 25.0 g √

M.mass of Na_2CO_3 = 106 g/mol √

$number\ of\ moles\ of\ solute$ (Na^+)
=2 ∗ $number\ of\ moles\ of\ solute$ (Na_2CO_3)

$$\frac{(mass\ of\ Na^+)}{(A.mass\ of\ Na^+)} = \frac{2*(mass\ of\ Na_2CO_3)}{(M.mass\ of\ Na_2CO_3)}$$

$$\frac{(mass\ of\ Na^+)}{(22.99\ \frac{g}{mol})} = \frac{2*(25.0\ g)}{(106\ \frac{g}{mol})}$$

→ $mass\ of$ Na^+ = 10.844 g

9-Suppose that we want to prepare 1 L of a 0.400 M HCL solution from → a solution of 1.00 M HCL?

$V_{dil.}$= 1 L √

√ $M_{dil.}$= 0.400 M

X $V_{conc.}$= ?

√ $M_{conc.}$= 1.00 M

HCL $_{conc.}$ = HCL $_{dil.}$

$M_1*V_1=M_2*V_2 \rightarrow$ 1 M $*V_1$=0.4 M*1 L

$\rightarrow V_1$= 0.4 L

10-How can you prepare 500 ml of 0.1 M HCl, if the percentage is 35%, → and density = 1.18 g/ml, M.mass= 36.5g/mol?

Vol.$_{dil.}$ = 500 ml. √

$M_{dil.}$ = 0.1 √

Sp. gr. =1.18 √

% = 35 % √

Vol.$_{conc.}$ = ? X

$Molar\ mass$= 1*1+ 1*35.5 = 36.5 g /mol.

$$M_{conc.} = \frac{density \times \% \times 1000}{Molar\ mass}$$

$$M_{conc.} = \frac{1.18 \times 35l100 \times 1000}{36.5}$$

$M_{conc.}$ = 11.315 M

$M_{dil.} * V_{dil.} = M_{conc.} * V_{conc.}$

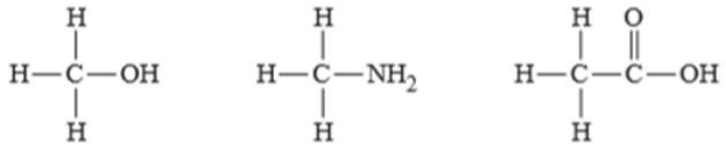
0.1 M * 500 ml.= 11.315 M * $V_{conc.}$

$V_{conc.}$ = 4.4 19 ml.

Lec.8

5.1 **Introduction to Organic Chemistry**

The branch of chemistry that deals with carbon compounds is **Organic Chemistry**. The simplest type of organic compounds is the hydrocarbons, which contain only carbon and hydrogen atoms. Carbon can form more compounds than most other elements because carbon atoms are able not only to form single, double, and triple carbon-carbon bonds, but also to link up with each other in chains and ring structures. The hydrocarbons are used as fuels for domestic and industrial heating, for generating electricity and powering internal combustion engines, and as starting materials for the chemical industry. One class of hydrocarbons is called the alkanes. The chemistry of organic compounds is largely determined by the functional groups, which consist of one or a few atoms bonded in a specific way. For example, when an H atom in methane is replaced by a hydroxyl group (OOH), an amino group (ONH_2), and a carboxyl group (OCOOH), the following molecules are generated:



Methanol Methylamine Acetic acid

Classes of Organic compounds

Classes of organic compounds can be distinguished according to functional groups they contain. A functional group is a group of atoms that is largely responsible for the chemical behaviour of the parent molecule. Different molecules containing the same kind of functional group or groups undergo similar reactions. Thus, by learning the characteristic properties of a

Few functional groups, we can study and understand the properties of many organic compounds. The functional groups known as alcohols, ethers, aldehydes and ketones, carboxylic acids, and amines. All organic compounds are derived from a group of compounds known as hydrocarbons because they are made up of only hydrogen and carbon. On the basis of structure, hydrocarbons are divided into two main classes; aliphatic and aromatic.

2.5 Aliphatic Hydrocarbons

Aliphatic hydrocarbons do not contain the benzene group, or the benzene ring.

2.5.1 Alkanes

Alkanes are hydrocarbons that have the general formula CnH_{2n+2}, where, 2, ...the essential characteristic of alkanes is that only single covalent bonds are present. The alkanes are known as saturated hydrocarbons because they contain the maximum number of hydrogen atoms that can bond with the number of carbon atoms present. The simplest alkane (that is, with) is methane CH_4, which is a natural product of the anaerobic bacterial decomposition of vegetable matter under water. Because it was first collected in marshes, methane became known as “marsh gas.” Commercially, methane is obtained from natural gas. The carbon atoms in all the alkanes can be assumed to be sp^3-

hybridized. The structures of ethane and propane are straight forward, for there is only one way to join the carbon atoms in these molecules. Butane, however, has two possible bonding schemes resulting in different compounds called n-butane (n stands for normal) and isobutane. n-butane is a straight-chain alkane because the carbon atoms are joined in a continuous chain. In a branched-chain alkane like isobutane, one or more carbon atoms are bonded to a nonterminal carbon atom. Isomers that differ in the order in which atoms are connected are called structural isomers. In the alkane series, as the number of carbon atoms increases, the number of structural isomers increases rapidly. For example, C_4H_{10} has two isomers; $C_{10}H_{22}$ has 75 isomers; and $C_{30}H_{62}$ has over 400 million possible isomers.

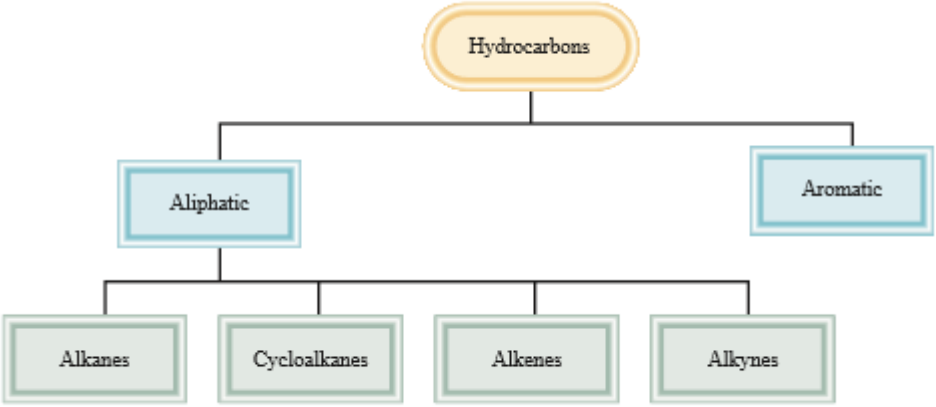


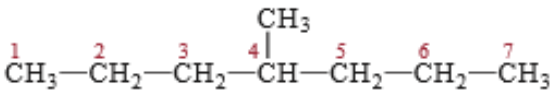
Table 5.1 the first 10 straight - chain alkane

Name of Hydrocarbon	Molecular Formula	Number of Carbon Atoms	Melting Point (°C)	Boiling Point (°C)
Methane	CH_4	1	−182.5	−161.6
Ethane	$CH_3—CH_3$	2	−183.3	−88.6
Propane	$CH_3—CH_2—CH_3$	3	−189.7	−42.1
Butane	$CH_3—(CH_2)_2—CH_3$	4	−138.3	−0.5
Pentane	$CH_3—(CH_2)_3—CH_3$	5	−129.8	36.1
Hexane	$CH_3—(CH_2)_4—CH_3$	6	−95.3	68.7
Heptane	$CH_3—(CH_2)_5—CH_3$	7	−90.6	98.4
Octane	$CH_3—(CH_2)_6—CH_3$	8	−56.8	125.7
Nonane	$CH_3—(CH_2)_7—CH_3$	9	−53.5	150.8
Decane	$CH_3—(CH_2)_8—CH_3$	10	−29.7	174.0

Alkane Nomenclature 2.5.1.1

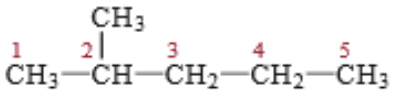
The nomenclature of alkanes and all other organic compounds is based on the recommendations of the International Union of Pure and Applied Chemistry (IUPAC). The first four alkanes (methane, ethane, propane, and butane) have non-systematic names. As table 5.1 shows, the number of carbon atoms is reflected in the Greek prefixes for the alkanes containing 5 to 10 carbons. We now apply the IUPAC rules to the following examples:

1. The parent name of the hydrocarbon is that given to the longest continuous chain of carbon atoms in the molecule such as; compound is heptane because there are seven carbon atoms in the longest chain:

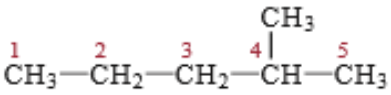


2. An alkane less one hydrogen atom is an alkyl group. For example, when a hydrogen atom is removed from methane, we are left with the CH_3 fragment, which is called a methyl group. Similarly, removing a hydrogen atom from the ethane molecule gives an ethyl group, or C_2H_5. Any chain branching off the longest chain is named as an alkyl group.

3. When one or more hydrogen atoms are replaced by other groups, the name of the compound must indicate the locations of carbon atoms where replacements are made. The procedure is to number each carbon atom on the longest chain in the direction that gives the smaller numbers for the locations of all branches. Consider the two different systems for the same compound shown below:

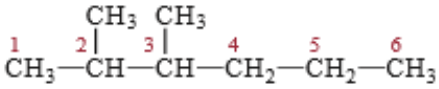


2-methylpentane

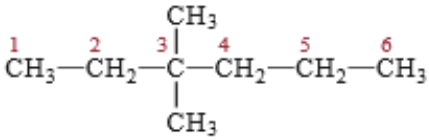


4-methylpentane

4. When there is more than one alkyl branch of the same kind present, we use a prefix such as di-, tri-, or tetra- with the name of the alkyl group. Consider the following examples:

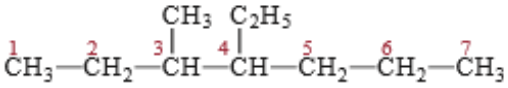


2,3-dimethylhexane



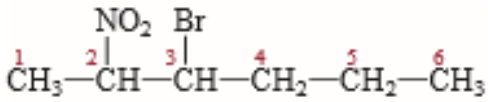
3,3-dimethylhexane

When there are two or more different alkyl groups, the names of the groups are listed alphabetically. For example,



4-ethyl-3-methylheptane

5. Of course, alkanes can have many different types of substituents. Thus, the compound is called 3-bromo-2-nitrohexane. Note that the substituent groups are listed alphabetically in the name, and the chain is numbered in the direction that gives the lowest number to the first substituted carbon atom.



Reactions of Alkanes 2.5.1.2

Alkanes are generally not considered to be very reactive substances. However, under suitable conditions they do react. For example, natural gas, gasoline, and fuel oil are alkanes that undergo highly exothermic combustion reactions:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \quad \Delta H^\circ = -890.4 \text{ kJ/mol}$$
$$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(l) \quad \Delta H^\circ = -3119 \text{ kJ/mol}$$

Halogenation of alkanes—that is, the replacement of one or more hydrogen atoms by halogen atoms—is another type of reaction that alkanes undergo. When a mixture of methane and chlorine is heated above 100 ^{0}C or irradiated with light of a suitable wavelength, methyl chloride is produced:

$$CH_4(g) + Cl_2(g) \longrightarrow \underset{\text{methyl chloride}}{CH_3Cl(g)} + HCl(g)$$

(Free radical mechanism)

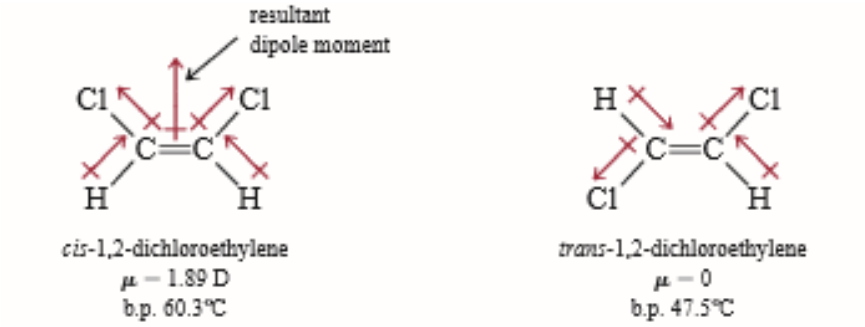
Lec.9

Alkenes

The alkenes (also called olefins) contain at least one carbon-carbon double bond. Alkenes have the general formula C_nH_{2n}, where n = 2, 3, ...The simplest alkene is C_2H_4, ethylene, in which both carbon atoms are sp^2-hybridized and the double bond is made up of a sigma bond and a pi bond.

Geometric Isomers of Alkenes 2.5.1.3

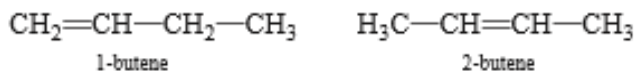
In a compound such as ethane, C_2H_6, the rotation of the two methyl groups about the carbon-carbon single bond (which is a sigma bond) is quite free. The situation is different for molecules that contain carbon-carbon double bonds, such as ethylene, C_2H_4. In addition to the sigma bond, there is a pi bond between the two carbon atoms. Rotation about the carbon-carbon linkage does not affect the sigma bond, but it does move the two 2pz orbitals out of alignment for overlap and, hence, partially or totally destroys the pi bond. This process requires an input of energy on the order of 270 kJ/mol. For this reason, the rotation of a carbon-carbon double bond is considerably restricted, but not impossible. Consequently, molecules containing carbon carbon double bonds (that is, the alkenes) may have geometric isomers, which have the same type and number of atoms and the same chemical bonds but different spatial arrangements. Such isomers cannot be interconverted without breaking a chemical bond. The molecule dichloroethylene, ClHCPCHCl, can exist as one of the two geometric isomers called cis-1,2-dichloroethylene and trans-1,2-dichloroethylene:



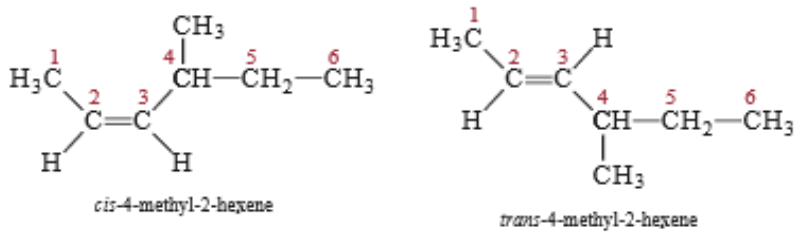
where the term cis means that two particular atoms (or groups of atoms) are adjacent to each other, and trans means that the two atoms (or groups of atoms) are across from each other. Generally, cis and trans isomers have distinctly different physical and chemical properties. Heat or irradiation with light is commonly used to bring about the conversion of one geometric isomer to another, a process called cis-trans isomerization, or geometric isomerization.

Alkene Nomenclature 2.5.1.4

In naming alkenes we indicate the positions of the carbon-carbon double bonds. The names of compounds containing C=C bonds end with -ene. As with the alkanes, the name of the parent compound is determined by the number of carbon atoms in the longest chain, as shown here:



The numbers in the names of alkenes refer to the lowest numbered carbon atom in the chain that is part of the C=C bond of the alkene. The name "butene" means that there are four carbon atoms in the longest chain. Alkene nomenclature must specify whether a given molecule is cis or trans if it is a geometric isomer, such as:



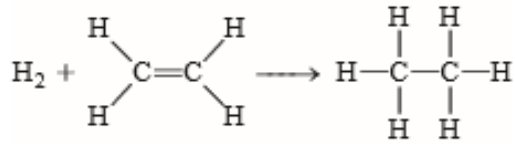
Properties and Reactions of Alkenes 2.5.1.5

Ethylene is an extremely important substance because it is used in large quantities in manufacturing organic polymers (very large molecules) and in preparing many other organic chemicals. Ethylene and other alkenes are prepared industrially by the cracking process, that is, the thermal decomposition of a large hydrocarbon into smaller molecules. When ethane is heated to about 800 ^{0}C in the presence of platinum, it undergoes the following reaction:

$$C_2H_6(g) \xrightarrow[\text{catalyst}]{\text{Pt}} CH_2{=}CH_2(g) + H_2(g)$$

The platinum acts as a catalyst, a substance that speeds up a reaction without being used up in the process and therefore does not appear on either side of the equation. Other alkenes can be prepared by cracking the higher members of the alkane family.

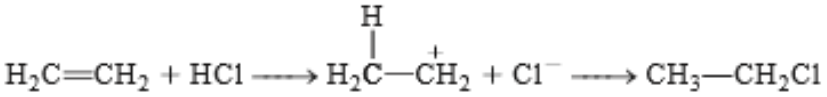
Alkenes are classified as unsaturated hydrocarbons, compounds with double or triple carbon-carbon bonds. Unsaturated hydrocarbons commonly undergo addition reactions in which one molecule adds to another to form a single product. An example of an addition reaction is hydrogenation, which is the addition of molecular hydrogen to compounds containing C=C and C triple C bonds.



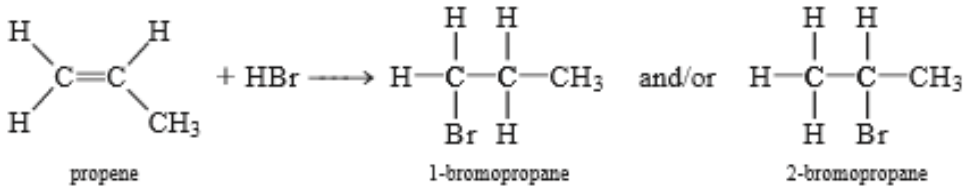
Other addition reactions to the C=C bond involve the hydrogen halides and halogens:

$$H_2C{=}CH_2 + X_2 \longrightarrow CH_2X{-}CH_2X$$
$$H_2C{=}CH_2 + HX \longrightarrow CH_3{-}CH_2X$$

in which X represents a halogen atom. The electron density maps of HCl and ethylene. When the two molecules react, the interaction is between the electron-rich region (pi electrons of the double bond) and the electron-poor region of HCl, which is the H atom. The steps are:



The addition of a hydrogen halide to an unsymmetrical alkene such as propene is more complicated because two products are possible:



Alkynes

Alkynes contain at least one carbon-carbon triple bond. They have the general formula C_nH_{2n-2}, where n= 2, 3, ...

Alkyne Nomenclature 2.5.1.6

Names of compounds containing triple bonds end with -yne. Again the name of the parent compound is determined by the number of carbon atoms in the longest chain. As in the case of alkenes, the names of alkynes indicate the position of the carbon-carbon triple bond, as, for example, in:

$HC{\equiv}C{-}CH_2{-}CH_3$ $\quad$ $H_3C{-}C{\equiv}C{-}CH_3$

1-butyne $\quad$ 2-butyne

Properties and Reactions of Alkynes 2.5.1.7

The simplest alkyne is ethyne, better known as acetylene (C_2H_2). Acetylene is a colorless gas (B.p. =84 ^{0}C) prepared in the laboratory by the reaction between calcium carbide and water:

$$CaC_2(s) + 2H_2O(l) \longrightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

Industrially, it is prepared by the thermal decomposition of ethylene at about 1100 ^{0}C:

$$C_2H_4(g) \longrightarrow C_2H_2(g) + H_2(g)$$

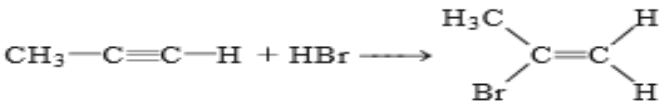
Acetylene has many important uses in industry. Because of its high heat of combustion:

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l) \quad \Delta H^\circ = -2599.2 \text{ kJ/mol}$$

It undergoes these addition reactions with hydrogen halides and halogens:

$$CH{\equiv}CH(g) + HX(g) \longrightarrow CH_2{=}CHX(g)$$
$$CH{\equiv}CH(g) + X_2(g) \longrightarrow CHX{=}CHX(g)$$
$$CH{\equiv}CH(g) + 2X_2(g) \longrightarrow CHX_2{-}CHX_2(l)$$

Methylacetylene (propyne), $CH_3{-}C{\equiv}C{-}H$, is the next member in the alkyne family. It undergoes reactions similar to those of acetylene. The addition reactions of propyne also obey Markovnikov's rule:

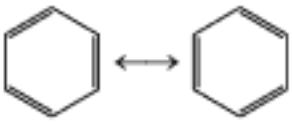


propyne $\quad$ 2-bromopropene

Lec. 10

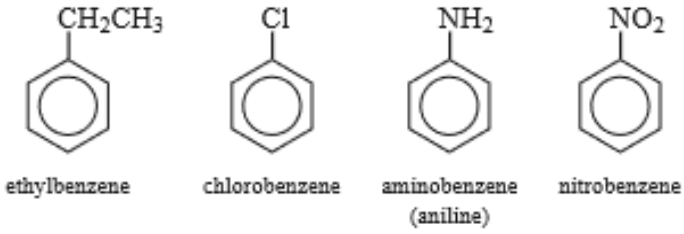
Aromatic Hydrocarbons 2.5.1.8

Benzene (C_6H_6) is the parent compound of this large family of organic substances. The properties of benzene are best represented by both of the following resonance structures.



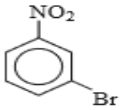
Nomenclature of Aromatic Compounds

The naming of mono substituted benzenes, that is, benzenes in which one H atom has been replaced by another atom or a group of atoms, is quite straightforward, as shown next:



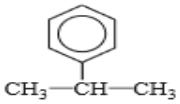
If more than one substituent is present, we must indicate the location of the second group relative to the first. The systematic way to accomplish this is to number the carbon atoms as follows:

The prefixes *o*- (*ortho*-), *m*- (*meta*-), and *p*- (*para*-) are also used to denote the relative positions of the two substituted groups, as just shown for the dibromobenzenes. Compounds in which the two substituted groups are different are named accordingly. Thus,

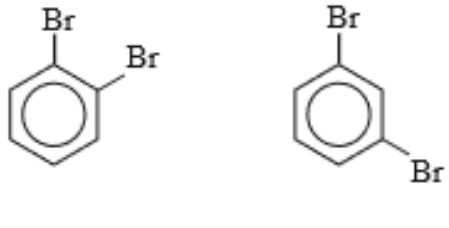


is named 3-bromonitrobenzene, or *m*-bromonitrobenzene.

Finally we note that the group containing benzene minus a hydrogen atom (C_6H_5) is called the *phenyl* group. Thus, the following molecule is called 2-phenylpropane:



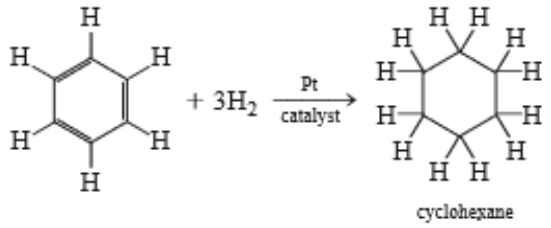
Three different dibromobenzenes are possible:



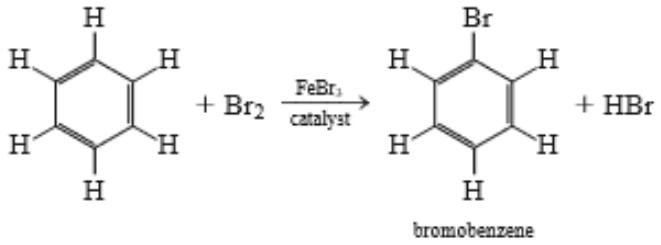
1,2-dibromobenzene (*o*-dibromobenzene) | 1,3-dibromobenzene (*m*-dibromobenzene) | 1,4-dibromobenzene (*p*-dibromobenzene)

Properties and Reactions of Aromatic Compounds 2.5.1.9

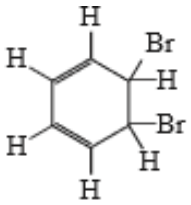
Benzene is a colorless, flammable liquid obtained chiefly from petroleum and coal tar. Perhaps the most remarkable chemical property of benzene is its relative inertness. Although it has the same empirical formula as acetylene (CH) and a high degree of unsaturation, it is much less reactive than either ethylene or acetylene. The stability of benzene is the result of electron delocalization. In fact, benzene can be hydrogenated, but only with difficulty. The following reaction is carried out at significantly higher temperatures and pressures than are similar reaction for the alkenes:



We saw earlier that alkenes react readily with halogens and hydrogen halides to form addition products, because the pi bond in C=C can be broken more easily. The most common reaction of halogens with benzene is substitution. For example,



Note that if the reaction were addition, electron delocalization would be destroyed in the product.

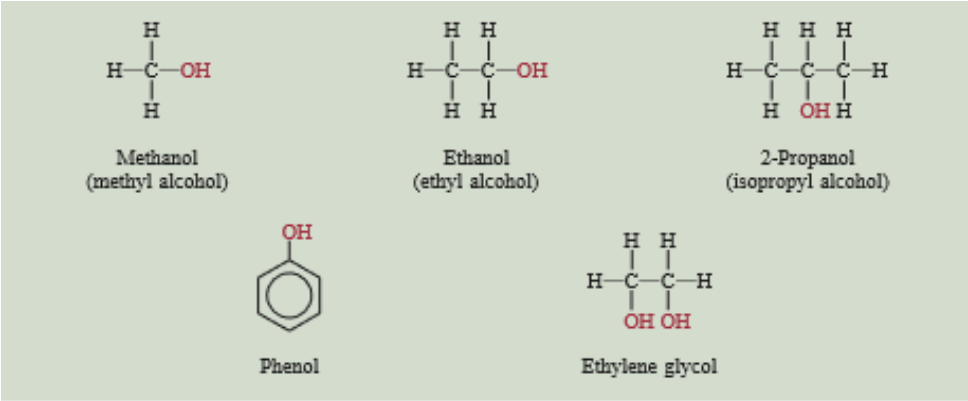


and the molecule would not have the aromatic characteristic of chemical unreactivity. Alkyl groups can be introduced into the ring system by allowing benzene to react with an alkyl halide using $AlCl_3$ as the catalyst:

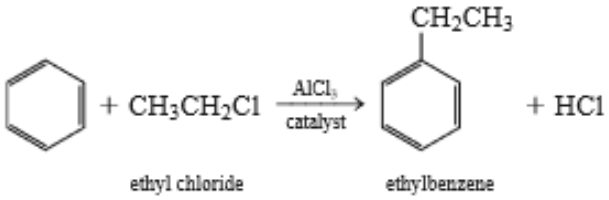
Chemistry of the Functional Groups

Alcohols

All alcohols contain the hydroxyl functional group, —OH.



Commercially, ethanol is prepared by an addition reaction in which water is combined with ethylene at about 280 ^{0}C and 300 atm:



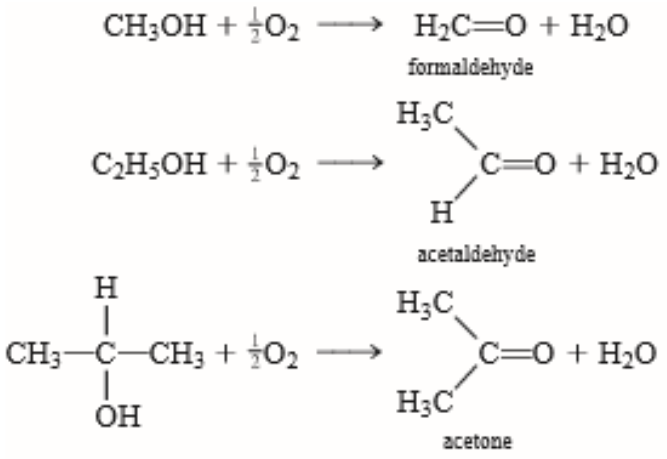
$$CH_2{=}CH_2(g) + H_2O(g) \xrightarrow{H_2SO_4} CH_3CH_2OH(g)$$

The alcohols are very weakly acidic; they do not react with strong bases, such as NaOH. The alkali metals react with alcohols to produce hydrogen:

$$2CH_3OH + 2Na \longrightarrow \underset{\text{sodium methoxide}}{2CH_3ONa} + H_2$$

Aldehydes and Ketones

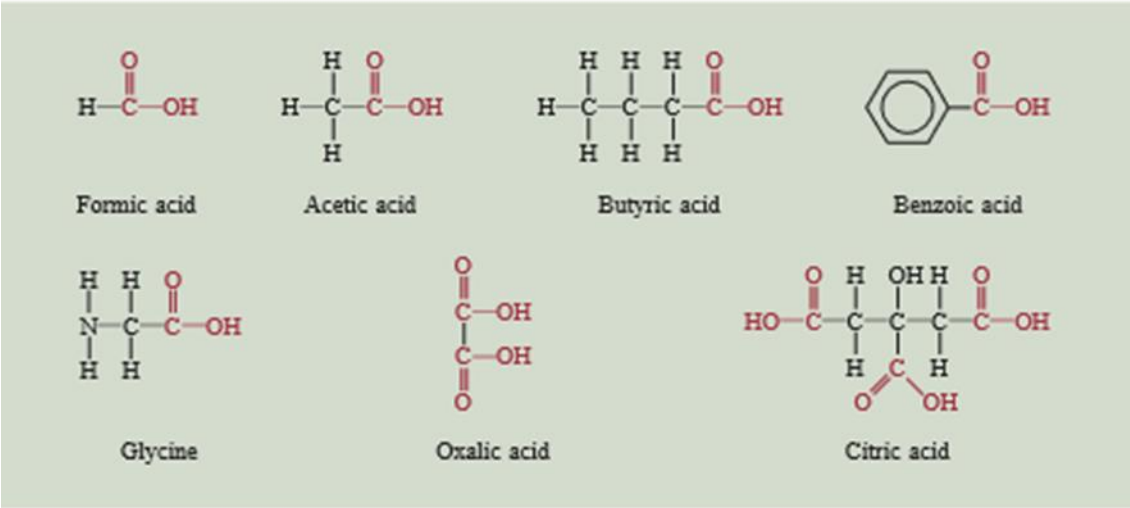
Under mild oxidation conditions, it is possible to convert alcohols to aldehydes and ketones:



The functional group in these compounds is the carbonyl group. In an aldehyde at least one hydrogen atom is bonded to the carbon in the carbonyl group. In a ketone, the carbon atom in the carbonyl group is bonded to two hydrocarbon groups.

Carboxylic Acids

Under appropriate conditions both alcohols and aldehydes can be oxidized to carboxylic acids, acids that contain the carboxyl group, —COOH. Carboxylic acids are widely distributed in nature; they are found in both the plant and animal kingdoms. All protein molecules are made of amino acids, a special kind of carboxylic acid containing an amino group (—NH_2) and a carboxyl group (—COOH).



Lec.11

Biochemistry

Biochemistry is the study of chemical processes that relating to living organisms

:The chemical elements of life .1

Although more than 25 types of elements can be found in biomolecules, six elements are most common. These are called the CHNOPS elements; the letters stand for the chemical abbreviations .of carbon, hydrogen, nitrogen, oxygen phosphorus, and sulfur

:2- Structure and composition of the cell

The cell is the basic structural, functional, and biological unit of all known organisms. A cell is the smallest unit of life. Cells are often called the "building blocks of life". Cells consist of cytoplasm enclosed within a membrane, which contains many biomolecules such as proteins and nucleic acids. Most plant and animal cells are only visible under a microscope, with dimensions between 1 and 100 micrometers. Organisms can be classified as unicellular (consisting of a single cell such as bacteria) or multicellular (including plants and animals). Most unicellular organisms are classed as microorganisms.

3. Biomolecules

3-1. The main classes of biomolecules are: Carbohydrates, Lipids , Proteins , Nucleic acid 3 . Carbohydrates Carbohydrates are polyhydroxy aldehydes or ketones. Carbohydrates consists of سكر كيتوني (يحوي Carbon, Hydrogen and Oxygen. The general formula of carbohydrates is CnH2nOn.

على مجموعة كاربونيل وسطية) سكر الديهايدي (يحوي مجموعة كاربونيل طرفية)

There are three major classes of carbohydrates that based on the number of forming units: 1) Monosaccharides: contain a single unit of polyhydroxyketone or aldehyde such as (glucose) 2) Oligosaccharides: contain 2-10 of monosaccharide units such as(succharose) 3) Polysaccharides: contain hundreds of monosaccharide units such as (starch).

3-2. Lipids:

In biology and biochemistry, a lipid is a biomolecule that is insoluble in aqueous solutions but soluble in nonpolar solvents such as benzene. Classification of lipids: 1) Simple lipids: are fats /oils & Waxes. 2) Compound or Complex lipids: are Phospholipid, Glycolipid & Lipoprotein. 3) Derived lipids are fatty acids, glycerol, fat soluble vitamins,etc.

3-3. Proteins

Proteins are large size molecules . The basic units of protein structure is called amino acids. An amino acid consists of an α- carbon atom attached to an amino group(–NH2), a carboxylic acid group(–COOH), a simple hydrogen atom, and a side chain commonly denoted as (–R). The side

chain (R) is different for each amino acid. A total of 20 amino acids exist in proteins. Amino acids can be released from proteins by hydrolysis. (Hydrolysis is the cleavage of a covalent bond by addition of water in adequate conditions.)

3-4. Nucleic acids

The two main types of nucleic acids are deoxyribonucleic acid (DNA) and ribonucleic acid (RNA). DNA is the genetic material found in all living organisms. RNA, is mostly involved in protein synthesis.

4. The chemical origin of life:

The chemical origin of life refers to the conditions that might have existed and therefore promoted the first replicating life forms. It considers the physical processes and chemical reactions that could have led to early replicator molecules.

Questions:

Q1/What is Biochemistry?

Q2/How can classify organisms according to their cell?

Q3/ What are the main classes of biomolecules?

Q4/ Write the types of carbohydrates and give example of each type.

Q5/ Is lipid soluble in water? What are the main classes of it? write examples for each one.

Q6/ What is an amino acid consists of ?and how can release it from protein?

Q7 What are the types of nucleic acids?