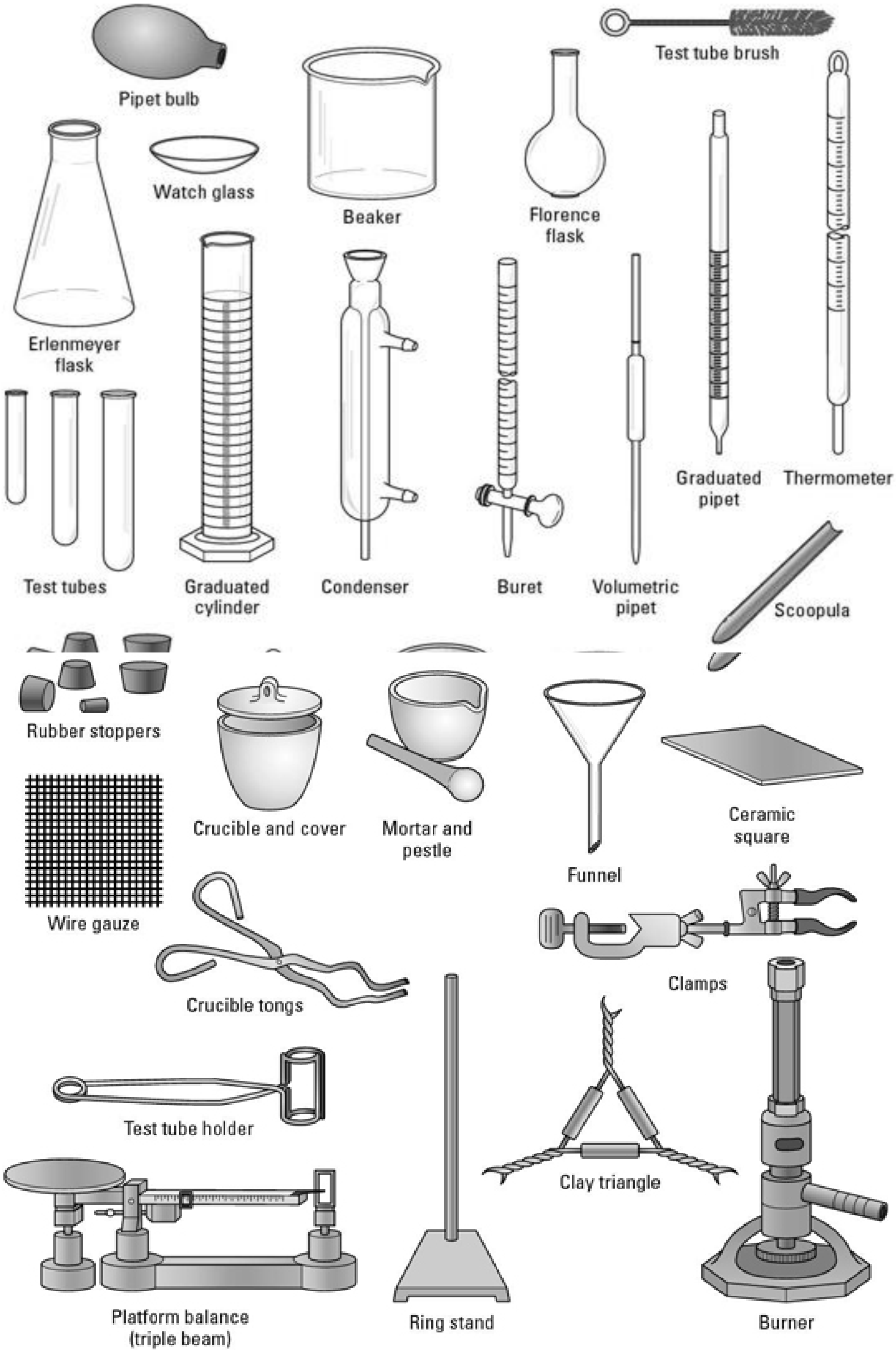
Tikrit University **College of Dentistry**

Chemistry lab:1 **First stage**

Lab Equipments



Lab Equipments

- **Balance:** Used for obtaining the masses of solid and liquid substances.
- **Beaker:** A flat-bottomed, cylindrical piece of glassware used for mixing and heating compounds.
- **Bunsen burner:** Attached to a gas line and lit to provide heat for your experiments.
- **Burette:** An extremely accurate device with a stopcock at the bottom used to measure volumes of reagents.
- **Ceramic Square:** Used to avoid burning the surface of your lab bench.
- **Clamps:** Used to hold a variety of things in place, especially test tubes.
- **Test tube:** Cylindrical open-topped piece of glassware that comes in varying sizes. Used to hold and mix small samples. Stirred by tapping the bottom with two fingers. Fit into both a centrifuge and test tube rack. **Graduated cylinder:** Used to precisely measure volumes.
- **Graduated cylinder:** Used to precisely measure volumes. They are available in various sizes.
- **Metal spatula:** Used to measure out solid substances.
- **Mortar and pestle:** Used to grind solid chemical compounds for chemistry experiments.

- **Pipette bulb:** Used to transfer accurately measured amounts of liquid from one container to another.
- **Rubber stoppers:** Used to close flasks or test tubes to prevent evaporation of liquids or escape of gases.
- **Crucible:** A cup-shaped container capable of sustaining high temperatures. It is used to heat chemical compounds.
- **Crucible tongs:** Used to handle the hot crucible.
- **Clay triangle:** Used to hold a crucible while it's being heated.
- **Condenser:** Used to collect vapors by condensing them into liquids as they contact the liquid-cooled inner surface of the condenser.
- **Erlenmeyer flask (Conical Flask):** Used to hold liquids. The small upper opening slows evaporation, so for some volatile liquids, a flask is a better choice than a beaker. The shape also makes it suitable for mixing and swirling liquids during a titration.
- **Florence flask (Volumetric Flask):** A type of flask, generally round-bottomed, usually suspended and heated from below. Its shape makes it easy to swirl and mix liquids inside of it.
- **Funnel:** Used together with filter paper to filter precipitates out of solutions.
- **Watch glass:** A piece of glassware in the shape of a large contact lens used for evaporating liquids.
- **Wire gauze:** Generally used as a surface for a beaker or flask to rest when being heated by a Bunsen burner.

- **Graduated Pipet**: Can be used to any of the given markings along its side.
- **Volumetric Pipet**: Extremely accurate, but only used for one volume.
- **Test Tube Brush:** An instrument used to clean the equipments by using water and purifier.
- **Test Tube Holder:** Used to hold a test tube while it's being heated.
- **Ring Stand**: Used in conjunction with clamps to hold equipment that cannot stand up on its own.
- **Scoopula:** Another instrument used to transfer solids from one place to another.
- **Thermometer:** Used to measure temperatures. Thermometer generally contain liquid mercury (Hg).

REACTIONS OF ACID RADICALS OR – ANIONS

We are studying some acid radicals such as:-

1- Chloride Cl^{-}

2- Iodide I^{-}

3- Phosphate PO_4^{-2}

4- Sulphate SO_4^{-2}

5- Thiocyanate SCN^{-}

6- Oxalate $C_2O_4^{-2}$

The reactions of these anions depend upon precipitation reaction, and their solubility's, or upon oxidation – reduction reaction as in a certain organic acid such as oxalate.

1-STUDY THE REACTIONS OF CHLORIDE ION Cl^{-}

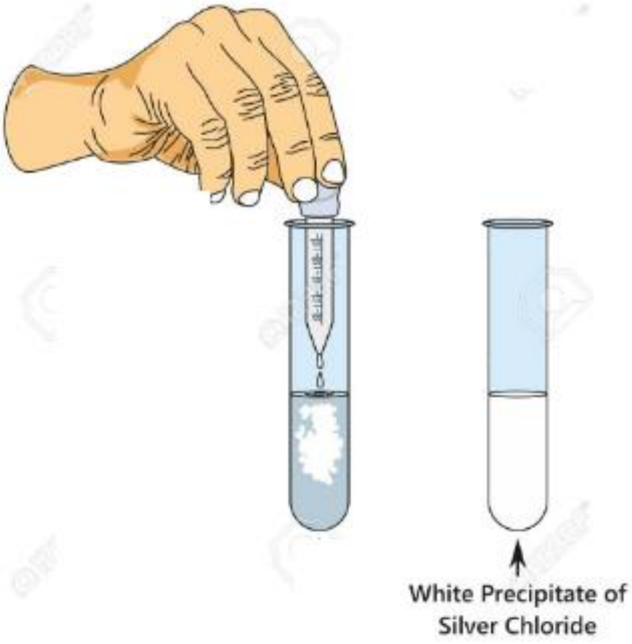
(USE NaCl SOLUTION 5%)

Chloride is constituent of hydrochloric acid (HCl) produced by parietal cells of stomach. It is more in extracellular fluid than intracellular fluid. It is required for proper erythrocyte (RBC) function by chloride shift. It is involved in maintenance of plasma volume. It is activator of some enzymes like amylase, angiotensin converting enzyme etc.

REACTION -1

-With $AgNO_3$ solution 5%

When an aqueous solution of silver nitrate ($AgNO_3$) is (10 drops) added to the aqueous solution of (10 drops) sodium chloride (NaCl), a white precipitate of silver chloride (AgCl) is formed that is indicated by the following chemical reaction.



$$NaCl + AgNO_3 \longrightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{AgCl}} + NaNO_3$$

REACTION -2

-Confirm by:

1- Try the solubility of AgCl in HNO_3 (insoluble).

2- Try the solubility of AgCl in dil NH_4OH (soluble).

In order to confirm that this precipitate is AgCl, 20 drops of NH_4OH is added to the same test tube. The white precipitate dissolves because AgCl salt forms a complex with NH_3.

$$AgCl_{(s)} + 2NH_3 \rightarrow [Ag(NH_3)_2]^+ + Cl^-$$

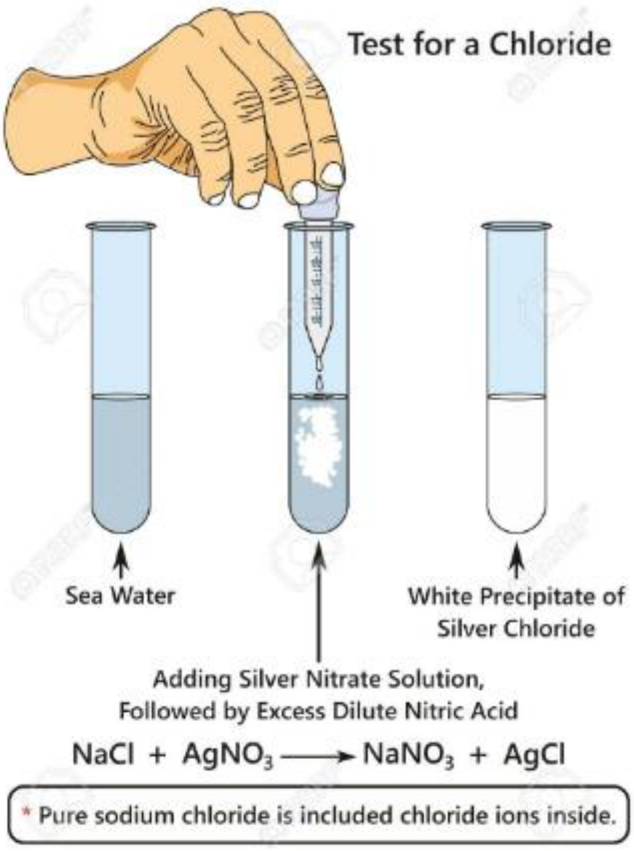
$$AgCl + 2NH_4OH \longrightarrow Ag(NH_3)_2Cl + 2H_2O$$

soluble complex

Added 15 drops of HNO_3 into the solvated complex and a white precipitate is re-constituted. Because with the addition of HNO_3, the silver diamine complex $[Ag(NH_3)_2]^+$ decomposes and re-precipitates as AgCl.

$$[Ag(NH_3)_2]^+ + Cl^- + 2HNO_3 \rightarrow 2NH_4^+ + 2NO_3^- + AgCl_{(s)}$$

Note:-Sea water and tap water Contains (Cl^-)



2-STUDY THE REACTIONS OF IODIDE ION I^-

(USE KI SOLUTION 5%)

Iodine is constituent of thyroid hormones. Thyroxine (T_4) and triiodothyronine (T_3) synthesis involves iodine utilization. Thyroid hormones are essential for mental development as well as physical development.

REACTION -1

-With $AgNO_3$ solution 5%

Add (10 drops) of $AgNO_3$ solution to (10 drops) of KI solution then you will see yellow precipitate of AgI in the tube which will be insoluble in NH_4OH and HNO_3.

$$KI + AgNO_3 \longrightarrow \underset{\downarrow}{AgI} + KNO_3$$

yell. ppt.
Insoluble in
NH_4OH and HNO_3

REACTION -2

-Confirm by starch test:

Iodides are readily oxides in acid solution to free iodine by a number of oxidizing agents, free iodine identified by the deep blue coloration produced with starch solution.

Add (5 drops) of KI solution to (5 drops) of HNO_3 solution then you will see brown solution of free iodine.

$$KI + HNO_3 \longrightarrow HI + KNO_3$$

3-STUDY THE REACTIONS OF PHOSPHATE ION PO_4^{-2}

(USE Na_2HPO_4 SOLUTION 15%)

Phosphate is present in humans to an extant of about 500-700gm. More of it is present in bone and teeth. In the body phosphate is present in two forms. An inorganic form found in bone and teeth which is complexes with calcium and magnesium. Another is organic form. It is present as organic compounds in the cells and cell membranes. In the bone and teeth it is present as hydroxyl apatite. Phosphate is component of blood buffers.

REACTION -1

-With $AgNO_3$ solution 5%

Add (10 drops) of $AgNO_3$ solution to (10 drops) of Na_2HPO_4 solution then you will see yellow precipitate of Ag_3PO_4 in the tube which will be sparingly soluble in HNO_3.

$$Na_2HPO_4 + 3AgNO_3 \longrightarrow \underset{\substack{\downarrow \\ \text{yell. ppt.}}}{Ag_3PO_4} + 2NaNO_3 + HNO_3$$

Try the solubility of the yellow ppt. Ag_3PO_4 in

1- In dilute NH_4OH (soluble).

2- In dilute HNO_3 (soluble).

REACTION -2

-Confirm by $FeCl_3$ test:

Pale buff $FePO_4$, soluble in excess of $FeCl_3$ and in HCl, but insoluble in acetic acid (CH_3COOH).

Add (2drops) of $FeCl_3$ solution to (10 drops) of Na_2HPO_4 solution then you will see pale Buff ppt. $FePO_4$ in the tube.

$$Na_2HPO_4 + FeCl_3 \longrightarrow \underset{\text{pale Buff. ppt.}}{FePO_4} + 2NaCl + HCL$$

4-STUDY THE REACTIONS OF SULPHATE ION $SO_4^{=}$

(USE Na_2SO_4 SOLUTION 10%)

Reaction -1

Add (10 drops) of $AgNO_3$ solution to (10 drops) of Na_2SO_4 solution then you will see white precipitate of Ag_2SO_4 in the tube which will be sparingly soluble in HNO_3.

$$2AgNO_3 + Na_2SO_4 \rightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{Ag_2SO_4}} + 2NaNO_3$$

Reaction -2

-Confirm by reaction with $BaCl_2$ solution

Add (10 drops) of $BaCl_2$ solution to (10 drops) of Na_2SO_4 solution then you will see white precipitate of $BaSO_4$ which be insoluble in dil HCl.

$$Na_2SO_4 + BaCl_2 \rightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{BaSO_4}} + 2NaCl$$

To 0.5 mL of sample in a test tube, add 7 drops of HNO_3 and 1 drop of $BaCl_2$, respectively. The white precipitate is $BaSO_4$ and the formation of this precipitate indicates the presence of SO_4^{2-} ion in the sample.

Sulphate ions SO_4^{2-}

Add acidified barium chloride

Barium chloride + potassium sulfate →
Barium sulfate + potassium chloride

$$BaCl_2 + K_2SO_4 \rightarrow BaSO_4 + 2KCl$$

$$Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4\,(s)}$$

White precipitate

5-STUDY THE REACTIONS OF THIOCYANATE ION SCN^- (USE KSCN SOLUTION 5%)

Reaction -1

Add (10 drops) of $AgNO_3$ solution to (10 drops) of KSCN solution then you will see white precipitate of AgSCN in the tube which will be soluble in NH_4OH but insoluble in dil HNO_3.

$$KSCN + AgNO_3 \longrightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{AgSCN}} + KNO_3$$

Reaction -2

-Confirm by $FeCl_3$ test:

Add (5 drops) of $FeCl_3$ solution to (5 drops) of KSCN solution the product solution will be red blood solution of $Fe(SCN)_3$

$$FeCl_3 + 3KSCN \rightarrow \underset{\text{red blood colour solution}}{Fe(SCN)_3} + 3KCl$$

6-STUDY THE REACTIONS OF OXALATE ION $C_2O_2^{=}$ (USE $K_2C_2O_4$ SOLUTION 5%)

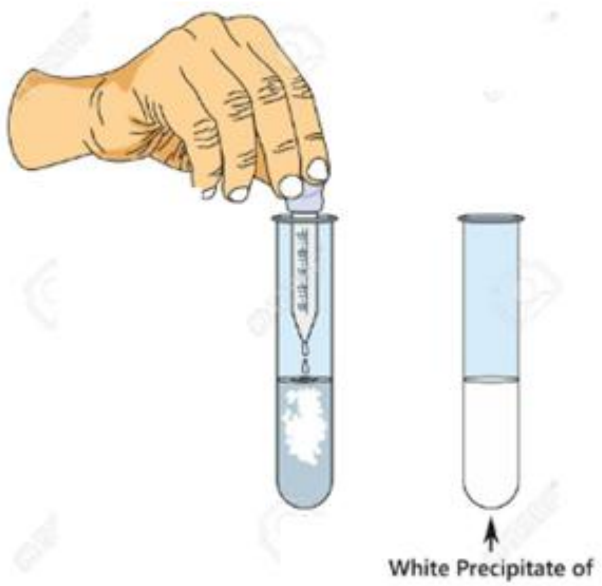
Many metal ions form insoluble precipitates with oxalate, a prominent example being calcium oxalate, the primary constituent of the most common kind of kidney stones.

Reaction -1

-With $AgNO_3$ solution (5%)

When an aqueous solution of silver nitrate (AgNO3) is (10 drops) added to the aqueous solution of (10 drops) potassium oxalate ($K_2C_2O_4$), a white precipitate of silver oxalate ($Ag_2C_2O_4$) is formed that is indicated by the following chemical reaction.

$$K_2C_2O_4 + 2AgNO_3 \rightarrow \underset{\text{Wt.PPt}}{\underset{\downarrow}{Ag_2C_2O_4}} + 2KNO_3$$



1- Try the solubility of Ag $_2C_2O_4$ in HNO_3 (soluble).
2- Try the solubility of Ag $_2C_2O_4$ in dil NH_4OH (soluble).

Reaction - 2

-Confirm by $KMnO_4$ test:

Decolourized the purpul colour of Permanganate (2 drops) when warmed in acid solution (2 drops) with oxalate (10 drops) in (60-70C)

$$K_2C_2O_4 + H_2SO_4 \rightarrow H_2C_2O_4 + K_2SO_4$$

$$2\,KMnO_4 + H_2C_2O_4 + 3\,H_2SO_4 \rightarrow 2MnSO_4 + 10CO_3 + 8H_2O + K_2SO_4$$

This reaction depended upon oxidation reduction reaction.

REACTIONS OF (METAL IONS) OR – CATIONS:

The compounds of these metals are characterized by their precipitation as chlorides by diluted hydrochloric acid (HCl) or by soluble chloride.

These metals or cations are:

1- Silver – Ag^+.
2- Lead – Pb^{+2}.
3- Mercury – Hg_2^{+2}.

1.Silver – Ag^+

Silver is a white mutable and ductile metal, it is insoluble in (H_2SO_4 and dil.HCl) but dissolves in nitric acid HNO_3 (2:1) and in boiling conc.H_2SO_4.

$$Ag^+ + 2HNO_3 \longrightarrow AgNO_3 + NO_2^- + H_2O$$

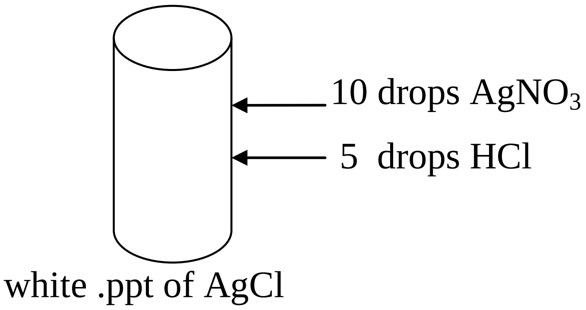
$$2Ag^+ + 2H_2SO_4 \longrightarrow Ag_2SO_4 + SO_4^{-2} + 2H_2O$$

1- REACTIONS OF SILVER ION Ag^+

(Use a solution of silver nitrate ($AgNO_3$ 5%)

REACTION -1

-With diluted hydrochloric acid (HCl)



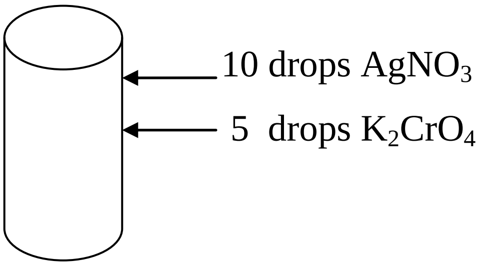
$$HCl + AgNO_3 \longrightarrow \underset{\substack{\downarrow \\ \text{wt. ppt.}}}{AgCl} + HNO_3$$

1- The white precipitate of AgCl is insoluble in water (H_2O) and acid (including nitric acid HNO_3).
2- The white precipitate of AgCl is soluble in dilute NH_4OH to formation the complex soluble salt $Ag(NH_3)_2Cl$.

$$AgCl + 2NH_4OH \longrightarrow \underset{\text{Complex soluble salt}}{Ag(NH_3)_2Cl} + 2H_2O$$

REACTION -2

-With potassium chromate K_2CrO_4



red ppt. of Ag_2CrO_4

$$2AgNO_3 + K_2CrO_4 \longrightarrow \underset{\substack{\downarrow \\ \text{red ppt.}}}{Ag_2CrO_4} + 2KNO_3$$

1- The red precipitate of silver chromate Ag_2CrO_4 is insoluble in dilute acetic acid (CH_3COOH).
2- The red precipitate of silver chromate Ag_2CrO_4 is soluble in NH_4OH and in dilute HNO_3.

2.Lead - Pb^{+2}

Lead is a bluish – grey metal, it is readily dissolved by dilute nitric acid (HNO_3).

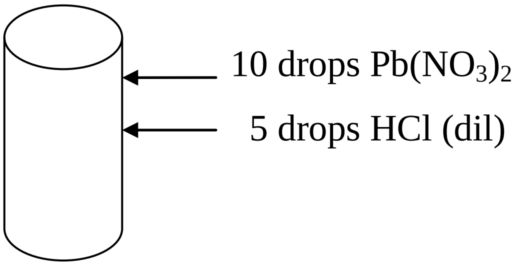
$$3Pb^{+2} + 8HNO_3 \longrightarrow 3Pb(NO_3)_2 + 2NO + 4H_2O$$

2- REACTIONS OF LEAD ION - Pb^{+2}

(Use a solution of lead nitrate $Pb(NO_3)_2$ 5%)

REACTION -1

-With diluted hydrochloric acid (HCl)



white ppt. of $PbCl_2$

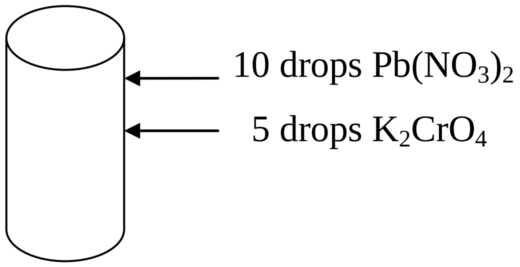
$$2HCl + Pb(NO_3)_2 \longrightarrow \underset{\substack{\downarrow \\ \text{wt. ppt.}}}{PbCl_2} + 2HNO_3$$

1- The white precipitate of $PbCl_2$ is soluble in hot water.

2- The white precipitate of $PbCl_2$ is separate in the solution cooled.

REACTION -2

-With potassium chromate K_2CrO_4



yell. ppt. of $PbCrO_4$

$$Pb(NO_3)_2 + K_2CrO_4 \longrightarrow \underset{\text{yell. ppt.}}{\underset{\downarrow}{PbCrO_4}} + 2KNO_3$$

1- The yellow precipitate of $PbCrO_4$ is insoluble in NH_4OH.
2- The yellow precipitate of $PbCrO_4$ is soluble in alkali hydroxide (NaOH).

$$PbCrO_4 + 4NaOH \longrightarrow \underset{\substack{\text{Complex soluble}\\ \text{salt}}}{Na_2(PbO_2)} + Na_2CrO_4 + 2H_2O$$

3.Mercury - Hg_2^{+2}

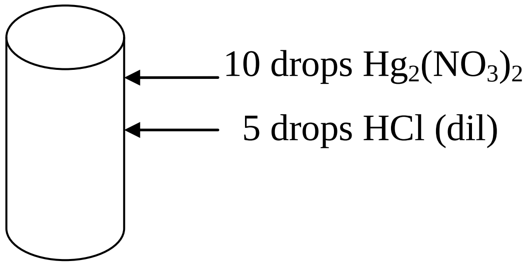
Mercury is a silvery – white, liquid metal at the ordinary temperature, it is unaffected by treatment with dilute HCL or dilute H_2SO_4 but react readily with nitric acid HNO_3. Cold dil. of nitric acid and excess of mercury yield mercurous nitrate $Hg_2(NO_3)_2$.

3- REACTIONS OF MERCUROUS ION – Hg_2^{+2}

(Use a solution of mercurous nitrate $Hg_2(NO_3)_2$ 5%)

REACTION -1

-With diluted hydrochloric acid (HCl)



white .ppt of Hg_2Cl_2

$$2HCl + Hg_2(NO_3)_2 \longrightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{Hg_2Cl_2}} + 2HNO_3$$

The white precipitate of Hg_2Cl_2 is insoluble in hot water.

REACTION -2

-With NH_3 solution:

The white precipitate of Hg_2Cl_2 react with NH_4OH to formation black precipitate $Hg(NH_2)Cl$.

$$Hg_2Cl_2 + 2NH_4OH \longrightarrow \underset{\text{black. ppt.}}{\underset{\downarrow}{Hg(NH_2)Cl}} + Hg + NH_4Cl + 2H_2O$$

Chemistry lab:4 **First stage**

Preparation and standardization of 0.1 M(HCl) hydrochloric acid solution

Hydrochloric acid is present in the digestive juices of the human stomach. Excessive secretion of the acid causes gastric ulcer, while a marked deficiency of it impairs the digestive process and is sometimes the primary cause of deficiency anemia. Concentrated hydrochloric acid causes burns and inflammation of the skin.

Theory:- Hydrochloric acid is produced in solutions up to 38% HCl (concentrated grade). Higher concentrations up to just over 40% are chemically possible, but the evaporation rate is then so high that storage and handling need extra precautions, such as pressure and low temperature. Laboratory grade hydrochloric acid is not sufficiently pure to be used as a primary standard, because it evaporates easily. In this experiment, a standard solution of sodium carbonate is used to determine the exact concentration of a hydrochloric acid solution. The neutralization reaction that occurs is as follows:

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

Methyl orange indicator solution is used. At the end-point – when neutralization just occurs – the indicator changes color from yellow to peach-pink.

Procedure:-

1. **Preparing (50 ml) 0.1 M HCl Solution**:38 % HCL shows density 1.19 g/mL and we can find M by next : -

$$M = \frac{sp.gr * \% * 1000}{M.wt}$$

Calculate the volume of HCl (conc.):- We must dilute it to preparing 0.1 M HCl in 50 ml from next: $(M* V)_{\text{conc.}} = (M * V)$ **dilute**

$$M * V_{ml} = 0.1 * 50_{ml}$$

Transfer V ml by cylinder to clean and dry beaker containing 30 ml D.W, transfer the solution to volumetric flask capacity 50 ml, and complete the volume to the mark by D.W.

2. **Preparing (50 ml) 0.1 M Na_2CO_3 Solution**:-calculate amount from sodium carbonate for prepare 0.1 M in 50 ml -

$$M = \frac{Wt.(gm)}{M.Wt.} * \frac{1000}{V(ml)}$$

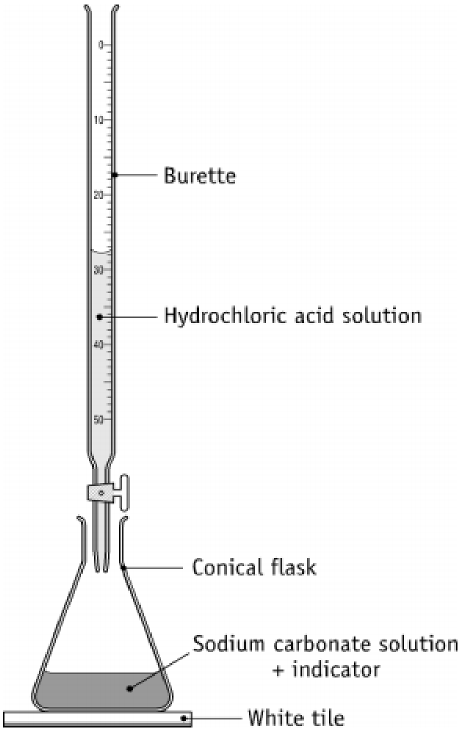
$$0.1 = \frac{Wt.(gm)}{106} * \frac{1000}{50}$$

$$Wt. = 0.53 \text{ gm}$$

Weigh 0.53 gm. from Na_2CO_3 in clean and dry beaker and dilute in 30 ml D.W, transfer solution to volumetric flask capacity 50 ml and complete the volume to the mark by D.W.

3. Transfer known volume of 5 ml the sodium carbonate solution, with a pipette, to a conical flask then add one or two drops of methyl orange to this solution.

4. Add the acid unknown solution from the burette gradually with continuous swirling of the solution in the conical flask and near the end point, the acid is added drop by drop. Continue the addition of the acid until the color of the solution passes from yellow to faint red.



5. Repeat the experiment three times and tabulate your results then take the mean of the three readings.

Calculations: Calculate the molarity of HCl:-

$$\text{m mol HCl} = \text{m mol } Na_2CO_3$$

$$(M * V)\,HCl = (M * V)\,Na_2CO_3 * \frac{1}{2}$$

$$(M * V_{\text{burette}}) = (0.1 * 5) * \frac{1}{2}$$

Discussion:-

1. What the difference between primary and secondary standard substances?

2. Calculate the volume of conc. HCl required for preparing 250 ml 0.1 M?

Chemistry lab:5 **First stage**

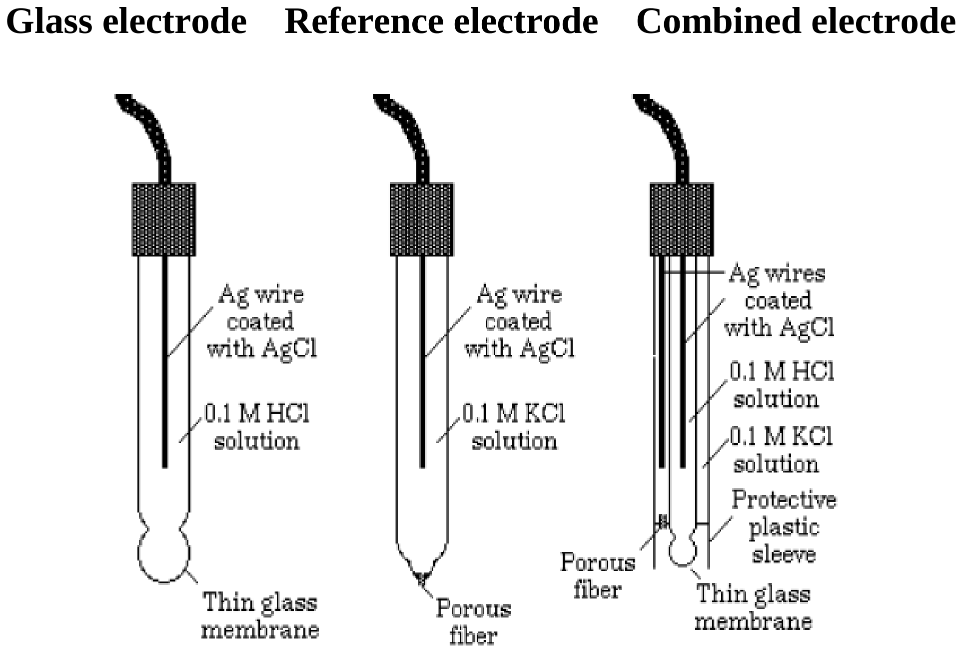
pH METER

Principles of operation of a pH meter:

A pH meter is essentially a voltmeter with high input impedance which measures the voltage of an electrode sensitive to the hydrogen ion concentration, relative to another electrode which exhibits a constant voltage. The key feature of the pH-sensitive electrode is a thin glass membrane that's outside surface contacts the solution to be tested. The inside surface of the glass membrane is exposed to a constant concentration of hydrogen ions (0.1 M HCl).

Inside the **glass electrode** assembly, a silver wire, coated with silver chloride and immersed in the HCl solution, is called an Ag/AgCl electrode. This electrode carries current through the half-cell reaction. The potential between the electrode and the solution depends on the chloride ion concentration, but, since this is constant (0.1 M), the electrode potential is also constant.

A **reference electrode** is needed to complete the electrical circuit. A common choice is to use another Ag/AgCl electrode as the reference. The Ag/AgCl electrode is immersed in a 0.1 M KCl solution which makes contact with the test solution through a porous fiber which allows a small flow of ions back and forth to conduct the current. The potential created at this junction between the KCl solution and the test solution is nearly zero and nearly unaffected by anything in the solution, including hydrogen ions.



Using the pH Meter:

Allow the meter a few minutes to stabilize after you plug it in. When you are not using the meter, keep the electrode immersed in pH 7.0 buffer to a depth of about one inch. The meter must be calibrated by using standards of known pH before an unknown is measured. Since the unknowns are acidic, the pH 4.00 and pH 7.00 standards should be used.

An accurate pH reading depends on standardization, the degree of static charge, and the temperature of the solution.

1. The pH meter should be standardized each time it is used with a buffer of known pH, preferably one closest to the desired final pH. To calibrate the pH meter, expose the hole in the electrode, rinse the electrode with deionized water, and place the electrode in a standard solution, e.g., pH 7.
2. Make sure the solution you are measuring is at room temperature since the pH can change with a change in temperature.

3. The pH-sensitive glass membrane is very thin and very easily broken. Therefore do not touch the membrane with anything harder and do that very gently; do not put the electrode down on the desk. Do not drop the electrode or bump it on the bottom of the beaker when immersing it in a solution.
4. The glass membrane must be thoroughly hydrated to work properly. Do not allow the electrode to remain out of water any longer than necessary. When the electrode is not in use, keep it immersed in the pH 7.00 buffer.

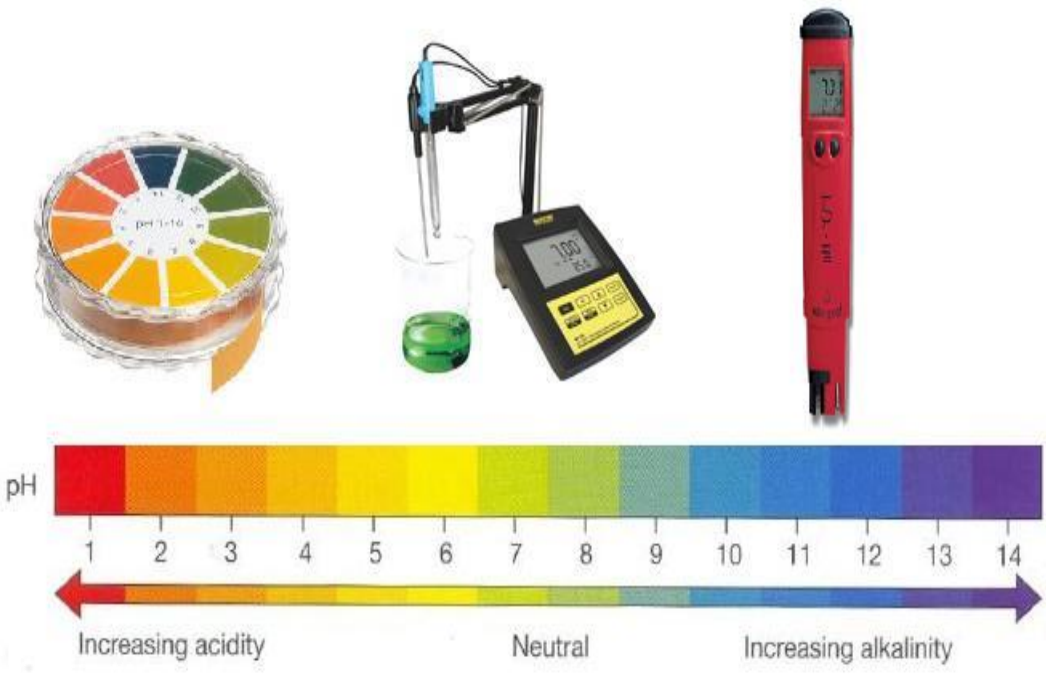
The importance of pH and pH control:

Solution pH and pH control play a major role in many facets of our lives. Consider a few examples:

- Agriculture: Crops grow best in a soil of proper pH. Proper fertilization involves the maintenance of a suitable pH.
- Physiology: If the pH of our blood were to shift by one unit, we would die. Many biochemical reactions in living organisms are extremely pH dependent.
- Industry: From manufacture of processed foods to the manufacture of automobiles, industrial processes often require rigorous pH control.
- Municipal services: Purification of drinking water and treatment of sewage must be carried out at their optimum pH.

pH meter

Litmus Paper



Chemistry lab:6 **First stage**

Action strong acid and base on Indicators

Indicators:

There are many substances that show one colour in an acidic medium and another colour in a basic medium. Such substances are called **acid-base indicators**.

An acid/base indicator is a weak organic acid or a weak organic base whose un dissociated form differs in color from its conjugate base or its conjugate acid form.

Litmus is a natural dye found in certain lichens. It was the earliest indicator to be used. It shows red colour in acidic solutions and blue colour in basic solutions. Phenolphthalein and methyl orange are some other indicators.

Procedure:

1- Take about 2 mL solution of acid or base in a test tube.

2- Add a drops of indicator to it and observe the colour. **What to observe?**

3- Tabulate your observations in Table.

Sample solution	**red litmus**	**blue litmus**	**phenolphthalein**	**methyl red**	**methyl orange**

Chemistry lab:7 First stage

"Experiment for preparing of methane gas in the laboratory"

Chemical properties of alkanes:

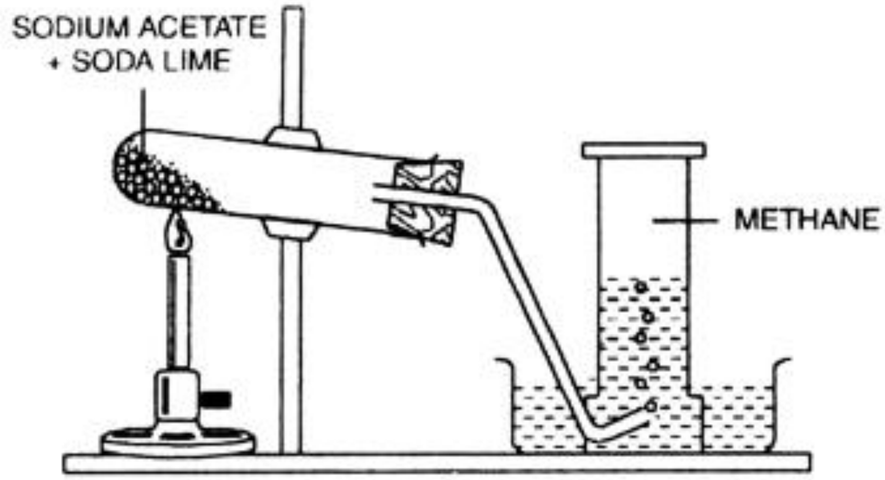
1. Alkanes have a formula C_nH_{2n+2}, n=1,2,3 …etc.
2. Alkane doesn't have functional group because it's formed from saturated hydrocarbon (C-C and C-H).
3. Alkane structure (R-R or R-H) R=alkyl group.
4. Alkane is unaffected by many common chemical reagents but reacts with chlorine or fluorine in the presence of light.
5. The alkanes can exist as gases, liquids, or solids at room temperature. The unbranched alkanes (methane CH_4, ethane C_2H_6, propane C_3H_8, and butane C_4H_{10}) are gases; while from (pentane C_5H_{12} to hexadecane $C_{17}H_{36}$) are liquids; and the larger than hexadecane are solid.

Procedure:

Methane can be prepared in the laboratory by heating a mixture of sodium acetate with sodium hydroxide according to the equation:

$$CH_3COONa + NaOH \xrightarrow{\Delta} CH_4 + Na_2CO_3$$

1- Firstly measure out 5 gm of sodium acetate or (sodium ethanoate).
2- Then add 2.5 gm of pulverized sodium hydroxide (or soda lime) and mix these two compounds well.
3- Transfer the mixture into a hard glass test tube / boiling tube and heat strongly.
4- (Sodium hydroxide NaOH and calcium oxide CaO) are called **soda lime**.
5- Methane can be collected by tube of the gas will replace the water.



After prepare methane we can study physical properties of methane or (alkanes):

1- Combustion: in combustion of methane gas is produce carbon dioxide, water vapor and heat.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + Heat\ \Delta$$

$$Alkane + O_2 \rightarrow CO_2 + H_2O + Heat\ \Delta$$

2- It is a colorless and odorless gas.

3- It is lighter than water and insoluble in water but soluble in non-polar solvents (organic solvents) like; alcohol (R-OH), carbon tetrachloride (CCl_4) … etc.

NOW:

1. Can you prepare hexane, butane and ethane?
2. Define sodium lime?
3. Write chemical equation for combustion of propane?

4- Ethane can be solved in ………. and ……… but not in …………..

Chemistry lab: 8 First stage

"Experiment for preparing of ethene(ethylene) gas in the laboratory"

Alkenes:

1-Alkenes are unsaturated hydrocarbons have (C=C) that consider the functional group.

2-Alkenes formula C_nH_{2n} where n= 2, 3, 4 ………

3-Alkene reacts with brom water (Br_2), when the red brom water is adding into alkene solution its color will disappear.

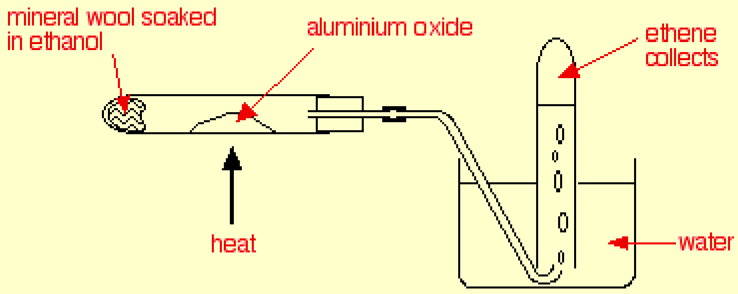
Procedure:

The dehydration of ethanol to give ethene (ethylene):

1-This is a simple way of making gaseous alkenes like ethene. If ethanol vapor is passed over heated aluminium oxide powder, the ethanol is essentially broken to give ethene and water vapor.

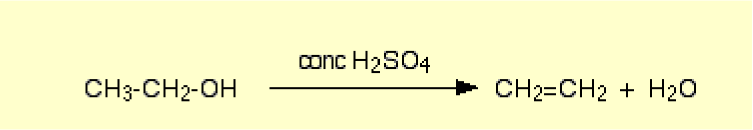
$$CH_3\text{-}CH_2\text{-}OH \xrightarrow{Al_2O_3} CH_2{=}CH_2 + H_2O$$

To make a few test tubes of ethene, you can use this apparatus:



It wouldn't be too difficult to imagine scaling this up by boiling some ethanol in a flask and passing the vapor over aluminium oxide heated in a long tube.

2-Ethanol is heated with an excess of concentrated sulphuric acid at a temperature of 170°C. The gases produced are passed through sodium hydroxide solution to remove the carbon dioxide and sulphur dioxide produced from side reactions.



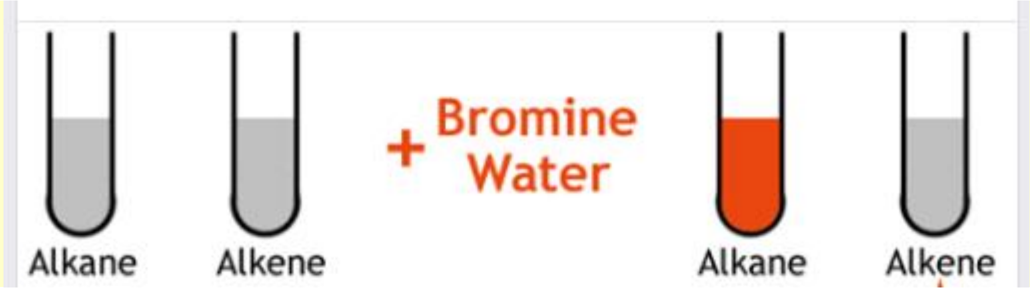
$$CH_3\text{-}CH_2\text{-}OH \xrightarrow{\text{conc } H_2SO_4} CH_2{=}CH_2 + H_2O$$

Do the following

1-Distinguish between ethane and ethene.

2-By the same way prepare propene.

3-see the diagram then explain why alkane has red color.



4- Dehydration reagents are ………. and ………. .

5- See the equation then complete

$CH_3\text{-}CH{=}CH\text{-}CH_3 + Br_2 \rightarrow$

$CH_3\text{-}CH_2\text{–}CH_2\text{-}CH_3 + Br_2 \rightarrow$

Chemistry lab:9 **First stage**

"Experiment for preparing of ethyne (acetylene) gas in the laboratory"

Chemical properties of alkynes or (acetylenes):

1. Acetylenes are unsaturated hydrocarbons that have carbon atoms in chains linked by one or more triple bond (C≡C).
2. Acetylenes have a formula C_nH_{2n-2}; R-C≡C-H and R-C≡C-R where R=alkyl group, acetylene formula C_2H_2 (H-C≡C-H).
3. Acetylenes have two functional groups (C≡C) and acidic hydrogen atom (R-C≡C-H).

Experiment:

Ethyne gas or (acetylene) is prepared by the reaction of water with calcium carbide (CaC_2). Acetylene is a colorless gas.

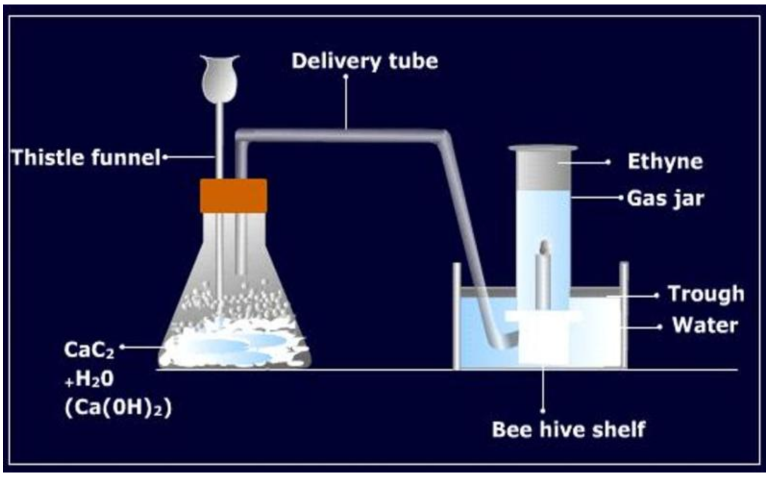
$$CaC_2 + 2H_2O \rightarrow H\text{-}C\equiv C\text{-}H + Ca(OH)_2$$

Main Chemicals required:

1. Calcium Carbide (CaC_2) [also known as calcium dicarbide] is a white/gray solid.
2. Water H_2O.

Procedure

1. Set up the apparatus as shown in the diagram.
2. Place few pieces of calcium carbide in a Büchner flash.
3. Add water from the dropping funnel, few drops at a time.
4. Collect the gas produced in the test tube.
5. Discard the first test tube as it a mixture of air and ethyne.



After preparing ethyne now we study some physical properties of ethyne

1-Combustion:

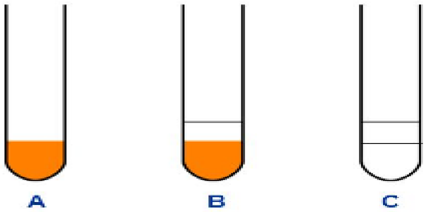
The ethyne gas burns with a **luminous smoky flame** as a result of unburned carbons.

$$2C_2H_2+5O_2 \longrightarrow 4CO_2+2H_2O$$

2-Addition of bromine test:

Addition of bromine water is presence of a triple bond C≡C (unsaturated compound), turns the color from ***yellow/red color*** to ***colorless***. The decolorisation of bromine is the standard test for **unsaturated** in a compound.

A. Place bromine water in the test tube.
B. Add the hydrocarbon to the test tube and shake
C. If the bromine color disappears then the hydrocarbon is unsaturated.



$$H\text{-}C{\equiv}C\text{-}H \quad + \quad 2Br_2 \longrightarrow CHBr_2\text{-}CHBr_2$$

Yellow –red ***Colorless***

3-Addition of acidified potassium permanganate ($KMnO_4$) test:

This is also used to test for presence of a triple bond C≡C (unsaturated compound), the color turn from purple to colorless.

A. Place the hydrocarbon to the test tube.

B. Add $KMnO_4$ solution in the test tube and shake.
C. If the $KMnO_4$ color disappears then the hydrocarbon is unsaturated.

$$R\text{-}C\equiv CH \xrightarrow[H_3O^+]{KMnO4} R\text{-}COOH + CO_2$$

$$R\text{-}C\equiv CR' \xrightarrow[H_3O^+]{KMnO4} R\text{-}COOH + R'\text{-}COOH$$

4-Addition of tollen reagent [$Ag(NH_3)_2OH$] test:

The tollen reagent $Ag(NH_3)_2OH$ turns from **colorless** to **white precipitate.**

A. Place the alkyne to the test tube.
B. Add some drops of $Ag(NH_3)_2OH$ solution to the test tube and shake .
C. If the solution colorless turns to white precipitate then the alkyne has acidic hydrogen atom (***terminal Hydrogen atom***)

$$H\text{-}C\equiv C\text{-}H + Ag(NH_3)_2OH \rightarrow H\text{-}C\equiv C\text{-}Ag + 2NH_3 + H_2$$

(Terminal Hydrogen atom)

$$R\text{-}C\equiv C\text{-}R + Ag(NH_3)_2OH \rightarrow N.R$$

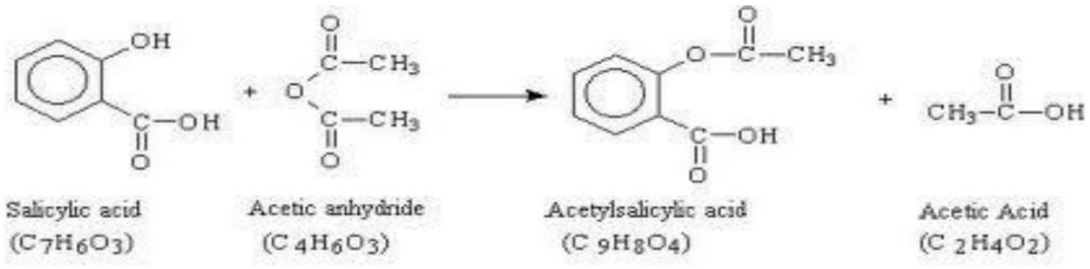
(No terminal Hydrogen atom)

Chemistry lab:10 **First stage**

Aspirin synthesis

Principle

Aspirin is formed by reaction of salicylic acid with acetic anhydride in the presence of sulfuric acid and heat according to the equation:



Acetic anhydride serves as both a reactant and a solvent.

Reagents:

Salicylic acid, acetic anhydride, sulfuric acid.

Procedures:

1-Weight 2 gm of salicylic acid (2- hydroxyl benzoic acid). Place the solid in a test tube and take it to the hood.

2-Carefully add 3 ml of acetic anhydride and then 5 drops of sulfuric acid to the solid, and shake well for 5 minutes. (sometimes product forms at this stage).

3-Place the test tube in a water bath (60 C) and heat, with intermetten shaking, for 10 minutes.

4-Place the tube in an ice water bath. Add 5 ml for distilled water and stir for 5 minutes (keep the tube in the bath while stirring)

6- Leave the crystals to dry overnight

7-Determine the weight of the crystals.

Radiation

Invisible energy waves or particles

The radioactivity is the property of some atoms to spontaneously give off energy as particles or rays. The atoms that make up the radioactive materials are the source of radiation.

Isotopes are variants of atoms of a particular chemical element, which have differing numbers of neutrons.

Radioactivity it is the emission of particles and energy by isotope.

Radioactive decay is the process by which an atomic nucleus of an unstable atom loses energy by emitting ionizing particles (ionizing radiation).

The radiation emitted can be one of three types, called alpha, beta, or gamma radiation

1. α (alpha) radiation- is a stream of particles moving at about 1/10 the speed of light that attrached toward the negative electrostatic field.The emission of an alpha particle (which contains 2 protons and 2 neutrons) from an atomic nucleus. When this occurs, the atom's atomic mass will decrease by 4 units and atomic No. will decrease by 2 particles.

Alpha Particles

- Two neutrons and two protons
- Charge of +2
- Emitted from nucleus of radioactive atoms

- Transfer energy in very short distances (10 cm in air)
- Shielded by paper or layer of skin

Alpha emitters can accumulate in tissue (bone, kidney, liver, lung, spleen) causing local damage

Alpha Radiation α has the same constitution as a helium nucleus. Alpha particles may be written as:

$$^{4}_{2}\mathrm{He} \text{ or } ^{4}_{2}\alpha$$

They have a double positive charge and a mass of 4 u

2. β (beta) radiation- Beta particles $^{0}_{-1}\beta$:- are stream of electrons that attached toward the positive electrostatic field.,which indicates that they consist of negatively charged particles. The electrons are produced within the nucleus by the transformation(transmutation)of a neutron into a proton and an electron $^{1}_{0}n \rightarrow ^{1}_{1}P + ^{0}_{-1}\beta$

After this happens, the electron is emitted from the nucleus into the electron cloud.

Beta Particles

Small electrically charged particles

- Similar to electrons
- Charge of -1
- Ejected from nuclei of radioactive atoms
- Shielded by wood, body penetration 0.2 to 1.3 cm depending on energy
- Can cause skin burns

Beta-minus Radiation Beta-minus particles are electrons

Beta-minus Radiation

β^- produced when a neutron decays Beta-minus particles may be written as

$${}^{0}_{-1}\beta^{-} \text{ or } {}^{0}_{-1}e^{-}$$

They have a negative charge and a mass of 1 /1800 u

Beta-minus Radiation

β^{--} produced when a neutron decays

3. ɣ (gama)radiation-it consists of photons with a frequency of greater than 1019 Hz. It is much like X-rays, and can pass completely through the human body. Thus gamma rays emitted from outside of the body may cause ionization, and possible health effects, in any organ in the body.

Gamma - Rays

➢ Electromagnetic photons or radiation (identical to x-rays)

➢ Emitted from nucleus of radioactive atoms – spontaneous emission

➢ Highly penetrating – extensive shielding required

➢ Serious external radiation hazard

Gamma Radiation

Gamma rays are a form of electro-magnetic radiation

ɣ release is often associated with α or β decay. Gamma rays remove energy from an unstable nucleus

Positron the same electron mass but opposite charge (+). Because positron is consider to β+ some references used negatron for β- and positron for β+ .

Beta-plus Radiation b + particles are positrons

Beta-plus Radiation

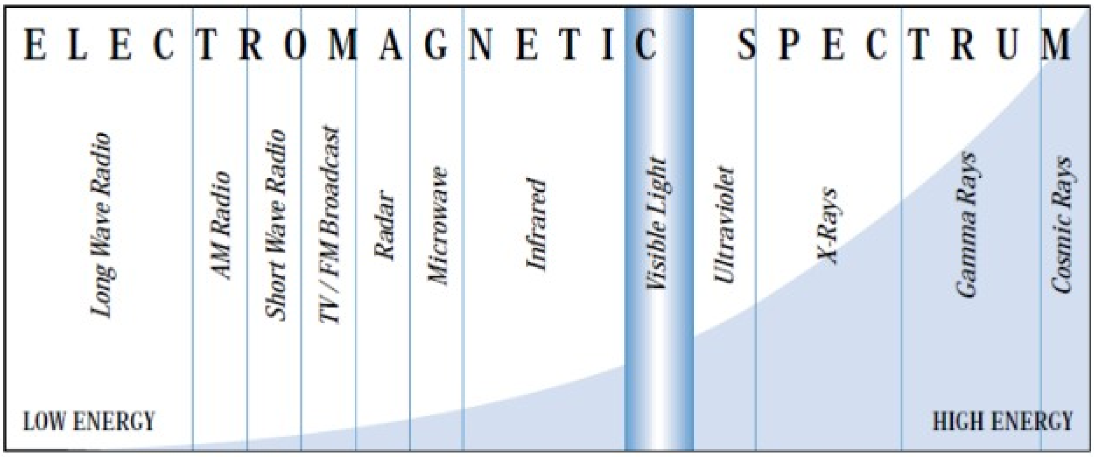
β+ is produced when a proton decays

$$ {}_{1}^{1}\mathbf{p} \rightarrow {}_{0}^{1}\mathbf{n} + {}_{+1}^{0}\beta^{+} $$

X-rays have smaller wavelengths and energy higher than ultraviolet waves. Xray light tends to act more like a particle than a wave. X-rays have a wavelength in the range of 0.01 to 10 nanometer, and energies in the range 120 eV to 120 keV.

X-ray can form ions in matter by knocking electrons off the atoms and molecules in its path creating ions, that’s why it is called ionization radiation.

X-ray cannot pass through bones and teeth as easily as through the tissues so dental X-rays show the presence of cavities and advanced bone destruction.



The Electromagnetic Spectrum

X-rays

- ➢ Overlap with gamma-rays
- ➢ Electromagnetic photons or radiation
- ➢ Produced from orbiting electrons or free electrons – usually machine produced
- ➢ Emitted with various energies & wavelengths
- ➢ Highly penetrating – extensive shielding required
- ➢ External radiation hazard

There are two less common but important types of nuclear radiation are Neutron and positron

Type of radiation	Composition	Symbol	Electric charge	Approximate penetration of skin (cm)
Alpha	Helium nucleus	α, $^{4}_{2}He$	+2	0.01
Beta	Electron	β, $^{0}_{-1}e$	-1	1
Gamma	Energy	ɤ	0	100
Neutron	Neutron	n, $^{1}_{0}n$	0	10
Positron	Positron	B^{+}, $^{0}_{1}e$	+1	1

●

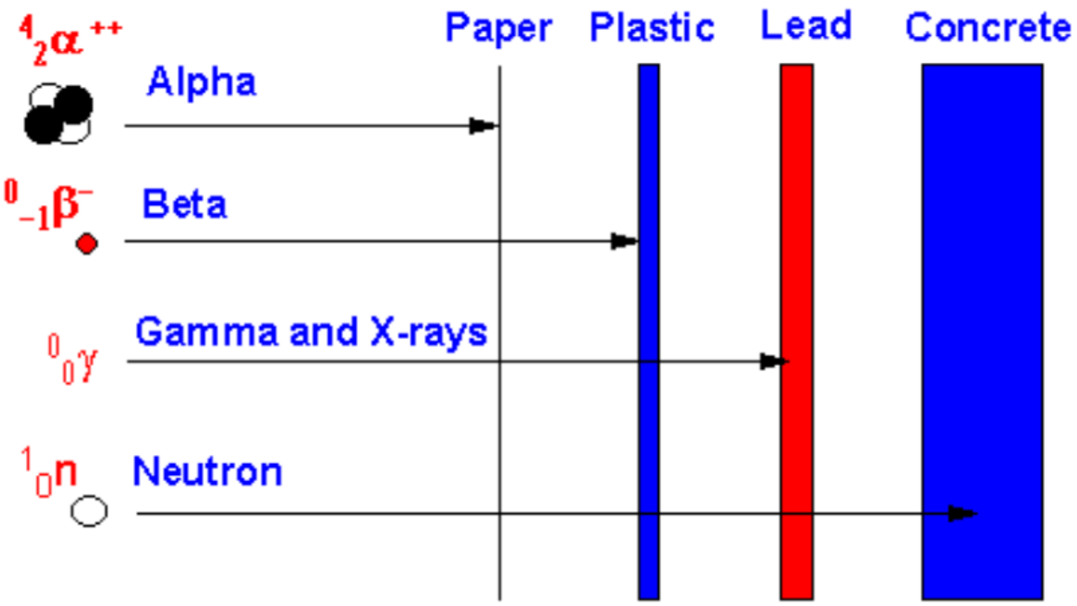
Penetrating power

α has a high mass, It is stopped by a few centimetres of air

β has a small mass, It is stopped by a few millimetres of aluminium

Ɣ has zero mass, It is stopped by thick lead or concrete

Penetrating power



Ionization

Ionizing radiation is produced by unstable atoms, because they have an excess of energy or mass or both. Unstable atoms are said to be radioactive. In order to reach stability, these atoms give off, or emit, the excess energy or mass. These emissions are called radiation.

Ionizing and Non-ionizing radiation?

Radiation carries a range of energy forming an electromagnetic spectrum.

Radiation that does not have enough energy to break chemical bonds but can vibrate atom is referred to as "Non-ionizing Radiations" e.g. radiowaves, microwaves, infrared, visible light etc.

Radiation that has enough energy to break chemical bonds is referred to as 'ionizing radiation, e.g. alpha particles, beta particles, gamma rays etc.

Detection ionizing radiation.

1- Scintillation counters are instruments that contain a surface coated with a special substance that gives off flashes of light when hit by ionizing radiation .The invisible ionizing radiation stricks the surface and some of its energy is transformed into light.Electronic devices magnify and record theses flashes.

2- Geiger-Muller counter the detecting part of the instrument is a metal tube.contains windows is made of a thin material to allow alpha and beta particles to enter .A large potential difference is maintained between tha metal walls of the tube and the center wire .when ionizing radiation enters the tube ,it forms ions that causes a pulse is counted by an electric device that produce either a meter reading or an audible click sound..

3- Film badges:X-ray technicians and other who work around radiation usually are required to wear film badge .These badges indicate the accumulated amount of radiation to which they have been exposed .They contain a piece of photographic film whose darkening is directly proportional to the amount of radiation received.

Nuclear reactions

Some radioisotopes contain nuclei which are highly unstable and decay spontaneously by splitting into 2 smaller nuclei. Such spontaneous decays are accompanied by the release of neutrons.

Nuclear reaction can be induced by bombarding atoms with neutrons. The nuclei of the atoms then split into 2 parts. Induced fission decays are also accompanied by the release of neutrons.

In nuclear reaction, two nuclei with low mass numbers combine to produce a single nucleus with a higher mass number

The followings rules must be observed when writing nuclear reactions

1-The sum of the atomic numbers on both sides of the equation must be the same.

2- The sum of the mass numbers on both sides of the equation must be the same.

$^{40}_{19}K \rightarrow {}^{40}_{20}Ca + {}^{0}_{-1}\beta$

$^{212}_{84}Po \rightarrow {}^{208}_{82}Pb + {}^{4}_{2}\alpha$

$^{197}_{79}Au + {}^{1}_{0}n \rightarrow {}^{198}_{79}Au$

$^{14}_{7}N + {}^{1}_{0}n \rightarrow {}^{14}_{6}C + {}^{1}_{1}H$ $({}^{1}_{1}H = p)$

$^{238}_{92}U \rightarrow {}^{234}_{90}{}^{Th} + {}^{4}_{2}He$ α – decay

$^{16}_{7}N \rightarrow {}^{16}_{8}O + {}^{0}_{-1}e$ β – decay

$^{99m}_{43}Tc \rightarrow {}^{99}_{43}Tc + \gamma$ metastable isotop

$^{131}_{53}I \rightarrow {}^{131}_{54}Xe + {}^{0}_{-1}\beta + {}^{0}_{0}\gamma$

$^{238}_{92}U \rightarrow {}^{0}_{0}\gamma + {}^{238}_{92}U^{*}$ (*) indicates slightly lower energy

Radioisotope (radisotope)

The isotopes produced artificially by bombardment with one of the various particles are called radisotopes.Radisotopes have the same chemical properties as non-radioactive isotopes of the same element because chemical properties are based upon electrons only. radisotopes have medical uses.

Half life

It is the time needed for one-half of the original nuclei of an isotope to decay to other substances and represent by the symbol $t_{1/2}$ which indicate half life. The importance of half life, it tell us how long a sample of isotope will exist. For medical work ,a radioisope must have a half-life such that it will remian in the body long enough to supply the radiation needed and yet will decrease in radioactivity within a reasonable period of time so that the body doesnot receive excess radiation . Radioistopes of long half-life are very dangerous to the body for example radium has a half-life of 1590years ,therefore ,if it is taken into the body it continues to give off its radiation with no change in amount during the life of that person.Some radioactive elements have a half-life measured in terms of billions of years ,whereas others are measured in fractions of a second.

ionization radiation has been known to cause:

- Skin cancer.
- Bone cancer. And damage to bone cells.
- Leukaemia.
- Other types of cancer.
- Damage to brain cells.
- Eye cells.
- Damage to other organ cells and tissues.
- Genetic risk (which is more difficult to determine because it may appear after several generations). This type of risk causes damage to genes of the cell nuclei).

X-ray is also harmful; person must take precaution to avoid exposure. The effect of being exposure to high level of radiation kill the cells (this fact is used to treat cancer), so normally cancer cells is exposed to high energy radiation (X-Ray, ɣ radiation...) to destroy these cells.

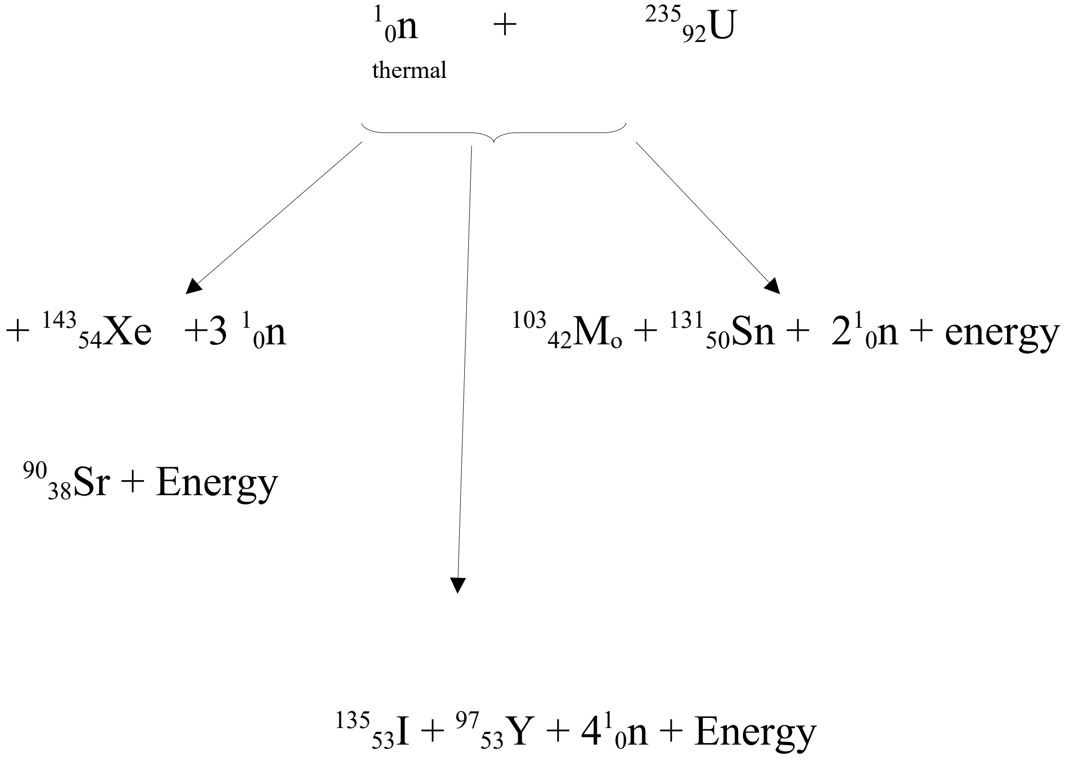
Energy and Nuclear reactions

1- Fission reaction: the nuclear reaction that causes an atom to split into several smaller parts is called fission reaction. This splitting process is accompanied by the release of large amount of energy.

A nuclear power uses a fissionable material (one that is capable of undergoing fission) such as U – 125, as fuel.

- U – 235 is the only naturally occurring isotope that undergoes this reaction.

- Two other isotopes formed by nuclear bombardment, PU – 239 and U – 233 also undergo fission reactions.



$^{1}_{0}n$ + $^{235}_{92}U$

thermal

$+ ^{143}_{54}Xe \ +3\ ^{1}_{0}n$

$^{90}_{38}Sr + Energy$

$^{103}_{42}M_o + ^{131}_{50}Sn + \ 2^{1}_{0}n + energy$

$^{135}_{53}I + ^{97}_{53}Y + 4^{1}_{0}n + Energy$

Each neutrons released react with other nuclei this is called chain reaction. The amount of material needed for a chain reaction to continue is called the "*Critical Mass*".

Nuclear fusion

Results from combination of two small nuclei to form a large nucleus with the release of large amount of energy.

$$^{2}_{1}H + ^{3}_{1}H \rightarrow ^{4}_{2}He + ^{1}_{0}n + energy$$

Such fusion reaction occur in two places in the world:-

1- In the sun.
2- In the hydrogen bomb.

The high temperature needed to carry out the fusion reaction is furnished by the explosion of an atomic bomb. The atomic bomb is set off first and provides the energy to trigger the hydrogen bomb.

The Effects of Radiation on the Cell at the Molecular Level

When radiation interacts with target atoms, resulting in ionization or excitation.

The absorption of energy produces damage to molecules by direct and indirect actions.

For direct action, damage occurs to atoms on key molecules in the biologic system for example radiation can strick molecule of water in the cells and may knock an electron from a water molecule or it is may remove a hydrogen ion from water molecule

$H_2O \rightarrow e - H_2O^+$

$H_2O \rightarrow H^+ - OH^-$

Indirect action involves the production of free radicals whose toxic damage on the key molecule results in a biologic effect.

Acids, Bases and Salts of Medical Interestes

Acids:

An acid is any coumpound that ionized in water to form hydrogen ions. Certain acids ionize completely when dissolved in water such acids are called **strong acids**such as (HCl, HBr, HI, HNO_3 and H_2SO_4 are strong acids).

Other acids ionized only slightly when dissolved in water such acids are called **weak acids**. An example is acetic acid (CH_3COOH), when dissolved in water 2% of the molecules ionize to form hydrogen and acetate ions according to the following equation:

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

(H_2CO_3, HF, and HCN are also weak acids).

Another way of classifying acids is by the number of hydrogen ions they form when dissolved in water.one H in a molecule that can form hydrogen ions is called a monoprotic acid, e.g. $H^+ Cl^-$ A diprotic acid will give two H^+ , e.g. H_2SO_4 . A triprotic acid will give three H^+ e.g acid is phosphoric acid, H_3PO_4.

$HCl \longrightarrow H^+ + Cl^-$ Monobrotic acid

$H_2SO_4 \longrightarrow 2H^+ + SO_4^{-2}$ Diprotic acid

$H_3PO_4 \longrightarrow 3H^+ + PO_4^{-3}$ Triprotic acid

The two hydrogen ions of H_2SO_4 are formed in two separate steps as shown by the following equations:

$$H_2SO_4 + H_2O \longrightarrow HSO_4^- + H_3O^+$$

Bisulfate ion

$$HSO_4^- + H_2O \longrightarrow SO_4^{-2} + H_3O^+$$

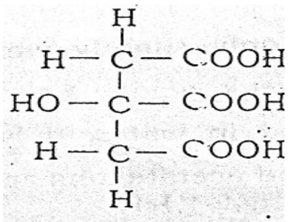
Sulfate ion

The strength of an acid depends on the concentration of H^+ ions formed in solution. Strong acids produce a high concentration of H^+ ions whereas weak acids produce a low H^+ concentration in solution.

Properties of Acids:

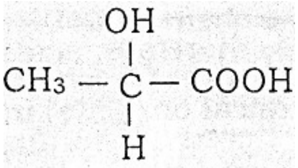
1-The hydrogen atoms of an acidic molecule that are released to form ions in water are called ***acidic hydrogens***. In acids such as HCl, H_2SO_4, and HNO_3, all the hydrogens in the compound are acidic. In acetic acid (CH_3COOH), only one of the hydrogens is acidic.

2-Not all the acids are liquids, also there are many solid acids, like boric acid (H_3BO_4) and citric acid.



Citric acid

3-The acid taste of the acids is due to the hydrogen ions produced when these acids dissolved in water. So citric acid is responsible for the acidictaste of lemon and grapfruit juices. Acitic acid is responsible for the acidity of vinegar. Lactic acid is responsible for the acidity of yogurt.



Lactic acid

4-When acids react with certain compounds, these compounds change in color. Substances that change in color in the presence of acids are called **indicators**.One of the most indicators for acids is **litmus**. Blue color of litmus paper converted to red color in the presence of an acid. Another common indicator – **phenonlphthalein** turns from red to colorless in the presence of an acid.

Chemical Reactions of Acids:

1-Acids react with metal oxides and hydroxides to form water and a salt. (in neutralization reaction solution is neither Basic)

$$2HCl + MgO \longrightarrow MgCl_2 + H_2O$$

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

2-Acid react with Na, Ca, Mg, Zn, Fe, Sn, and Pb to produce hydrogen gas and a salt.

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$$

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2 \uparrow$$

So acids cannot be stored in containers made of these metals.

Iron can react with acids, therefore, acids slould not be allowed to come in the content of sugrical and dental insutruments.

3-Acids react with carbonate and bicarbonate to produce carbon dioxide, water, and salt.

$$2HCl + CaCO_3 \longrightarrow CaCl_2 + H_2CO_3$$

$$H_2SO_4 + 2NaHCO_3 \longrightarrow Na_2SO_4 + 2H_2CO_3$$

Where,

$$H_2CO_3 \longrightarrow CO_2 + H_2O$$

Ionization of acid

The degree of ionization of any acid is given by its *ionization* constant (K_a). The equilibrium constant for the ionization of an acid in water is defined as its ionization constant.

$$C_2H_3O_2H + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

$$K = \frac{[H_3O^+][C_2H_3O_2^-]}{[H_2O][C_2H_3O_2H]} \qquad [H_3O^+] = [H^+]$$

$$K[H_2O] = K_a = \frac{[H^+][C_2H_3O_2^-]}{[C_2H_3O_2H]} \qquad \text{"ionization constant"}$$

$$H_2CO_3 + H_2O \rightleftharpoons H_3O^+ + HCO_3^- \qquad K_{a1} = \frac{[H^+][HCO_3]}{[H_2CO_3]}$$

$$HCO_3^- + H_2O \rightleftharpoons H_3O^+ + CO_3^- \qquad K_{a2} = \frac{[H^+][CO_3^-]}{[H_2CO_3]}$$

- Strong acids have large value of K_a.
- Weak acids have small value of K_a.

  $PK_a = -\log K_a$

Uses of Acids:

1- HCl: Pateint who have a lower than normal amount of hydrochloric acid in the stomach, condition called hypoacidity, are given dilute hydrochloric acid orally before meals to overcome this deficiency.

2- Acetylsalisalic acid (aspirin): Aspirin is frequently taken by people with a cold to relieve headache, muscle pain, and fever.

3- Ascorbic acid (Vitamin C): Is found in oranges and used mainly in the treatment of scurvy,absorption of iron and the maintenance of bones, and **teeth**. Vitamin C is one of many ***antioxidants*** that can protect against damage caused by *free radicals*, as well as toxic chemicals and pollutants like cigarette smoke. Deficiency of vitamin C causes loose **teeth**

Acids Effect on Teeth

1-Dental erosion is the loss of the surface of the teeth due to acidic food and acidic drink or acids coming up from the stomach. These acids can dissolve the crystals that make up the teeth, leading to tooth surface loss. These acids can also soften the tooth surface

2-After consuming something sugary or starchy, oral bacteria eats whatever food remains in the mouth after meals and produces acid as a waste product.

3- people who suffer from strong acid reflux or who vomit regularly also expose their teeth to their own strong stomach acid, which can increase tooth decay.

4- Enamel and hard tooth structures demineralize at pH of 5 or 5.5. During the demineralization process, calcium ions from our teeth leave enamel and make them weak or porous.

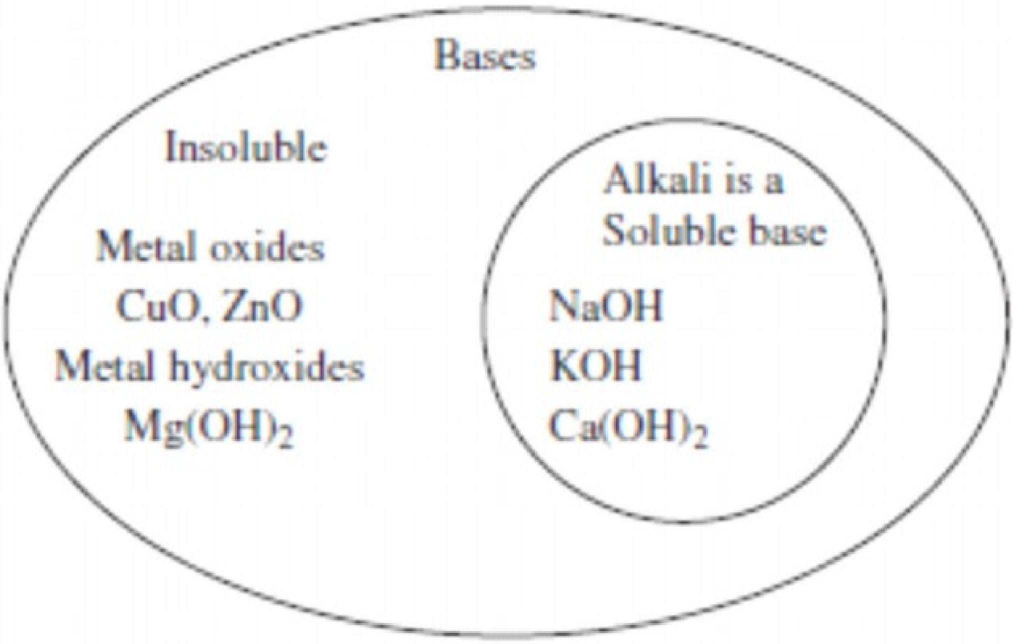
Arrhenius's base**:**

A base is any compund that increase the hydroxide ion concentration in water. Bases are also called alkaline substances(An alkali is a solution of a base in water). Sodium hydroxide (NaOH) is a common example of a base. When dissolved in water, hydrated sodium ions and hydroxide ions are formed.

$$NaOH \longrightarrow Na^+ + OH^-$$

Other common bases are potassium hydroxide (KOH), Calcium hydroxide ($Ca(OH)_2$) and magnesium hydroxide ($Mg(OH)_2$). These bases all exist as ions in aqueous solution and are all called strong bases.

Calcium hydroxide and magnesium hydroxide differ from Sodium hydroxide and potassium hydroxide, they are both only slightly soluble in water.



Weak bases such as

$$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$C_5H_5N + H_2O \rightleftharpoons C_5H_5NH^+ + OH^-$$

Properties and Reactions of Bases:

1-Soluble bases are called alkalis, e.g. sodium hydroxide, NaOH , potassium hydroxide, KOH, and ammonium hydroxide, NH4OH.

2-. Bases are oxides or hydroxides of metallic elements.

3-Bases and alkalis will react with acids to neutralize them, forming salts plus water:

Acid + Base → Salt + water
Acid + alkali → Salt + water

4-All alkalis contain a hydroxide ion, OH^- , that will react with and ‘pick up’ a H^+ ion to form a water molecule.

5-Bases are produced when metalic oxides are dissolved in water.

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Calicium oxide   Calcium hydroxide

6-Bases react with indicators. They turn litmus color from red to blue, turn methyl orange from red to yallow, and turn phenolphthalein from colorless to red.

7-Bases neutralize acids to form water and salt.

$$Ca(OH)_2 + H_2SO_4 \longrightarrow CaSO_4 + H_2O$$

8- Strong bases react with certain metals to produce hydrogen gas.

$$2Al^{+3} + 6NaOH + 6H_2O \longrightarrow 3H_2 + 2Na_3Al(OH)_6$$

Sodium aluminate

Thus a strong base such as (NaOH) can not stored in an aluminum container because it will react with it and dissolve the container.

9-Strong bases can dissolve lipids and proteins.

Ionization of Base

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$K = \frac{[NH_4^+][OH^-]}{[H_2O][NH_3]}$$

$$K[H_2O] = K_b = \frac{[NH_3][OH^-]}{[NH_3]}$$

$$PK_b = -\log K_b$$

- Weak bases have small K_b.
- Strong bases have large K_b.

Uses of Bases:

1-Calcium hydroxide solution $Ca(OH)_2$: It is used to overcome excess acidity in the stomach.

2-Magnesium hydroxide $Mg(OH)_2$: It is used as a laxative.

3-Bases referred to in medicine and biology usually contain nitrogen atoms (nitrogenous bases', which is a term often used in relation to amines, amino acids and proteins)that have the ability to pick up a proton and become a positive ion:

$$NH_3 + H^+ \rightarrow NH_4^+$$

Ammonia

$$HOOCCH_2NH_2 + H^+ \rightarrow HOOCCH_2NH_3^+$$

Amphoteric substances: *are substances which can react as both an acid .and a base like water (H_2O)*

$$HNO_2 + H_2O \longrightarrow NO^-_2 + H_3O^+.$$

Acid1 base2 base1 acid2

$$NH_3 + H_2O \longrightarrow NH^-_4 + OH^-$$

Base1 acid2 acid1 base2

Bronsted Acids and Bases

- Acid any compound or ion that donates a proton.
- Base any compound or ion that accepts a proton.

Hydrogen chloride HCl/ conjugate acid- base pair

$$HCl + H_2O \rightleftharpoons H_3O + Cl^-$$

Conjugate acid-base pair

Conjugate acid-base pair

$$NH_3 + H_3O \rightleftharpoons H_2O + NH_4^+$$

Conjugate acid-base pair

$$H_2O + H_2O \leftrightarrows H_3O^+ + OH^-$$

Acid1 base2 acid2 base1

$$HCO_3^- + H_3O^+ \rightleftharpoons H_2O + H_2CO_3$$ "Amphoteric behavior "

Base1 acid2 base2 acid1

$$HCO_3^- + OH^- \rightleftharpoons H_2O + CO_3^{-2}$$

Acid1 base2 acid2 base1

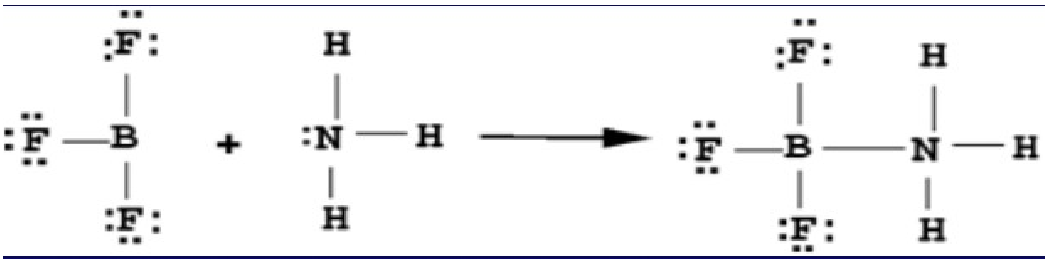
$$H_2O + HCl \rightleftharpoons H_3O^+ + Cl^-$$

The stronger the acid, the weaker is its conjugate base. The weaker an acid, the stronger is its conjugate base.

Lewis acids and bases

A Lewis acid is defined to be any species that accepts one pair electrons.

- A Lewis base is any species that donates one pair electrons.



salts

A salt is any ionic compound composed of positively charged cations and negatively charged anions, so that the product is neutral and without a net charge. These ions can be inorganic (Cl^-) as well as organic ($CH3COO^-$) and monoatomic (F^-) as well as polyatomic ions ($SO4\ 2^-$). Salts are formed (as well as water) when acids and bases react.

Classification of salts

1- Neutral salts:When strong acid reacts with strong base the products will be neutral salt and water for that the solution pH will 7 such as NaCl and Na_2SO_4
$HCl + NaOH \rightarrow NaCl$ (neutral)

2- Basic salts: When weak acid reacts with strong base the products will be basic salt and water for that the solution pH will more than 7 such as $C_2H_3O_2Na$ (weak acid and strong base)

3- Acidic salts: When strong acid reacts with weak base the products will be acidic salt and water for that the solution pH will less than 7 such as NH_4Cl

Ionization of water

$2H_2O \leftrightharpoons H_3O^+ + OH^-$

Ionization

$[H_3O^+] = [H^+] = 1 \times 10^{-7}$ M and $[OH^-] = 1 \times 10^{-7}$ M

K= ¿¿

$K[H_2O]^2 = [H_3O^+][OH^-] = K`$

$H_2O \leftrightharpoons H^+ + OH^-$ $[H_3O^+] = [H^+]$

$K = K_w = [H^+][OH^-]$

$= [1X10^{-7}][1X10^{-7}] = 1X10^{-14}$

An aqueous solution in which $[H^+]$ is greater than $[OH^-]$ is "acidic solution". In a basic solution $[OH^-]$ is greater than $[H^+]$.

$pH=-\log[H^+]$

$pOH=-\log[OH^-]$

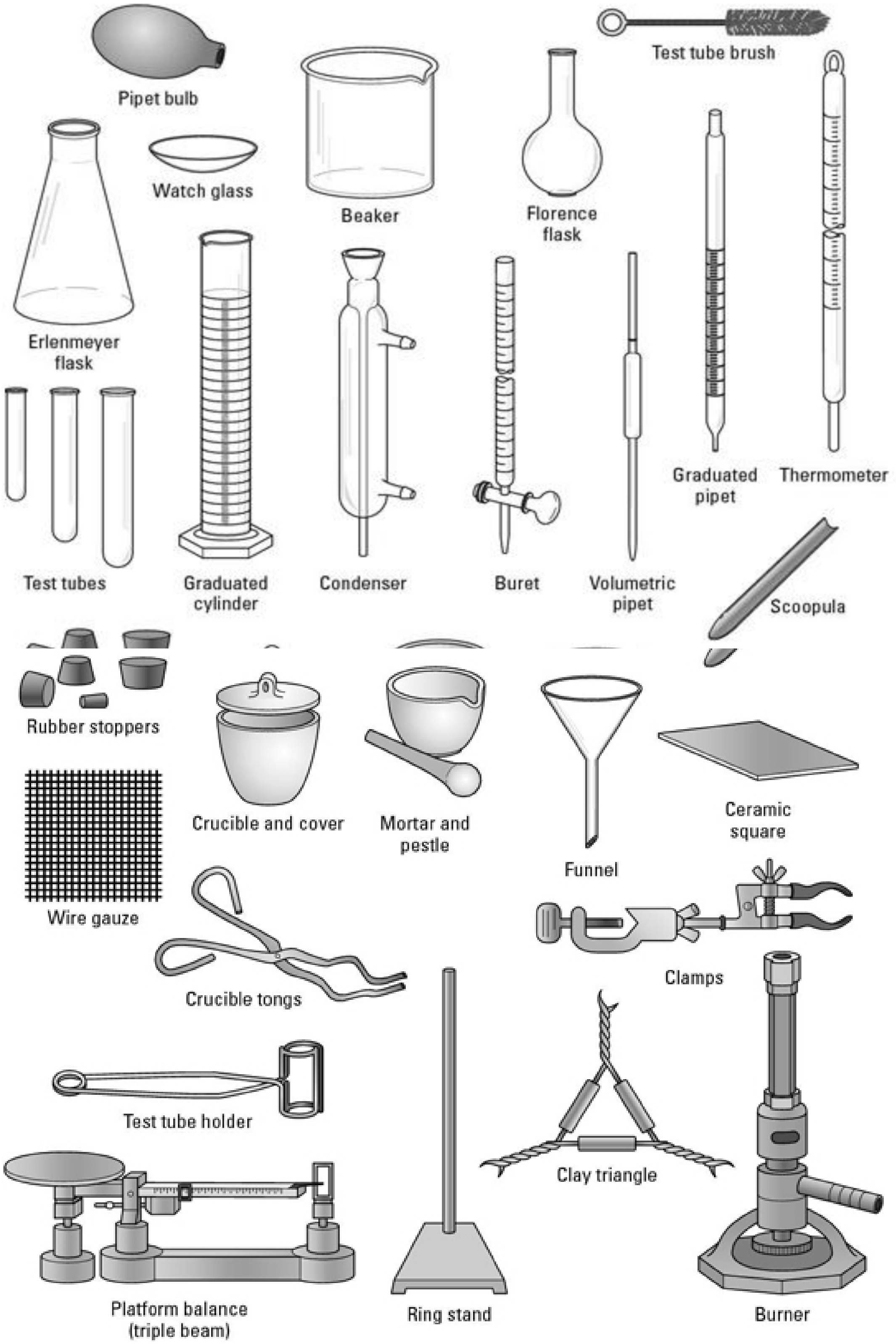
$pH+pOH=14$ (neutral solution)

An aqueous solution in which pH is less than 7 is "acidic solution". In a basic solution pH is greater than 7.

Tikrit University **College of Dentistry**

Chemistry lab:1 **First stage**

Lab Equipments



Lab Equipments

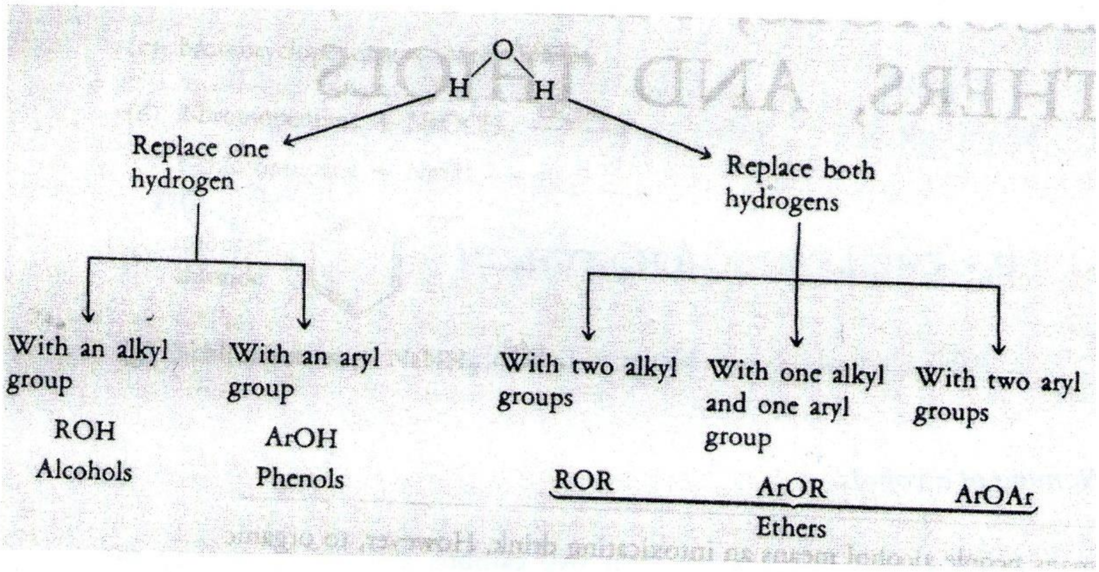
- **Balance:** Used for obtaining the masses of solid and liquid substances.
- **Beaker:** A flat-bottomed, cylindrical piece of glassware used for mixing and heating compounds.
- **Bunsen burner:** Attached to a gas line and lit to provide heat for your experiments.
- **Burette:** An extremely accurate device with a stopcock at the bottom used to measure volumes of reagents.
- **Ceramic Square:** Used to avoid burning the surface of your lab bench.
- **Clamps:** Used to hold a variety of things in place, especially test tubes.
- **Test tube:** Cylindrical open-topped piece of glassware that comes in varying sizes. Used to hold and mix small samples. Stirred by tapping the bottom with two fingers. Fit into both a centrifuge and test tube rack. **Graduated cylinder:** Used to precisely measure volumes.
- **Graduated cylinder:** Used to precisely measure volumes. They are available in various sizes.
- **Metal spatula:** Used to measure out solid substances.
- **Mortar and pestle:** Used to grind solid chemical compounds for chemistry experiments.

- **Pipette bulb:** Used to transfer accurately measured amounts of liquid from one container to another.
- **Rubber stoppers:** Used to close flasks or test tubes to prevent evaporation of liquids or escape of gases.
- **Crucible:** A cup-shaped container capable of sustaining high temperatures. It is used to heat chemical compounds.
- **Crucible tongs:** Used to handle the hot crucible.
- **Clay triangle:** Used to hold a crucible while it's being heated.
- **Condenser:** Used to collect vapors by condensing them into liquids as they contact the liquid-cooled inner surface of the condenser.
- **Erlenmeyer flask (Conical Flask):** Used to hold liquids. The small upper opening slows evaporation, so for some volatile liquids, a flask is a better choice than a beaker. The shape also makes it suitable for mixing and swirling liquids during a titration.
- **Florence flask (Volumetric Flask):** A type of flask, generally round-bottomed, usually suspended and heated from below. Its shape makes it easy to swirl and mix liquids inside of it.
- **Funnel:** Used together with filter paper to filter precipitates out of solutions.
- **Watch glass:** A piece of glassware in the shape of a large contact lens used for evaporating liquids.
- **Wire gauze:** Generally used as a surface for a beaker or flask to rest when being heated by a Bunsen burner.

- **Graduated Pipet**: Can be used to any of the given markings along its side.
- **Volumetric Pipet**: Extremely accurate, but only used for one volume.
- **Test Tube Brush:** An instrument used to clean the equipments by using water and purifier.
- **Test Tube Holder:** Used to hold a test tube while it's being heated.
- **Ring Stand**: Used in conjunction with clamps to hold equipment that cannot stand up on its own.
- **Scoopula:** Another instrument used to transfer solids from one place to another.
- **Thermometer:** Used to measure temperatures. Thermometer generally contain liquid mercury (Hg).

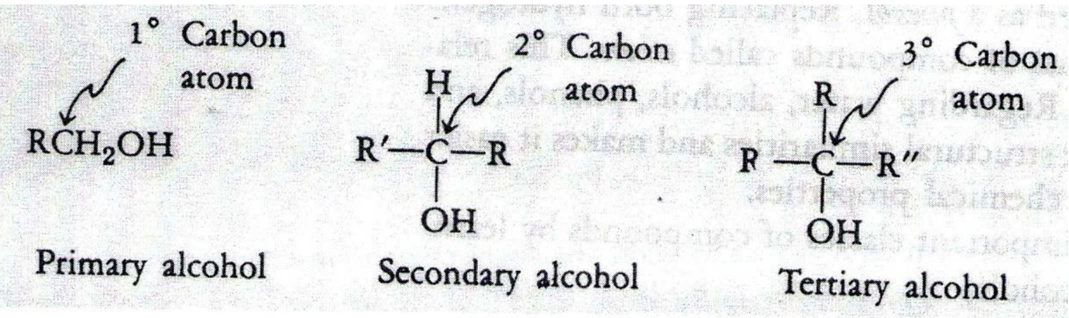
Alchols

Alcohols can be viewed as organic analogues of water in which one hydrogen is replaced by an alkyl group(R-OH)



Classification of alcohols

Primary alcohol is a compound in which the hydroxyl group is bonded to a primary carbon.In a secondary alcohol the hydroxyl group is bonded to secondary carbon .In the tertiary alcohol the hydroxyl group is bonded to tertiary carbon.



Primary alcohol · Secondary alcohol · Tertiary alcohol

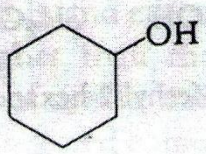
CH_3OH
Methyl alcohol

CH_3CH_2OH
Ethyl alcohol

$CH_3CH_2CH_2CH_2OH$
Butyl alcohol

$CH_2{=}CHCH_2OH$
Allyl alcohol

$(CH_3)_3CCH_2OH$
Neopentyl alcohol



Cyclohexyl alcohol

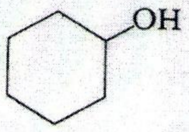
Naming of alcohols

Alcohols containing more than five carbons are best named by the IUPAC rules. The following additions to the rules set out in Chapter 11 are needed to deal with alcohols:

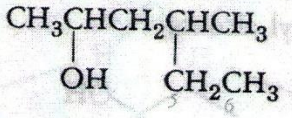
1. The longest continuous chain of carbon atoms that contains the hydroxy group is taken as the parent chain.
2. The parent name is obtained by substituting the ending *-ol* for the ending *-e* of the corresponding alkane.
3. The chain is numbered to give the lowest number to the carbon bonded to the hydroxyl group.
4. The suffix *-ol* is used for one hydroxy group; *-diol,* for two; *-triol,* for three, and so forth.

CH_3CH_2OH

Ethanol



Cyclohexanol



$$\underset{OH}{CH_3\underset{|}{C}HCH_2\underset{CH_2CH_3}{\underset{|}{C}}HCH_3}$$

4-Methyl-2-hexanol

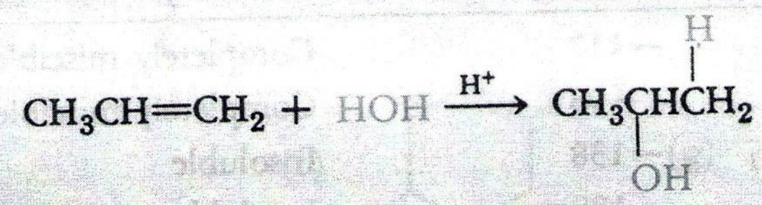
$CH_2{=}CHCH_2OH$

2-Propen-1-ol

Notice that when a double bond (or a triple bond) is part of the parent chain, the hydroxy group is given the lowest number.

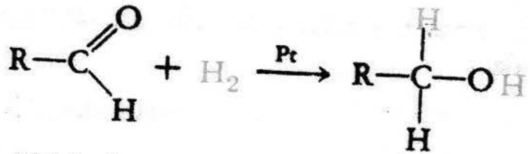
Preparing Alcohols

1-hydration of alkene

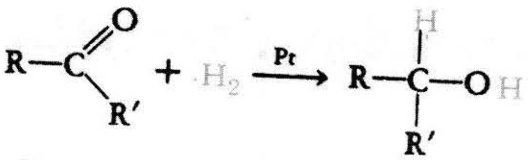


$$CH_3CH{=}CH_2 + HOH \xrightarrow{H^+} CH_3\underset{OH}{C}H\overset{H}{C}H_2$$

2-reduction of carboxyl compounds



Aldehyde 1° Alcohol

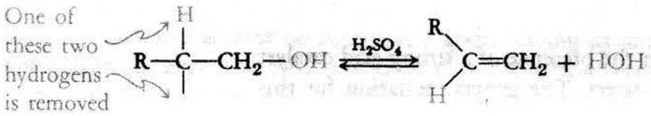


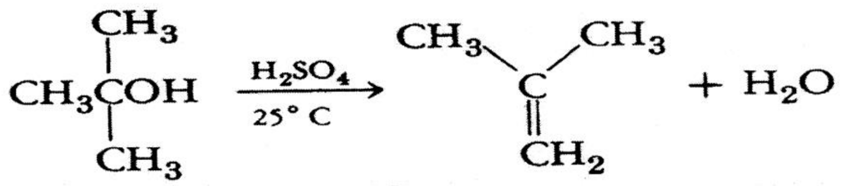
Ketone 2° Alcohol

Reactions of Alcohols

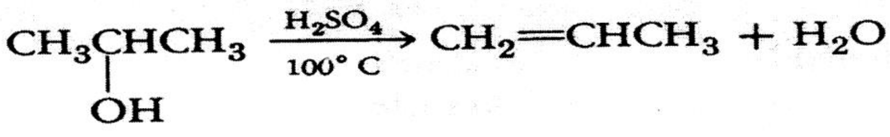
Dehydration

Alcohols react with concentrated sulfuric acid (H_2SO_4) to eliminate a molecule of water to form an alkene, as shown by the following equation:





3° Alcohol



$$CH_3CH(OH)CH_3 \xrightarrow[100°\,C]{H_2SO_4} CH_2{=}CHCH_3 + H_2O$$

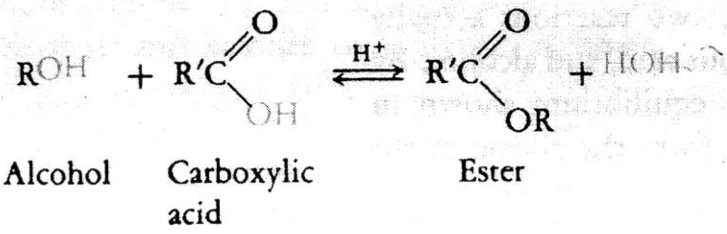
2° Alcohol

$$CH_3CH_2OH \xrightarrow[150°\,C]{H_2SO_4} CH_2{=}CH_2 + H_2O$$

1° Alcohol

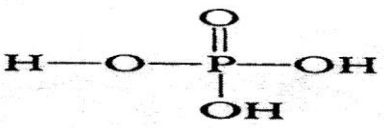
Ester Formation

Alcohols react with carboxylic acids in the presence of a strong acid catalyst to form a class of compounds called esters. The general equation for this reaction is the following:



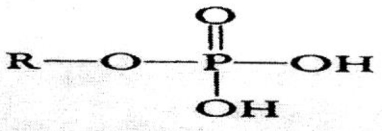
Alcohols also react with mineral acids to form compounds that are called esters of inorganic acids. In these compounds the acidic hydrogen of the inorganic acid is replaced by an alkyl group. Alkyl phosphates are an important class of such compounds in living systems. Their relation to phosphoric acid is clear from the following example:

Phosphoric acid



$\mathrm{H{-}O{-}P({=}O)(OH){-}OH}$

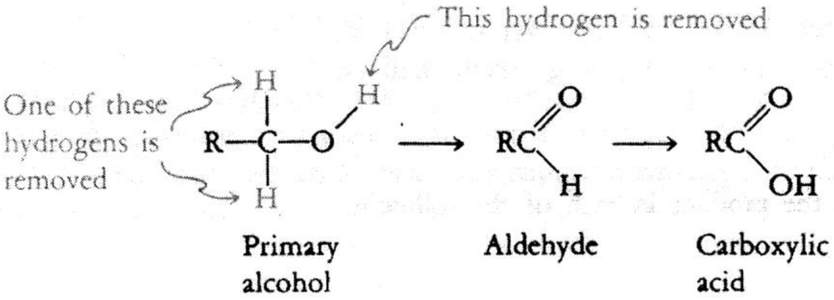
Alkyl phosphate



$\mathrm{R{-}O{-}P({=}O)(OH){-}OH}$

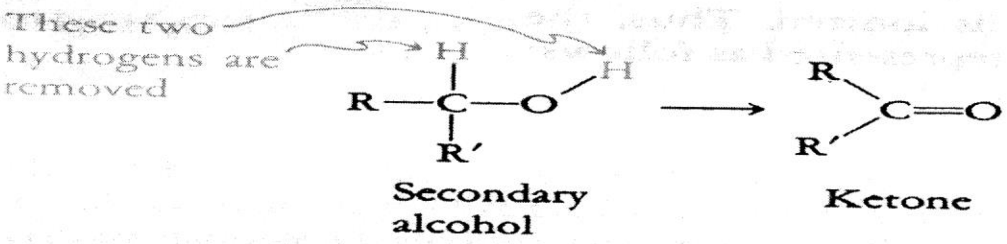
Oxidation

The products of oxidation of alcohols depend on the structure of the alcohol. Primary alcohols are first oxidized to aldehydes, which are further oxidized to carboxylic acids. Secondary alcohols form ketones on oxidation, and tertiary alcohols are not easily oxidized.

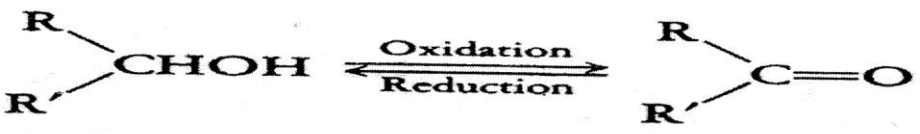


This hydrogen is removed

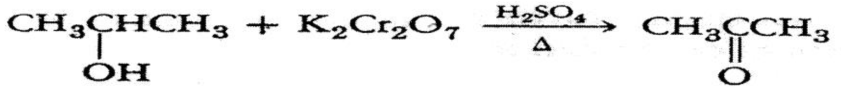
One of these hydrogens is removed



These two hydrogens are removed



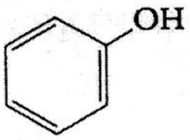
$$CH_3CH_2CH_2OH + KMnO_4 \xrightarrow[\Delta]{KOH} CH_3CH_2CO_2H$$



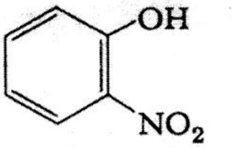
$$CH_3CH(OH)CH_3 + K_2Cr_2O_7 \xrightarrow[\Delta]{H_2SO_4} CH_3C(=O)CH_3$$

Phenols

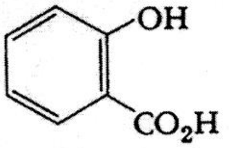
Phenols are compounds that contain a hydroxy group bonded to a benzene ring. The following are examples of phenols:



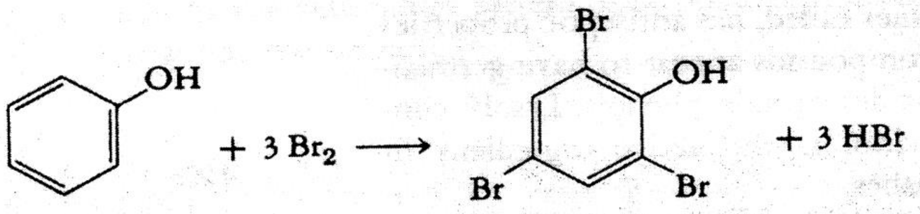
Phenol



2-Nitrophenol



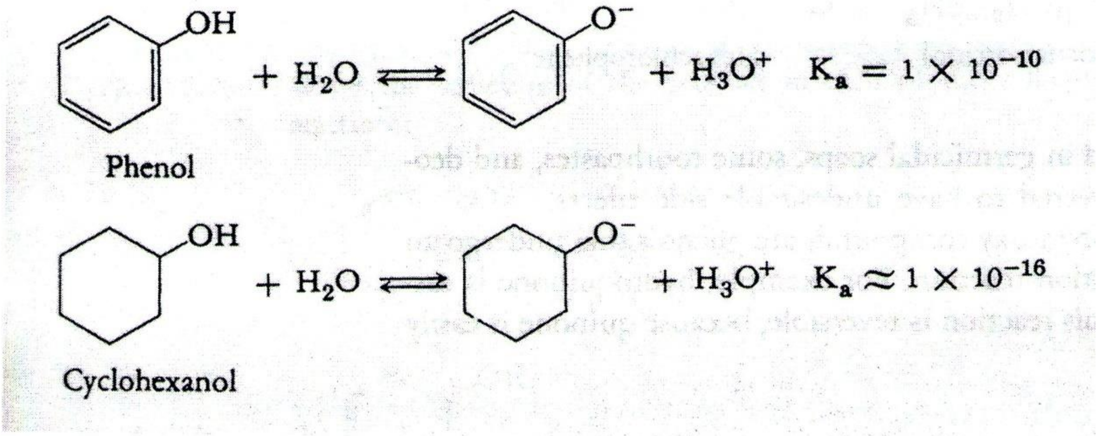
Salicylic acid



Phenol

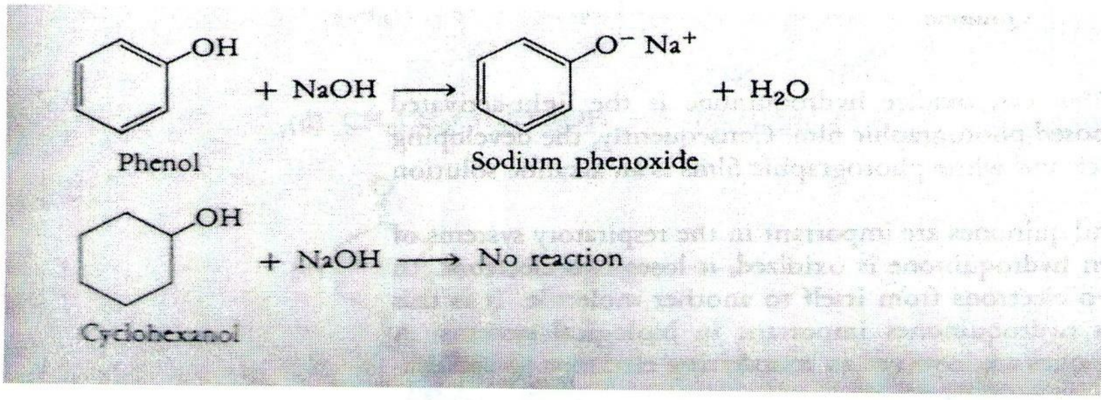
2,4,6-Tribromophenol

Phenols differ from alcohols in one important way.Phenols are much stronger acids than are alcohols.

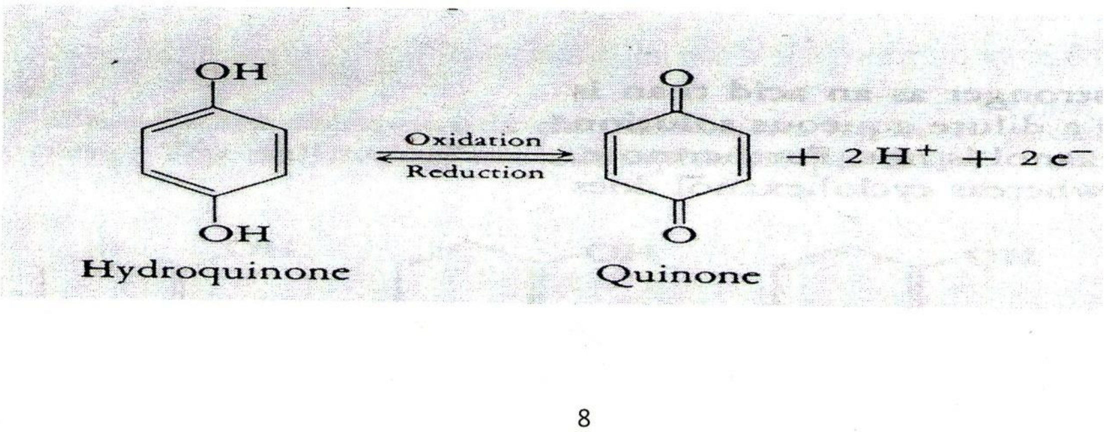


Dilute aqueous solution of phenol is acidic, whereas cyclohexanol not.

Phenol reacts with bases such as hydroxide ion , whereas cyclohexanol not.



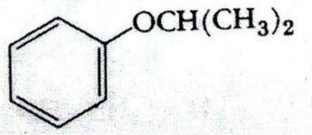
Reduction of quinone



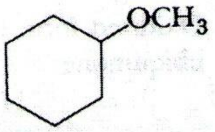
Ethers are compounds that contain an oxygen atom bonded to two alkyl groups, two aryl groups, or one aryl and one alkyl group. Ethers, like most other classes of organic compounds, can be named either by a common method or by IUPAC rules. The simpler ethers are usually known by their common names. By this method, the word ether is used as a root, and the names of the two groups bonded to the oxygen are prefixed alphabetically. The following examples illustrate this method:

$CH_3CH_3OCH_2CH_3$

Diethyl ether



Isopropyl phenyl ether

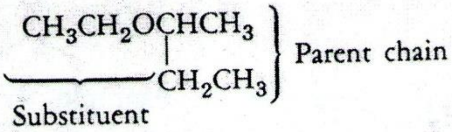


Cyclohexyl methyl ether

Notice that the individual parts of the name are separated by a space.

To name ethers by the IUPAC rules, the more complicated group attached to the oxygen is chosen as the parent. By complicated, we mean the group that is the most branched or the most substituted or has the longest chain. The other group and the oxygen are considered as substituents on this chain. For example:

$$\underbrace{\mathrm{CH_3O}}_{\text{Substituent}}\underbrace{\mathrm{CH_2CH{=}CH_2}}_{\text{Parent chain}}$$

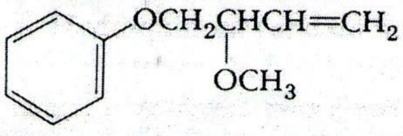


$$\mathrm{CH_3CH_2O\underset{\displaystyle CH_2CH_3}{\underset{|}{C}}HCH_3}$$

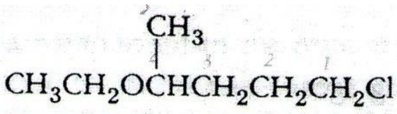
Parent chain

Substituent

The substituent is named as a prefix, and its name is constructed by replacing the *-yl* of the alkyl name or phenyl by *-oxy*. In this way, CH_3O- is called methoxy; C_6H_5O-, phenoxy; CH_3CH_2O-, ethoxy, and so forth. The following examples illustrate the application of these rules:



3-Methoxy-4-phenoxy-1-butene

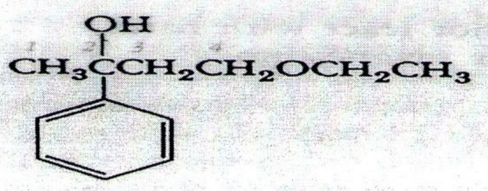


$CH_3CH_2O\overset{CH_3}{\overset{|}{C}}HCH_2CH_2CH_2Cl$

1-Chloro-4-ethoxypentane

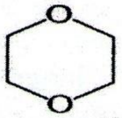
$CH_3OCH_2CH_3$

Methoxyethane

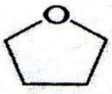


4-Ethoxy-2-phenyl-2-butanol

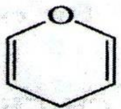
The following cyclic ethers have widely used common names:



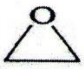
Dioxane



Tetrahydrofuran

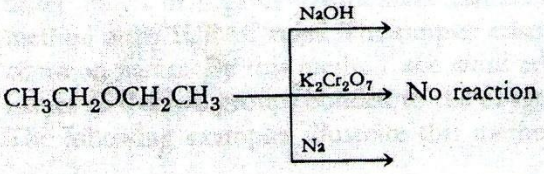


Pyran



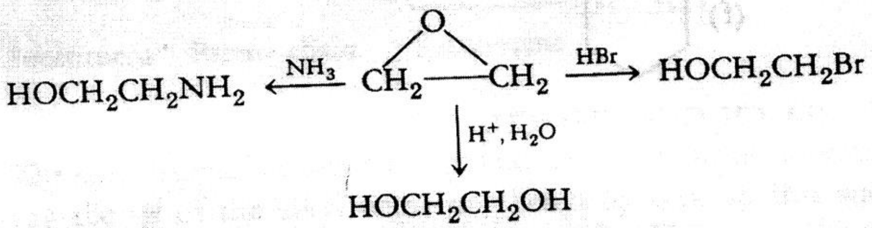
Ethylene oxide

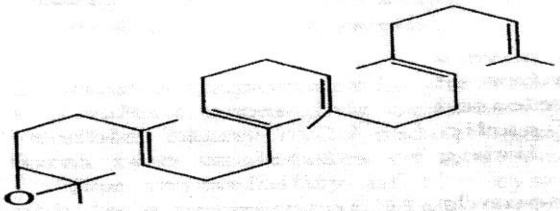
Ethers are quite unreactive. For example, they do not react with bases, most oxidizing reagents, or metals such as sodium or magnesium.



However, ethers do react with acids. The products of the reaction with hydroiodic acid (HI) are an alkyl iodide and an alcohol. For example:

$$\underset{\text{Diethyl ether}}{CH_3CH_2OCH_2CH_3} + \underset{\text{Hydroiodic acid}}{HI} \xrightarrow{\Delta} \underset{\text{Ethyl iodide}}{CH_3CH_2I} + \underset{\text{Ethyl alcohol}}{CH_3CH_2OH}$$





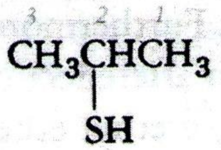
Squalene 2,3-epoxide

Thiols

Thiols are the sulfur analogues of alcohols and contain the $-\underset{|}{\text{C}}-\text{SH}$ functional group. The IUPAC names of thiols are formed by adding the suffix *-thiol, -dithiol,* and so forth to the name of the parent hydrocarbon. Common names are obtained by first naming the alkyl group followed by the word *mercaptan.* It is sometimes necessary to name the functional group by the prefix *mercapto-*. The following examples illustrate these rules:

CH_3CH_2SH

Ethanethiol
(ethyl mercaptan)

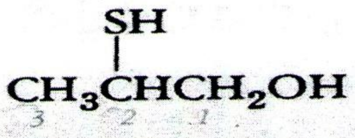


$CH_3CH(SH)CH_3$

2-Propanethiol
(isopropyl mercaptan)

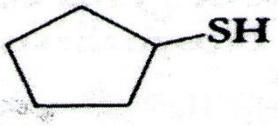
$HSCH_2CH_2SH$

Ethanedithiol



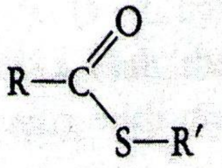
$CH_3CH(SH)CH_2OH$

2-Mercapto-1-propanol



Cyclopentanethiol

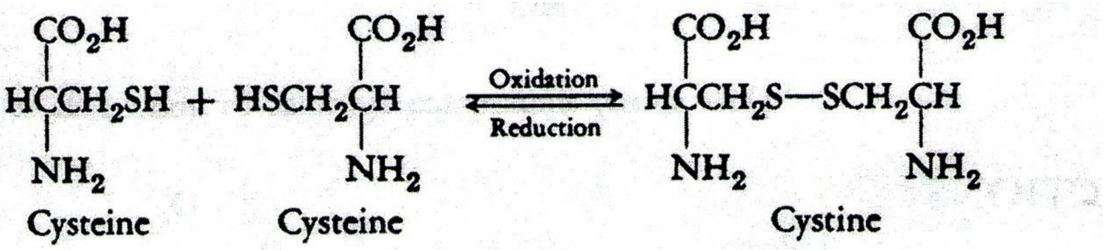
Thiols are more volatile than are the corresponding alcohols and have a very disagreeable odor. Various thiols are found in nature. A number of thiols are responsible for the odor of a skunk; 1-propanethiol is released when an onion is peeled, and thiols are responsible for the odor of garlic. Thiols are highly reactive and are present in small quantities in living systems, where they play key roles.



$$4\,RCH_2SH + O_2 \xrightarrow{Fe} 2\,RCH_2SSCH_2R + 2\,H_2O$$

Thiol Disulfide

Notice that the oxidation of a thiol occurs at the sulfur atom, whereas oxidation of an alcohol occurs at the carbon atom.



$$\underset{NH_2}{\overset{CO_2H}{HC}}CH_2SH + HSCH_2\underset{NH_2}{\overset{CO_2H}{CH}} \xrightleftharpoons[\text{Reduction}]{\text{Oxidation}} \underset{NH_2}{\overset{CO_2H}{HC}}CH_2S{-}SCH_2\underset{NH_2}{\overset{CO_2H}{CH}}$$

Cystine is important in protein structure because its disulfide bond serves as a covalent cross-link between two polypeptide chains. This reaction is reversible. Many mild reducing agents break the disulfide bond of cystine to form two molecules of cysteine.

Chemistry lab:5 **First stage**

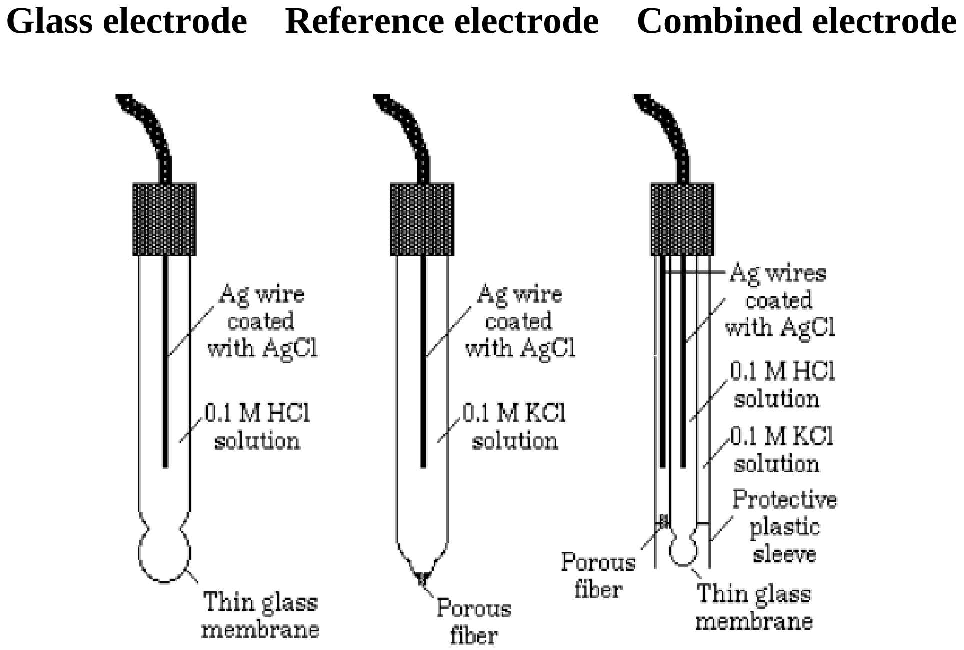
pH METER

Principles of operation of a pH meter:

A pH meter is essentially a voltmeter with high input impedance which measures the voltage of an electrode sensitive to the hydrogen ion concentration, relative to another electrode which exhibits a constant voltage. The key feature of the pH-sensitive electrode is a thin glass membrane that's outside surface contacts the solution to be tested. The inside surface of the glass membrane is exposed to a constant concentration of hydrogen ions (0.1 M HCl).

Inside the **glass electrode** assembly, a silver wire, coated with silver chloride and immersed in the HCl solution, is called an Ag/AgCl electrode. This electrode carries current through the half-cell reaction. The potential between the electrode and the solution depends on the chloride ion concentration, but, since this is constant (0.1 M), the electrode potential is also constant.

A **reference electrode** is needed to complete the electrical circuit. A common choice is to use another Ag/AgCl electrode as the reference. The Ag/AgCl electrode is immersed in a 0.1 M KCl solution which makes contact with the test solution through a porous fiber which allows a small flow of ions back and forth to conduct the current. The potential created at this junction between the KCl solution and the test solution is nearly zero and nearly unaffected by anything in the solution, including hydrogen ions.



Glass electrode Reference electrode Combined electrode

Using the pH Meter:

Allow the meter a few minutes to stabilize after you plug it in. When you are not using the meter, keep the electrode immersed in pH 7.0 buffer to a depth of about one inch. The meter must be calibrated by using standards of known pH before an unknown is measured. Since the unknowns are acidic, the pH 4.00 and pH 7.00 standards should be used.

An accurate pH reading depends on standardization, the degree of static charge, and the temperature of the solution.

1. The pH meter should be standardized each time it is used with a buffer of known pH, preferably one closest to the desired final pH. To calibrate the pH meter, expose the hole in the electrode, rinse the electrode with deionized water, and place the electrode in a standard solution, e.g., pH 7.
2. Make sure the solution you are measuring is at room temperature since the pH can change with a change in temperature.

3. The pH-sensitive glass membrane is very thin and very easily broken. Therefore do not touch the membrane with anything harder and do that very gently; do not put the electrode down on the desk. Do not drop the electrode or bump it on the bottom of the beaker when immersing it in a solution.
4. The glass membrane must be thoroughly hydrated to work properly. Do not allow the electrode to remain out of water any longer than necessary. When the electrode is not in use, keep it immersed in the pH 7.00 buffer.

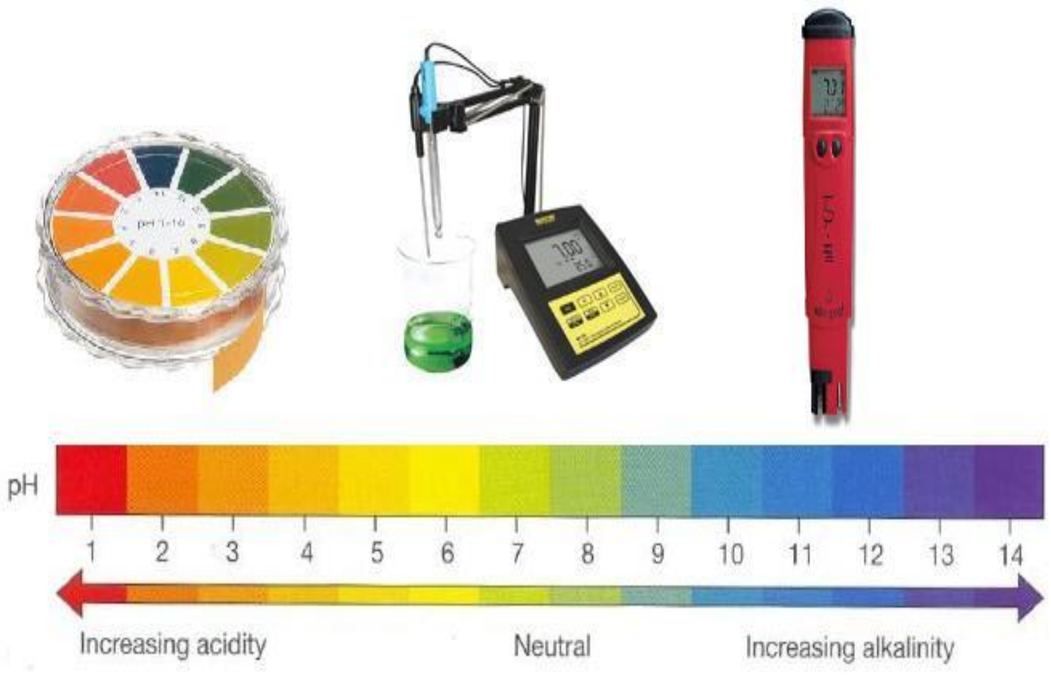
The importance of pH and pH control:

Solution pH and pH control play a major role in many facets of our lives. Consider a few examples:

- Agriculture: Crops grow best in a soil of proper pH. Proper fertilization involves the maintenance of a suitable pH.
- Physiology: If the pH of our blood were to shift by one unit, we would die. Many biochemical reactions in living organisms are extremely pH dependent.
- Industry: From manufacture of processed foods to the manufacture of automobiles, industrial processes often require rigorous pH control.
- Municipal services: Purification of drinking water and treatment of sewage must be carried out at their optimum pH.

pH meter

Litmus Paper



Chemistry lab:3 **First stage**

REACTIONS OF (METAL IONS) OR – CATIONS:

The compounds of these metals are characterized by their precipitation as chlorides by diluted hydrochloric acid (HCl) or by soluble chloride.

These metals or cations are:

1- Silver – Ag^{+}.
2- Lead – Pb^{+2}.
3- Mercury – Hg_2^{+2}.

1.Silver – Ag^{+}

Silver is a white mutable and ductile metal, it is insoluble in (H_2SO_4 and dil.HCl) but dissolves in nitric acid HNO_3 (2:1) and in boiling conc.H_2SO_4.

$$Ag^{+} + 2HNO_3 \longrightarrow AgNO_3 + NO_2^{-} + H_2O$$

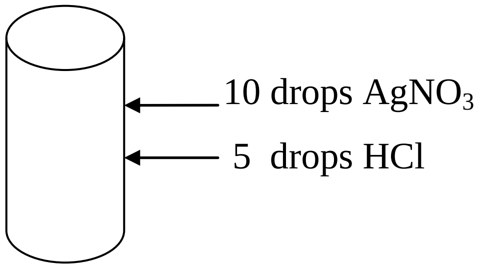
$$2Ag^{+} + 2H_2SO_4 \longrightarrow Ag_2SO_4 + SO_4^{-2} + 2H_2O$$

1- REACTIONS OF SILVER ION Ag^{+}

(Use a solution of silver nitrate ($AgNO_3$ 5%)

REACTION -1

-With diluted hydrochloric acid (HCl)



white .ppt of AgCl

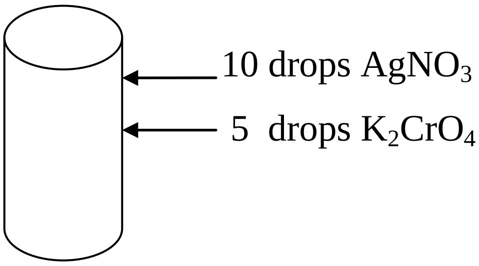
$$HCl + AgNO_3 \longrightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{AgCl}} + HNO_3$$

1- The white precipitate of AgCl is insoluble in water (H_2O) and acid (including nitric acid HNO_3).
2- The white precipitate of AgCl is soluble in dilute NH_4OH to formation the complex soluble salt $Ag(NH_3)_2Cl$.

$$AgCl + 2NH_4OH \longrightarrow \underset{\text{Complex soluble salt}}{Ag(NH_3)_2Cl} + 2H_2O$$

REACTION -2

-With potassium chromate K_2CrO_4



red ppt. of Ag_2CrO_4

$$2AgNO_3 + K_2CrO_4 \longrightarrow \underset{\substack{\downarrow \\ \text{red ppt.}}}{Ag_2CrO_4} + 2KNO_3$$

1- The red precipitate of silver chromate Ag_2CrO_4 is insoluble in dilute acetic acid (CH_3COOH).
2- The red precipitate of silver chromate Ag_2CrO_4 is soluble in NH_4OH and in dilute HNO_3.

2.Lead - Pb^{+2}

Lead is a bluish – grey metal, it is readily dissolved by dilute nitric acid (HNO_3).

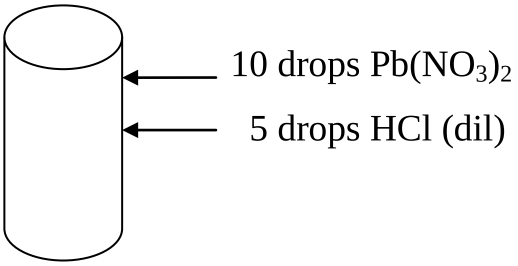
$$3Pb^{+2} + 8HNO_3 \longrightarrow 3Pb(NO_3)_2 + 2NO + 4H_2O$$

2- REACTIONS OF LEAD ION - Pb^{+2}

(Use a solution of lead nitrate $Pb(NO_3)_2$ 5%)

REACTION -1

-With diluted hydrochloric acid (HCl)



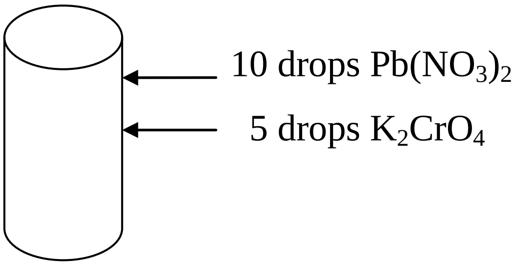
white ppt. of $PbCl_2$

$$2HCl + Pb(NO_3)_2 \longrightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{PbCl_2}} + 2HNO_3$$

1- The white precipitate of $PbCl_2$ is soluble in hot water.
2- The white precipitate of $PbCl_2$ is separate in the solution cooled.

REACTION -2

-With potassium chromate K_2CrO_4



yell. ppt. of $PbCrO_4$

$$Pb(NO_3)_2 + K_2CrO_4 \longrightarrow \underset{\text{yell. ppt.}}{\underset{\downarrow}{PbCrO_4}} + 2KNO_3$$

1- The yellow precipitate of $PbCrO_4$ is insoluble in NH_4OH.
2- The yellow precipitate of $PbCrO_4$ is soluble in alkali hydroxide (NaOH).

$$PbCrO_4 + 4NaOH \longrightarrow \underset{\substack{\text{Complex soluble} \\ \text{salt}}}{Na_2(PbO_2)} + Na_2CrO_4 + 2H_2O$$

3.Mercury - Hg_2^{+2}

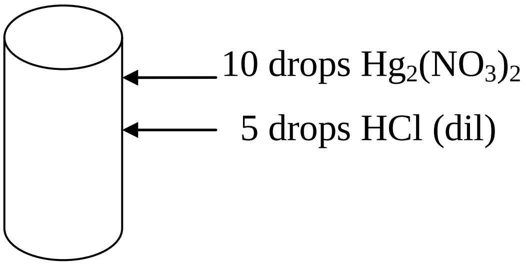
Mercury is a silvery – white, liquid metal at the ordinary temperature, it is unaffected by treatment with dilute HCL or dilute H_2SO_4 but react readily with nitric acid HNO_3. Cold dil. of nitric acid and excess of mercury yield mercurous nitrate $Hg_2(NO_3)_2$.

3- REACTIONS OF MERCUROUS ION – Hg_2^{+2}

(Use a solution of mercurous nitrate $Hg_2(NO_3)_2$ 5%)

REACTION -1

-With diluted hydrochloric acid (HCl)



white .ppt of Hg_2Cl_2

$$2HCl + Hg_2(NO_3)_2 \longrightarrow \underset{\underset{\text{wt. ppt.}}{\downarrow}}{Hg_2Cl_2} + 2HNO_3$$

The white precipitate of Hg_2Cl_2 is insoluble in hot water.

REACTION -2

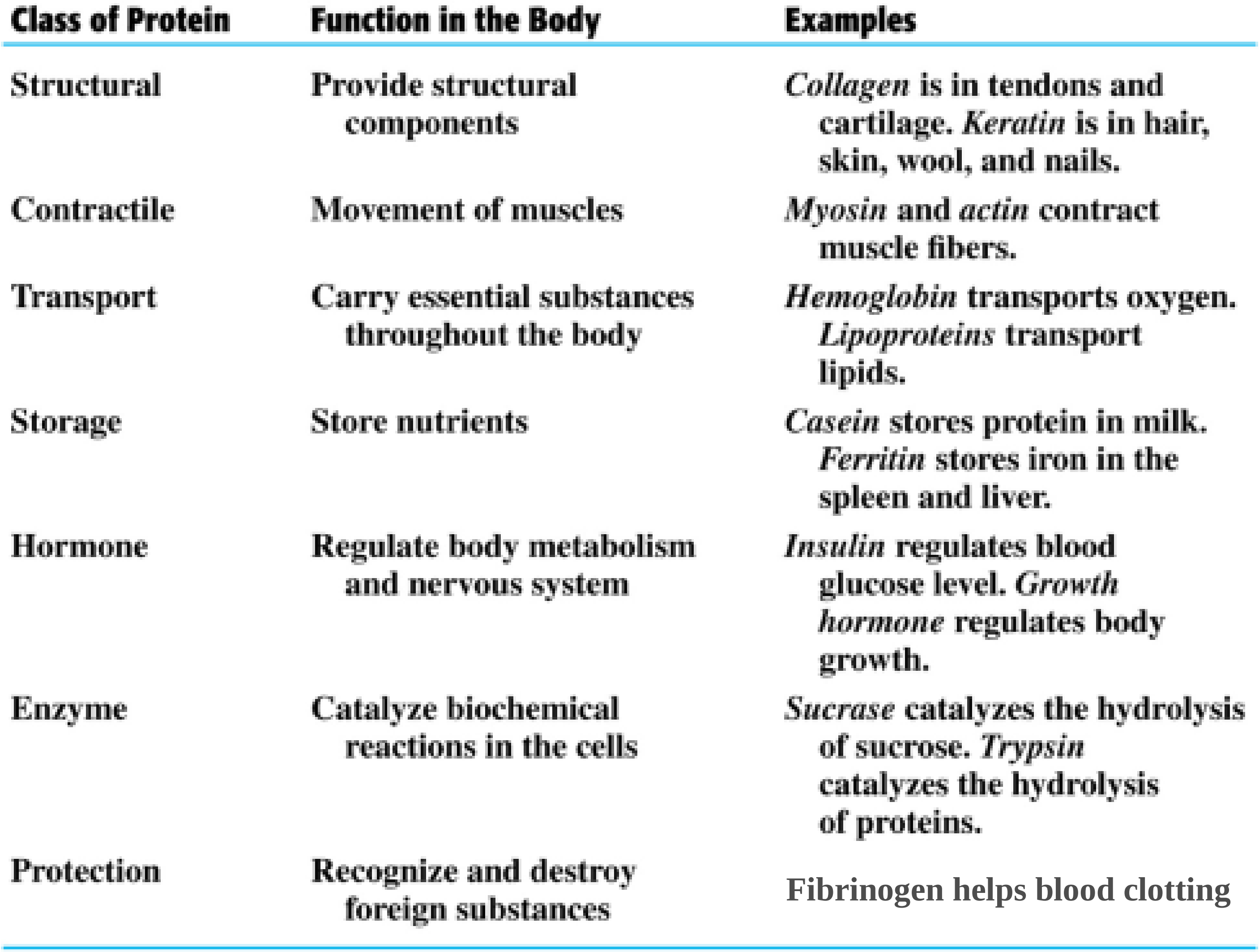
-With NH_3 solution:

The white precipitate of Hg_2Cl_2 react with NH_4OH to formation black precipitate $Hg(NH_2)Cl$.

$$Hg_2Cl_2 + 2NH_4OH \longrightarrow \underset{\underset{\text{black. ppt.}}{\downarrow}}{Hg(NH_2)Cl} + Hg + NH_4Cl + 2H_2O$$

Proteins

Function of proteins



Class of Protein	Function in the Body	Examples
Structural	Provide structural components	*Collagen* is in tendons and cartilage. *Keratin* is in hair, skin, wool, and nails.
Contractile	Movement of muscles	*Myosin* and *actin* contract muscle fibers.
Transport	Carry essential substances throughout the body	*Hemoglobin* transports oxygen. *Lipoproteins* transport lipids.
Storage	Store nutrients	*Casein* stores protein in milk. *Ferritin* stores iron in the spleen and liver.
Hormone	Regulate body metabolism and nervous system	*Insulin* regulates blood glucose level. *Growth hormone* regulates body growth.
Enzyme	Catalyze biochemical reactions in the cells	*Sucrase* catalyzes the hydrolysis of sucrose. *Trypsin* catalyzes the hydrolysis of proteins.
Protection	Recognize and destroy foreign substances	Fibrinogen helps blood clotting

Proteins

- Proteins account for 50% of the dry weight of the human body.

- Unlike lipids and carbohydrates, proteins are not stored, so they must be consumed daily.

- Current recommended daily intake for adults is 0.8 grams of protein per kg of body weight (more is needed for children).

- Dietary protein comes from eating meat and milk.

Proteins

100,000 different proteins in human body

Fibrous proteins:

Insoluble in water – used for structural purposes (Keratin & Collagen).

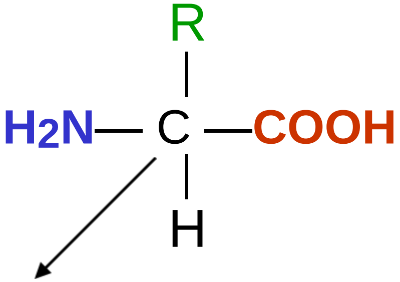
Globular proteins:

More or less soluble in water – used for nonstructural purposes.

Amino acids

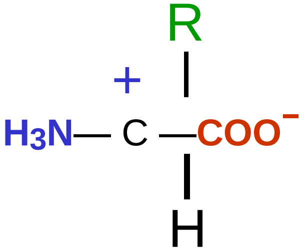
- Are the building blocks of proteins.
- Contain carboxylic acid and amino groups.
- Are ionized in solution (soluble in water).
- They are ionic compounds (solids-high melting points).
- Contain a different side group (R) for each.

side chain



$$H_2N-\overset{\displaystyle R\atop |}{\underset{\displaystyle | \atop H}{C}}-COOH$$

α-carbon



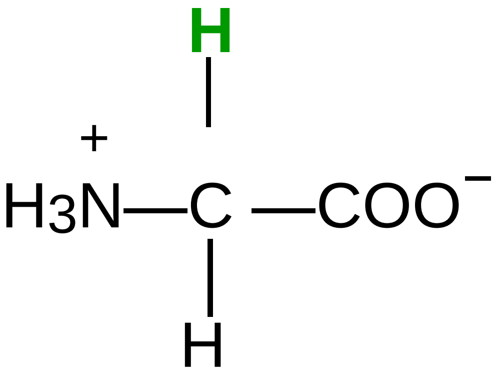
$$H_3\overset{+}{N}-\overset{\displaystyle R\atop |}{\underset{\displaystyle | \atop H}{C}}-COO^-$$

Zwitterion

Ionized form (Salt)

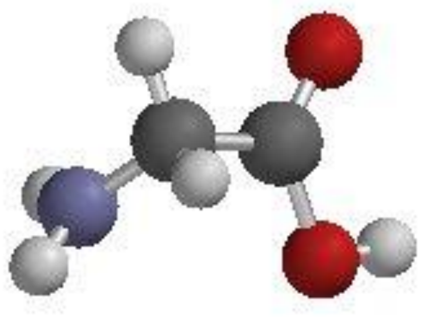
Amino acids

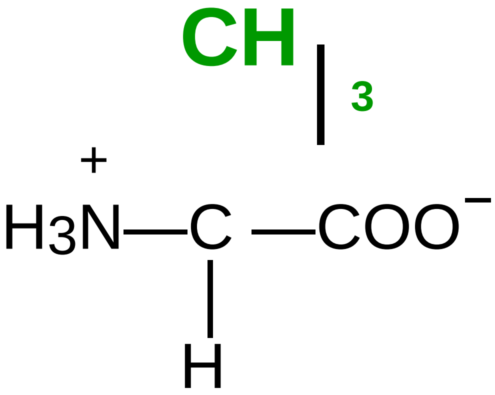
Only difference: containing a different side chain (R) for each.



$$\overset{+}{H_3N}-\overset{\displaystyle \mathbf{H}}{\underset{\displaystyle H}{\overset{|}{\underset{|}{C}}}}-COO^{-}$$

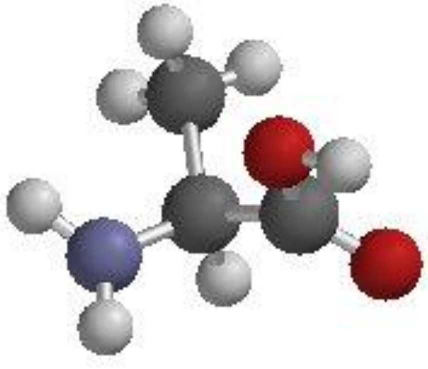
glycine





$$\overset{+}{H_3N}-\overset{\displaystyle \mathbf{CH_3}}{\underset{\displaystyle H}{\overset{|}{\underset{|}{C}}}}-COO^{-}$$

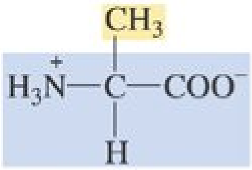
alanine



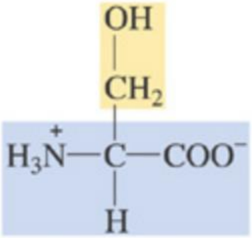
Amino acids

Amino acids are classified as:

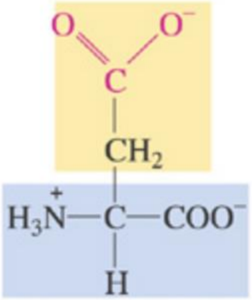
- Nonpolar (Neutral) amino acids (hydrophobic) with hydrocarbon (alkyl or aromatic) sides chains.
- Polar (Neutral) amino acids (hydrophilic) with polar or ionic side chains.
- Acidic amino acids (hydrophilic) with acidic side chains (-COOH).
- Basic amino acids (hydrophilic) with –NH_2 side chains.



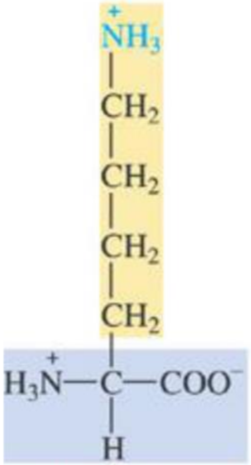
Alanine (Ala)



Serine (Ser)



Aspartic acid (Asp)



Lysine (Lys)

Amino acids

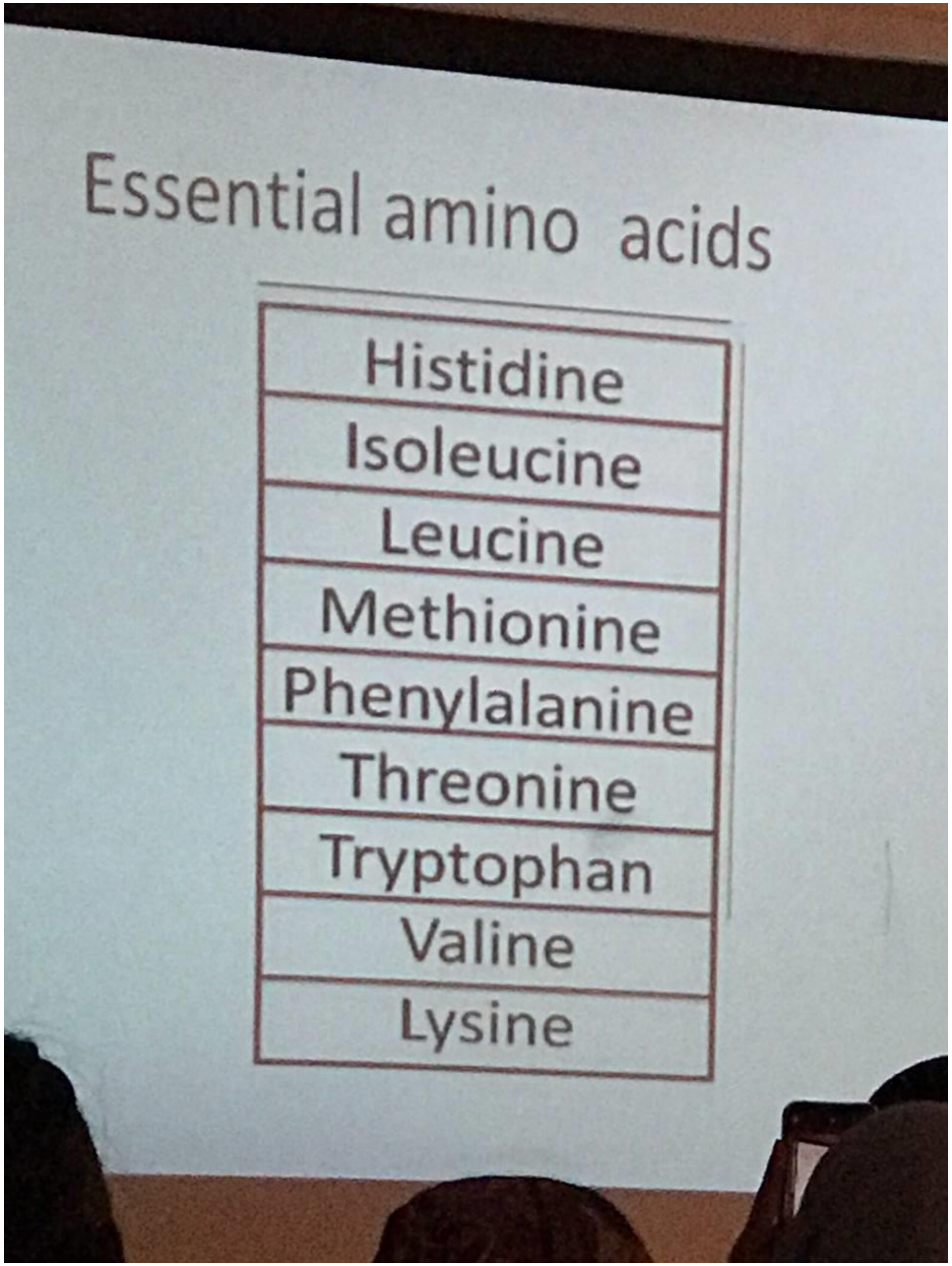
There are many amino acids.

There are only 20 different amino acids in the proteins in humans.

They are called α amino acids.

- Humans cannot synthesize 10 of these 20 amino acids.
(Essential Amino Acids)

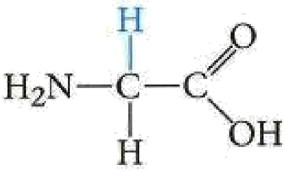
- They must be obtained from the diet (almost daily basis).

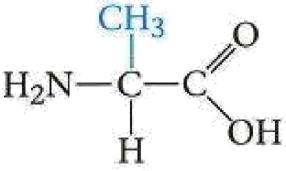


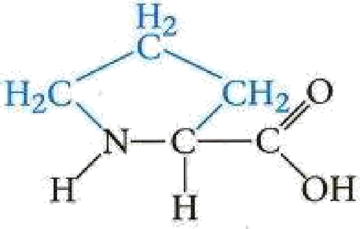
Essential amino acids

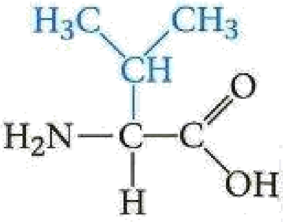
- Histidine
- Isoleucine
- Leucine
- Methionine
- Phenylalanine
- Threonine
- Tryptophan
- Valine
- Lysine

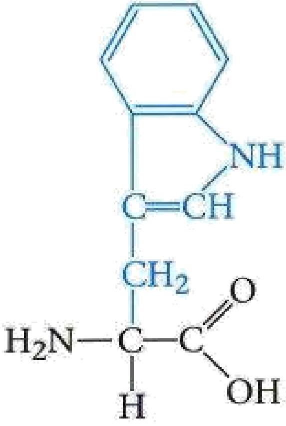
Nonpolar R Groups (hydrophobic)

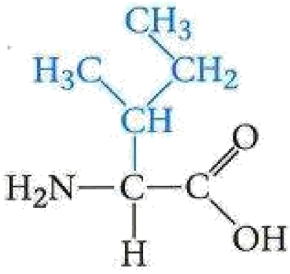

Glycine (gly)

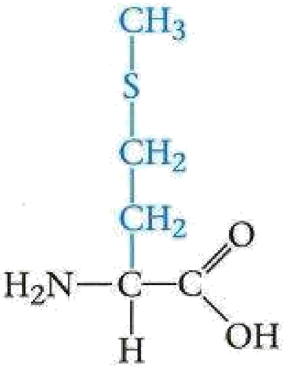

Alanine (ala)

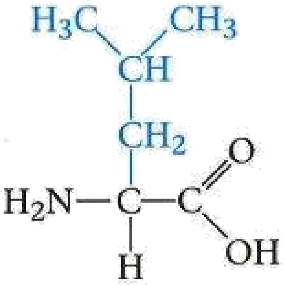

Proline (pro)

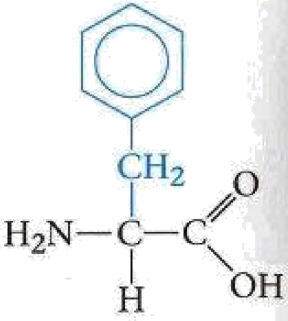

Valine (val)


Tryptophan (trp)

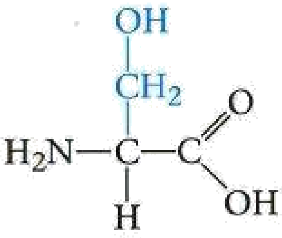

Isoleucine (ile)


Methionine (met)

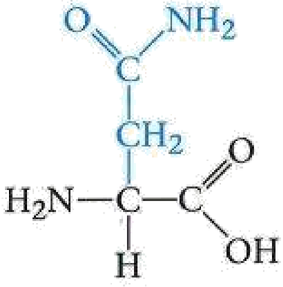

Leucine (leu)


Phenylalanine (phe)

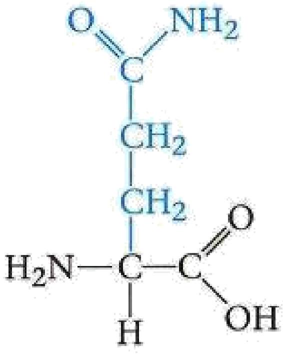
Polar R Groups (hydrophilic)



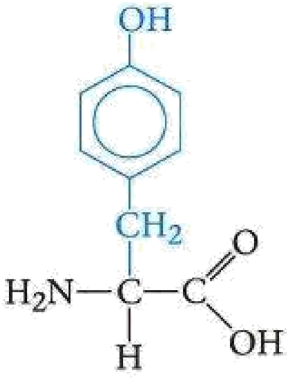
Serine (ser)



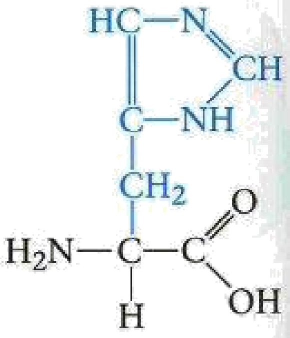
Asparagine (asn)



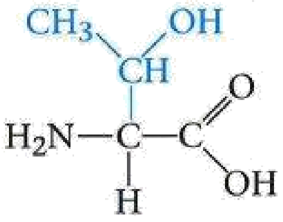
Glutamine (gln)



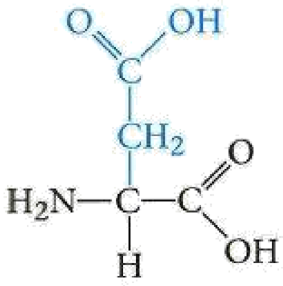
Tyrosine (tyr)



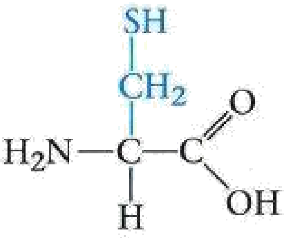
Histidine (his)



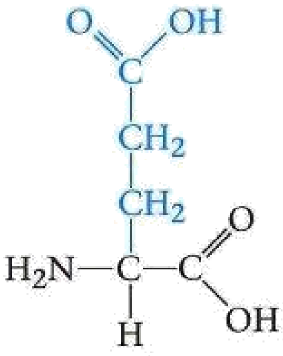
Threonine (thr)



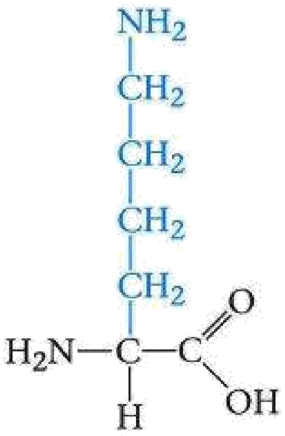
Aspartic acid (asp)



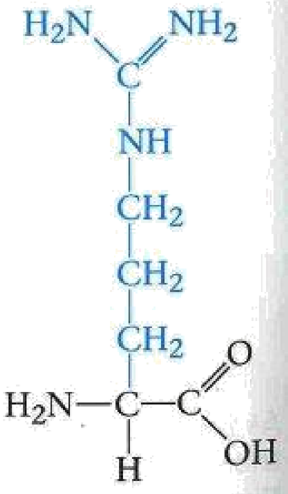
Cysteine (cys)



Glutamic acid (glu)



Lysine (lys)

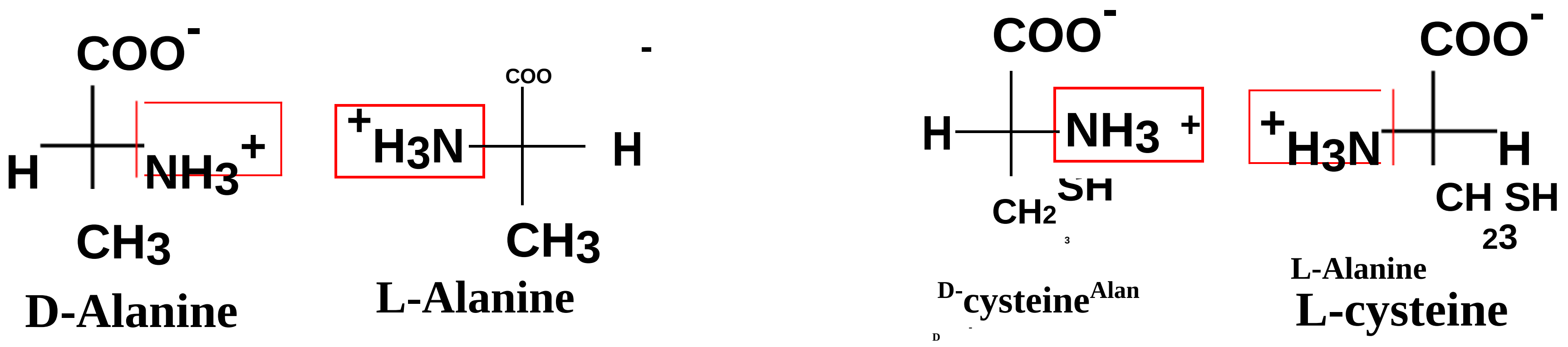


Arginine (arg)

All of the α-amino acids are chiral (except glycine)

↓

Four different groups are attached to central carbon (α-carbon).



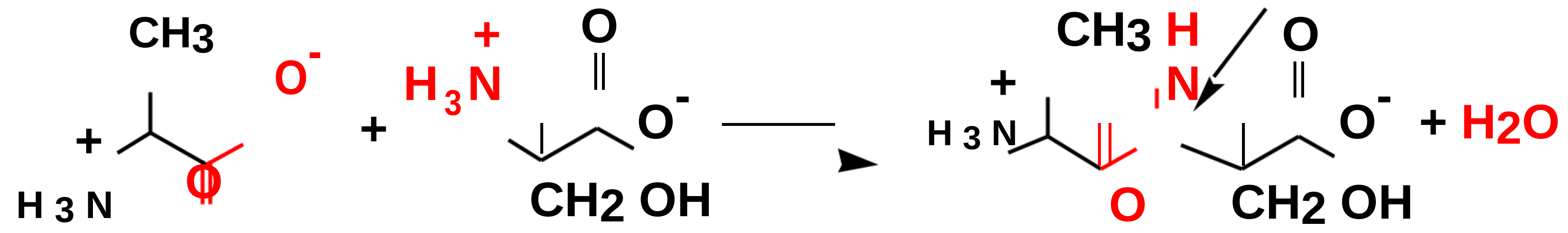
D-Alanine L-Alanine D-cysteine L-cysteine

L isomers is found in the body proteins and in nature.

Peptide

A dipeptide forms:

- When an amide links two amino acids (Peptide bond).
- Between the COO^- of one amino acid and the NH_3^+ of the next amino acid.



peptide (amide bond)
bond

Alanine (Ala)

Serine (Ser)

Alanylserine
(Ala-Ser)

Peptide

- **Dipeptide:** A molecule containing two amino acids joined by a peptide bond.

- **Tripeptide:** A molecule containing three amino acids joined by peptide bonds.

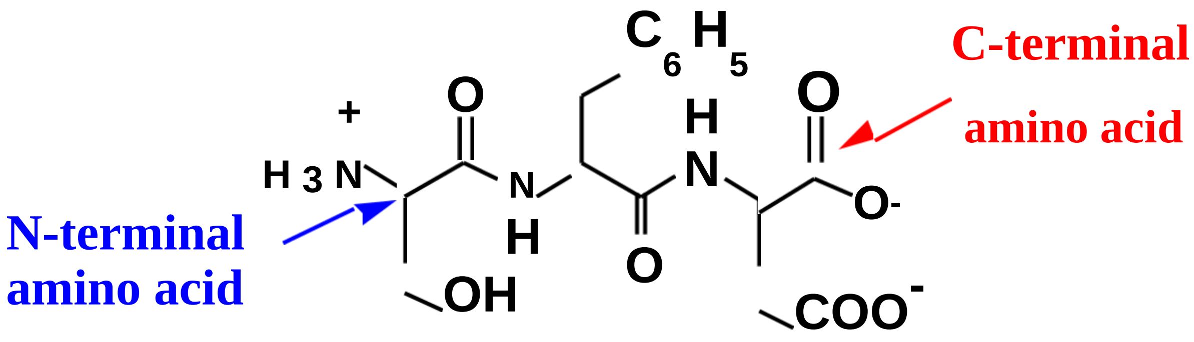
- **Polypeptide:** A macromolecule containing many amino acids joined by peptide bonds.

- **Protein:** A biological macromolecule containing at least 40 amino acids joined by peptide bonds.

Naming of peptides

C-terminal amino acid: the amino acid at the end of the chain having the free -COO$^-$ group (always written at the left).

N-terminal amino acid: the amino acid at the end of the chain having the free -NH$_3^+$ (always written at the right). group



N-terminal amino acid

C-terminal amino acid

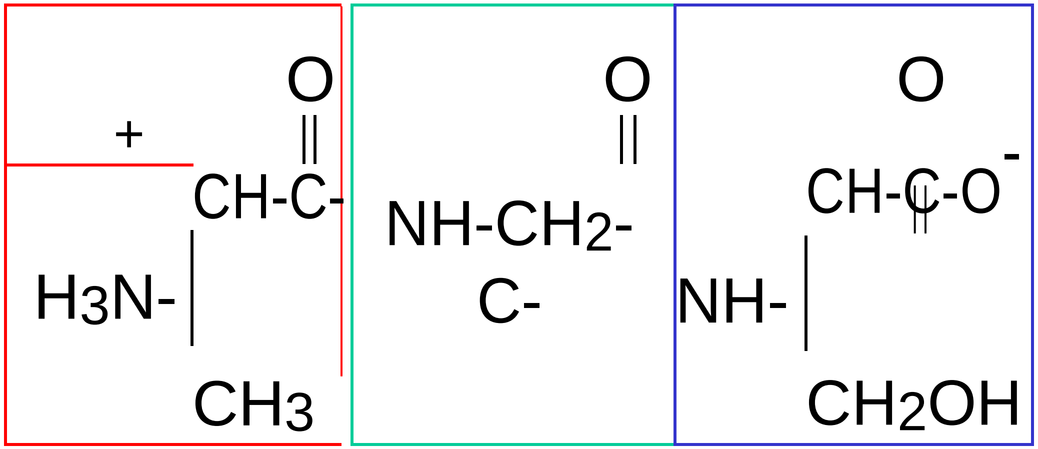
Ser-Phe-Asp

Naming of peptides

- Begin from the N terminal.

- Drop “-ine” or “-ic acid” and it is replaced by “-yl”.

- Give the full name of amino acid at the C terminal.



$$H_3\overset{+}{N}-\underset{CH_3}{CH}-\overset{O}{\overset{\|}{C}}-NH-CH_2-\overset{O}{\overset{\|}{C}}-NH-\underset{CH_2OH}{CH}-\overset{O}{\overset{\|}{C}}-O^-$$

From alanine
alanyl

From glycine
glycyl

From serine
serine

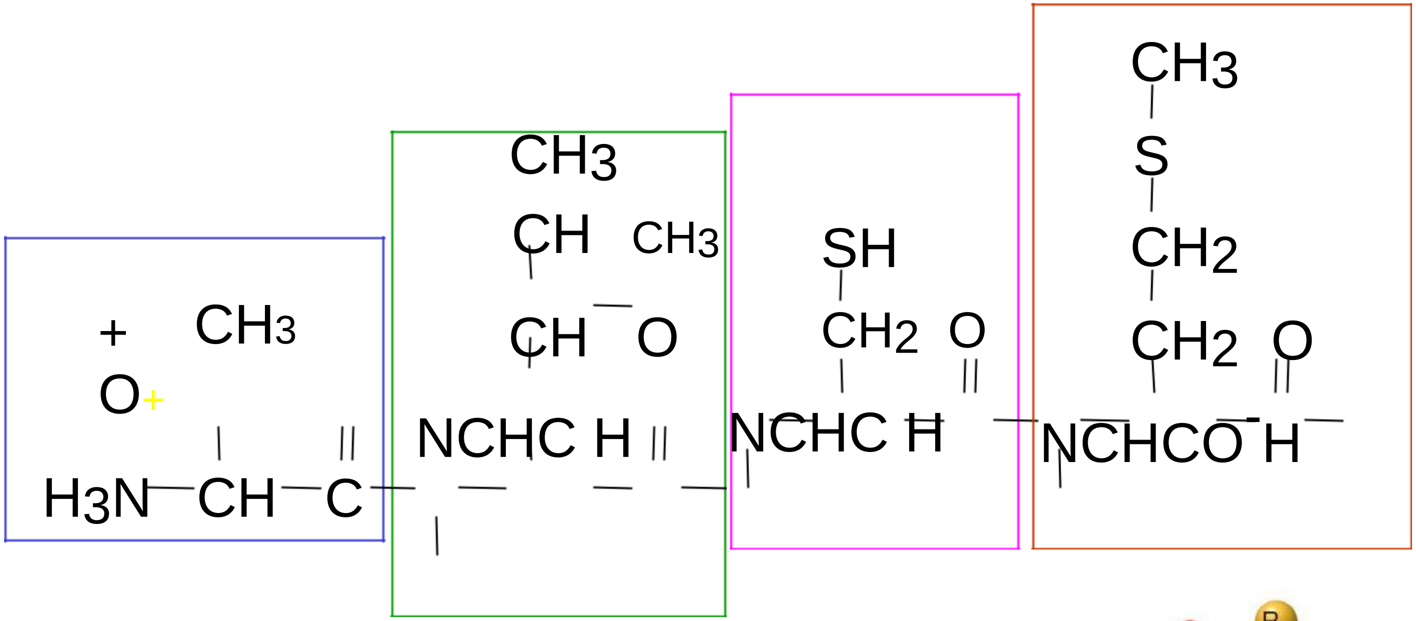
Alanylglycylserine
(Ala-Gly-Ser)

Structure of proteins

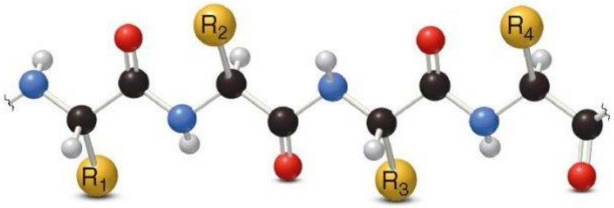
1. Primary structure
2. Secondary structure
3. Tertiary structure
4. Quaternary structure

Primary Structure of proteins

- The order of amino acids held together by peptide bonds.
- Each protein in our body has a unique sequence of amino acids.
- The backbone of a protein.
- All bond angles are 120^{o}, giving the protein a zigzag arrangement.

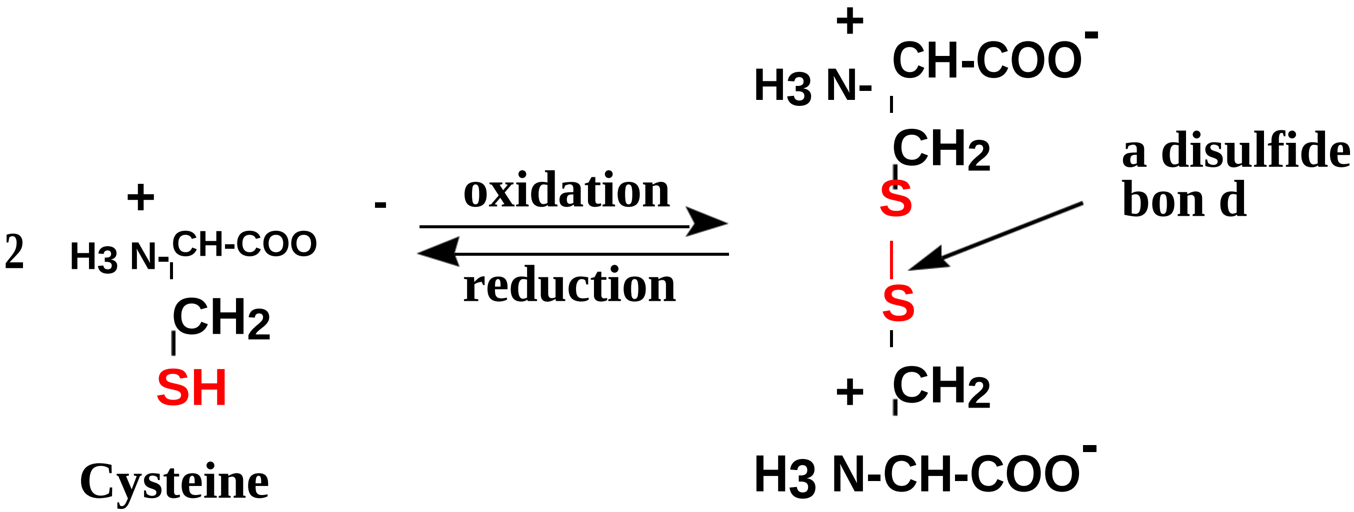


Ala–Leu–Cys–Met



Cysteine

The -SH (sulfhydryl) group of cysteine is easily oxidized to an -S-S- (disulfide).



$$2\ \begin{array}{l} H_3\overset{+}{N}-CH-COO^- \\ \qquad\quad | \\ \qquad\ \ CH_2 \\ \qquad\quad | \\ \qquad\ \ SH \end{array} \ \underset{\text{reduction}}{\overset{\text{oxidation}}{\rightleftharpoons}}\ \begin{array}{l} H_3\overset{+}{N}-CH-COO^- \\ \qquad\quad | \\ \qquad\ \ CH_2 \\ \qquad\quad | \\ \qquad\quad S \\ \qquad\quad | \\ \qquad\quad S \\ \qquad\quad | \\ \qquad\ \ CH_2 \\ \qquad\quad | \\ H_3\overset{+}{N}-CH-COO^- \end{array}$$

Cysteine

Cystine

Primary Structure of proteins

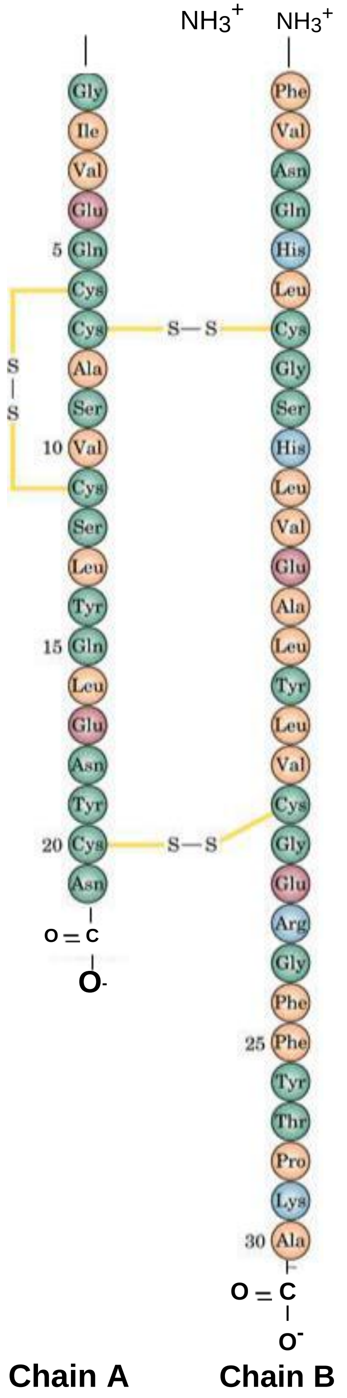
The primary structure of insulin:

- Is a hormone that regulates the glucose level in the blood.

- Was the first amino acid order determined.

- Contains of two polypeptide chains linked by disulfide bonds (formed by side chains (R)).

- Chain A has 21 amino acids and chain B has 30 amino acids.



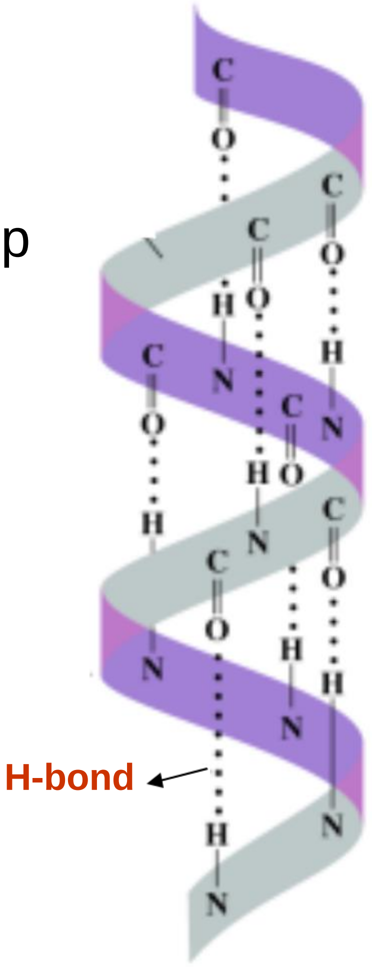
Secondary Structure of proteins

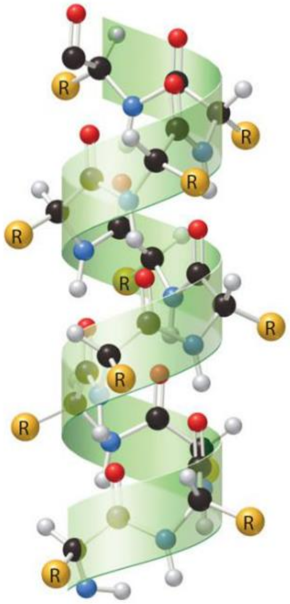
Describes the way the amino acids next to or near to each other along the polypeptide are arranged in space.

1. Alpha helix (α helix)
2. Beta-pleated sheet (-pleated sheet)
3. Triple helix (found in Collagen)

Secondary Structure - α-helix

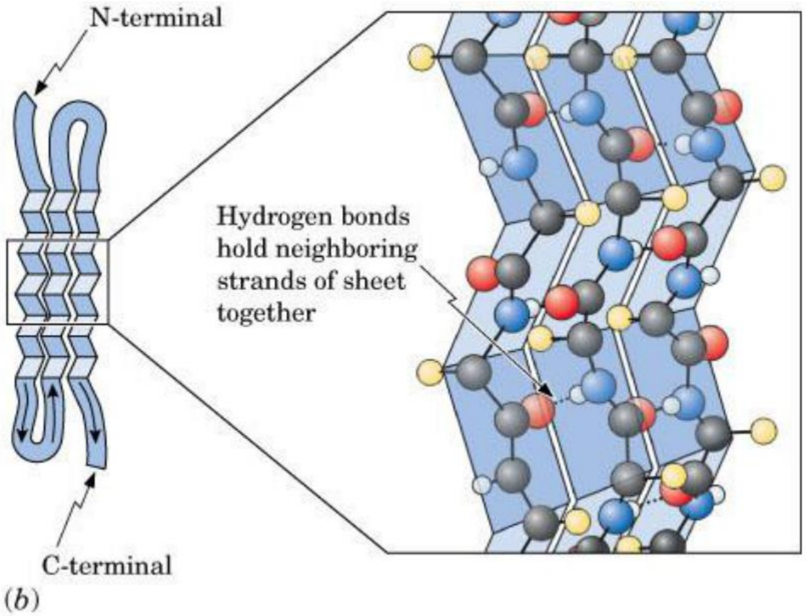
- A section of polypeptide chain coils into a rigid spiral.
- Held by H bonds between the H of N-H group and the O of C=O of the fourth amino acid down the chain (next turn).
- looks like a coiled
- All R- groups point outward from the helix.
- Myosin in muscle and α-Keratin in hair have this arrangement.

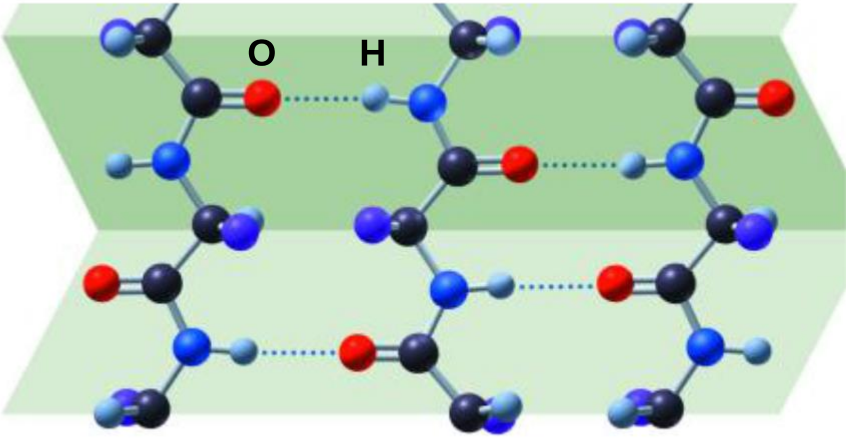




Secondary Structure - -pleated sheet

- Consists of polypeptide chains (strands) arranged side by side.
- Has hydrogen bonds between the peptide chains.
- Has R groups above and below the sheet (vertical).
- Is typical of fibrous proteins such as silk.



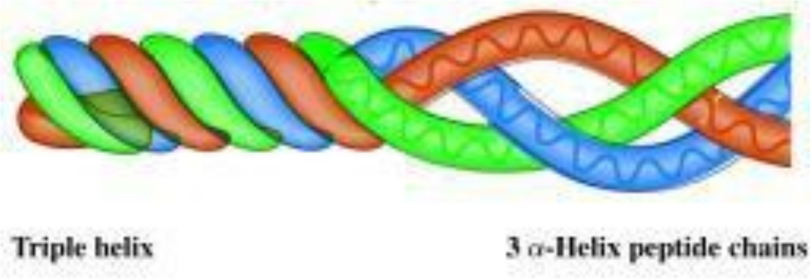


Secondary Structure – Triple helix (Superhelix)

- Collagen is the most abundant protein.

- Three polypeptide chains (three α-helix) woven together.

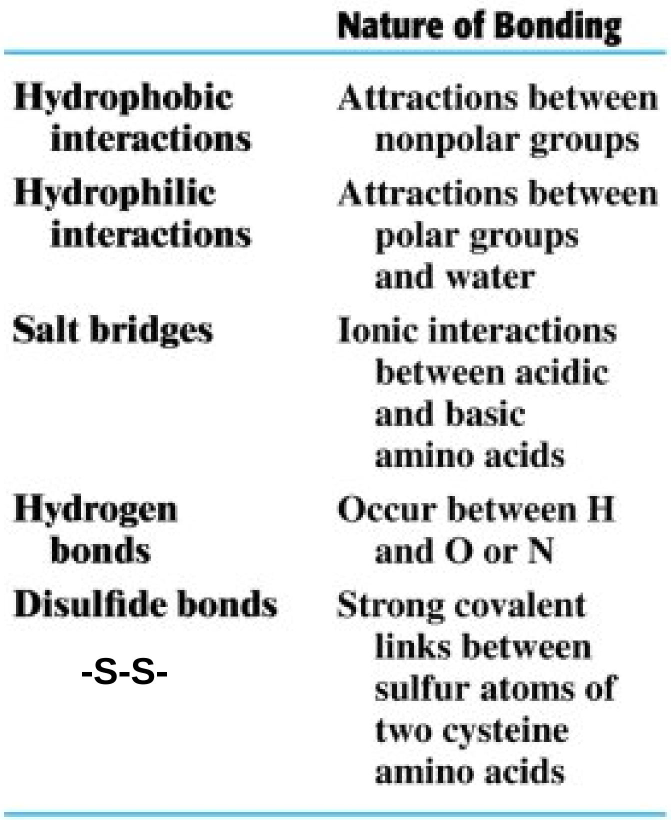
- It is found in connective tissues: bone, teeth, blood vessels, tendons, and cartilage.



Tertiary Structure

The tertiary structure is determined by attractions and repulsions between the side chains (R) of the amino acids in a polypeptide chain.

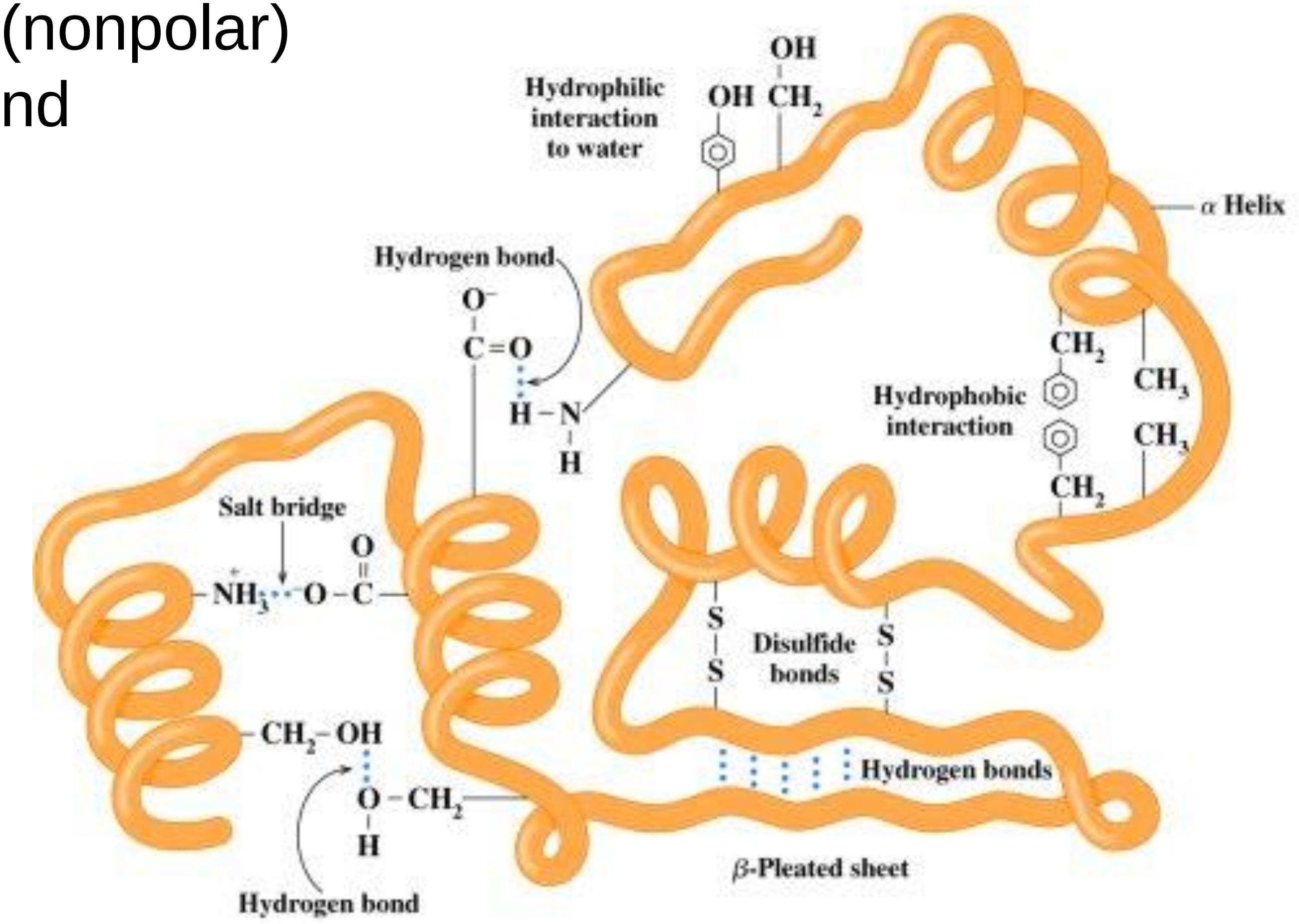
Interactions between side chains of the amino acids fold a protein into a specific three-dimensional shape.



	Nature of Bonding
Hydrophobic interactions	Attractions between nonpolar groups
Hydrophilic interactions	Attractions between polar groups and water
Salt bridges	Ionic interactions between acidic and basic amino acids
Hydrogen bonds	Occur between H and O or N
Disulfide bonds **-S-S-**	Strong covalent links between sulfur atoms of two cysteine amino acids

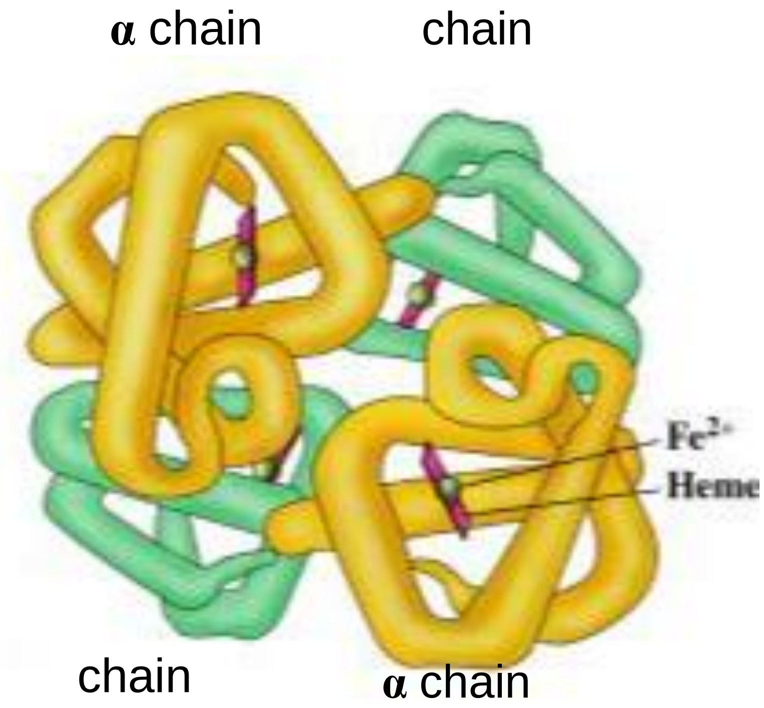
Tertiary Structure

(1) Disulfide (-S-S-)
(2) salt bridge (acid-base)
(3) Hydrophilic (polar)
(4) hydrophobic (nonpolar)
(5) Hydrogen bond



Quaternary Structure

- Occurs when two or more protein units (polypeptide subunits) combine.
- Is stabilized by the same interactions found in tertiary structures (between side chains).
- Hemoglobin consists of four polypeptide chains as subunits.
- Is a globular protein and transports oxygen in blood (four molecules of O_2).



Hemoglobin

Alkenes

Alkenes

Alkenes are alphatic (branched or unbranched) hydrocarbons having one carbon-to-carbon double bond (C=C) and the general molecular formula C_nH_{2n} . Because alkenes contain less than the maximum possible number of hydrogen atoms per carbon atom, they are said to be unsaturated. An older is olefins.

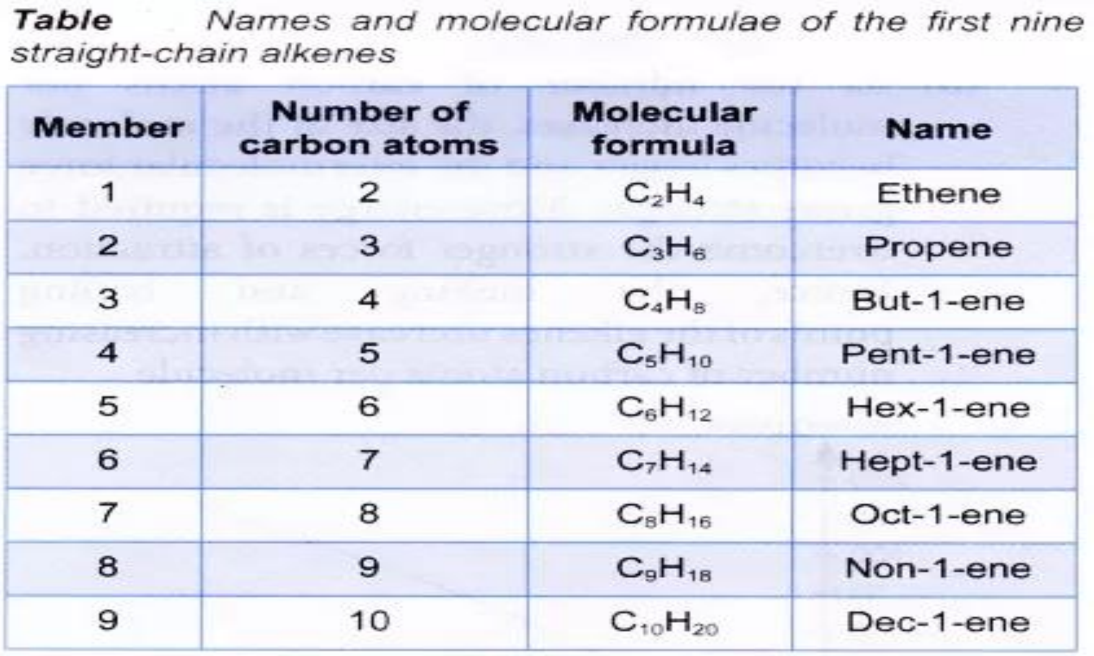


Table Names and molecular formulae of the first nine straight-chain alkenes

Member	Number of carbon atoms	Molecular formula	Name
1	2	C_2H_4	Ethene
2	3	C_3H_6	Propene
3	4	C_4H_8	But-1-ene
4	5	C_5H_{10}	Pent-1-ene
5	6	C_6H_{12}	Hex-1-ene
6	7	C_7H_{14}	Hept-1-ene
7	8	C_8H_{16}	Oct-1-ene
8	9	C_9H_{18}	Non-1-ene
9	10	$C_{10}H_{20}$	Dec-1-ene

Naming of alkene

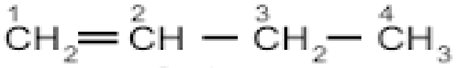
1-The longest chain selected as the parent hydride must include the double bond.

2-The suffix *-ane* in the corresponding alkane parent name is replaced with the suffix *-ene* if the compound contains only one double bond, *-adiene* for two double bonds, *-atriene* for three double bonds, etc.

3-Carbons in the parent chain are numbered beginning with the chain end nearest a double bond.

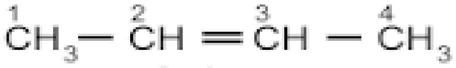
4-The stem name (e.g., pentene or hexene) is affixed with a numerical locant identifying the lowest carbon number associated with the double bond (e.g., pent-2-ene for $CH_3CH=CHCH_2CH_3$).

1-butene , 2-butene and 2-methyl-1-butene



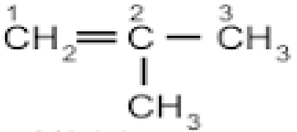
$\overset{1}{C}H_2{=}\overset{2}{C}H-\overset{3}{C}H_2-\overset{4}{C}H_3$

But-1-ene



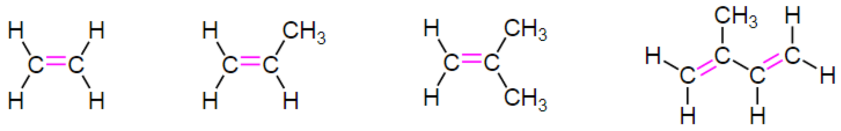
$\overset{1}{C}H_3-\overset{2}{C}H{=}\overset{3}{C}H-\overset{4}{C}H_3$

But-2-ene

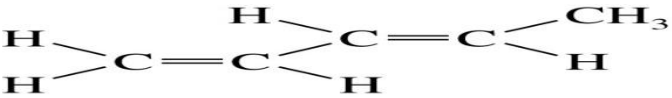


$\overset{1}{C}H_2{=}\underset{CH_3}{\overset{2}{C}}-\overset{3}{C}H_3$

2-Methylprop-1-ene



Common Name:	**Ethylene**	**Propylene**	**Isobutylene**	**Isoprene**
IUPAC Name:	**Ethene**	**Propene**	**2-Methylpropene**	**2-Methyl-1,3-butadiene**

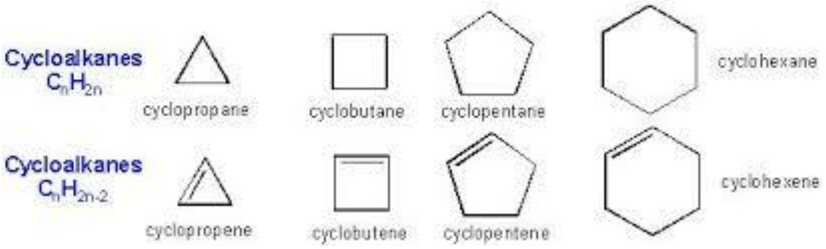


1,3-pentadiene

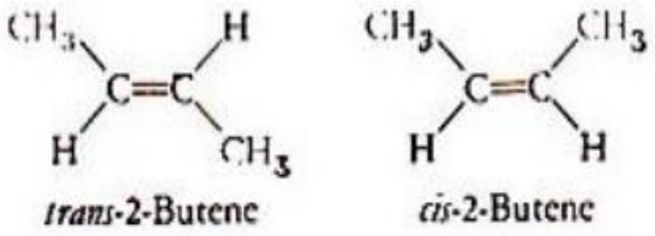
1,4-pentadiene

Cycloalkene

A cycloalkene or cycloolefin is a type of alkene hydrocarbon which contains a closed ring of carbon atoms,. Some cycloalkenes, such as cyclobutene and cyclopentene



Alkenes can form geometric isomers. The trans isomer has the two methyl groups across from each other, and the cis isomer has the groups adjacent to each other(other example trans-2-hexene,cis-2-hexene)



Addition Reactions OF Alkenes

Alkenes and alkanes undergo different types of reactions. Alkanes react by substitution, whereas addition to the double bond is the reaction of alkenes.

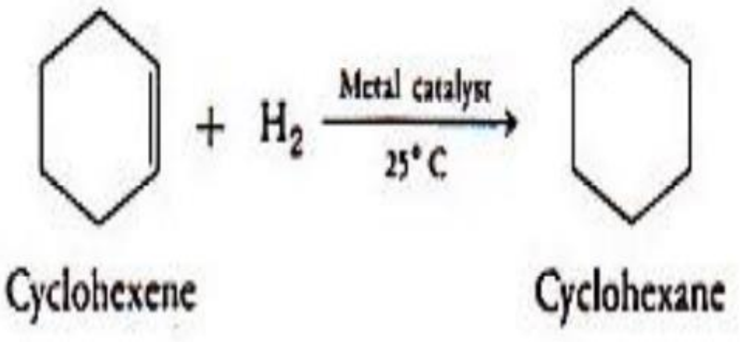
$$CH_3CH_3 + Cl_2 \xrightarrow{\text{Light}} ClCH_2CH_3 + HCl \quad \text{Substitution}$$

$$CH_2{=}CH_2 + Cl_2 \longrightarrow ClCH_2CH_2Cl \quad \text{Addition}$$

1- Addition of Hydrogen:

Reduction of the carbon – carbon double bond The addition of H_2 to alkenes in the presence of a metal catalyst converts them to alkanes.

Some specific examples:

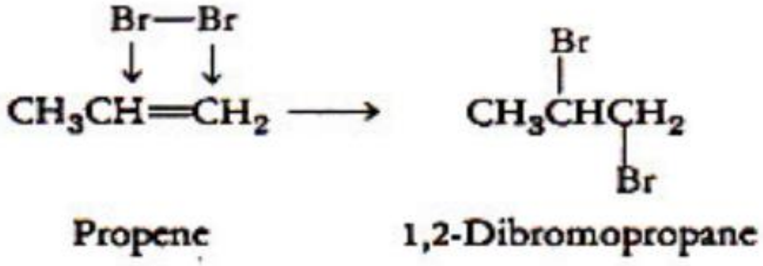


$$CH_2{=}CHCH{=}CHCH_3 + 2\,H_2 \xrightarrow[25^\circ C]{\text{Metal catalyst}} CH_3(CH_2)_3CH_3$$

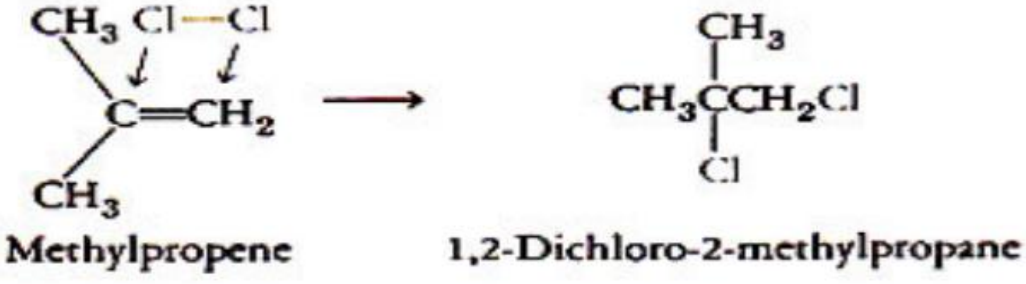
1,3-Pentadiene Pentane

2- Addition of Halogens

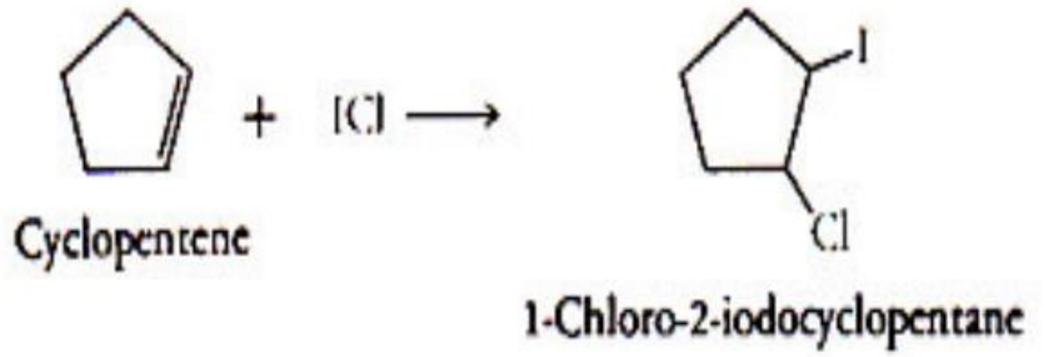
The addition of bromine and chlorine to alkenes occurs readily. Examples



$$CH_3CH{=}CH_2 + Br{-}Br \longrightarrow CH_3CHBrCH_2Br$$

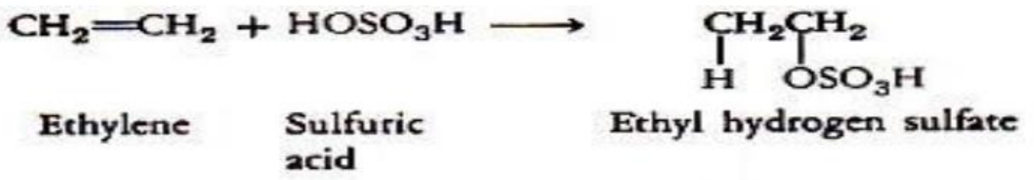


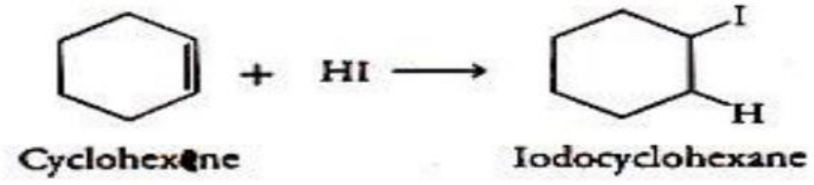
Iodine usually does not react with alkenes but the interhalogens iodine monochloride (ICl) and iodine mono bromide (IBr) are added readily.

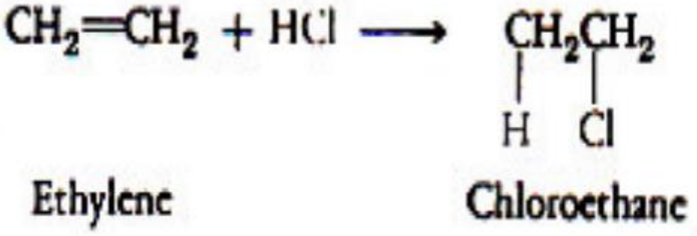


3- **Addition of Acides**

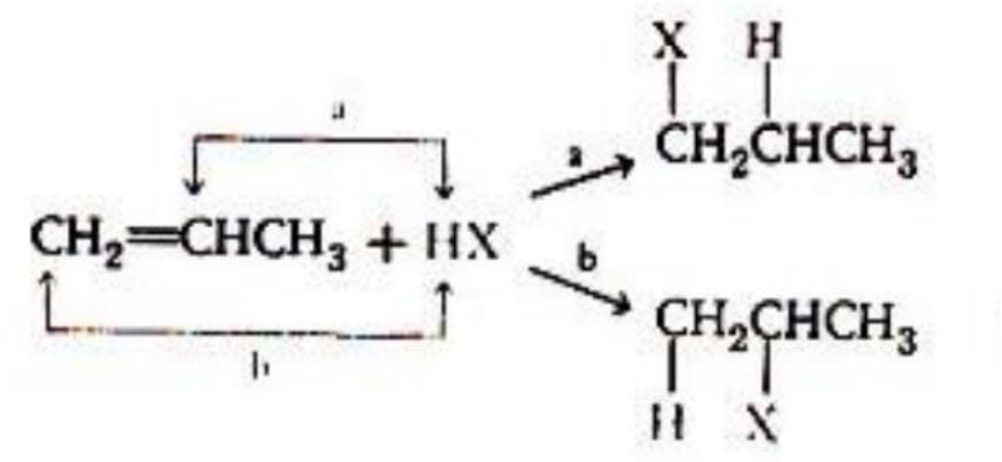
Acids such as sulfuric acid and the hydrogen halides are readily added to alkenes. Examples



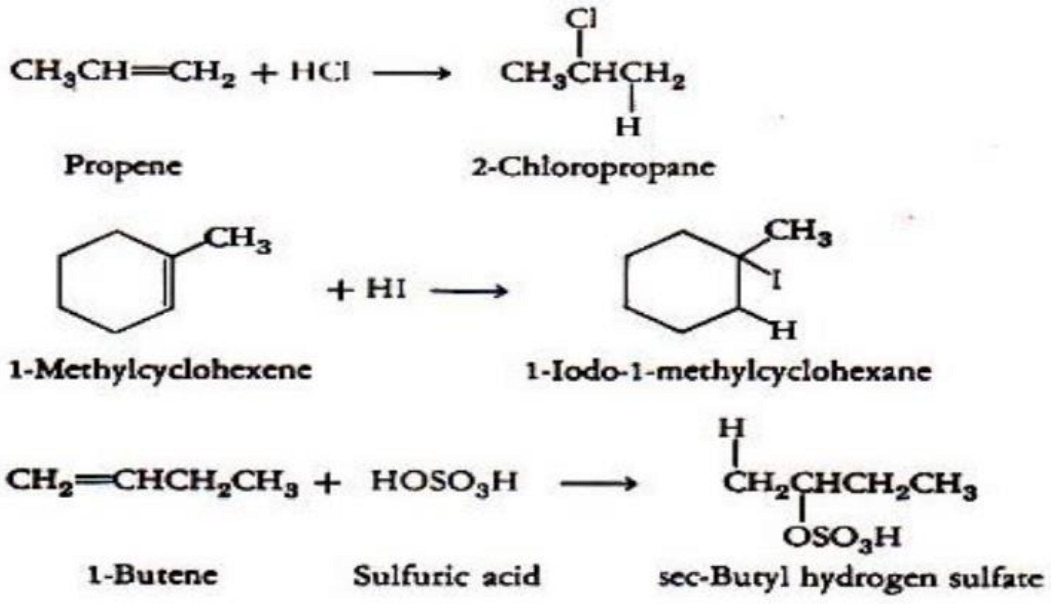




The addition of an unsymmetrical reagent such as HX to an unsymmetrical alkene can form two isometric products:

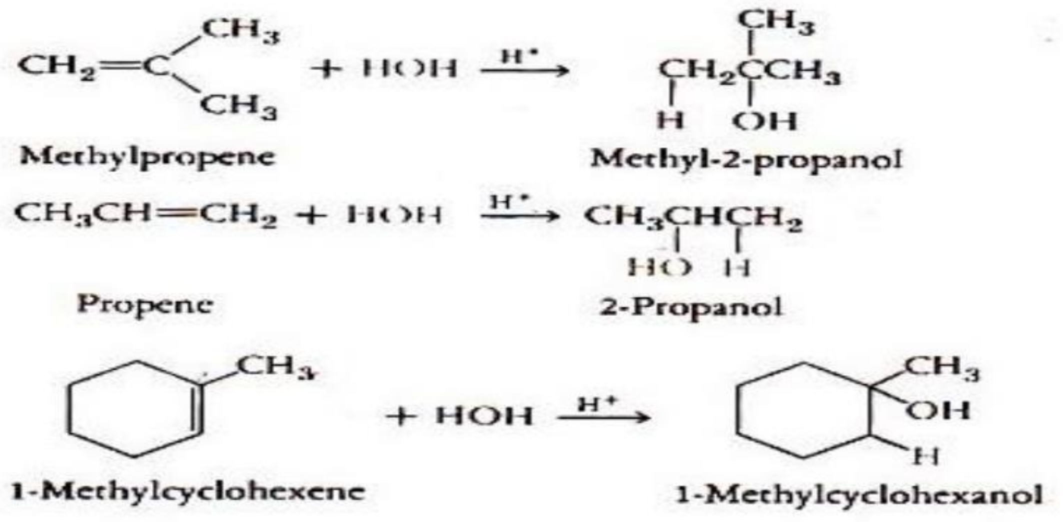


Actually, only one product is obtained, the one formed when the hydrogen of the acid is added to the carbon of the double bond containing the greatest number of hydrogens.This rule is called the Markownikoff rule. Examples

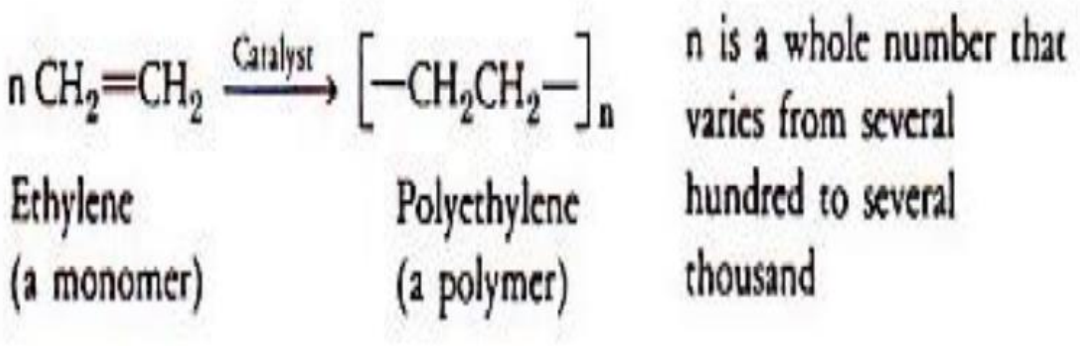


4- Addition of Water

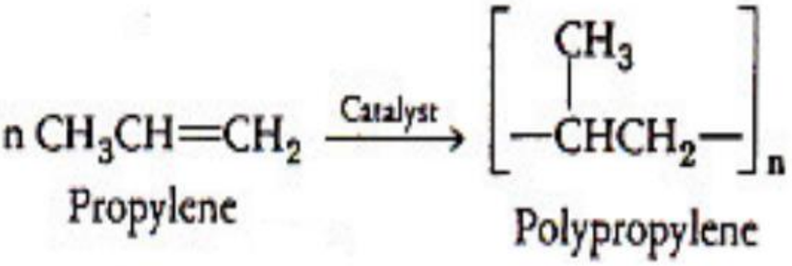
The addition of water to alkenes, called Hydration,riquires the presence of a strong acid catalyst such as sulfuric or phosphoric acid.



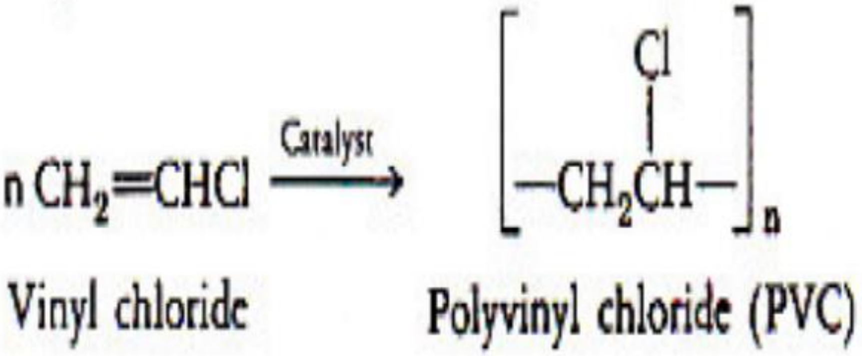
Polymerization: One alkene adding to another Polymer: A huge molecule with a high molecular weight formed by combining a large number of monomers(monomer:one molecule of an alkene),for example, ethylene polymerizes to form polyethylene with the a catalyst.



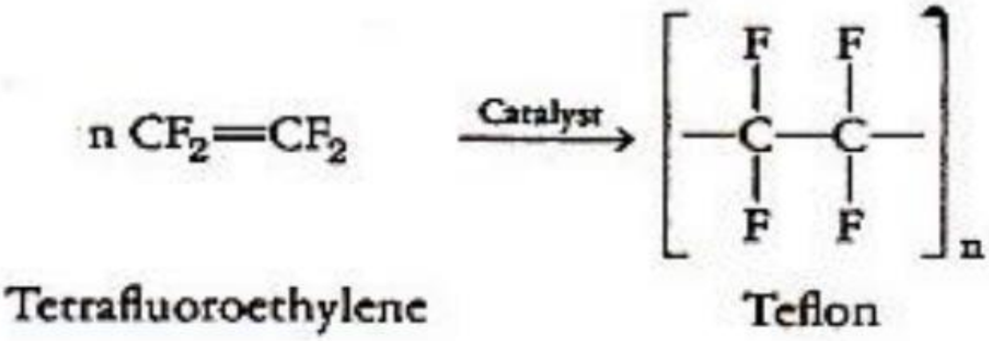
n is a whole number that varies from several hundred to several thousand



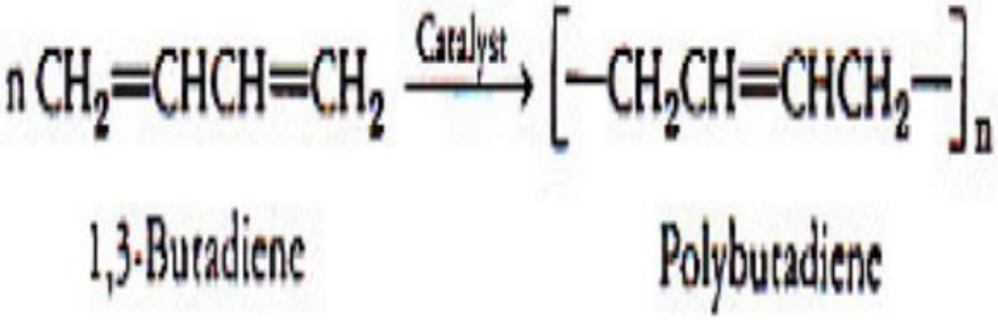
$$n\,CH_3CH{=}CH_2 \xrightarrow{Catalyst} \left[-\overset{\displaystyle CH_3}{\overset{|}{C}}HCH_2-\right]_n$$



$$n\,CH_2{=}CHCl \xrightarrow{Catalyst} \left[-CH_2\overset{\displaystyle Cl}{\overset{|}{C}}H-\right]_n$$



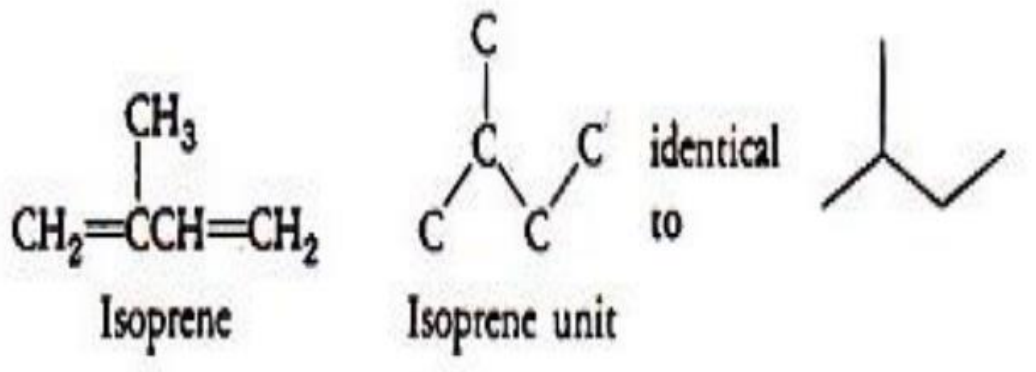
$$n\,CF_2{=}CF_2 \xrightarrow{Catalyst} \left[-\underset{\displaystyle F}{\overset{\displaystyle F}{C}}-\underset{\displaystyle F}{\overset{\displaystyle F}{C}}-\right]_n$$



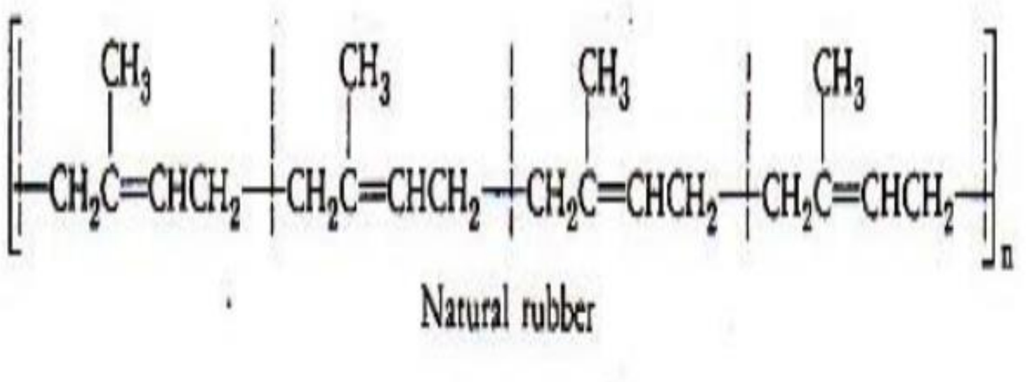
$$n\,CH_2{=}CHCH{=}CH_2 \xrightarrow{Catalyst} \left[-CH_2CH{=}CHCH_2-\right]_n$$

Polymers Formed By Living Systems

Polymerization also occurs in living systems. A large number of compounds called terpenes are found in living systems. These compounds are all polymers made of a repeating five-carbon unit that is structurally related to isoprene

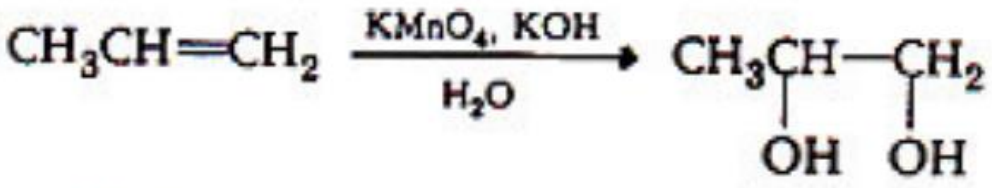


Natural rubber is an example of a polymer made by combining thousands of isoprene unit.



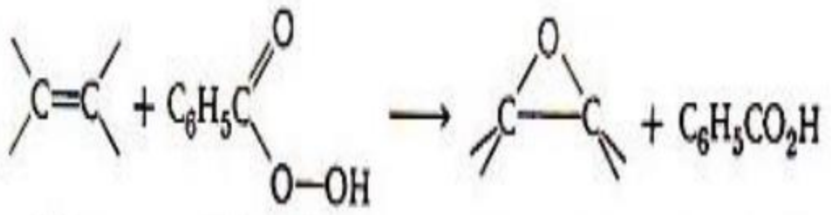
Oxidation of carbon-carbon Double Bonds

The carbon-carbon double bond of an alkene reacts readily with a number of oxidizing reagents such as potassium permanganate, peracids and ozone. The product of the reaction depends on the reagent and the experimental conditions

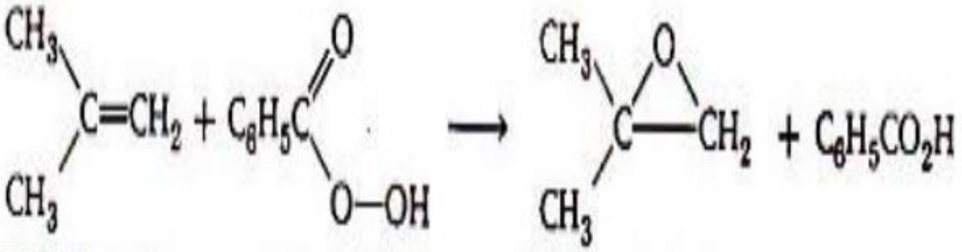


$$\mathrm{CH_3CH{=}CH_2} \xrightarrow[\mathrm{H_2O}]{\mathrm{KMnO_4,\ KOH}} \mathrm{CH_3\underset{\displaystyle OH}{\underset{|}{C}H}{-}\underset{\displaystyle OH}{\underset{|}{C}H_2}}$$

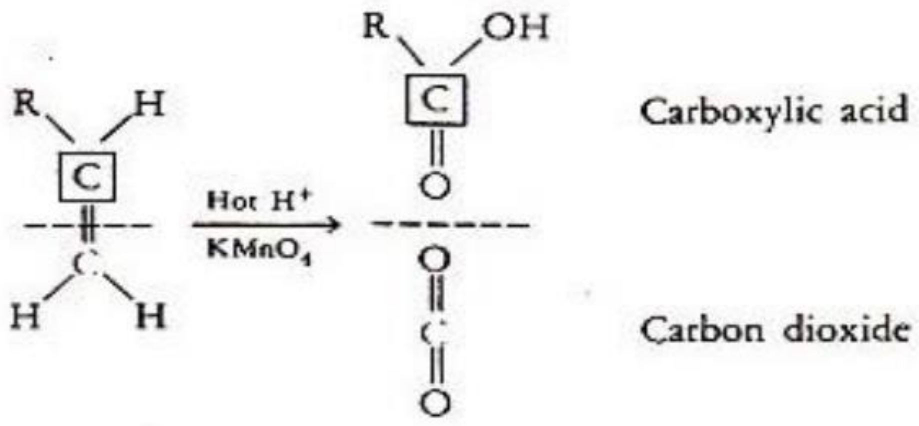
Propene 1,2-Propanediol

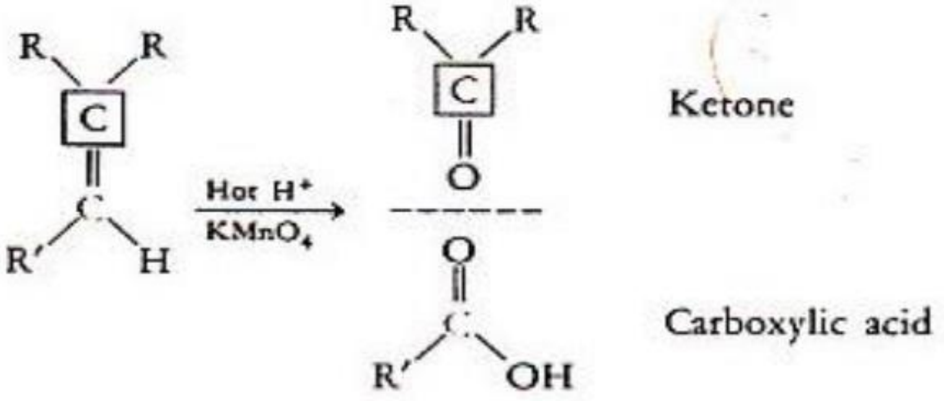


Alkene Perbenzoic acid An oxirane (an epoxide) Benzoic acid

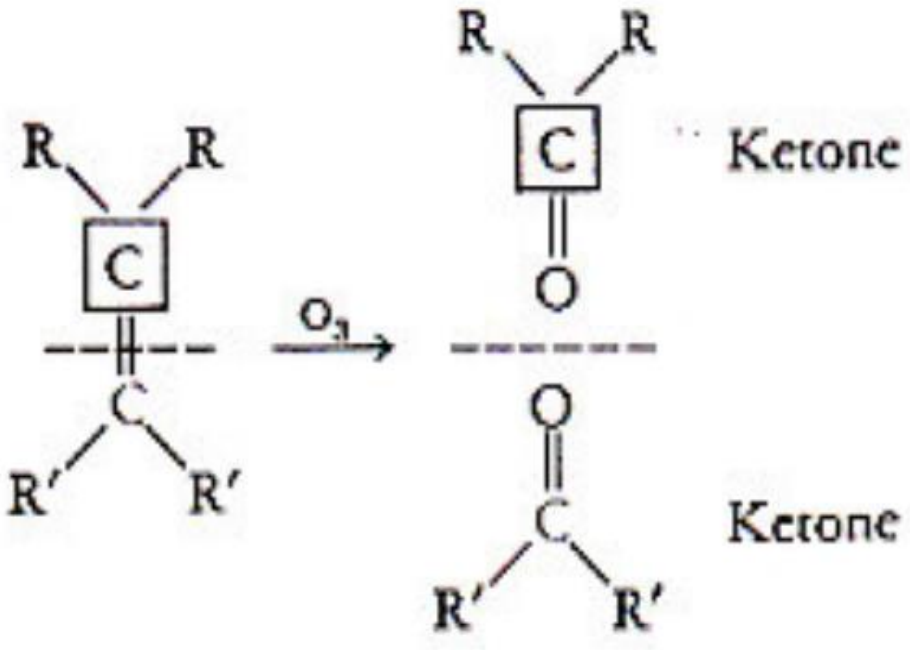


Methylpropene Perbenzoic acid 1,1-Dimethyloxirane Benzoic acid





The reaction with O3 called ozonolysis,also breaks the double bond.



Alkynes

Any of a group of unsaturated hydrocarbons that have carbon atoms in chains linked by one or more triple bonds and that have the general formula **C_nH_{2n-2}** Alkynes can be solid, liquid, or gaseous and include acetylene. The group of alkynes as a whole is called the **alkyne series** or the acetylene series. The first five members of the alkyne series are acetylene (or ethyne), propyne, butyne, pentyne, and hexyne.H-C≡C-H ,R-C≡C-H and R-C≡C-R .

CH_3-C≡C-H methyl acetylene , CH_3-CH_2-C≡C-H ethylacetylene

CH_3-C≡C- CH_3 dimethyl acetylene CH_3-CH_2-C≡C-H CH_3 methyl ethyl acetylene

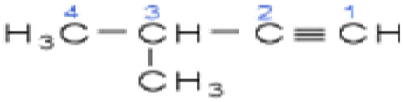
Naming of Alkynes

1- The -yne suffix (ending) indicates an alkyne.

2- The longest chain chosen for the root name must include both carbon atoms of the triple bond.

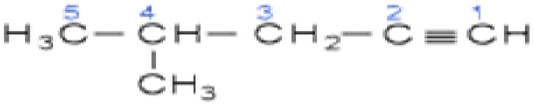
3- The root chain must be numbered from the end nearest a triple bond carbon atom. If the triple bond is in the center of the chain, the nearest substituent rule is used to determine the end where numbering starts.

4- The smaller of the two numbers designating the carbon atoms of the triple bond is used as the triple bond locator.



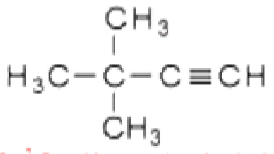
$$H_3\overset{4}{C}-\underset{CH_3}{\overset{3}{C}H}-\overset{2}{C}\equiv\overset{1}{C}H$$

3-methyl-1-butyne



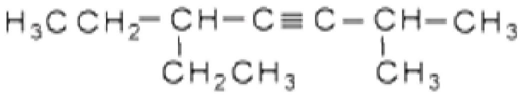
$$H_3\overset{5}{C}-\underset{CH_3}{\overset{4}{C}H}-\overset{3}{C}H_2-\overset{2}{C}\equiv\overset{1}{C}H$$

4-methyl-1-pentyne



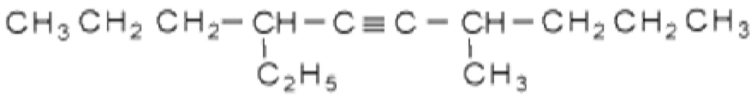
$H_3C-\underset{\underset{\displaystyle CH_3}{|}}{\overset{\overset{\displaystyle CH_3}{|}}{C}}-C\equiv CH$

3, 3-dimethyl-1-butyne



$H_3CCH_2-\underset{\underset{\displaystyle CH_2CH_3}{|}}{CH}-C\equiv C-\underset{\underset{\displaystyle CH_3}{|}}{CH}-CH_3$

5-ethyl-2-methyl-3-heptyne



$CH_3CH_2CH_2-\underset{\underset{\displaystyle C_2H_5}{|}}{CH}-C\equiv C-\underset{\underset{\displaystyle CH_3}{|}}{CH}-CH_2CH_2CH_3$

4-ethyl-7-methyl-5-decyne

Physical Properties of Alkynes

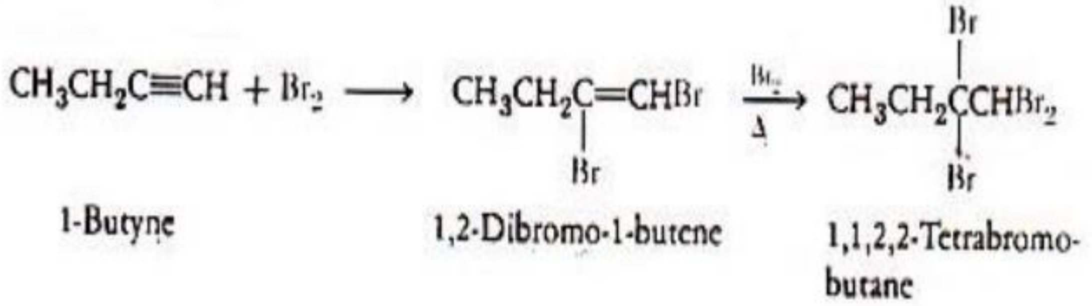
The properties of alkynes pretty much follow the same pattern of those of alkanes and alkenes.Alkynes are unsaturated carbon that shares a triple bond at **the** carbon site.All alkynes are odourless and colourless with the exception of ethylene which has a slight distinctive odour.

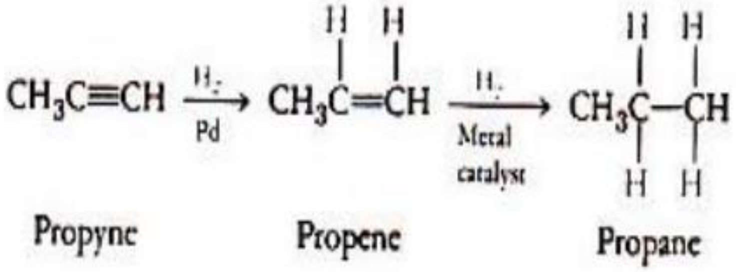
The first three alkynes are gases, and the next eight are liquids. All alkynes higher than these eleven are solids .Alkynes are slightly polar in nature

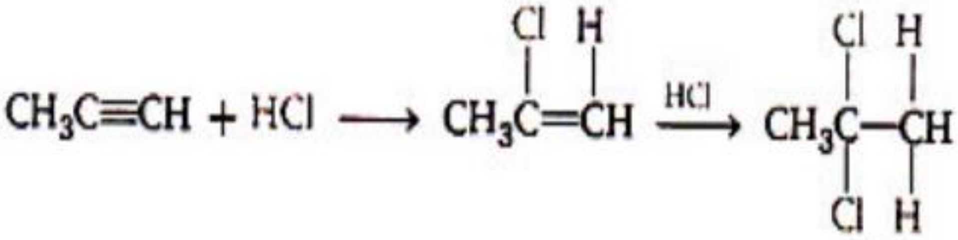
The boiling point and melting point of alkynes increases as their molecular structure grows bigger. The boiling point increases with increase in their molecular mass

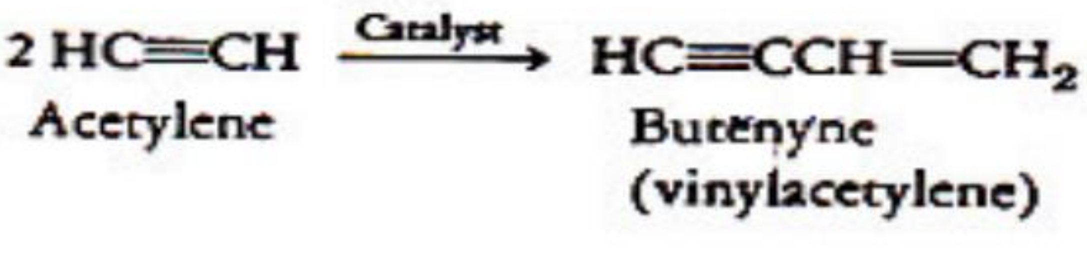
Also, the boiling points of alkynes are slightly higher than those of their corresponding alkenes, due to the one extra bond at the carbon site.

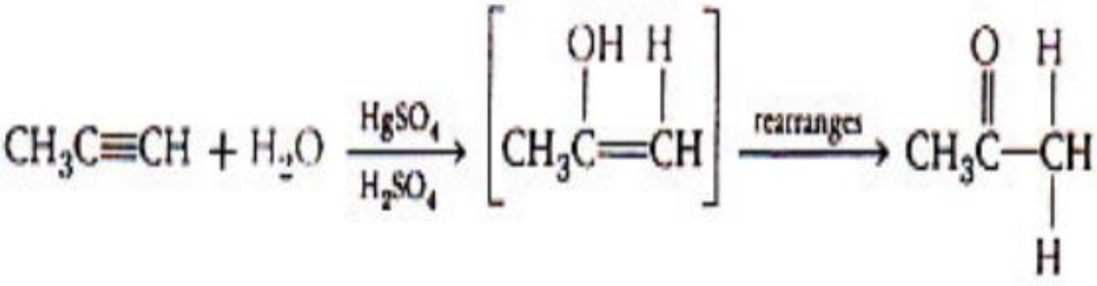
The reactions of alkynes are similar to those of alkenes.





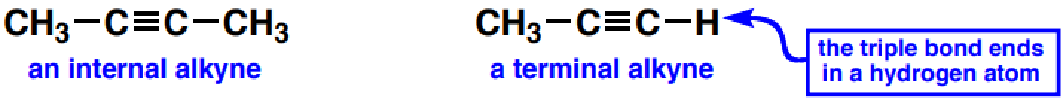


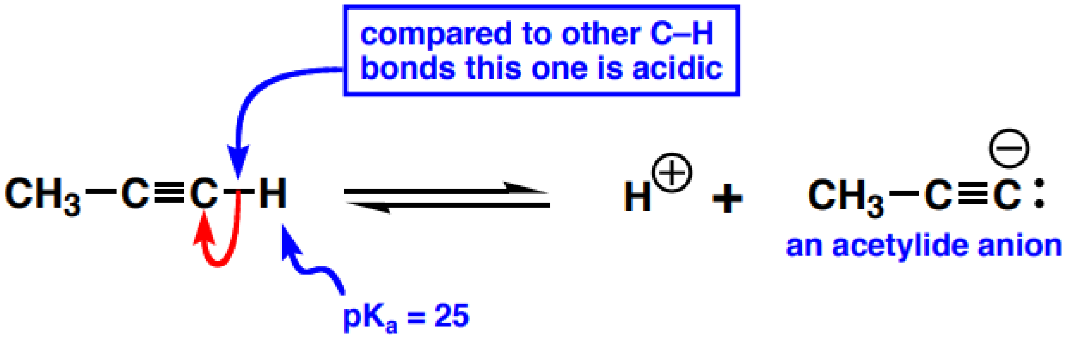


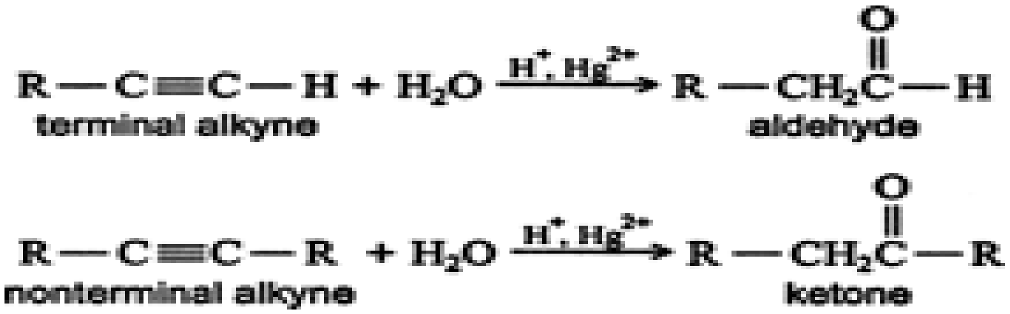


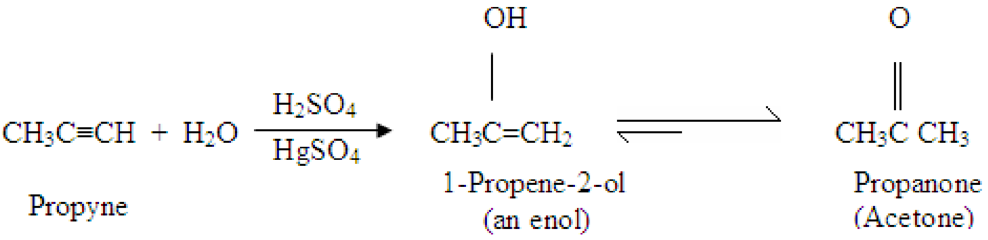
Alkyne has two functional groups (carbon triple bond) and(acidic hydrogen atom). are referred to as **terminal alkynes.** These types of alkynes are weakly acidic. Exposure to a strong base, such as sodium amide, produces an acid-base reaction

The **acidity** of a terminal **alkyne** is due to the high level of s character in the sp hybrid orbital, which bonds with the s orbital of the **hydrogen atom** to form a single covalent bond. ... This slight positive charge makes the **hydrogen atom** a weak **proton**, which can be removed by a strong base







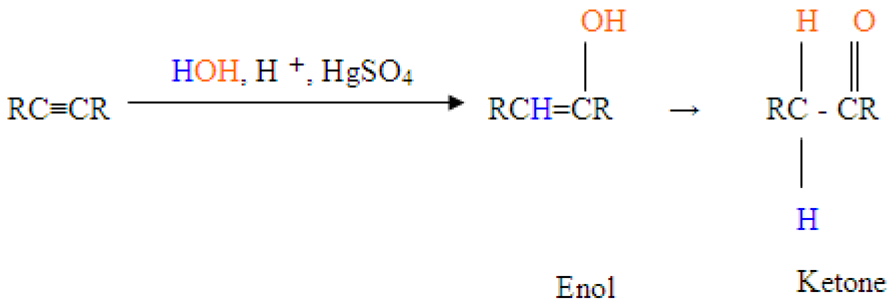


$$CH_3C\equiv CH + H_2O \xrightarrow[HgSO_4]{H_2SO_4} \overset{OH}{\overset{|}{CH_3C}}{=}CH_2 \rightleftharpoons \overset{O}{\overset{\|}{CH_3C}}\,CH_3$$

Propyne

1-Propene-2-ol
(an enol)

Propanone
(Acetone)



Preparing of alkynes

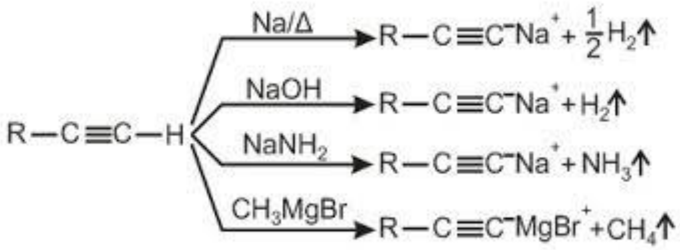
1-Methan converts into acetylene

$$2CH_4 \xrightarrow[0.01-0.1\ \text{sec.}]{1500^\circ C} HC\equiv CH + 3H_2$$

2-Calcium carbide (CaC_2) reacts with water to produce acetylene (C_2H_2):

$CaC_2 + 2\ H_2O \rightarrow Ca(OH)_2 + HC\equiv CH$

3-Substitution of acidic hydrogen by sodium atom



All the salts can react with alkylhalide (R-X) to lengthen alkyne chain.

$H\text{-}C\equiv C\text{-}H + Na \longrightarrow H\text{-}C\equiv CNa + 1/2\ H_2$

$H\text{-}C\equiv CNa + CH_3Cl \longrightarrow H\text{-}C\equiv C\text{-}CH_3 + NaCl$

$H\text{-}C\equiv C\text{-}CH_3 + Na \longrightarrow Na\text{-}\ C\equiv C\text{-}CH_3 + 1/2\ H_2$

$Na\text{-} C\equiv C\text{-}CH_3 + CH_3Cl \rightarrow CH_3\text{-}C\equiv C\text{-}CH_3 + NaCl$

Carbohydrates

Carbohydrates are defined chemically as aldehyde or ketone derivatives of the higher polyhydric alcohols, or compounds which yield these derivatives on hydrolysis

CLASSIFICATION

1. MONOSACCHARIDES(simple sugars) General formula $C_nH_{2n}O_n$ They can be subdivided:

Depending upon the number of carbon atoms , Trioses, pentoses, hexoses, etc

- Depending upon whether aldehyde or ketone groups are present as aldoses or ketoses

2. Disaccharides: sugars which yield two molecules of the same or different molecules of monosaccharide on hydrolysis general formula $C_n(H_2O)_{n-1}$ Example :

maltose(glucose+glucose),

Lactose(glucose+galactose),

Sucrose(glucose+ fructose)

3. Oligosaccharides: are those which yield 3 to 10 monosaccharides units on hydrolysis

4. polysaccharides(glycans): are those which yield more than 10 molecules of monosaccharides on hydrolysis general formula $(C_6H_{10}O_5)_n$ Polysaccharides are further divided into:

- Homopolysaccharides

- Hetropolysaccharides

BIOMEDICAL IMPORTANCE

- Chief source of energy
- Constituents of compound lipids and conjugated proteins
- Degradation products act as promoters or catalysts
- Are used as drugs like cardiac glycosides/antibiotics
- Lactose: principal sugar of milk , in lactating mammary gland
- Degradation products utilized for synthesis of other substances such as fatty acid , cholesterol, amino acid,etc
- Inherited deficiency of certain enzymes in metabolic pathways of different carbohydrates can cause diseases

Carbohydrates

DEFINITION

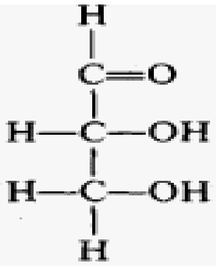
• Carbohydrates are defined chemically as aldehyde or ketone derivatives of the higher polyhydric alcohols, or compounds which yield these derivatives on hydrolysis

CLASSIFICATION

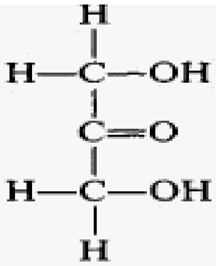
1. MONOSACCHARIDES(simple sugars) General formula $C_nH_{2n}O_n$ They can be subdivided:

A- • Depending upon the number of carbon atoms , Trioses, pentoses, hexoses, etc

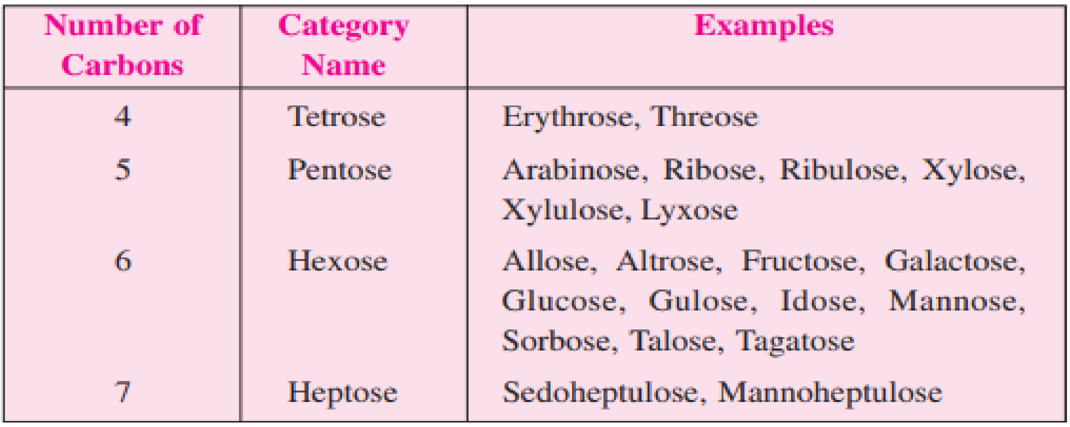
B- • Depending upon whether aldehyde or ketone groups are present as aldoses or ketoses



D = glyceraldehyde (aldose)



Dihydroxyacetone (ketose)



Number of Carbons	Category Name	Examples
4	Tetrose	Erythrose, Threose
5	Pentose	Arabinose, Ribose, Ribulose, Xylose, Xylulose, Lyxose
6	Hexose	Allose, Altrose, Fructose, Galactose, Glucose, Gulose, Idose, Mannose, Sorbose, Talose, Tagatose
7	Heptose	Sedoheptulose, Mannoheptulose

2. Disaccharides: sugars which yield two molecules of the same or different molecules of monosaccharide on hydrolysis general formula $Cn(H_2O)_{n-1}$

Example:

maltose(glucose+glucose),

Lactose(glucose+galactose)

,Sucrose(glucose+ fructose) •

3. Oligosaccharides: are those which yield 3 to 10 monosaccharides units on hydrolysis •

4.polysaccharides(glycans): are those which yield more than 10 molecules of monosaccharides on hydrolysis general formula $(C_6H_{10}O_5)n$ Polysaccharides are further divided into:

A• Homopolysaccharides

B• Hetropolysaccharides

BIOMEDICAL IMPORTANCE

- Chief source of energy
- Constituents of compound lipids and conjugated proteins
- Degradation products act as promoters or catalysts
- Are used as drugs like cardiac glycosides/antibiotics
- Lactose: principal sugar of milk , in lactating mammary gland
- Degradation products utilized for syntesis of other substances such as fatty acid , cholesterol, amino acid,etc
- Inherited deficiency of certain enzymes in metabolic pathways of different carbohydrates can cause diseases

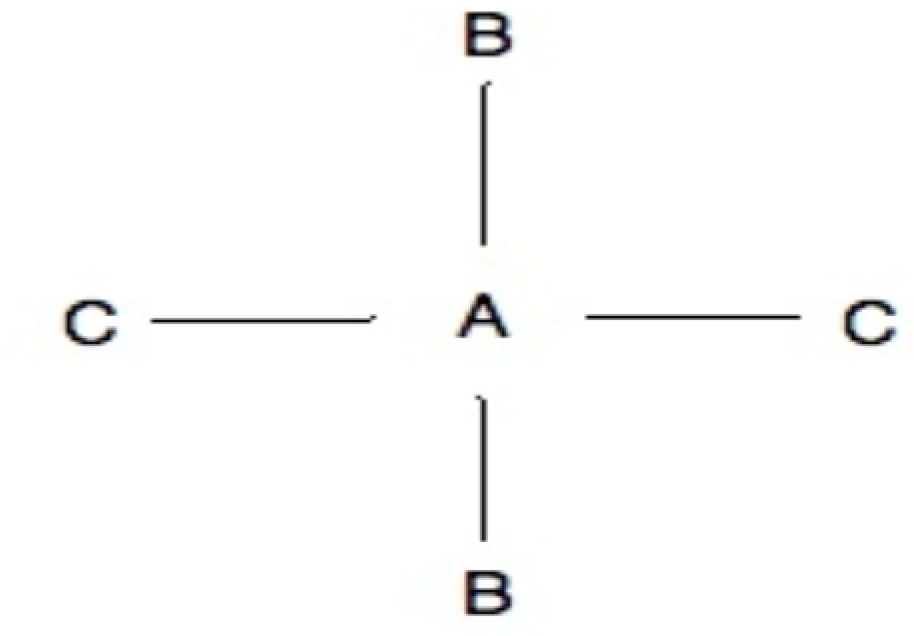
Isomerism in monosaccharides

- Isomers are the molecules that have the same molecular formula

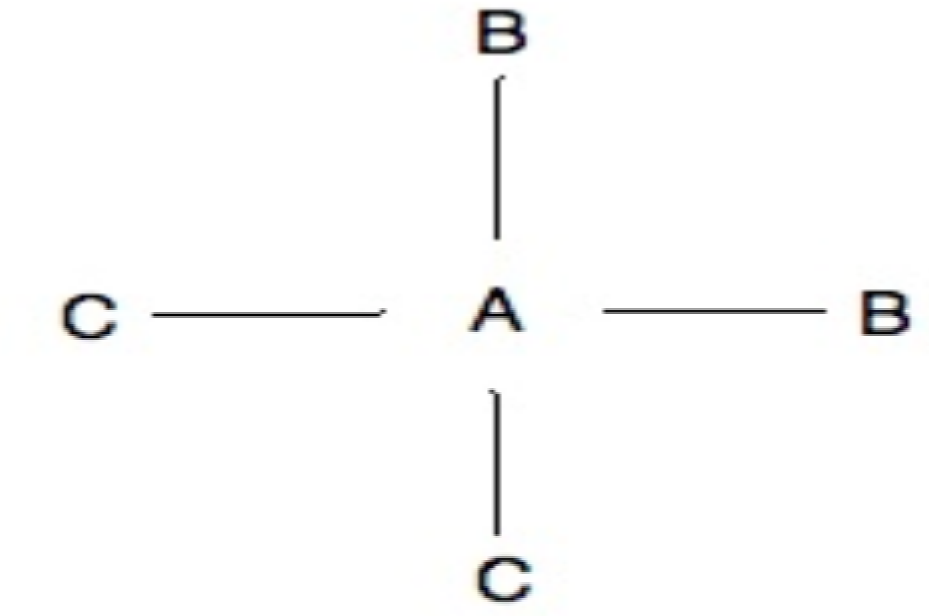
Stereoisomers •

Stereoisomers are molecules that have the same molecular formula, but have a different arrangement of the atoms in space. (different structures).
• For example, a molecule with the formula AB_2C_2 , has two ways it can be drawn:

Isomer 1



Isomer2



Examples of isomers:

1. Glucose

2. Fructose

3. Galactose

4. Mannose

Same chemical formula $C_6 H_{12} O_6$

EPIMERS

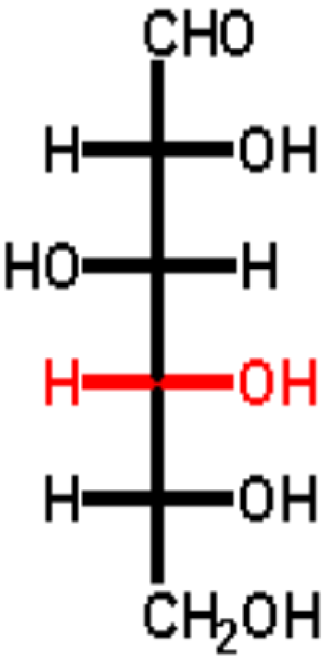
- EPIMERS are sugars that differ in configuration at ONLY 1 specific POSITION, other than the carbonyl carbon.

- Examples of epimers :

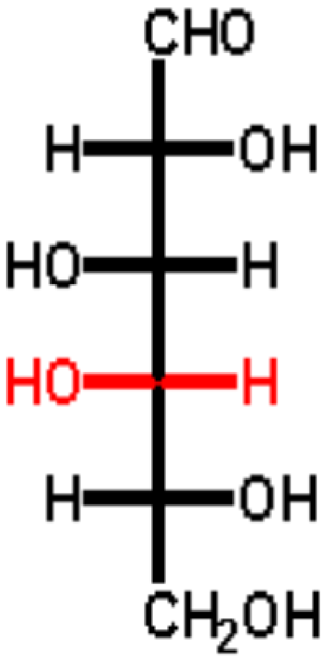
- D-glucose & D-galactose (epimeric at C_4)

- D-glucose & D-mannose (epimeric at C_2)

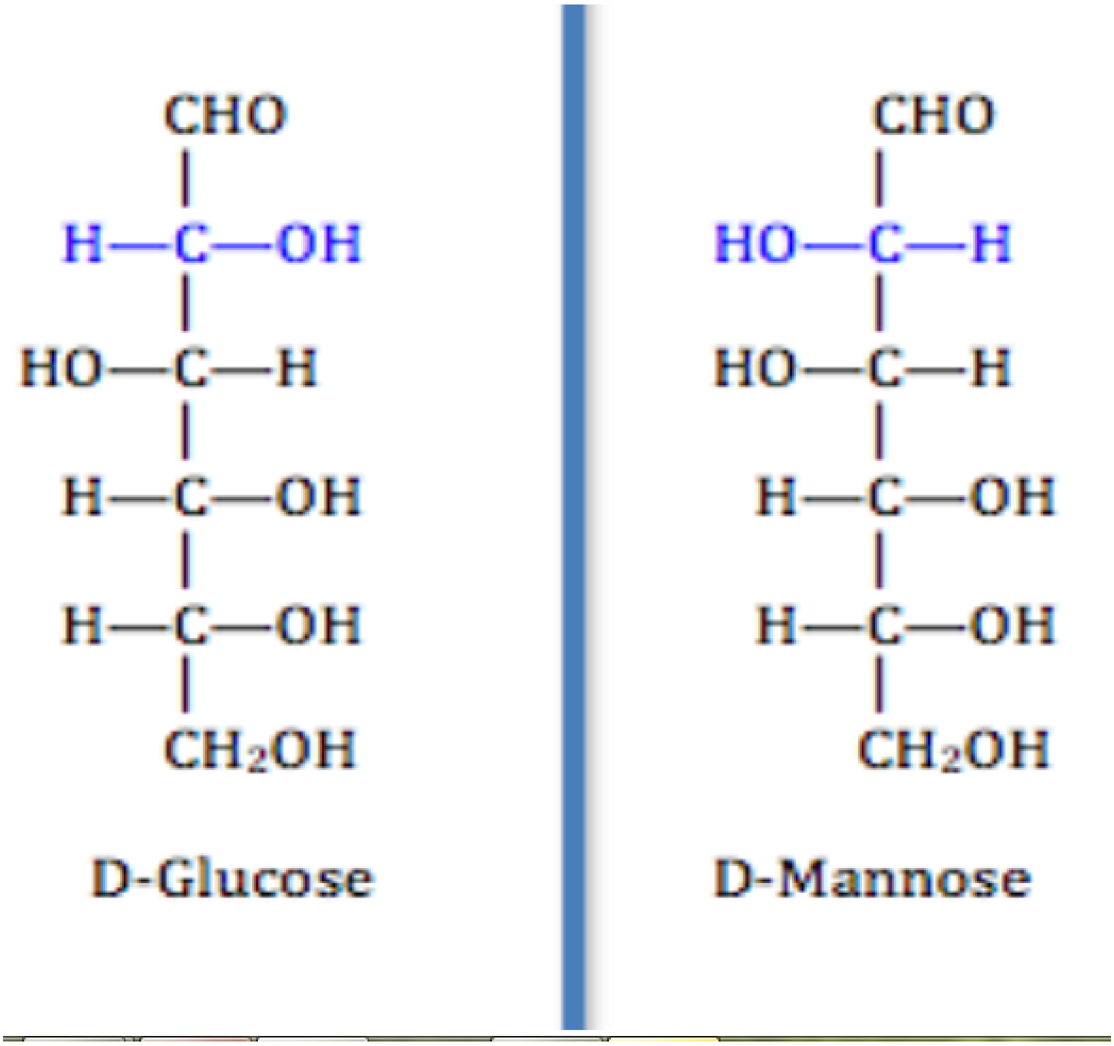
- D-idose & L-glucose (epimeric at C_5)



D-glucose

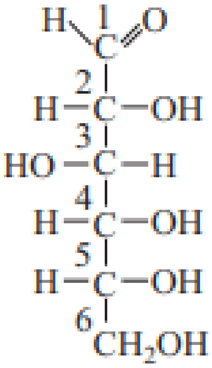


D-galactose

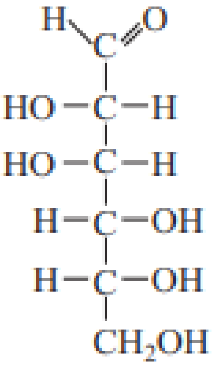


D-Glucose

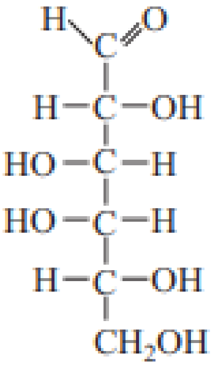
D-Mannose



D-Glucose



D-Mannose



D-Galactose

ENANTIOMERS

Non-Superimposable COMPLETE mirror image (differ in configuration at EVERY CHIRAL CENTER.

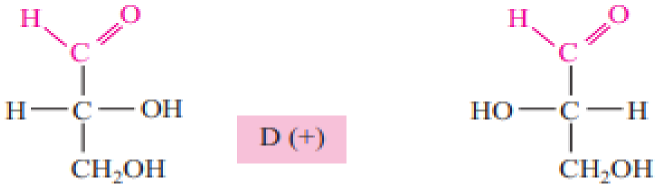
- The two members of the pair are designated as D and L forms.

- This designation is based on the presence of OH group on the penultimate asymmetric carbon.

- In D form the OH group on the asymmetric carbon is on the right.

- In L form the OH group is on the left side.

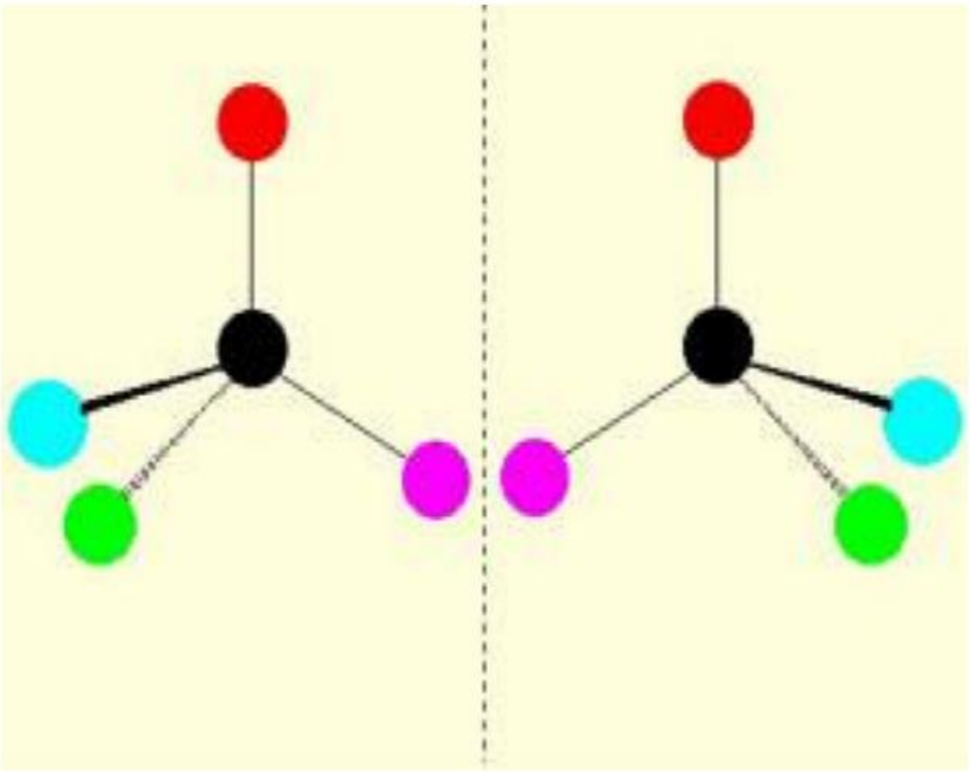
- D-glucose and L-glucose are enantiomers.

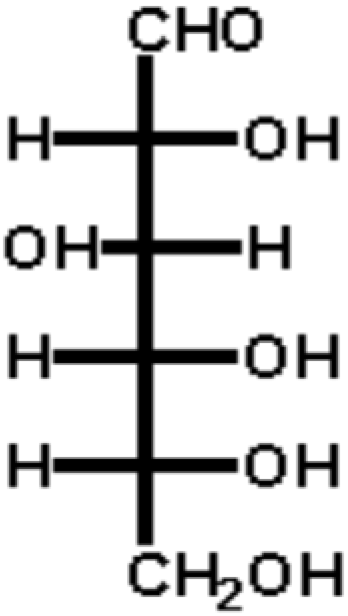


D (+)

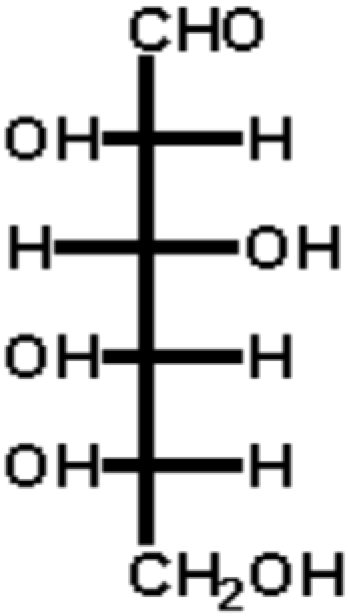
L (–)

glyceraldehyde





D-glucose



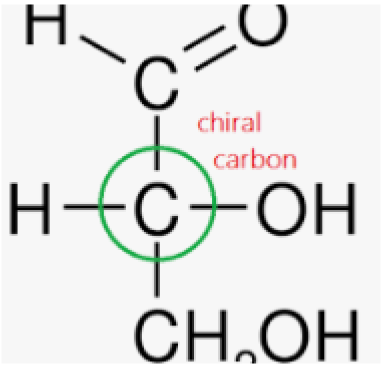
L-glucose

- D and L monosaccharides are metabolized by specific enzymes.

- In mammals most are D monosaccharides.

Asymmetric carbon •

A carbon linked to four different atoms or groups farthest from the carbonyl carbon. • Also called Chiral carbon.



cyclization

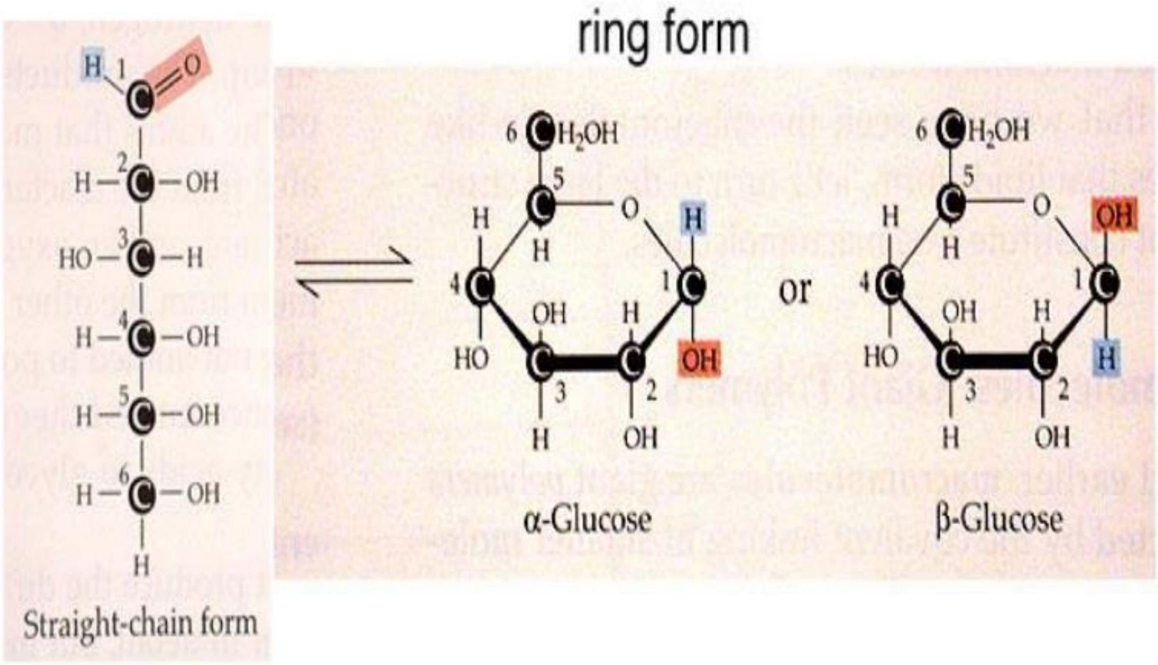
• Less then 1%of CHO exist in an open chain form

. • Predominantly found in ring form.

• involving reaction of C-5 OH group with the C-1 aldehyde group or C-2 of keto group.

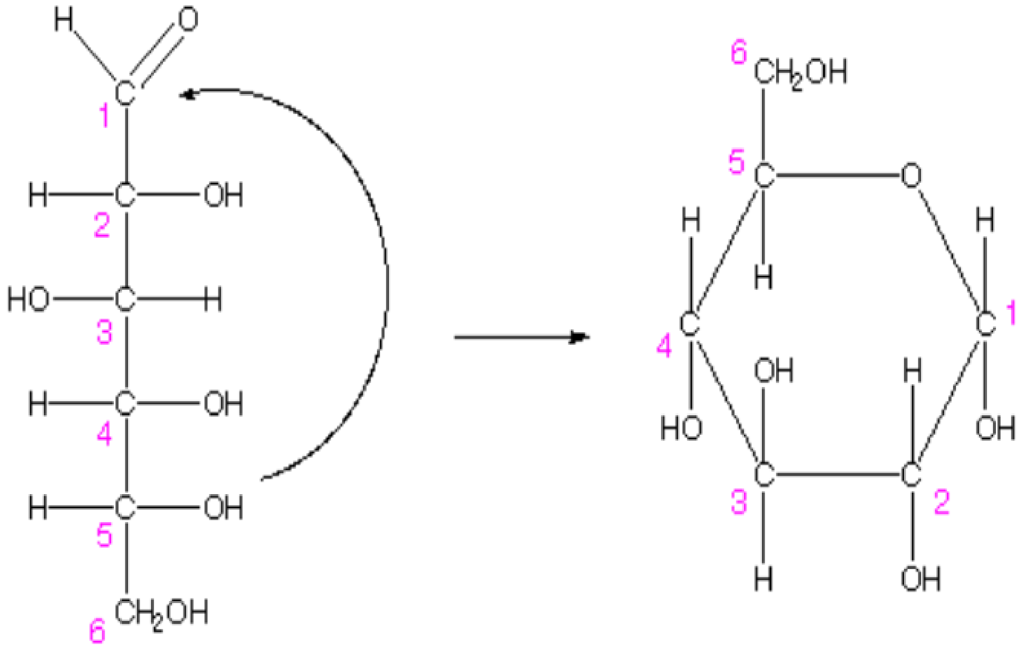
• Six membered ring structures are called Pyranoses

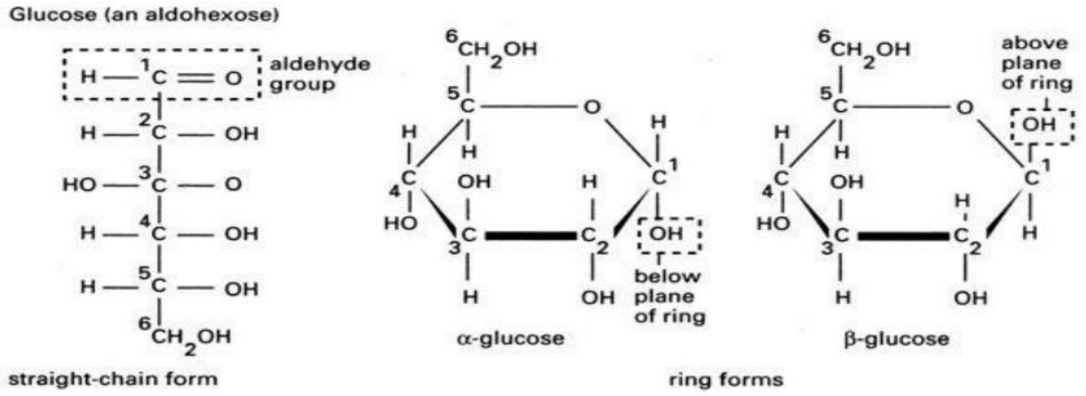
. • five membered ring structures are called Furanoses .

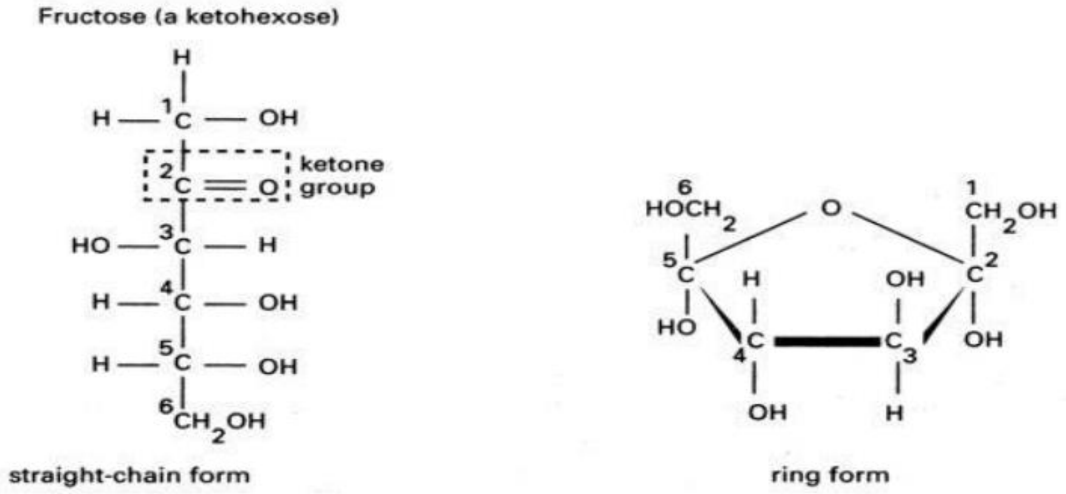


Anomeric carbon

• The carbonyl carbon after cyclization becomes the **anomeric carbon**. • This creates α and β configuration.





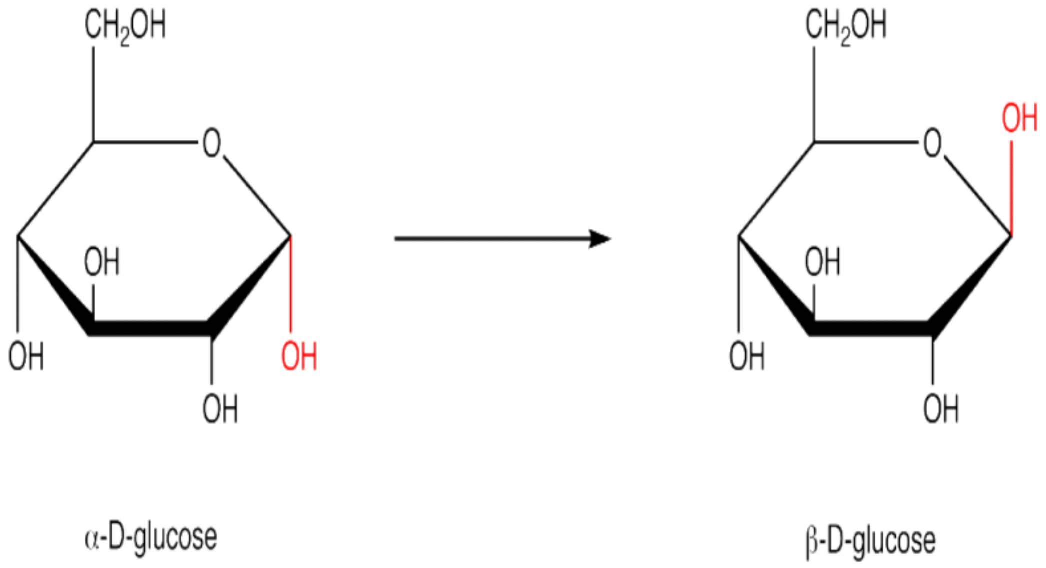


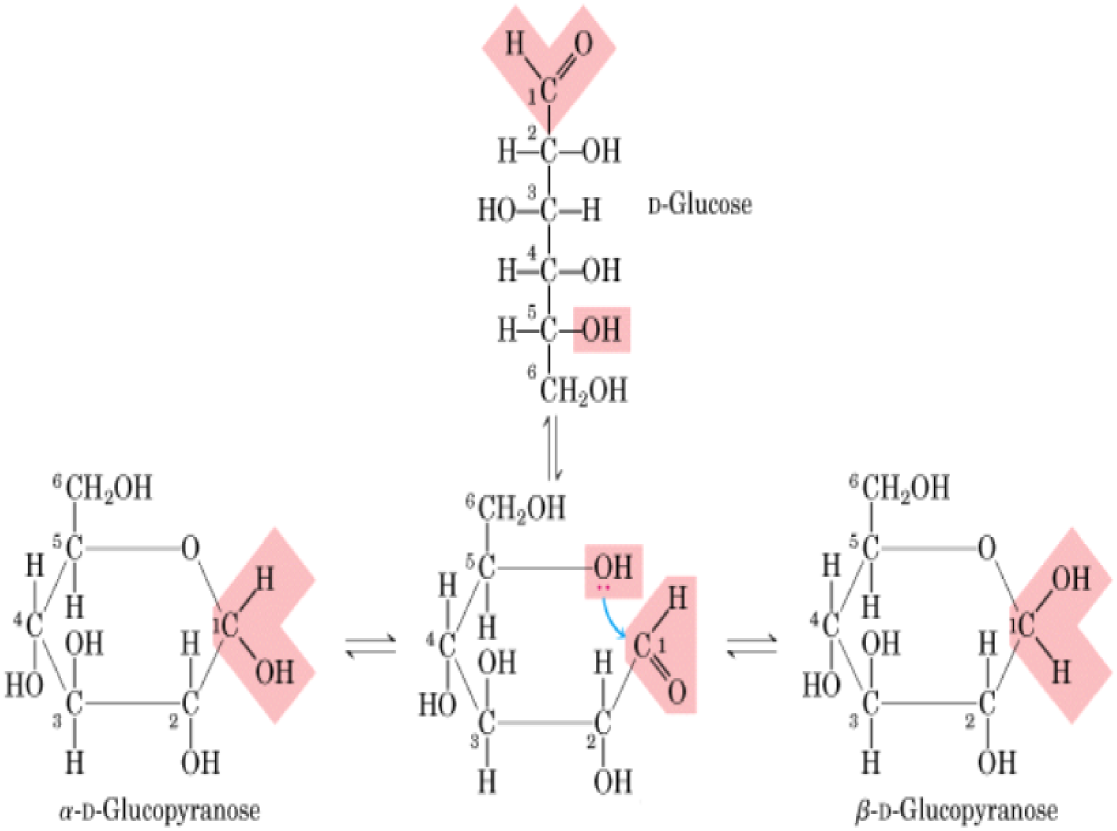
Such α and β configuration are called diastereomers and they are not mirror images.

Enzymes can distinguished between these two forms: • Glycogen is synthesized from α-D glucopyranose • Cellulose is synthesized from β -D glucopyranose

MUTAROTATION

• Unlike the other stereoisomeric forms, α and β anomers spontaneously interconvert in solution. • This is called mutarotation.





Optical Activity

• When a plane polarized light is passed through a solution containing monosaccharides the light will either be rotated towards right or left. • This rotation is because of the presence of asymmetric carbon atom.

• If it is rotated towards left- levorotatory (-)

• If it is rotated towards right- dextrorotatory(+)

Vant Hoff's Law:

• Aldohexoses have 4 asymmetric carbon atoms i.e. from 2-4 .

• This law states that the number of possible optical isomers is equal to 2^n where n is the number of asymmetric carbon atoms. • $2^4 = 16$ possible isomers

The reason that this is so important, however, is not because these particular compounds can rotate light. Instead, it has to do with the way that biological reactions take place. Biological reactions generally are catalyzed by enzymes and the enzymes are set up to specifically handle a particular orientation of atoms within a molecule.

An enzyme that would be able to take D-glucose, and hydrolyze it, or oxidize it, or do something else to it would not be able to carry out that same reaction on the mirror image of D-glucose

Plane polarized light will neither be rotated to left or right when;

1. No asymmetric carbon atom is present.

2. Equal amounts of dextro and levorotatory isomers are present (racemic mixture).

3. Meso compounds – they have asymmetric carbon atom but do not rotate light. Due to internal compensation. (Each molecule two halves, rotation of light is equal by both halves, so no rotation).

Reducing sugar

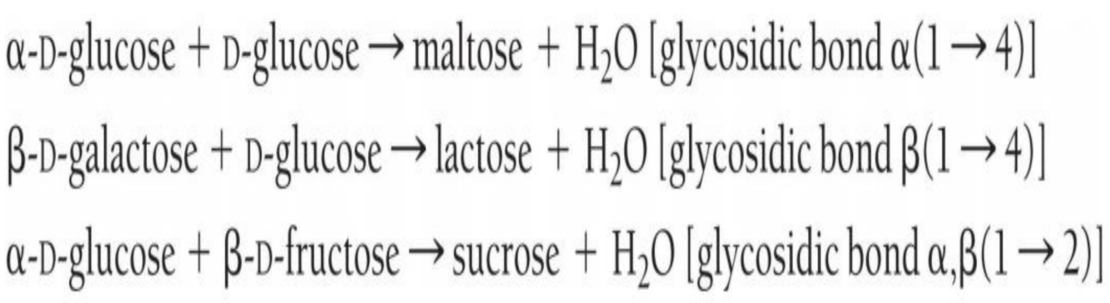
• Sugars in which the oxygen of the anomeric carbon is free and not attached to any other structure, such sugars can act as reducing agents and are called reducing sugars. • Sucrose is a non reducing sugar

Condensation and Hydrolysis—Forming and Breaking Glycosidic Bonds • The –OH group that is most reactive in a monosaccharide is the one on the anomeric carbon.

• When this hydroxyl group reacts with another hydroxyl group on another monosaccharide a glycosidic bond is formed.

Three Important Disaccharides—Maltose, Lactose, and Sucrose The formation of these three common disaccharides is outlined below

Disaccharides



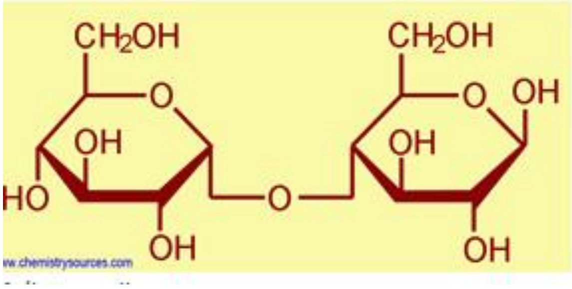
α-D-glucose + D-glucose $\rightarrow$ maltose + H_2O [glycosidic bond $\alpha(1 \rightarrow 4)$]

β-D-galactose + D-glucose $\rightarrow$ lactose + H_2O [glycosidic bond $\beta(1 \rightarrow 4)$]

α-D-glucose + β-D-fructose $\rightarrow$ sucrose + H_2O [glycosidic bond $\alpha,\beta(1 \rightarrow 2)$]

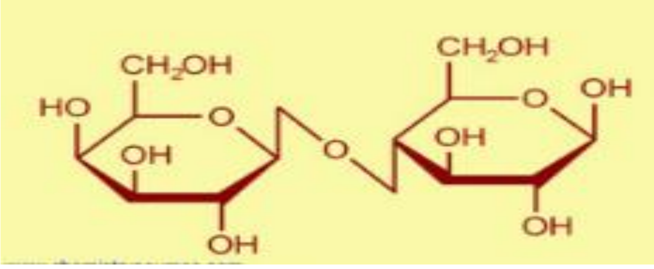
Maltose

- Maltose is known as malt sugar. • It is formed by the breakdown of starch. • One of the anomeric carbons is free, so maltose is a reducing sugar. (α 1→4 glycoside bond)



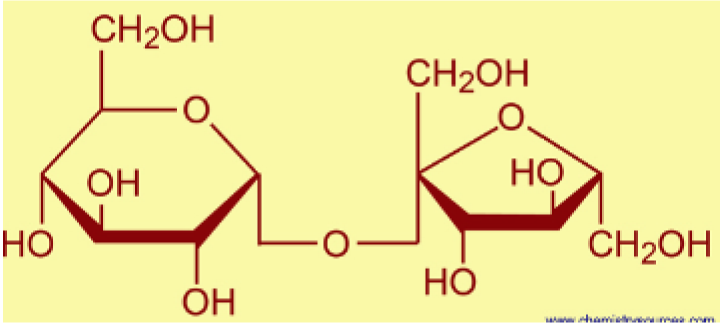
Lactose

- Lactose is known as milk sugar.
- It is found in milk and milk products.
- An intolerance to lactose can occur in people who inherit or lose the ability to produce the enzyme lactase that hydrolyzes lactose into its monosaccharide units. .(β 1→4 glycoside bond)
- The glycosidic bond is
- One of the anomeric carbons is free, so lactose is a reducing sugar.



Sucrose

- Sucrose is known as table sugar.
- It is the most abundant disaccharide found in nature.
- Sucrose is found in sugar cane and sugar beets. ***(α1→2 glycoside bond)**,*
- The glycosidic bond is
- Both anomeric carbons of the monosaccharides in sucrose are bonded, therefore, sucrose is not a reducing sugar.



Polysaccharides

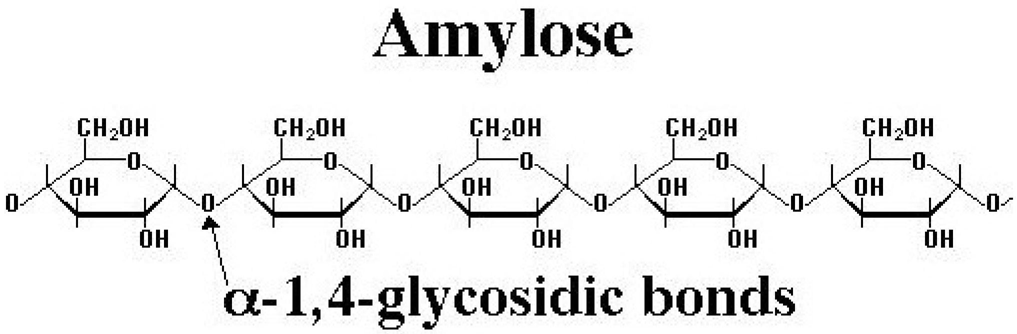
Polysaccharides are polymerized products of many monosaccharide units. They may be homo or hetero polysaccharides. Many polysaccharides, unlike sugars, are insoluble in water. Dietary fiber includes polysaccharides and oligosaccharides that are resistant to digestion and absorption in the human small intestine but which are completely or partially fermented by microorganisms in the large intestine.

1-Homopolysaccharides

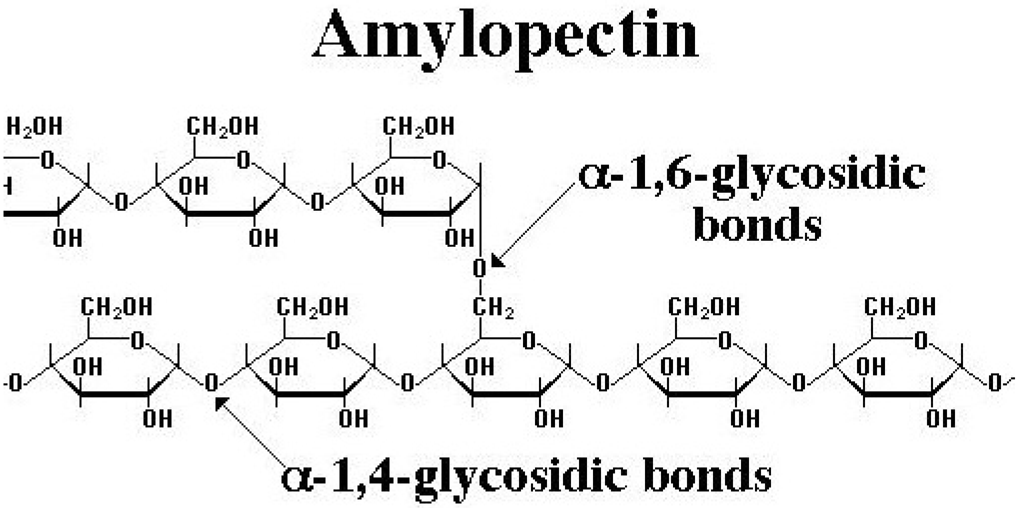
They have only one type of monosaccharide units such as α-D-Glucose and β-D-Glucose . Starch is the major form of stored carbohydrate in plants.

Starch

Starch is composed of a mixture of two substances:**amylose**, an essentially linear polysaccharide, and **amylopectin**, a highly branched polysaccharide. Both forms of starch are polymers of α-D-Glucose. Natural starches contain 10-20% amylose and 80-90% amylopectin. Amylose forms a colloidal dispersion in hot water (which helps to thicken gravies) whereas amylopectin is completely insoluble. ***Amylose*** molecules consist typically of 200 to 20,000 glucose units which form a helix as a result of the bond angles between the glucose units.the bonds type (α 1→4 glycoside bond)



Amylopectin differs from amylose in being highly branched. Short side chains of about 30 glucose units are attached with α1→6 linkages(α1→6 glycoside bond) approximately every twenty to thirty glucose units along the chain. Amylopectin molecules may contain up to two million glucose units.



Structure Of Amylopectin

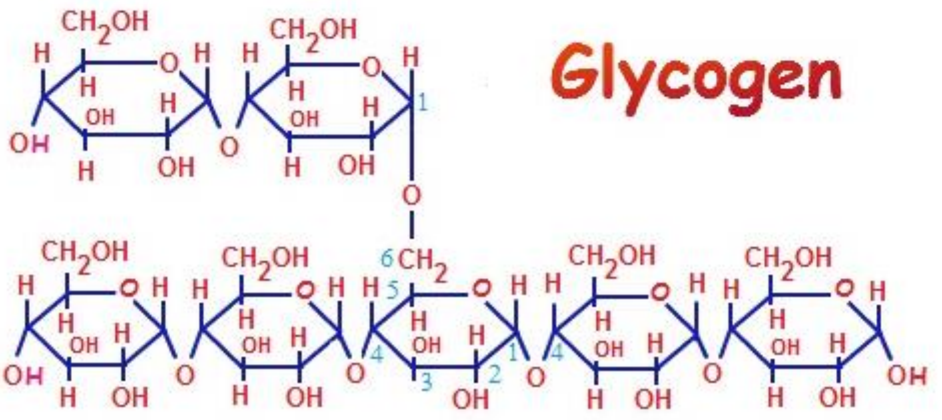
The side branching chains are clustered together within the amylopectin molecule

During fruit ripening, starch undergoes hydrolysis of the α(1→4) bonds to produce glucose and maltose, which are sweet.

- When we consume starch, our digestive system breaks it down into glucose units for use by our bodies.

Glycogen

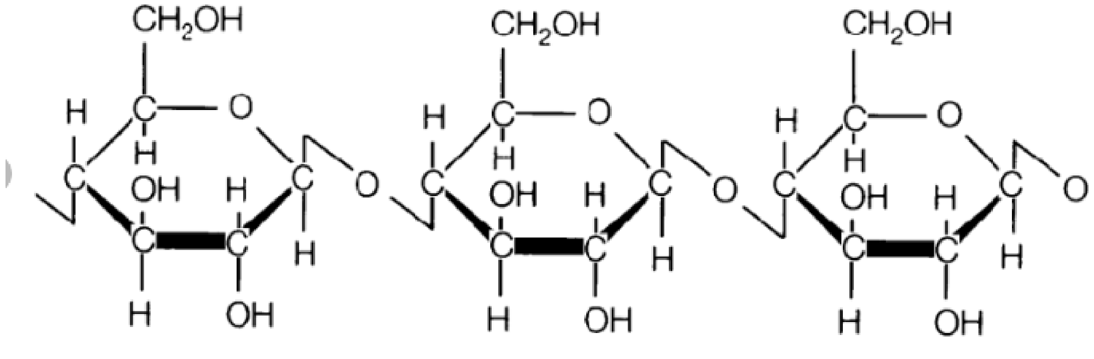
Glucose is stored as glycogen in animal tissues by the process of **glycogenesis**. When glucose cannot be stored as glycogen or used immediately for energy, it is converted to fat. Glycogen is a polymer of α-D-Glucose identical to amylopectin, but the branches in glycogen tend to be shorter (about 12 glucose units) and more frequent. The glucose chains are organized globularly like branches of a tree originating from a pair of molecules of **glycogenin**, a protein with a molecular weight of 38,000 that acts as a primer at the core of the structure. Glycogen is easily converted back to glucose to provide energy



Cellulose

Cellulose

Cellulose is a polymer of β-D-Glucose by (β1→4 glycoside bond), which in contrast to starch, is oriented with $–CH_2OH$ groups alternating above and below the plane of the cellulose molecule thus producing long, unbranched chains. The absence of side chains allows cellulose molecules to lie close together and form rigid structures. Cellulose is the major structural material of plants. Wood is largely cellulose, and cotton is almost pure cellulose. Cellulose can be hydrolyzed to its constituent glucose units by microorganisms that inhabit the digestive tract of termites and ruminants.but human cannot digest cellulose because the human body has not enzyme can hydrolyse (β1→4 glycoside bond)



2-HETEROPOLYSACCHARIDES

Heparin

- Heparin is a medically important polysaccharide because it prevents clotting in the bloodstream.
- It is a highly ionic polysaccharide of repeating disaccharides units of an oxidized monosaccharide and D-glucosamine. Heparin also contains sulfate groups that are negatively charged.
- It belongs to a group of polysaccharides called glycosaminoglycans.

Solutions

A solution is a homogeneous mixture created by dissolving one or more solutes in a solvent.

The component present in excess(larger amount) is called **Solvent**.

The other component(smaller amount) is called **Solute**.

Solutions that contain liquids as solvents are the types of solutions most familiar to use. They have the following properties:

1- Consist of a soluble material (the solute) dissolved in a liquid (the solvent).

2- Are clear

3- Are homogenous.

4- Don't settle.

5- May be separated by physical means.

6- Pass through filter papers.

Concentration of solution

Expressed as a ratio of the amount of solute to the total amount of solution

1-Weight/Weight percent (W/W %)

To specify the concentration of a solute in a solution.

$$W/W\ \% = \frac{\text{Weight of solute (g)}}{\text{weight of solute (g) + weight of solvent (g)}} \times 100$$

2-Volume/Volume percent (V/V %)

A convenient way of expressing the concentration of a liquid solute dissolved in a liquid.

$$V/V\ \% = \frac{\text{Volume of solute (ml)}}{\text{Total volume of solution (ml)}} \times 100$$

3-Weight/Volume percent (W/V)

Widely used method to expressing concentration is a combination of weight and volume.

$$W/V\ \% = \frac{\text{Weight of solute (g)}}{\text{Total volume of solution (ml)}} \times 100$$

4-Parts per million (p.p.m)

To express of small concentrations.

$$\text{Milligram per 100 ml} = \frac{\text{Weight of solute (mg)}}{\text{(100) ml of solution}}$$

Parts per million(ppm) and Parts per billion(ppb)

ppm=µg solute/ ml solution = mg solute / L solution

ppb=ng solute/ ml solution = µg solute / L solution

1L=1000ml =10dl , dl=100ml , µl=10^{-6} L

Chemical Concentrations

1-Molarity (M):Is number of moles of solute per liter of solution

$$M = \frac{\text{Number of moles of solute}}{\text{Number of liter of solution}}$$

$$M = \frac{\text{Mole}}{\text{L}} \quad , \quad \text{mole} = \frac{\text{weight}}{\text{molecular weight}} = \frac{\text{wt.}}{\text{M. Wt.}}$$

$$M = \frac{\frac{\text{Wt.}}{\text{M. Wt.}}}{\text{L}}$$

$$M = \frac{\text{Wt.}}{\text{M. Wt x L}}$$

$$M = \frac{\text{Wt.}}{\text{M.Wt.}} \times \frac{1000}{\text{V (ml.)}}$$

Molecular weight =Gram-molecular weight=1 mole of compound

2-Normality(N):

$$\text{normality} = N = \frac{\text{number of equivalents of solute}}{\text{1 liter of solution}} = \frac{\text{equivalents}}{\text{liter}}$$

where

$$\text{number of equivalents of solute} = \frac{\text{grams of solute}}{\text{equivalent weight of solute}}$$

then

$$N = \frac{\text{grams of solute}}{\text{eq wt solute} \times \text{L solution}} = \frac{\text{grams}}{\text{eq wt} \times \text{L}}$$

$$\text{Equivalent weight} = \frac{\text{Molecular weight}}{\text{Valence}}$$

$$N = \frac{\text{Wt.}}{\text{Eq. Wt.}} \times \frac{1000}{\text{V (ml)}}$$

Valance (n):number of hydrogen ions in acid or number of hydroxide ions in base or number of electrons transport in reaction.

$n=3H^{+}=3$

$$H_3PO_4 \longrightarrow 3H^{+} + PO_4^{-3}$$

$n=3H^{+}=3$

$$NaOH \longrightarrow Na^{+} + OH^{-}$$

$n=1OH^{-}=1$

$$Na_{(s)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H^{\circ}_{(g)}$$
$$Ca_{(s)} + 2\ HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + 2\ H^{\circ}_{(g)}$$
$$Al_{(s)} + 3\ HCl_{(aq)} \longrightarrow AlCl_{3(aq)} + 3\ H^{\circ}_{(g)}$$

$$\text{eq wt} = \frac{\text{at wt Na}}{1} = \frac{\text{at wt Ca}}{2} = \frac{\text{at wt Al}}{3} = \frac{\text{at wt H}}{1}$$

Dilution

Preparation of diluted acid from concentrated acid by using **dilution low**

$$\mathbf{M1 \times V1 = M2 \times V2}$$

Concentrated acid diluted acid

Solubility

Defined as the amount of solute that dissolved in given quantity of solvent to form a saturated solution.

We call the substances being dissolved the **solute,** whereas the **solvent** is the substances in which the solute being dissolved. The **aqueous** solutions, which are the solutions in which water is the solvent.

Saturated Solution

Defined as a solution in which no more solute will dissolve in the solvent at a temperature.

We may say that a solution is **dilute** if there are only a new solute particles dissolved in it, or **concentrated** if there are many solute particles dissolved.

Factors affecting the solubility of a solute

1-The nature of solvent and solute (like dissolve like) polar solvent dissolve polar solute.

2-Tempreture, for most solids increase the ~~solubility~~ will increase the solubility except $Ca(OH)_2$, $CaCr_2O_7$, $CaSo_4$, ~~the solubility decrease with increase the temperature~~ and gas in liquid the solubility decrease with increase the temperature.

3-The pressure, gas in liquid, the solubility will increase with the increase the pressure.

General Rule

"Like dissolves like"

*Polar solvent (water) is a good solvent for ionic compounds (NaCl).

*Gasoline (non polar compounds is a good solvents for other non polar organic compounds (oil). Some substances, like water and alcohol, can be mixed together and create a homogenous phase in any proportion. A solubility measure cannot be applied to such two substances. Such substances are called miscible. On the other hand if two substances cannot be mixed together (like water and oil), they are called immiscible.

Types of Solutions according to Physical State:

1- Solid solute in solid solvent (Ag in Ni or Cu in Au).

2- Solid solute in liquid solvent (NaCL in water or sugar in water).

3- Liquid solute in liquid solvent (ethanol in water).

4- Gas solute in liquid solvent (carbon dioxid CO_2 in water).

5- Gas solute in gas solvent (oxygen in acetylene).

Degree of saturation

1-Saturated solution

A **solution in which the maximum amount of solute has been dissolved**. Any more solute added will sit as crystals on the bottom of the container

2-Unsaturated Solution

solution (with less solute than the saturated solution) **that completely dissolves**, leaving no remaining substances

3-Supersaturated

A **solution that contains more than the maximum amount of solute that is capable of being dissolved at a given temperature**.

Types of Solutions according to tonicity:

1-Isotonic Solutions:

The normal concentration of NaCL in blood is 0.9% and this usually called normal saline solution.

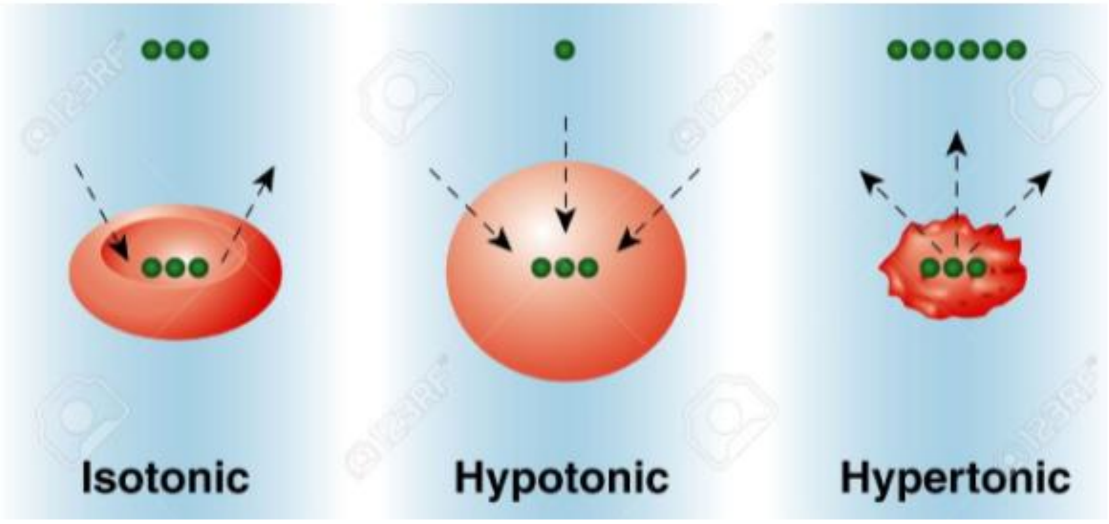
No effect on the red blood cell (RBC). The cell membrane behaves as semipermeable membrane, and since the concentration of NaCL inside and outside the RBC is the same then no flow of water will occur.

2- Hypertonic solutions:

That solutions with contain more than 0.9% of NaCL. If RBC put in this solution, then it will shrink since water will flow from the inside to outside the cell.

3- Hypotonic solutions:

That solutions with contain less than 0.9% of NaCL. If RBC put in this solution, then it will swell or burst, since water will flow from the outside to inside the cell.



Electrolytes and non electrolytes

An electrolyte is a substances that undergo either a physical or a chemical change in solution to yield ions that can conduct electricity are called electrolytes. If a substance yields ions in solution, that is, if the compound undergoes 100% dissociation, then the substance is a strong electrolyte.

A non electrolyte Substances that do not yield ions when dissolved in water. that does not conduct electricity when dissolved in water. For example: sugar molecular ($C_6H_{12}O_6$).Aqueous solution that conducts electricity is called an electrolytic solution. One that does not is called a non electrolytic solution.

Classification of electrolyte

Strong electrolyte	weak electrolyte	non electrolyte
HCl	CH_3COOH	$(NH_2)_2CO$ Urea
HNO_3	HF	CH_3OH
$HClO_4$	HNO_2	Glucose
H_2SO_4	NH_3	Sucrose

NaOH	H_2O	Nitrogen
$Ba(OH)_2$		Oxygen
NaCl		Carbon dioxide
$CaCl_2$		

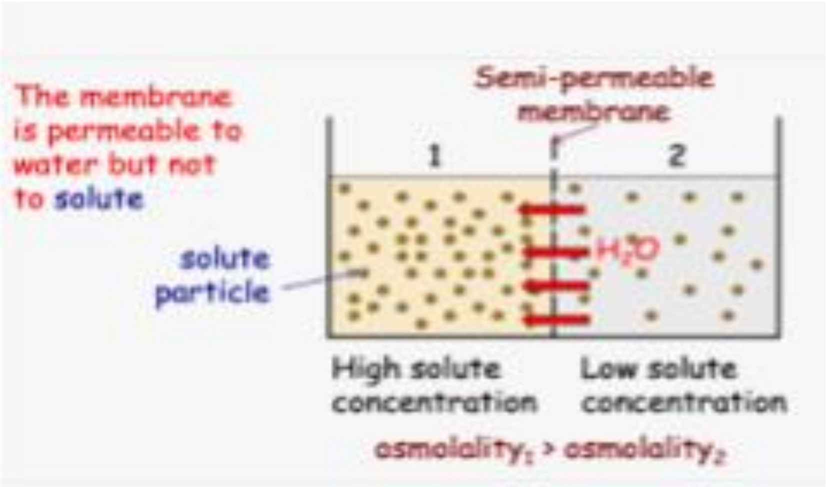
The mode of electrolytic and non-electrolytic solutions has been used to explain all the physical properties of solutions (osmosis and dialysis).

Osmosis and osmotic pressure

Cells have limiting boundary membranes that are called plasma membranes. These allow the exchange of materials back and forth between the interior of the cell and its exterior materials. Osmosis and dialysis are two ways that such an exchange of materials occurs.

Osmosis is the movement of water though an osmotic membrane from an aqueous solution that is less concentrated to one that is more concentrated.

If we have two sugar solutions one has a strong concentration and the other has a weak concentration and the two solutions separated by semipermeable membrane, as shown:



The two solutions try to be the same in concentration.

Preventing osmosis from occurring applying pressure is needed to stop water movement from place to another (osmotic pressure). The greater the number of particles, whether ions or molecules, in a solution, the greater it's osmotic pressure.

Suspensions

When some powdered clay is placed in water and strongly shaken, a suspension of clay in water will be produced.

This suspension will not be clear opaque upon standing, the clay will slowly settle.

When the suspension is poured into a funnel containing a piece of filter paper, only water passes through the filter paper, the clay doesn't of same medications, such as milk of magnesia, are administered as a suspension. Many bottles of medications state on the label shake before using. A mist is a suspension of liquid in a gas, water droplets suspension in air are one example of a mist.

Colloids and colloidal Dispersions

The particles in a solution are the size of atoms and molecules (0.05-0.25 nm). Sometimes intermolecular attractions between molecules cause several hundred or thousand of them to cluster together. The size of these clusters range from 1-100 nm. Matter containing particles of this size is called a "Colloid".A uniform dispersion of a colloid in water is called a "Colloidal-dispersion". A colloidal dispersion usually appears cloudy. The colloid in a colloidal dispersion is called the dispersed substance. Proteins form colloidal dispersion in water being colloidal so they can pass through a filter paper but no in membrane. Proteins present in blood stream cannot pass through the cell membranes and should remain in blood stream. Since the present of protein in the urine indicates damage to the membranes in the kidneys.

Solution	Suspension	Colloid
Pass through a filter paper and membrane	Don't pass through filter paper and membrane	Pass through a filter paper but not through membrane

Dialysis and Living Systems

Cell membranes that allow small molecules and ions to pass while holding back large molecules and colloidal particles are called "Dialyzing membranes". The selective passage of small molecules and ions in either direction by a dialyzing membrane is called "Dialysis". "Dialysis" differs from "Osmosis" in that osmotic membranes allow only solvent molecules to pass.

The kidneys are an example of organ in the body that uses dialysis to maintain the solute and electrolytes balance of the blood.

The main purpose of the kidneys is to clean the blood by removing the waste products of metabolism and control the concentrations of electrolytes, by allow blood cells and plasma proteins to pass. A180L of blood are purified daily in adult (68Kg). 1 percent is eliminated as urine.

Lipids are defined by their physical behavior rather than their chemical structure. Lipids are classified according to their (Solubility). Lipids are readily soluble in non-polar organic solvents such as diethyl ether, chloroform or benzene but are insoluble in water. A major portion of lipids structure is like a hydrocarbon. This is the reason that lipids are soluble in nonpolar solvents.

Important functions of lipids:-

1- Lipids serve as the main energy reserve for living system.

2- Form parts of cell members.

3- Regulate the activities of cell and tissues.

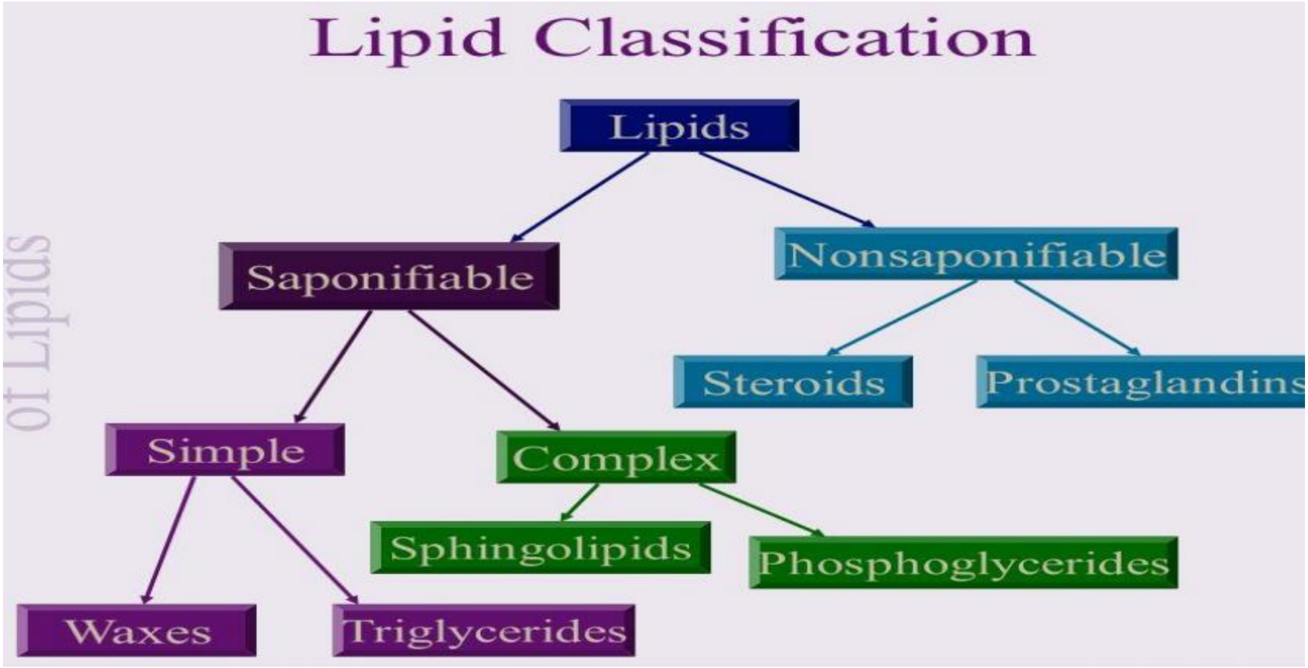
4- Provide insulation against changes in external temperature (subcutaneous fat)

5- Give shape and contour to the body

6- Help in absorption of fat soluble vitamins (A, D, E & K)

Classification of Lipids

Lipids can be classified according to whether or not they undergo saponification. Lipids that undergo alkaline hydrolysis are called (Saponifiable lipids). Lipids that do not undergo such a reaction are called (non-saponifiable lipids)



Fatty Acids

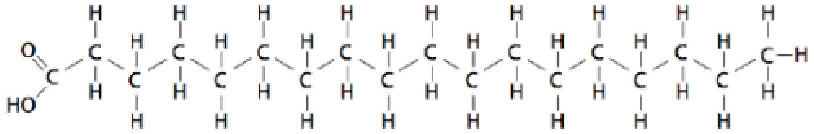
Fatty acids are biomolecules containing a (Polar) carboxyl functional group(COOH) connected to unbranched aliphatic chain(nonpolar).

$$CH_3(CH_2)_nCOOH$$

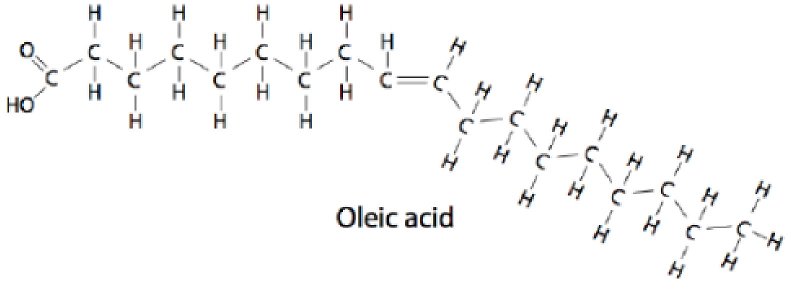
Fatty acids are rarely found in a free form in cells and tissues but are most often bound in fats (Triacyglycerols). The number of carbon atoms in a fatty acid can range from 4 (as found in butter) to as many as 36 (found in the brain). Fatty acids whose hydrocarbons chain contains one or more double bonds are called (unsaturated fatty acids). Those whose hydrocarbon contains no double bonds are called (Saturated fatty acids).

Palmitic acid ($C_{16} H_{32} O_2$) and Stearic acid ($C_{18}H_{36}O_2$) are the most common saturated fatty acids. Oleic acid ($C_{18} H_{34}O_2$) is the most common unsaturated fatty acid. Most of the naturally occurring lipids contain even chain fatty acids while odd numbered fatty acids are seen in microbial cell walls. Linoleic(Omega 6)and Linolenic (Omega 3) acids are polyunsaturated fatty acids and they are called essential fatty acids because they cannot be synthesized by the body and have to be supplied in the diet.

* Double bonds in naturally occurring fatty acids are almost always of the Cis configuration.



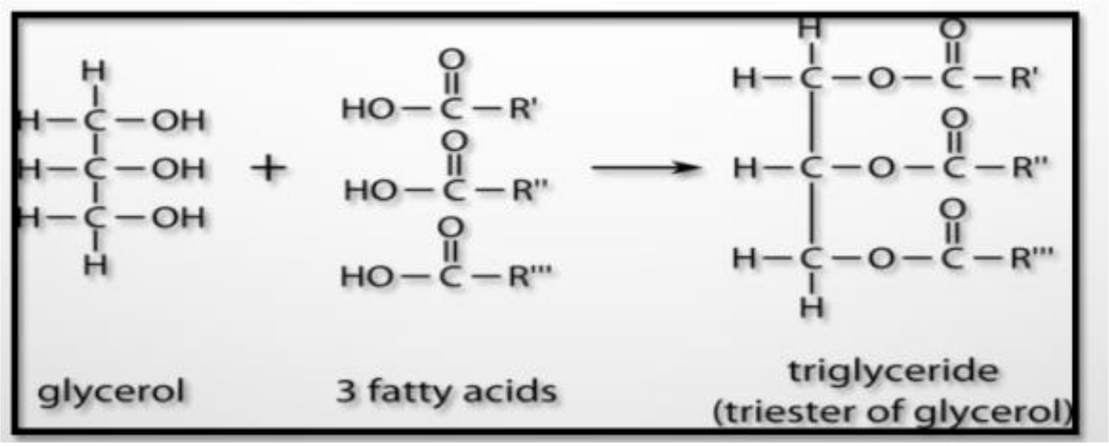
Stearic acid



Oleic acid

Triacyl glycerols

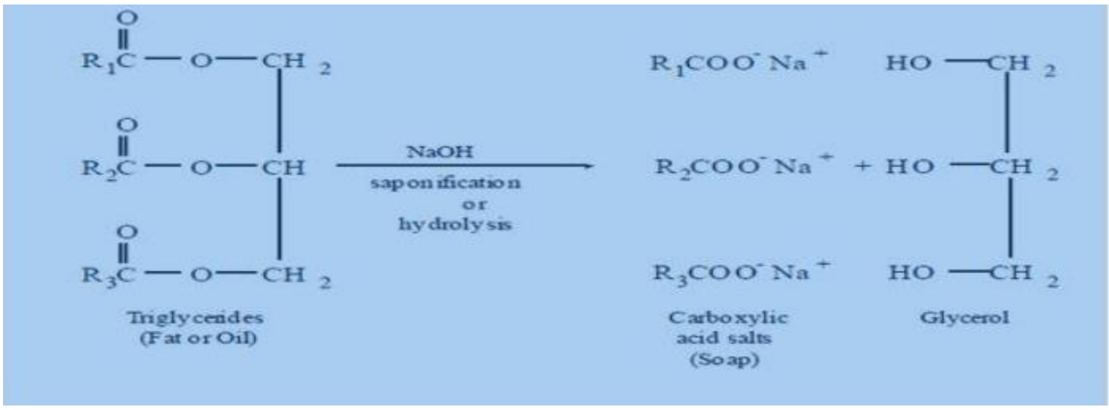
Almost all fatty acids present in nature are found as constituents of the nonpolar lipids called triacylglycerols. The basic foundation molecule of the triacylglycerols is the trihydroxyl compound glycerol. Each hydroxyl group can be linked to a fatty acid by esterification . Thus, triacylglycerols are triesters formed from three fatty acids and glycerol.

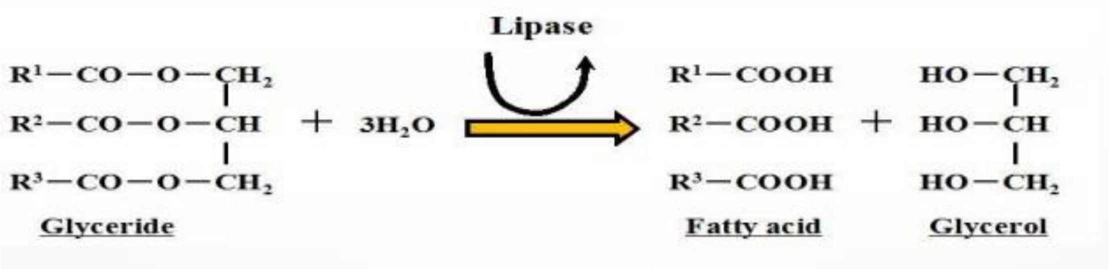


Triacylglycerol that are solid at room temp. are called Fats those that are liquid at room temp. are called Oils.

Generally, plant oils contain more unsaturated fatty acids than animal fats. A natural fat or oil is not a pure compound, but a mixture of triacylglycerols. This is because of the difference in the type of acyl group in the triacylglycerols. Triacylglycerols undergo the usual reactions of esters.

A particulary important reaction is alkaline or enzyme catalyzed hydrolysis.





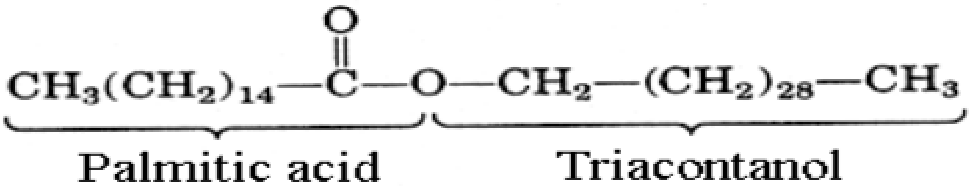
Catalytic hydrogenation of the unsaturated acyl groups of the triacylglecrols in vegetable oils converts them to saturated triacylglycerls. This reaction has been carried out industrially to produce oleomargarine.

Q/ In the industrial process, not all the double bonds are hydrogenated, Why?

The triacylglycerols have two primary biological roles, energy metabolism and insulation.

(Adipocytes) are animal cells specialized for fat storage. Enzymes called (Lipases) are present in adipocytes to catalyze the release of fatty acids. Triacylglycerols are also present below the skin layer of animals especially those living in polar regions, to insulate against temperature extremes. The fatty acids, in addition to being found in triacylglycerols are also present in other derivative forms called waxes, which are nonpolar lipids.

Waxes are esters of fatty acids and long chain monohydroxy alcohols. Bees wax is an ester composed of palmitic acid and the alcohol triacontanol.



$$\underbrace{\mathrm{CH_3(CH_2)_{14}}-\overset{\mathrm{O}}{\overset{\|}{\mathrm{C}}}}_{\text{Palmitic acid}}-\underbrace{\mathrm{O}-\mathrm{CH_2}-\mathrm{(CH_2)_{28}}-\mathrm{CH_3}}_{\text{Triacontanol}}$$

Polar Lipids

Class of lipids with chemical structures similar to the triacylglycerols, but quite different in biological function. Polar lipids are combined with protein molecules for the construction of biological membranes , which provide a protective shield around cells. Two classes of lipids, glycerophospholipids & sphingolipids are introduced in polar lipids.

Glycerophospholipids

The foundation molecule for the glcerophospholipids is 1,2-diacylglycerol 3- phosphate.

The term phosphatidic acid refers to a group of many molecules that differ by the identity of two fatty acids.

- In the glycerophospholipids, a second alcohol is esterified .with the phosphate group. That alcohol is usually one of the following the aminoalcohol; ethanolamine or choline, the amino acids; serine; or the polyhydroxyl compound inositol.

- With the exception of inositol , all of the alcohols have an amino or other functional group that becomes ionic at physiological pH.

- The major structured difference between triacylglycerols and glycerophospholipids is: triacylglycerol s are nonpolar and hydrophobic.

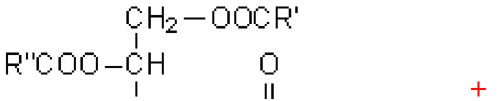
glycerophospholipids are; a polar head and nonpolar tails. These two characteristics are essential for membrane structure.

Sphingolipids The foundation molecule is the 18-carbon alcohol sphingosine.

Sphingosine has two functional groups (amino & hydroxyl)

- The molecule called ceramide has a polar head (hydroxyl group on C1of sphingosine) and two nonpolar tails.

- The sphingomyeline contain a fatty acid on the amino group and a phosphocholine unit esterified with hydroxyl group. This results in a molecule with a polar (Ionic) head and two non polar tails similar to the glycerophospholipids.



phosphatidylcholine

- Carbohydrate containing lipids use the ceramide as foundation molecules. The cerebroside consists of ceramide with a monosaccharide unit in glycosidic linkage at the hydroxyl group at C1.

The carbohydrates commonly found are glucose , galactose and N-acetyl galactosamine.

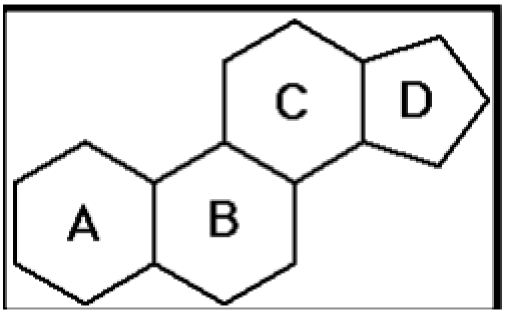
Function of glycosphingolipids

1- play role in biological membrane structure.

2- Involved in other specialized cellular functions including recognition events at cell surfaces, tissue specifity of cell association and the transmission of nerve impulses.

Steroids

The steroids all have the characteristics fused-ring system of three six-membered rings labeled A,B,C and one five-membered ring called the D ring.

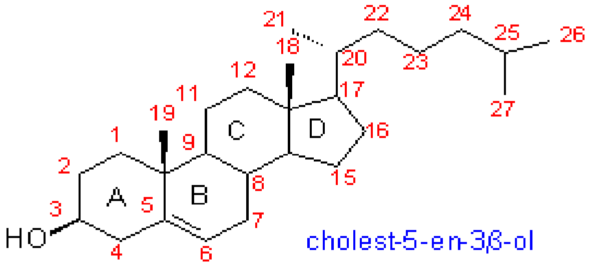


Ketones, alcohols, double bonds & hydrocarbon chains decorate the ring system in various types of steroids. Cholesterol, the best known steroids has a hydroxyl group on the A ring, a double bond in Ring B, and hydrocarbon chains attached at several locations.

- Cholesterol has a polar head (hydroxyl group) and a nonpolar tail (hydrocarbon Skeleton)

- Cholesterol and some of its derivatives accompany the glycerophospholipids and sphingolipids in biological membranes.

- Cholesterol and its ester derivatives are abundant in plasma proteins called (Lipoproteins) whose function is to transport the cholesterol to peripheral tissue or use in construction of membranes and as a biosynthetic precursor for steroid hormones and other biologically active products.



Functions of Cholesterol

1-Cell membranes: cholesterol is a component of membranes and has a modulating effect on the fluid state of the membrane
2-Nerve conduction: cholesterol has an insulating effect on nerve fibers

3-Bile acids and bile salts: are derived from cholesterol. Bile salts are important for fat absorption

4-Steroid hormones: Glucocorticoids, androgens and estrogens are from cholesterol

5-Vitamin D2: is from 7-dehydro-cholesterol

6-Esterification: the OH group of cholesterol is esterified to fatty acids to form cholesterol esters. This esterification occurs in the body by transfer of a poly unsaturated fatty acid moity by Lecithin cholesterol acyl transferase.

Lipoprotein

can be classified according to their density to Chylomicrons, Very low density lipoprotein (VLDL), Low density lipoprotein (LDL) and High density lipoprotein (HDL).

Cholesterol is the starting point for biosynthesis of the steroid hormones(such as Estradiol ,Testosterone and cortisol) and bile acids.

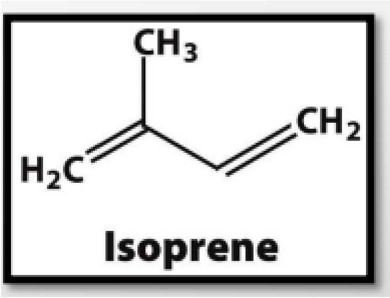
Cholate & its glycine derivative.

- These compounds have acidic functional groups with protons that dissociate at physiological PH to produce ionic structures called (bile salts).

- The bile salts are stored in the gall bladder and secreted into the intestines to help solubilize , digest, and absorb dietary fats.

Terpenes

Terpenes contain carbon carbon double bonds and are made up of two or more isoprene units.



They are designated according to the number of isoprene units in the molecule, if no. of isoprene units is 2 called (Monoterpenes) 4 (Diterpenes), 6 (Triterpenes), 8 (Tetraterpenes).

- Important terpenes in plants and animals include Limonene, ß-carotene, gibberellic acid, squalene and lycopene.

Fat- soluble Vitamins

This class of vitamins can be classified as terpenes and includes vitamins A, D, E and K. Vitamin A, or (retinol) , is a terpenes that also contains a primary hydroxyl group.

- Vitamin A is obtained from cod liver oil or animal liver. - Breaking the central double bond of β-carotene forms the correct carbon structure for two molecules of vitamin A.

Vitamin D have several forms and the most important are vitamin D_2 (ergocalciferol) and Vit D_3 (Cholecalciferol)

Formed from cholesterol present in the skin by ultraviolet radiation and its function to regulate of calcium and phosphorus metabolism.

Vitamin E, which called also α-Tocopherol is aromatic ring with long hydrocarbon chain. The known biological function of vitamin E is antioxidant prevents oxidation damage to cellular membranes.

Vitamin K , is a quinone with side chain made up of several isoprene units. The length of the side chain varies depending on the source of the vitamin. The biological function of vitamin K is to regulate blood clotting.

Tikrit University **College of Dentistry**

Chemistry lab:2 **First stage**

REACTIONS OF ACID RADICALS OR – ANIONS

We are studying some acid radicals such as:-

1- Chloride Cl^-
2- Iodide I^-
3- Phosphate PO_4^{-2}
4- Sulphate SO_4^{-2}
5- Thiocyanate SCN^-
6- Oxalate $C_2O_4^{-2}$

The reactions of these anions depend upon precipitation reaction, and their solubility's, or upon oxidation – reduction reaction as in a certain organic acid such as oxalate.

1-STUDY THE REACTIONS OF CHLORIDE ION Cl^-

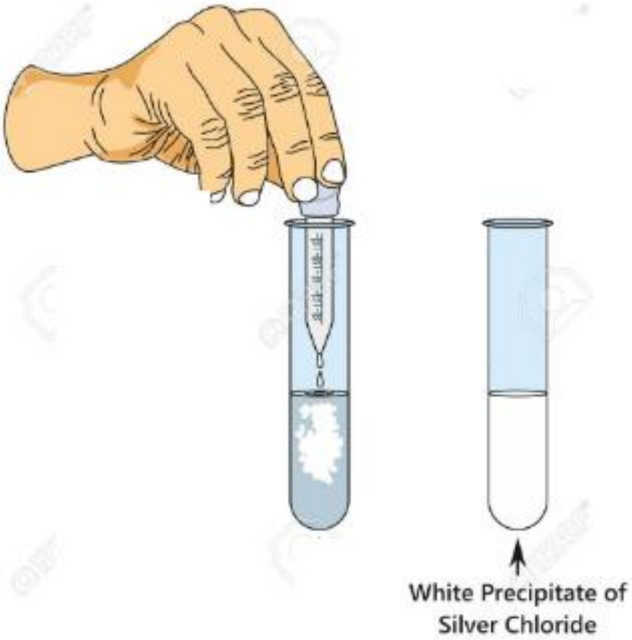
(USE NaCl SOLUTION 5%)

Chloride is constituent of hydrochloric acid (HCl) produced by parietal cells of stomach. It is more in extracellular fluid than intracellular fluid. It is required for proper erythrocyte (RBC) function by chloride shift. It is involved in maintenance of plasma volume. It is activator of some enzymes like amylase, angiotensin converting enzyme etc.

REACTION -1

-With $AgNO_3$ solution 5%

When an aqueous solution of silver nitrate ($AgNO_3$) is (10 drops) added to the aqueous solution of (10 drops) sodium chloride (NaCl), a white precipitate of silver chloride (AgCl) is formed that is indicated by the following chemical reaction.



$$\mathrm{NaCl} + \mathrm{AgNO_3} \longrightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{\mathrm{AgCl}}} + \mathrm{NaNO_3}$$

REACTION -2

-Confirm by:

1- Try the solubility of AgCl in HNO_3 (insoluble).

2- Try the solubility of AgCl in dil NH_4OH (soluble).

In order to confirm that this precipitate is AgCl, 20 drops of NH_4OH is added to the same test tube. The white precipitate dissolves because AgCl salt forms a complex with NH_3.

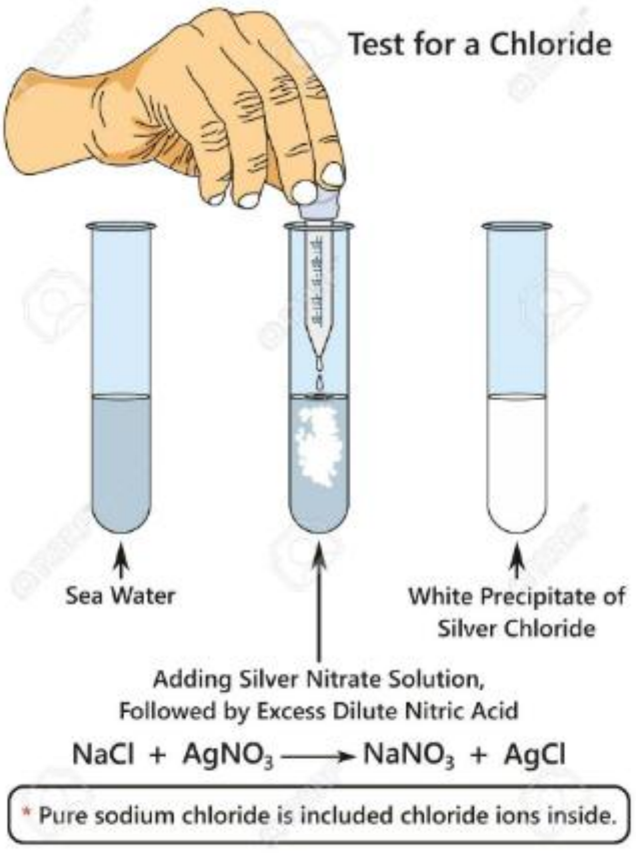
$$AgCl_{(s)} + 2NH_3 \rightarrow [Ag(NH_3)_2]^+ + Cl^-$$

$$AgCl + 2NH_4OH \longrightarrow \underset{\text{soluble complex}}{Ag(NH_3)_2Cl} + 2H_2O$$

Added 15 drops of HNO_3 into the solvated complex and a white precipitate is re-constituted. Because with the addition of HNO_3, the silver diamine complex $[Ag(NH_3)_2]^+$ decomposes and re-precipitates as AgCl.

$$[Ag(NH_3)_2]^+ + Cl^- + 2HNO_3 \rightarrow 2NH_4^+ + 2NO_3^- + AgCl_{(s)}$$

Note:-Sea water and tap water Contains (Cl^-)



2-STUDY THE REACTIONS OF IODIDE ION I^-

(USE KI SOLUTION 5%)

Iodine is constituent of thyroid hormones. Thyroxine (T_4) and triiodothyronine (T_3) synthesis involves iodine utilization. Thyroid hormones are essential for mental development as well as physical development.

REACTION -1

-With $AgNO_3$ solution 5%

Add (10 drops) of $AgNO_3$ solution to (10 drops) of KI solution then you will see yellow precipitate of AgI in the tube which will be insoluble in NH_4OH and HNO_3.

$$KI + AgNO_3 \longrightarrow \underset{\substack{\text{yell. ppt.} \\ \text{Insoluble in} \\ NH_4OH \text{ and } HNO_3}}{AgI\downarrow} + KNO_3$$

REACTION -2

-Confirm by starch test:

Iodides are readily oxides in acid solution to free iodine by a number of oxidizing agents, free iodine identified by the deep blue coloration produced with starch solution.

Add (5 drops) of KI solution to (5 drops) of HNO_3 solution then you will see brown solution of free iodine.

$$KI + HNO_3 \longrightarrow HI + KNO_3$$

3-STUDY THE REACTIONS OF PHOSPHATE ION PO_4^{-2}

(USE Na_2HPO_4 SOLUTION 15%)

Phosphate is present in humans to an extant of about 500-700gm. More of it is present in bone and teeth. In the body phosphate is present in two forms. An inorganic form found in bone and teeth which is complexes with calcium and magnesium. Another is organic form. It is present as organic compounds in the cells and cell membranes. In the bone and teeth it is present as hydroxyl apatite. Phosphate is component of blood buffers.

REACTION -1

-With $AgNO_3$ solution 5%

Add (10 drops) of $AgNO_3$ solution to (10 drops) of Na_2HPO_4 solution then you will see yellow precipitate of Ag_3PO_4 in the tube which will be sparingly soluble in HNO_3.

$$Na_2HPO_4 + 3AgNO_3 \longrightarrow \underset{\text{yell. ppt.}}{\underset{\downarrow}{Ag_3PO_4}} + 2NaNO_3 + HNO_3$$

Try the solubility of the yellow ppt. Ag_3PO_4 in

1- In dilute NH_4OH (soluble).

2- In dilute HNO_3 (soluble).

REACTION -2

-Confirm by $FeCl_3$ test:

Pale buff $FePO_4$, soluble in excess of $FeCl_3$ and in HCl, but insoluble in acetic acid (CH_3COOH).

Add (2drops) of $FeCl_3$ solution to (10 drops) of Na_2HPO_4 solution then you will see pale Buff ppt. $FePO_4$ in the tube.

$$Na_2HPO_4 + FeCl_3 \longrightarrow \underset{\text{pale Buff. ppt.}}{FePO_4} + 2NaCl + HCL$$

4-STUDY THE REACTIONS OF SULPHATE ION $SO_4^{=}$

(USE Na_2SO_4 SOLUTION 10%)

Reaction -1

Add (10 drops) of $AgNO_3$ solution to (10 drops) of Na_2SO_4 solution then you will see white precipitate of Ag_2SO_4 in the tube which will be sparingly soluble in HNO_3.

$$2AgNO_3 + Na_2SO_4 \rightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{Ag_2SO_4}} + 2NaNO_3$$

Reaction -2

-Confirm by reaction with $BaCl_2$ solution

Add (10 drops) of $BaCl_2$ solution to (10 drops) of Na_2SO_4 solution then you will see white precipitate of $BaSO_4$ which be insoluble in dil HCl.

$$Na_2SO_4 + BaCl_2 \rightarrow \underset{\text{wt. ppt.}}{\underset{\downarrow}{BaSO_4}} + 2NaCl$$

To 0.5 mL of sample in a test tube, add 7 drops of HNO_3 and 1 drop of $BaCl_2$, respectively. The white precipitate is $BaSO_4$ and the formation of this precipitate indicates the presence of SO_4^{2-} ion in the sample.

Sulphate ions SO_4^{2-}

Add acidified barium chloride

Barium chloride + potassium sulfate →
Barium sulfate + potassium chloride

$$BaCl_2 + K_2SO_4 \rightarrow BaSO_4 + 2KCl$$

$$Ba^{2+}_{(aq)} + SO_4{}^{2-}_{(aq)} \rightarrow BaSO_{4\,(s)}$$

White precipitate

5-STUDY THE REACTIONS OF THIOCYANATE ION SCN^- (USE KSCN SOLUTION 5%)

Reaction -1

Add (10 drops) of $AgNO_3$ solution to (10 drops) of KSCN solution then you will see white precipitate of AgSCN in the tube which will be soluble in NH_4OH but insoluble in dil HNO_3.

$$KSCN + AgNO_3 \longrightarrow \underset{\substack{\downarrow \\ \text{wt. ppt.}}}{AgSCN} + KNO_3$$

Reaction -2

-Confirm by $FeCl_3$ test:

Add (5 drops) of $FeCl_3$ solution to (5 drops) of KSCN solution the product solution will be red blood solution of $Fe(SCN)_3$

$$FeCl_3 + 3KSCN \rightarrow \underset{\text{red blood colour solution}}{Fe(SCN)_3} + 3KCl$$

6-STUDY THE REACTIONS OF OXALATE ION $C_2O_2^{=}$ (USE $K_2C_2O_4$ SOLUTION 5%)

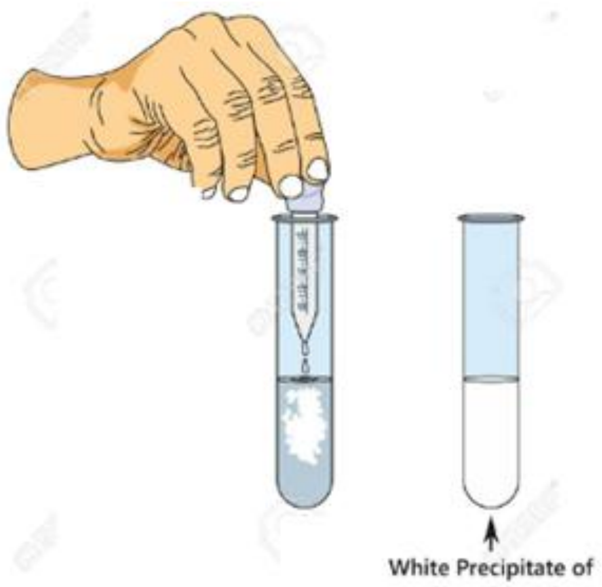
Many metal ions form insoluble precipitates with oxalate, a prominent example being calcium oxalate, the primary constituent of the most common kind of kidney stones.

Reaction -1

-With $AgNO_3$ solution (5%)

When an aqueous solution of silver nitrate (AgNO3) is (10 drops) added to the aqueous solution of (10 drops) potassium oxalate ($K_2C_2O_4$), a white precipitate of silver oxalate ($Ag_2C_2O_4$) is formed that is indicated by the following chemical reaction.

$$K_2C_2O_4 + 2AgNO_3 \rightarrow \underset{\text{Wt.PPt}}{\underset{\downarrow}{Ag_2C_2O_4}} + 2KNO_3$$



1- Try the solubility of $Ag_2C_2O_4$ in HNO_3 (soluble).
2- Try the solubility of $Ag_2C_2O_4$ in dil NH_4OH (soluble).

Reaction - 2

-Confirm by $KMnO_4$ test:

Decolourized the purpul colour of Permanganate (2 drops) when warmed in acid solution (2 drops) with oxalate (10 drops) in (60-70C)

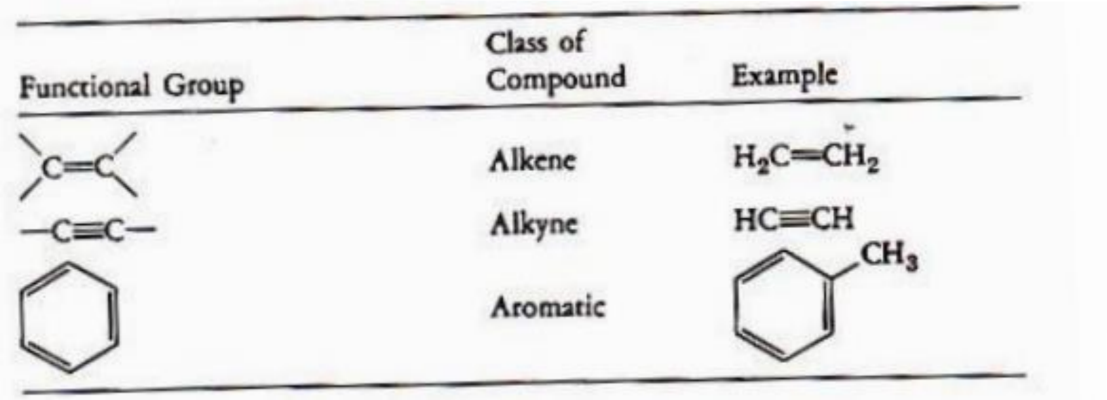
$$K_2C_2O_4 + H_2SO_4 \rightarrow H_2C_2O_4 + K_2SO_4$$

$$2\ KMnO_4 + H_2C_2O_4 + 3\ H_2SO_4 \rightarrow 2MnSO_4 + 10CO_3 + 8H_2O + K_2SO_4$$

This reaction depended upon oxidation reduction reaction.

Functional Groups

The chemical reactions of organic compounds occurred at specific sites in the molecules. These sites usually contained specific atoms or groups of atoms bonded to carbon. These groups at which reactions occur are called functional group.



Functional Group	Class of Compound	Example
>C=C<	Alkene	$H_2C=CH_2$
—C≡C—	Alkyne	HC≡CH
	Aromatic	CH_3

Alkanes

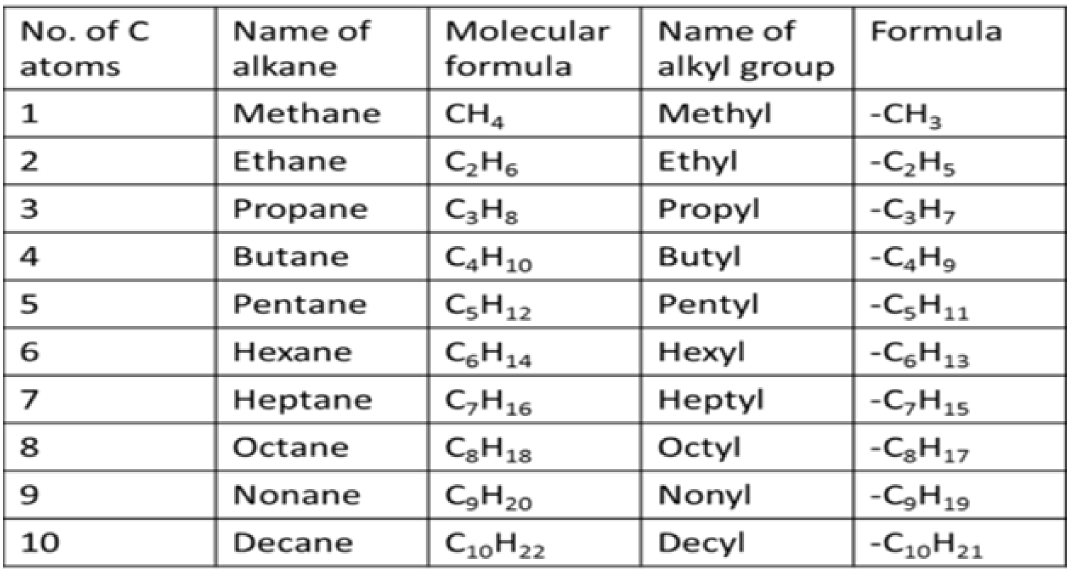
Alkanes are alphatic organic compounds that consist of single-bonded carbon(C-C) and hydrogen atoms(H-C) and lack any other functional groups. Alkanes have the general formula (CnH2n+2). Alkanes are also *saturated hydrocarbons*. Alkanes are the simplest and least reactive hydrocarbon species containing only carbons and hydrogens.

CnH2n+2 (n=1) → $C_1H_{2x1+2}=CH_4$ methane , (n=2) $C_2H_{2x2+2}=C_2H_6$ ethane

Alkyl Groups(R)

An alkyl group(R) is formed by removing one hydrogen from the alkane chain and is described by the formula (C_nH_{2n+1}). The removal of this hydrogen results in a stem change from **-ane** to **-**

yl. CnH2n+1 (n=1) → $C_1H_{2x1+1}=CH_3$ methyl , (n=2) $C_2H_{2x2+1}=C_2H_5$ ethyl . for that the alkane is (R-R or R-H)



No. of C atoms	Name of alkane	Molecular formula	Name of alkyl group	Formula
1	Methane	CH_4	Methyl	$-CH_3$
2	Ethane	C_2H_6	Ethyl	$-C_2H_5$
3	Propane	C_3H_8	Propyl	$-C_3H_7$
4	Butane	C_4H_{10}	Butyl	$-C_4H_9$
5	Pentane	C_5H_{12}	Pentyl	$-C_5H_{11}$
6	Hexane	C_6H_{14}	Hexyl	$-C_6H_{13}$
7	Heptane	C_7H_{16}	Heptyl	$-C_7H_{15}$
8	Octane	C_8H_{18}	Octyl	$-C_8H_{17}$
9	Nonane	C_9H_{20}	Nonyl	$-C_9H_{19}$
10	Decane	$C_{10}H_{22}$	Decyl	$-C_{10}H_{21}$

Naming Alkanes and Cycloalkanes

1-Identify the longest carbon chain. This chain is called the parent chain.

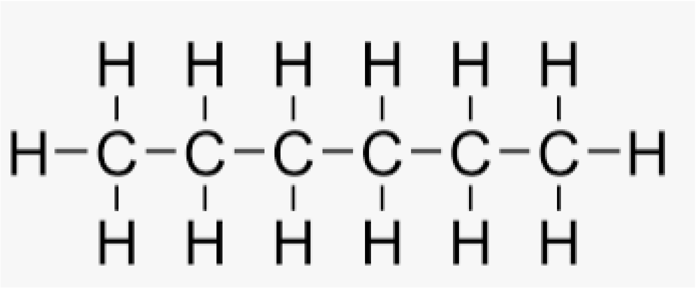
2-Identify all of the substituents (groups appending from the parent chain).

3-Number the carbons of the parent chain from the end that gives the substituents the lowest numbers. When compairing a series of numbers, the series that is the "lowest" is the one which contains the lowest number at the occasion of the first difference. If two or more side chains are in equivalent positions, assign the lowest number to the one which will come first in the name.

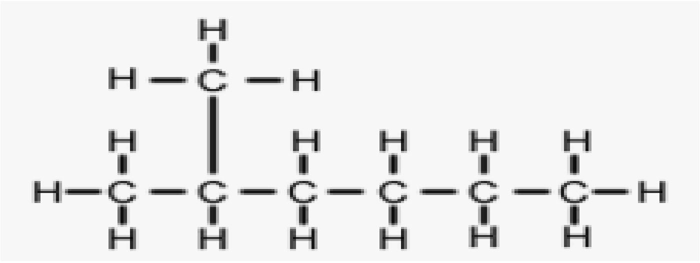
4-If the same substituent occurs more than once, the location of each point on which the substituent occurs is given. In addition, the number of times the substituent group occurs is indicated by a prefix (di, tri, tetra, etc.).

5-If there are two or more different substituents they are listed in alphabetical order using the base name (ignore the prefixes). The only prefix which is used when putting the substituents in alphabetical order is iso as in isopropyl or isobutyl. The prefixes sec- and tert- are not used in determining alphabetical order except when compared with each other.

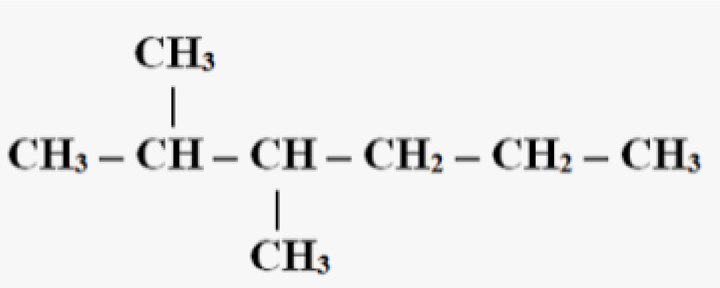
Hexane (Hex + ane)



2-methyl hexane

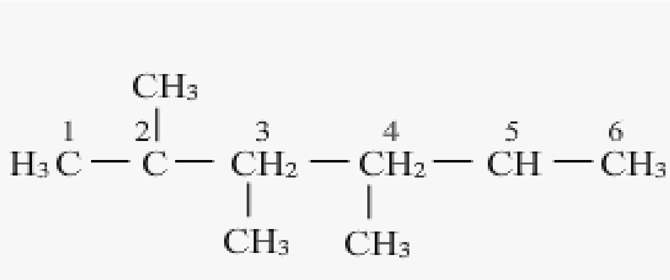


2,3-dimethyl hexane



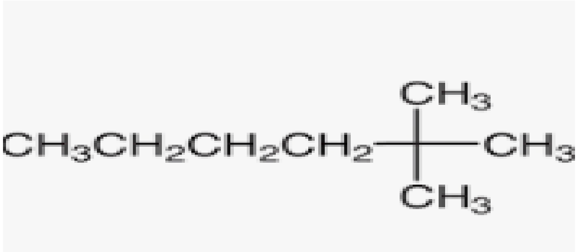
$$CH_3-\overset{\overset{CH_3}{|}}{CH}-\underset{\underset{CH_3}{|}}{CH}-CH_2-CH_2-CH_3$$

2,3,4-trimethylhexane



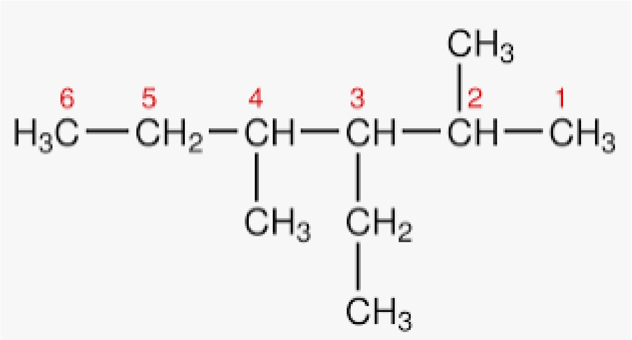
$$\overset{1}{H_3C}-\overset{\overset{CH_3}{2|}}{C}-\underset{\underset{CH_3}{|}}{\overset{3}{CH_2}}-\underset{\underset{CH_3}{|}}{\overset{4}{CH_2}}-\overset{5}{CH}-\overset{6}{CH_3}$$

2,2-dimethyl hexane



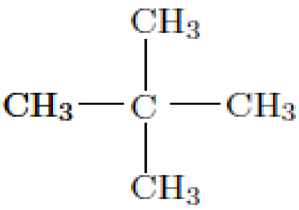
$$CH_3CH_2CH_2CH_2-\underset{CH_3}{\overset{CH_3}{+}}-CH_3$$

3-ethyl-2,4-dimethylhexane



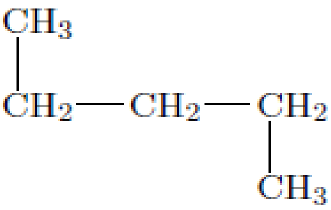
$$\overset{6}{H_3C}-\overset{5}{CH_2}-\underset{\underset{CH_3}{|}}{\overset{4}{CH}}-\underset{\underset{\underset{CH_3}{|}}{CH_2}}{\overset{3}{CH}}-\overset{\overset{CH_3}{|2}}{CH}-\overset{1}{CH_3}$$

neopentane



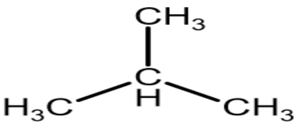
$$CH_3-\underset{\underset{CH_3}{|}}{\overset{\overset{CH_3}{|}}{C}}-CH_3$$

n-pentane

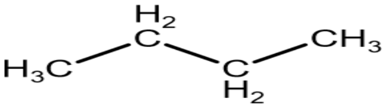


$$\overset{\overset{CH_3}{|}}{CH_2}-CH_2-\underset{\underset{CH_3}{|}}{CH_2}$$

Common naming for alkane



isobutane

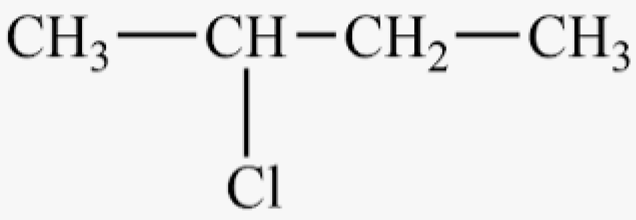


n-butane
butane

Alkyl halides

The halogen is treated as a substituent on an alkane chain. The halo- substituent is considered of equal rank with an alkyl substituent in the numbering of the parent chain. The halogens are represented as follows:

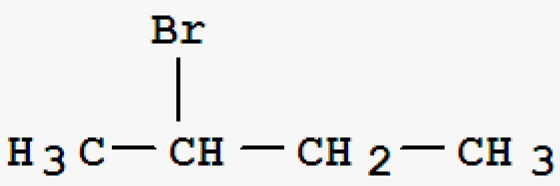
(F) fluoro- , (Cl) chloro- , (Br) bromo- , (I) iodo-



$$CH_3-\underset{\displaystyle Cl}{\underset{|}{CH}}-CH_2-CH_3$$

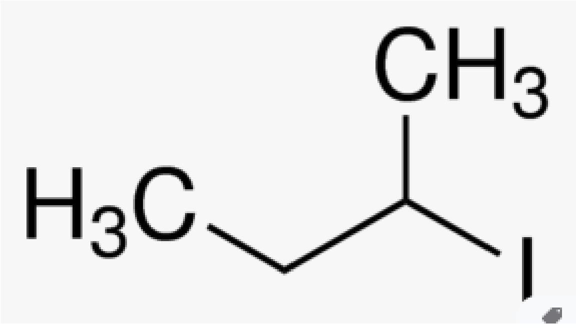
2-chlorobutane

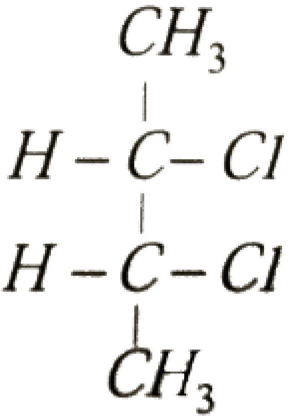
2-bromobutane



$$H_3C-\overset{\displaystyle Br}{\overset{|}{CH}}-CH_2-CH_3$$

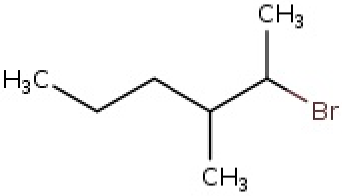
2-Iodobutane





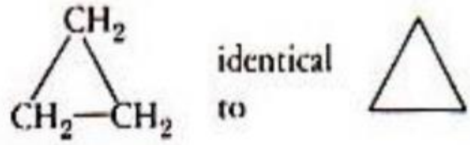
2, 3-dichlorobutane

2-Bromo-3-methylhexane

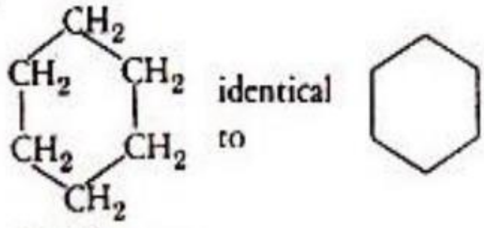


- Saturated hydrocarbons can also exist as rings. Cyclic compounds of carbon containing only single bonds are called Cycloalkanes.If The compounds contain only one ring, they have the general formula C_nH_{2n}

Cycloalkanes are named by adding the prefix cyclo- to the name of the straight – chain hydrocarbon containing the same number of carbon atoms. For example



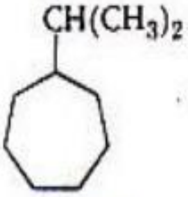
Cyclopropane



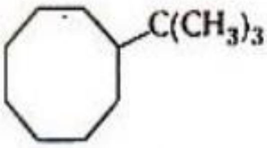
Cyclohexane

All the carbon atoms of a cycloalkane are equivalent. Therefore, no number prefix is needed for monosubstituted cycloalkanes. For example:

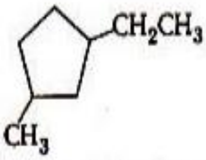
Methylcyclopropane



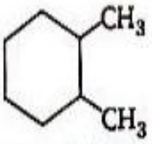
Isopropylcycloheptane



t-Butylcyclooctane

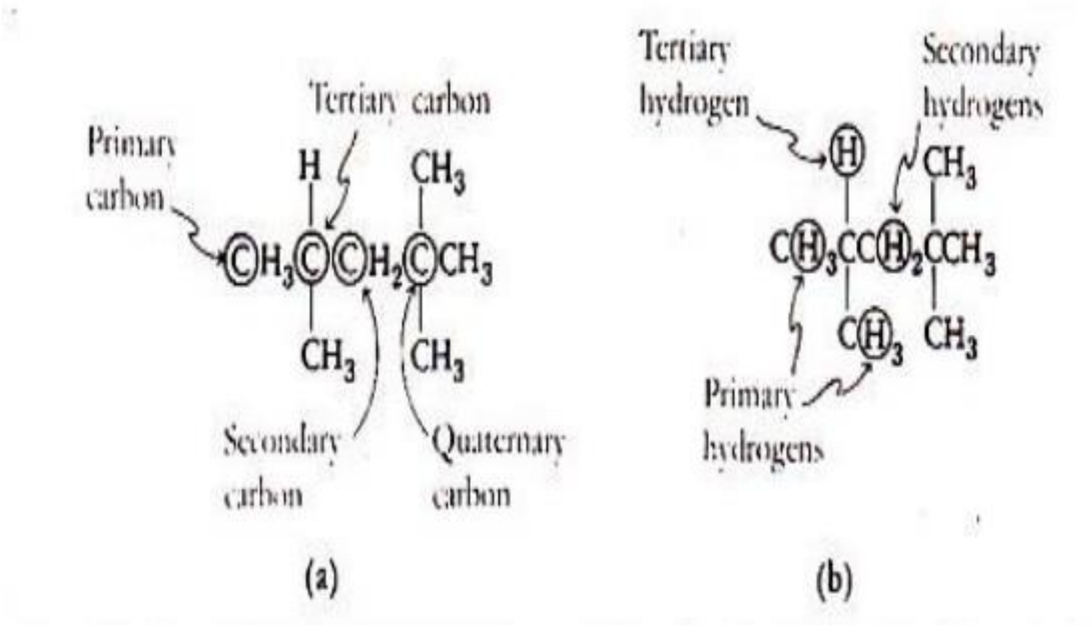


1-Ethyl-3-methylcyclopentane



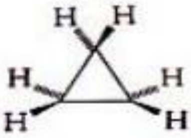
1,2-Dimethylcyclohexane

A carbon atom is designated: primary,secondary,or tertiary according to the number of carbon atoms bonded to it.

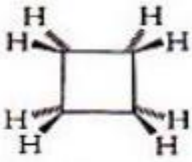


Geometric Isomers

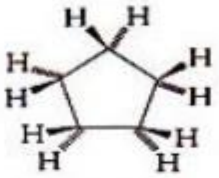
Molecules that differ in the three – dimensional arrangements of their atoms in space.



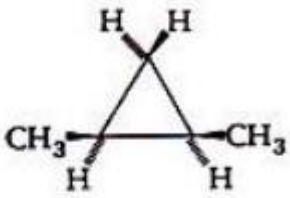
Cyclopropane



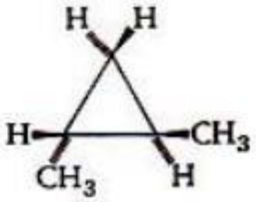
Cyclobutane



Cyclopentane



cis-1,2-Dimethylcyclopropane



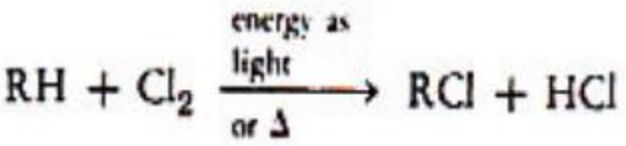
trans-1,2-Dimethylcyclopropane

Physical Properties

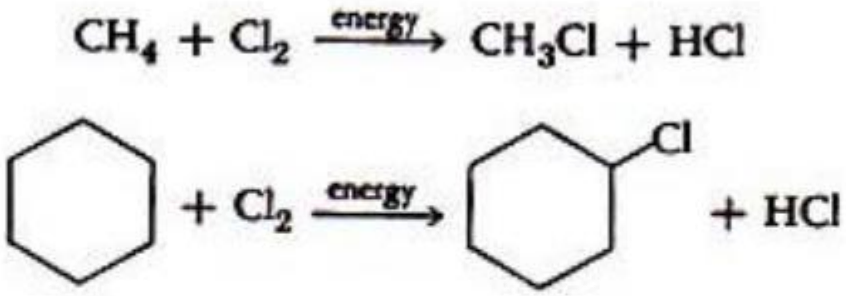
The alkanes can exist as gases, liquids, or solids at room temperature. The unbranched alkanes methane, ethane, propane, and butane are gases. Branched alkanes normally exhibit lower boiling points than unbranched alkanes of the same carbon content. This occurs because of the greater van der Waals forces that exist between molecules of the unbranched alkanes. The unbranched alkanes have greater van der Waals forces of attraction because of their greater surface areas. alkanes are almost completely insoluble in water. Alkanes are **non-polar solvents**. Since only C and H atoms are present, alkanes are nonpolar. Alkanes are immiscible in water but freely miscible in other non-polar solvents. Alkanes consisting of weak dipole dipole bonds can not break the strong hydrogen bond between water molecules hence it is not miscible in water.

Chemical Reactions

- Chlorination



$$RH + Cl_2 \xrightarrow[\text{or } \Delta]{\text{energy as light}} RCl + HCl$$



$$CH_4 + Cl_2 \xrightarrow{\text{energy}} CH_3Cl + HCl$$

Oxidation:

In the presence of excess oxygen,alkanes burn to form carbon dioxide,water and energy

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O + \Delta$$

If insufficient oxygen ,alkanes form carbon monoxide or carbon

$$2CH_4 + 3O_2 \longrightarrow 2CO + 4H_2O$$

$$CH_4 + O_2 \longrightarrow C + 2H_2O$$

Preparing alkane

$$\underset{\text{Sodium salt of carboxylic acid}}{RCOONa} + NaOH \xrightarrow[\text{Heat}]{CaO} \underset{\text{Alkane}}{RH} + Na_2CO_3$$

$$\underset{\text{Sodium ethanoate}}{CH_3COONa} + NaOH \xrightarrow[\text{Heat}]{CaO} \underset{\text{Methane}}{CH_4} + Na_2CO_3$$

$$\underset{\text{Sodium propanoate}}{C_2H_5COONa} + NaOH \xrightarrow[\text{Heat}]{CaO} \underset{\text{Ethane}}{C_2H_6} + Na_2CO_3$$

Tikrit University **College of Dentistry**

Chemistry lab:11 **First stage**

Carbohydrates

Carbohydrates define as carbon compounds containing hydrogen and oxygen in the ratio of 2:1, so that their empirical formula is $C_X\ (H_2O)_y$, with some exception to this rule. Carbohydrates could also be define as a keton and aldehyde derivation of poly hydro alcohols. Thus carbohydrate contain an aldehyde and keton group , giving rise to the reducing ability, and is known as an aldose or ketose.

The basic units of carbohydrates are monosaccarides which cannot be split further by hydrolysis. These monomers are named according to the number of carbon atoms in the chain; thus tetroses, pentoses and hexoses contain four, five and six carbon atoms, respectively.

Qualitative test for Carbohydrates :

It is suggested that the following tests be carried out on a number of carbohydrates considered typical of their group.

Solution of 0.5 should be tested to form an idea of the sensitivity of the tests.

Monosaccaride :	Glucose,	Fructose,	Galactose
Disaccharides:	Lactose,	Maltose,	Sucrose
Polysaccharides:	Starch,	Glycogen,	Inulin

1- Molish test:

Principle

Concentrated sulfuric acid hydrolysis glycosidic bonds to give monosaccarides which are then hydrated to furfural and its derivative (hydroxyl methyl furfural) which, in turn, compine with sulfonated

ά-naphthol to give a reddish- violet color. This is a general reaction for the detection carbohydrates.

reagents

1- Sulfuric acid (conc.)
2- ά- naphthol in ethanol.

Procedure

1- Add 2 drops of ά- naphthol 2 ml of the sugar (test) solution, in test tube.
2- Carefully, pour about 2 ml of H_2SO_4 (conc.) down the side of the tube to form 2 layer. Observe any change in color at the junction of the two liquids. (violet ring).

Reactions of reducing sugars

If a suspension of copper hydroxide in alkaline solution is heated, black cupric oxide is formed:

$$Cu(OH)_2 \longrightarrow CuO + H_2O$$

However, if a reducing substances is present, thus brown cuprous oxide is precipitated:

$$2Cu(OH)_2 \longrightarrow Cu_2O + H_2O + O_2$$

Carbohydrates, with a free aldehyde or keton group, have reducing properties in alkaline solution. In addition, monosaccarides act as reducing agents in weakly acid solution.

2- Benedict test

Principle

Benedict test is used to detect reducing abilities of sugars.

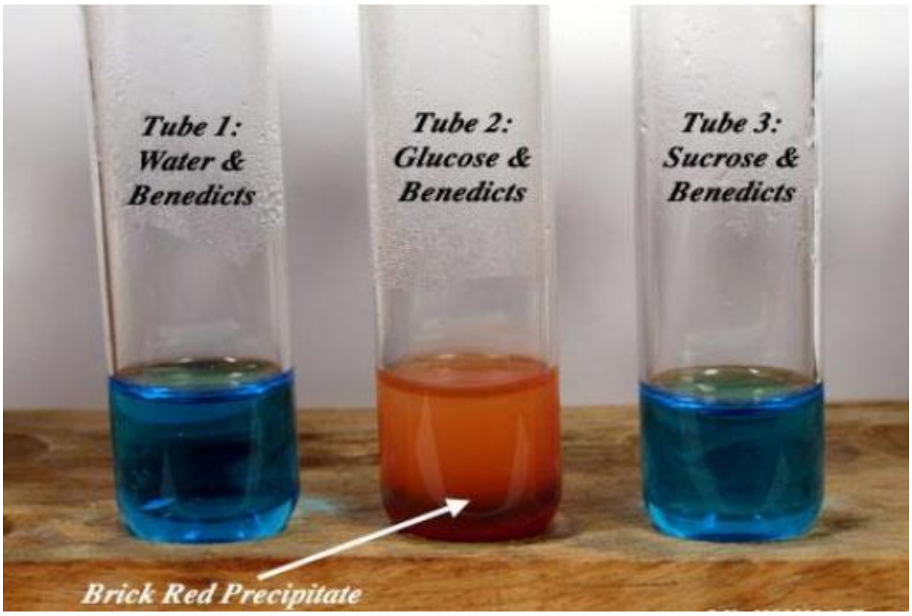
Reagent

Benedict reagent: (sodium citrate, sodium carbonate in warm water)
(copper sulfate in about 100 ml water)
Pour the first solution into a 2 liter beaker ad slowly add the copper sulfate solution with stirring.

Procedure

1- Add 5 drops of the sugar (test) solution to 2 ml of Benedict reagent, in a test tube.

2- Place the test tube in a boiling water bath for 5 minutes. Observe ang change in color or precipitation. (brick red precipitate)



3-Barfoed test

Principle

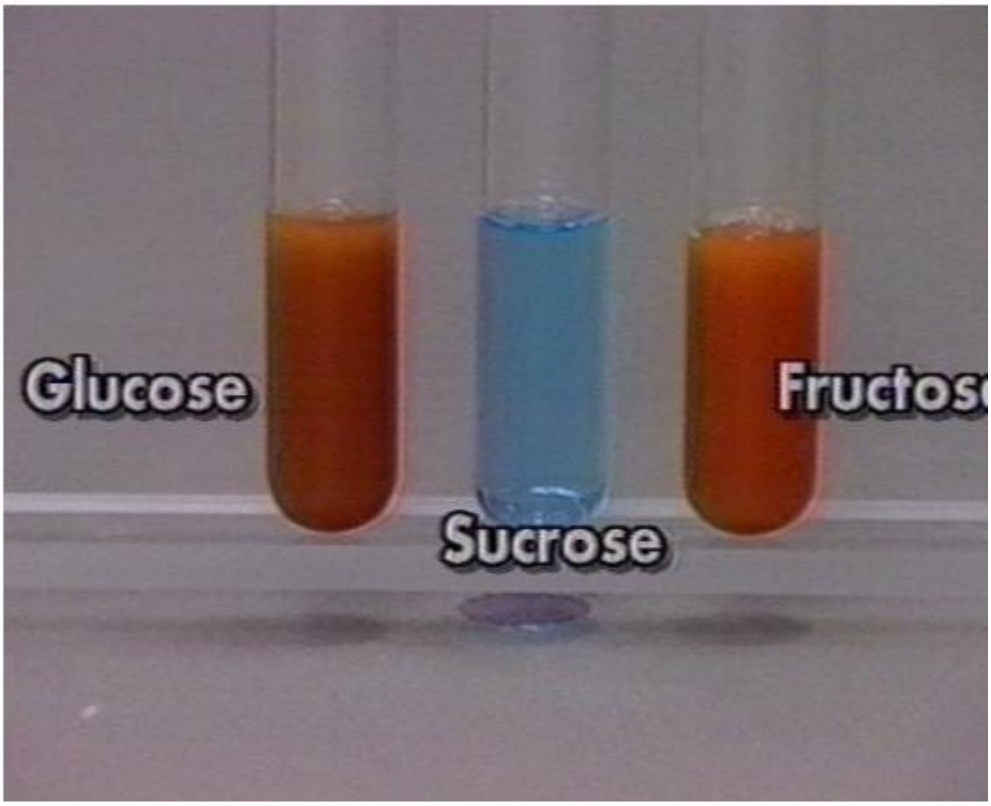
Barfoed reagent is weakly acidic and is only reduced by monosaccarides. Prolonged boiling may hydrolyse disaccharides to give a false positive reaction.

Reagents

1- Barfoed reagent: (copper acetate in water and add glacial acetic acid).

Procedure

1- Add 2 ml of the sugar (test) solution to 2 ml of barfoed reagent, in a test tube.
2- Boil for 3 minutes, and allow to stand.
(brick red).



4-Bial test for pentoses

Principle:

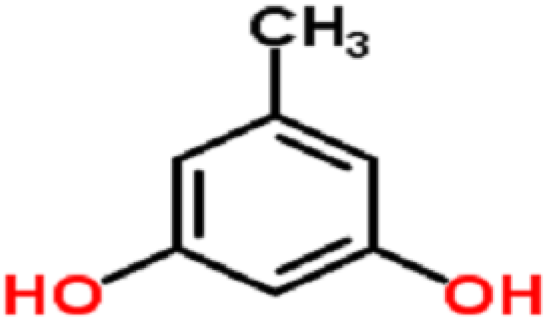
When pentoses are heated with concentrated HCl, furfural is formed which condenses with orcinol, in the presence of ferric ion, to give a blue-green color. The reaction is not absolutely specific for pentoses since prolonged of some hexoses yields hydroxyl methyl furfural, which also reacts with orcinol, to give colored complexes.

Reagents

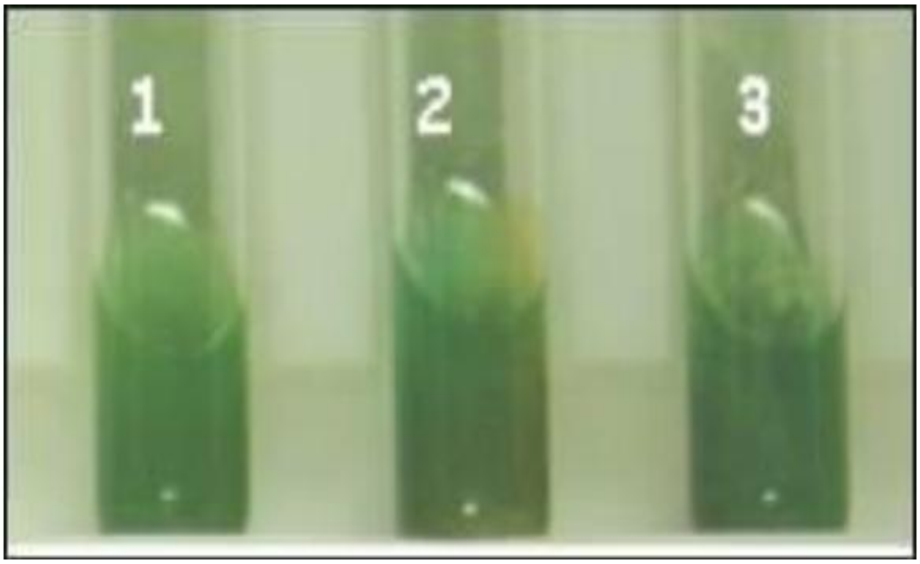
Bial orcinol reagent (orcinol in 500 ml of conc. HCl and add drops of $FeCl_3$ solution)

Procedure:

1- Add about 1 ml of the sugar (test) solution to 2.5 ml of Bial reagent, in a test tube, and heat until boiling commences.
2- Report the resulting color formed.



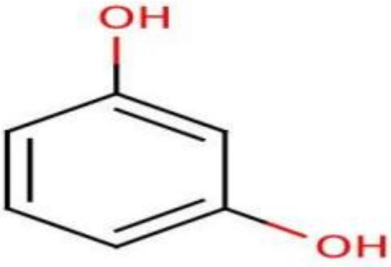
Orcinol



5-Seliwanoff test of ketoses:

Principle:

Ketoses are dehydrated more rapidly than aldoses to give furfural derivatives, which then condense with resorcinol, to form a red complex.



Resorcinol

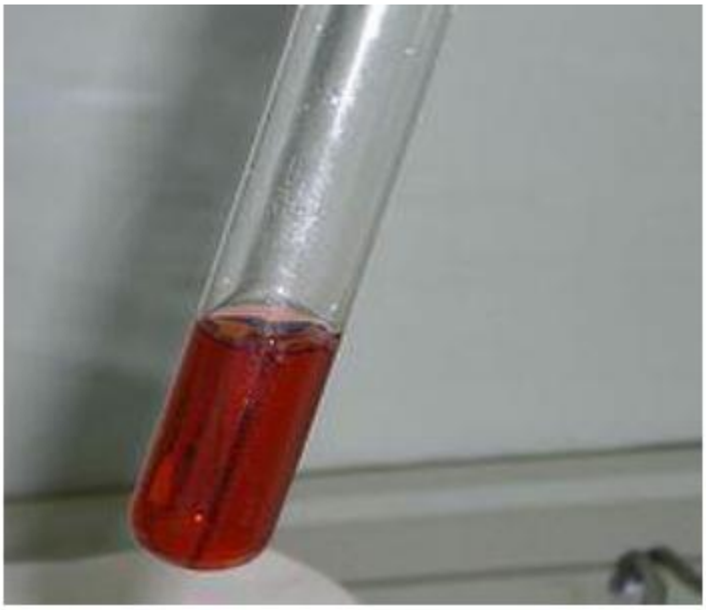
Reagents

Seliwanoff reagent (resorcinol in HCl.) .

Procedure

1-Add 2 drops of the sugar (test) solution, to 2 ml of the Seliwanoff reagent.

2-Heat in a boiling water bath for 5 minutes.



6-Iodine test

Principle

Iodine forms colored adsorption complexes with poly saccharides.

Reagents

1-Iodine solution (I_2 in KI)

2-1% solution of glycogen, starch, inulin.

Procedures

1-To 1 ml of carbohydrates solution add few drops of iodine. If the color is deep, add distilled water.

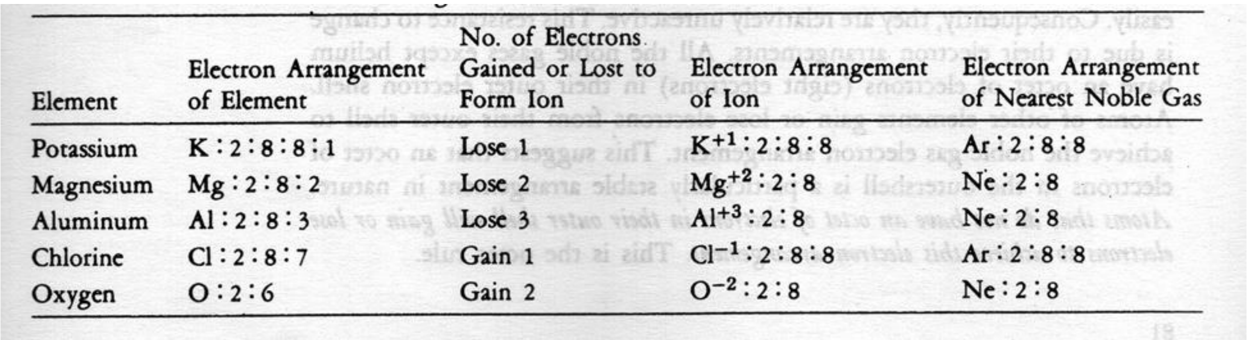
2-Heat the colored solution to boiling and note that the color disappears when hot and reappear upon cooling. If boiling is prolonged, the color disappears and dose not reappear by cooling duo to the escape of iodine (volatile).

Chemical Bond

Molecules are made up of atoms joined together .The links that join these atom are called chemical bonds.The properties of a molecule are determined by its chemical bonds.

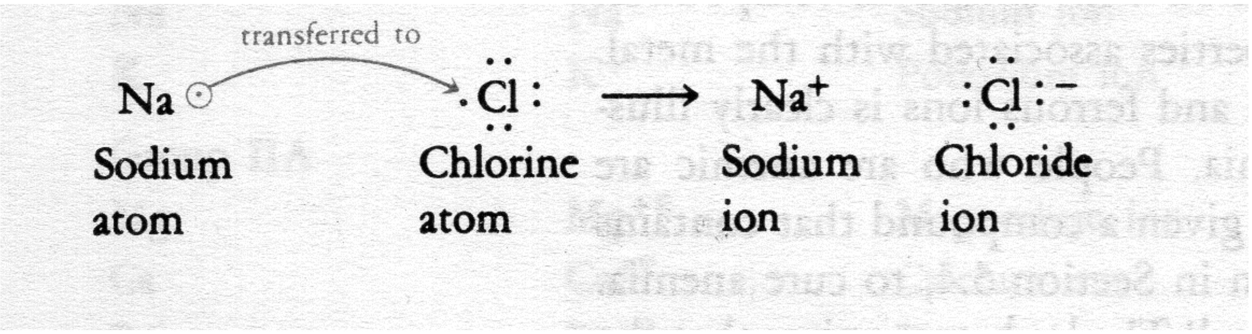
Types of Chemical Bonds

1-Ionic bonding is a chemical bond that involves the electrostatic attraction between oppositely charged ions, and is the primary interaction occurring in ionic compounds. The ions are atoms that have gained one or more electrons (known as anions, which are negatively charged) and atoms that have lost one or more electrons (known as cations, which are positively charged). In the simplest case, the cation is a metal atom and the anion is a nonmetal atom.



Element	Electron Arrangement of Element	No. of Electrons Gained or Lost to Form Ion	Electron Arrangement of Ion	Electron Arrangement of Nearest Noble Gas
Potassium	K : 2 : 8 : 8 : 1	Lose 1	K^{+1} : 2 : 8 : 8	Ar : 2 : 8 : 8
Magnesium	Mg : 2 : 8 : 2	Lose 2	Mg^{+2} : 2 : 8	Ne : 2 : 8
Aluminum	Al : 2 : 8 : 3	Lose 3	Al^{+3} : 2 : 8	Ne : 2 : 8
Chlorine	Cl : 2 : 8 : 7	Gain 1	Cl^{-1} : 2 : 8 : 8	Ar : 2 : 8 : 8
Oxygen	O : 2 : 6	Gain 2	O^{-2} : 2 : 8	Ne : 2 : 8

An ionic bond is the transfer of electrons from a metal to a non-metal in order to obtain a full valence shell for both atoms.



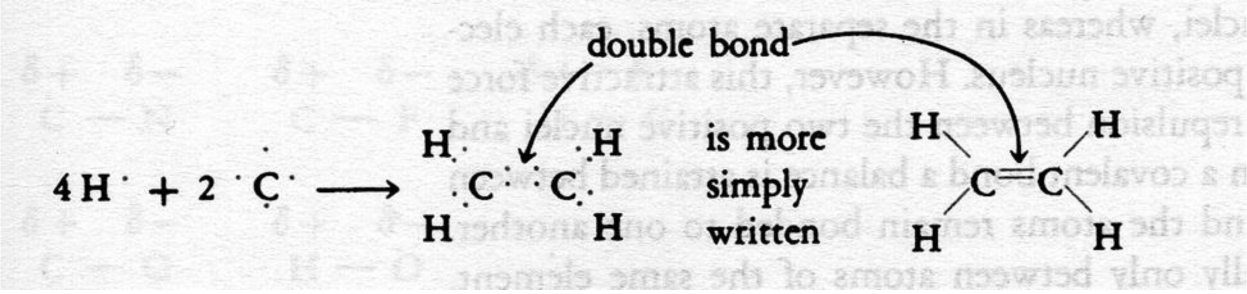
2- covalent bond(molecular bond) is a chemical bond that involves the sharing of electron pairs between atoms. These electron pairs are known as ***shared pairs*** or ***bonding pairs,*** and the stable balance of attractive forces between atoms, when they share electrons, is known as covalent bonding for many molecules, the sharing of electrons allows each atom to attain the equivalent of a full outer shell, corresponding to a stable electronic configuration(to achieve noble gas electron arrangement for each atom).

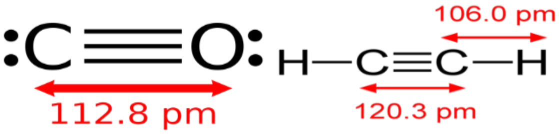
For getting stability, hydrogen atom shares its valence electron with other hydrogen atom to form single covalent bond results the formation of hydrogen molecule($H:H$ or $H-H$). So each bonded hydrogen atom gets two electrons and stabilized in hydrogen molecule.

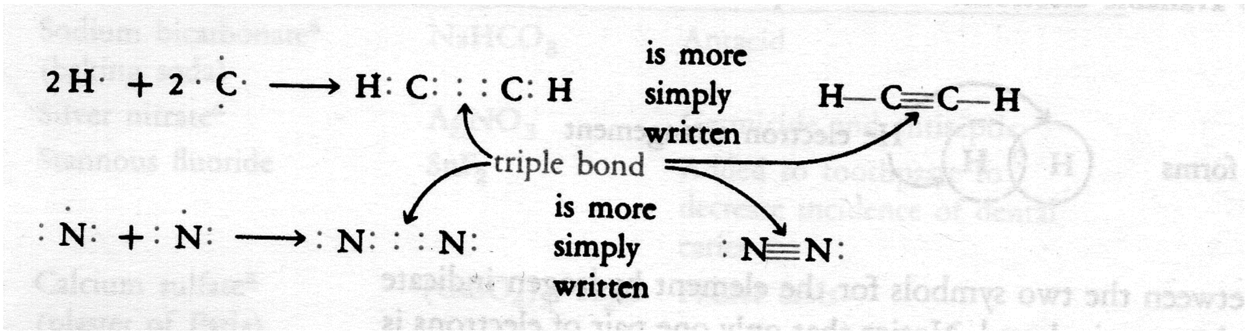
For getting stability, hydrogen atom shares its valence electron with other hydrogen atom to form single covalent bond results the formation of hydrogen molecule. So each bonded hydrogen atom gets two electrons and stabilized in hydrogen molecule.

Another example of covalent bonding is carbon dioxide molecule in which center carbon atom is bonded with two oxygen atoms through double covalent bonds($O=C=O$ or $O::C::O$). The center atom contains four valence electrons and to get the octet configuration, it tends to form four covalent bonds with same or different atoms. Similarly oxygen atom contains 6 valence electrons and needs two electrons to get octet configuration therefore it can form two covalent bonds. Therefore four covalent bonds by carbon atoms and two covalent bonds by oxygen atoms form carbon dioxide

A single line indicates a single bond between two atoms (i.e., involving one electron pair), double lines (=) indicate a double bond between two atoms (i.e., involving two electron pairs), and triple lines (≡) represent a triple bond, as found, for example, in carbon monoxide ($C\equiv O$). Single bonds consist of one sigma (σ) bond, double bonds have one σ and one pi (π) bond, and triple bonds have one σ and two π bonds.



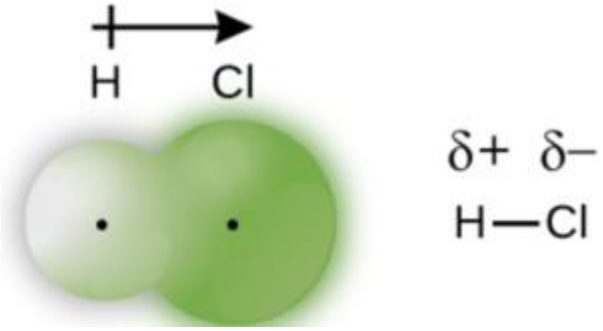




Covalent Polar bonds

A bond between 2 nonmetal atoms that have different electronegativities and therefore have unequal sharing of the bonding electron pair

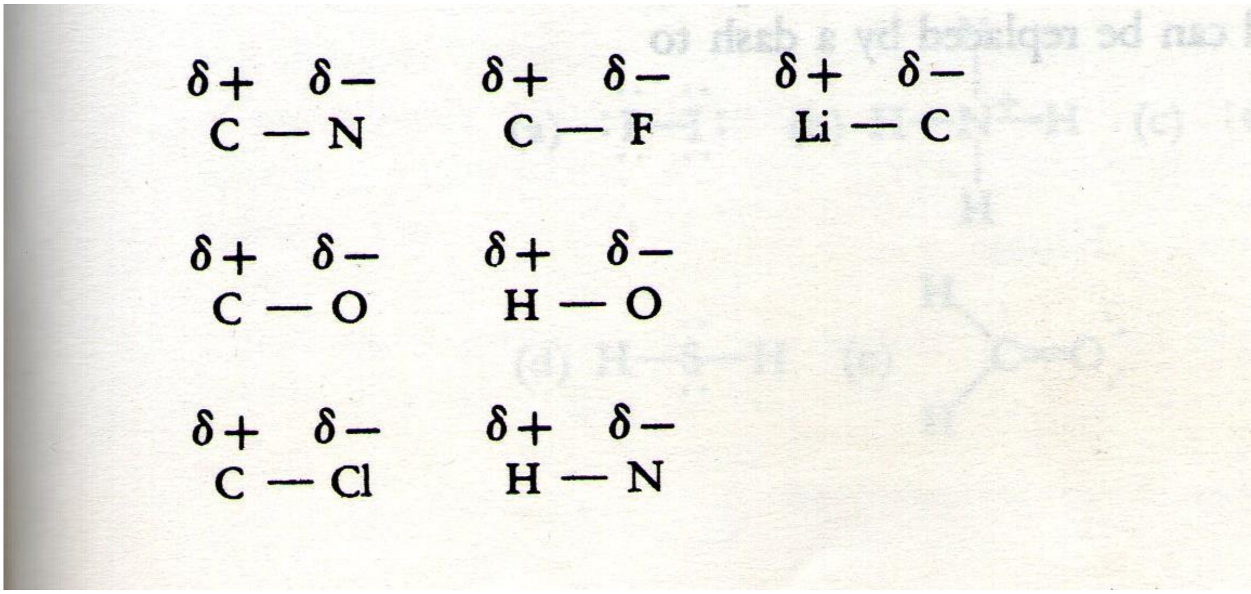
Example: In H-Cl, the electronegativity of the Cl atom is 3.0, while that of the H atom is 2.1



The result is a bond where the electron pair is displaced toward the more electronegative atom. This atom then obtains a partial-negative charge while the less electronegative atom has a partial-positive charge.This separation of charge or bond dipole can be illustrated using an arrow with the arrowhead directed toward the more electronegative atom. Within a molecule each polar bond has a bond dipole.A polar molecule always contains polar bonds, but some molecules with polar bonds are nonpolar.

Polar Molecule

A molecule in which the bond dipoles present do not cancel each other out and thus results in a molecular dipole.Cancellation depends on the shape of the molecule or Stereochemistry and the orientation of the polar bonds.



$\overset{\delta+}{C}-\overset{\delta-}{N}$ $\overset{\delta+}{C}-\overset{\delta-}{F}$ $\overset{\delta+}{Li}-\overset{\delta-}{C}$

$\overset{\delta+}{C}-\overset{\delta-}{O}$ $\overset{\delta+}{H}-\overset{\delta-}{O}$

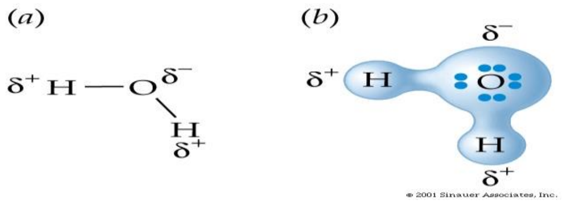
$\overset{\delta+}{C}-\overset{\delta-}{Cl}$ $\overset{\delta+}{H}-\overset{\delta-}{N}$

Molecular Dipole

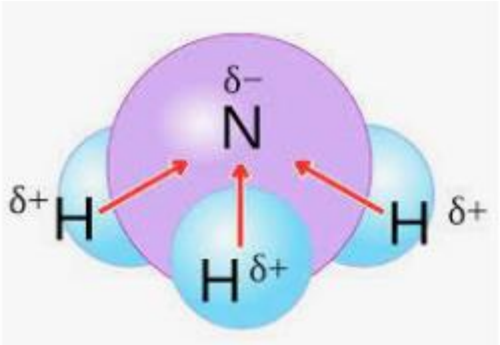
A result of the bond dipoles in a molecule.

Bond dipoles may or may not cancel out thereby producing either molecules that are nonpolar, if they cancel, or polar, if they do not cancel

Examples:H_2O is polar molecule its shape(V) molecule with 2 bond dipoles that are equal and same direction therefore the bond polarities not cancel and the molecule is polar.



© 2001 Sinauer Associates, Inc.



CO_2 is a linear molecule with 2 bond dipoles that are equal and oppositely directed therefore the bond polarities cancel and the molecule is nonpolar.

Carbon Dioxide (CO_2)

Carbon dioxide is a nonpolar compound. Both ends of the molecule are slightly negative in charge.

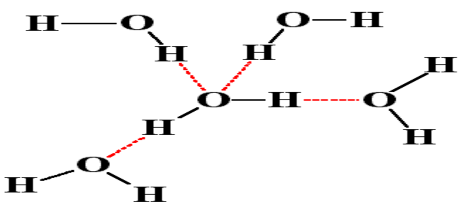
HCN is a linear molecule with 2 bond dipoles that are in the same direction and are not equal therefore the bond polarities do not cancel and the molecule is polar

3-Hydrogen bond

A hydrogen bond is an electrostatic attraction between two polar groups that occurs when a hydrogen (H) atom, covalently bound to a highly electronegative atom such as nitrogen (N), oxygen (O), or fluorine (F), experiences the electrostatic field of another highly electronegative atom nearby.

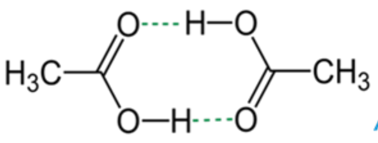
Hydrogen bonds can occur between molecules (intermolecular) or within different parts of a single molecule (intramolecular). Depending on the nature of the donor and acceptor atoms which constitute the bond. Hydrogen bonding is important in many chemical processes.

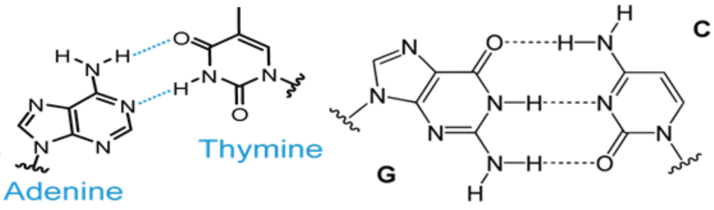
Hydrogen Bonding in Water

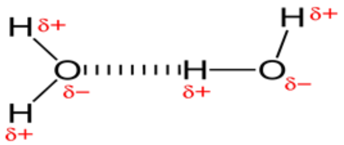


Hydrogen bonding is responsible for water's unique solvent capabilities. Hydrogen bonds hold complementary strands of DNA together(Guanine with Cytosine G≡C,and Thymine with Adenine T=A) and they are responsible for determining the three-dimensional structure of folded proteins including enzymes and antibodies.when the numbers of hydrogen bond had increased hydroge

Bond power.



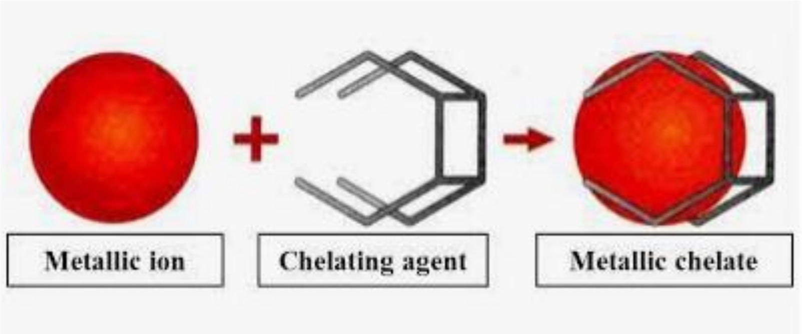




Chelation

Chelation is a type of bonding of ions and molecules to metal ions. It involves the formation or presence of two or more separate coordinate bonds between a polydentate ligand and a single central metal atom.

These ligands are called chelants, chelators or chelating agents.

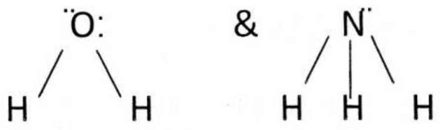


1- The case of zinc and its use as a maintenance therapy to prevent the absorption of copper in people with Wilson's disease.

2-

ligand

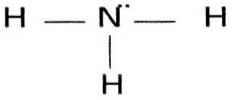
A ligand is an ion or molecule, which donates a pair of electrons to the central metal atom or ion to form a coordination complex



The atom in a ligand that is bound directly to the metal atom is known as the "donor atom". Depending on the number of donor atoms, ligands are classified as:

Monodentate, bidentate, or polydentate.

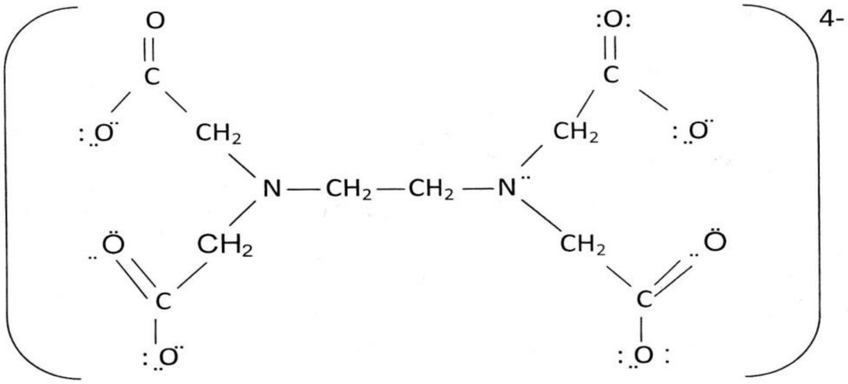
H_2O, NH_3,



$$H-\dot{N}-H \quad (\text{with } H \text{ below } N)$$

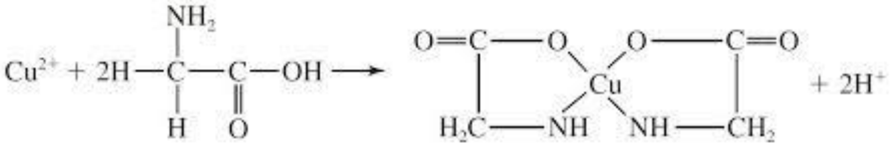
$H_2\dot{N}-CH_2-CH_2-\dot{N}H_2$

Monodentate Bidentate ligand (Ethtlenediamine)

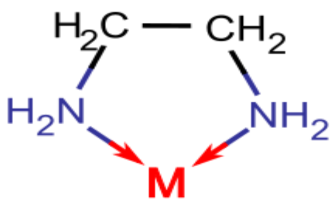


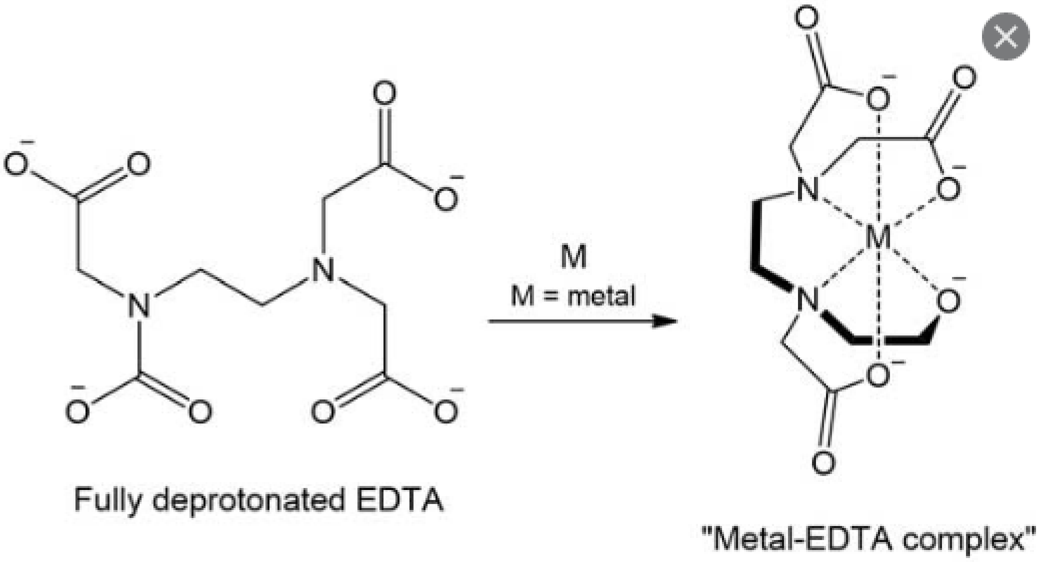
EDTA (Polydentate ligand)

Glycine is amino acid consider bidentate ligand



Glycine Cu²⁺ complex with glycine.



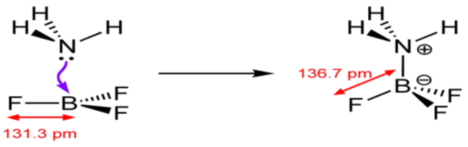


While individual hydrogen bonds are not very strong, a series of hydrogen bonds is very secure. When one molecule hydrogen bonds through two or more sites with another molecule, a ring structure known as a chelate is formed. Chelating compounds are useful for removing or mobilizing molecules and atoms such as metals. **chelation therapy to remove toxic metals from the body,EDTA (ethylenediaminetetra aceticacid) is injected into the bloodstream to remove heavy metals and/or minerals from the body. Chelation means "to grab" or "to bind." When EDTA is injected into the veins, it "grabs" heavy metals and minerals such as lead, mercury, copper, iron, arsenic, aluminum, and calcium and removes them from the body.**

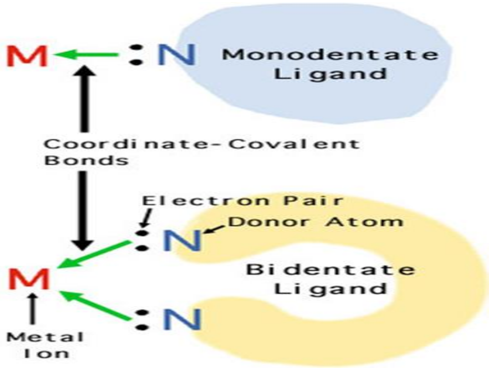
4-coordinate bond

A covalent chemical bond between two atoms that is produced when one atom shares a pair of electrons with another atom lacking such a pair. A type of covalent bond in which both the shared electrons are contributed by one of the two atoms(first atom with pair of electron and the another atom with empty orbital)

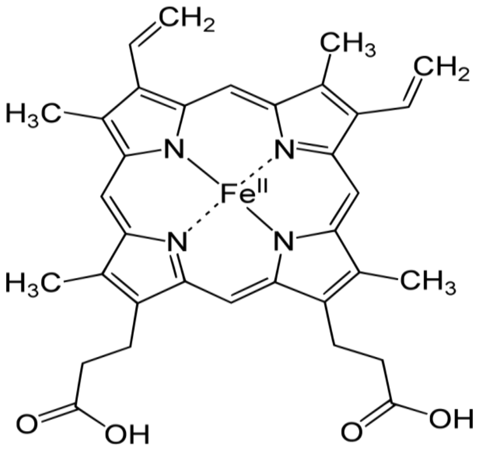
$_5$B $1s^2\ 2s^2\ 2p^1$ $_7$N $1s^2\ 2s^2\ 2p^4$

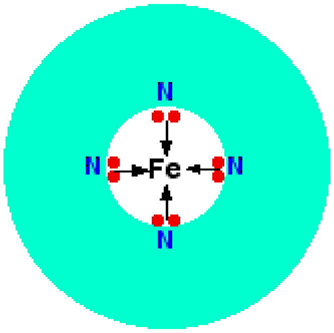


A compound that donates one electron pair is monodentate ligand while any compound donates two electron pairs is didentateligand.this compounds donate the electron pairs for empty orbital to hold coordinate bonds.



Hemoglobin (***Hb)*** (contains 4 subunits protein), each subunit contains a heme group. Each heme group contains an iron atom that is able to bind to one oxygen (O_2) molecule. Therefore, each hemoglobin protein can bind four oxygen molecules. One of the most important classes of chelating agents in nature are the porphyrins. A porphyrin molecule can coordinate to a metal using the four nitrogen atoms as electron-pair donors, and hence is a polydentate ligand . Heme is a porphyrin that is coordinated with Fe(II).In the body, the iron in the heme is coordinated to the four nitrogen atoms of the porphyrin of the hemoglobin protein .Hemoglobin is the protein that transports oxygen (O_2) in human blood from the lungs to the tissues of the body and transports carondioxide (CO_2) in human blood from the tissues of the body to the lungs.



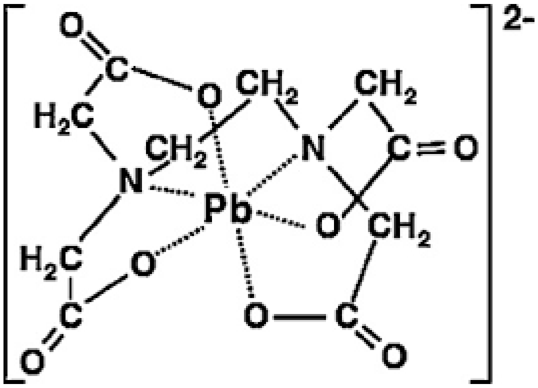


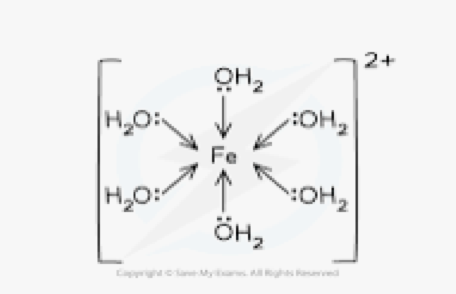
Example:Binding of EDTA with Pb^{+2}

$_{82}Pb$ $[Xe]\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^2$

Pb^{+2} $[Xe]\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^0$

Pb^{+4} $[Xe]\ 6s^0\ 4f^{14}\ 5d^{10}\ 6p^0$




2+
Fe
OH_2
H_2O
OH_2
H_2O
OH_2
OH_2
Copyright © Save My Exams. All Rights Reserved

طب اسنان
مرحلة الاولى

نظري

< ٥ -

Chemistry – Lecture (1) *Dr-Mahdi S. Hamad*

Atom:Smallest representative sample an element

Periodic table

The arrangement of element into two directions 1-(***groups***)that means arrangement of element in colums by this way the element in each group have similar chemical and physical properties as well as each element in the group have same number of electrons in outer shell(ns and np).2- (***periods***)that means arrangement of element in horizontal rows by this way the element in each period have same number of outer shell(*ns* and *np*).The elements are divided into three classes based on their physical properties.Elements that show metallic luster when polished,are capable of being drawn out into wire,can be hammered into sheets and are good conductors of heat and electricity are classed as *metal*.Element that do not have these properties are classed as *nonmetal*.A class between these two is called *metalloids*

Parts of atoms

Electron is a particle responsible for electric current .It is negatively charged particle(-),Proton is a particle positively charged particle(+) Neutron has no charge

Table . Three Important Subatomic Particles

Name	Charge	Mass (amu)	Symbol
Electron	−1	$\frac{1}{1837}$	e^-
Proton	+1	1.007	H^+, p^+, p
Neutron	0	1.004	n

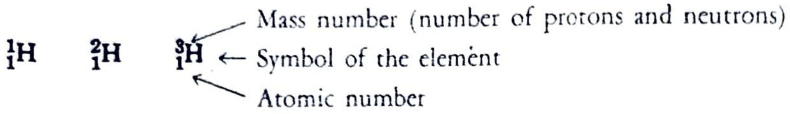
Atom has nucleus which contains protons and neutrons as well as electrons round out of nucleus in shells.Atom is electrically neutral ,therefore the charge created by protons in nucleus must be balanced exactely by equal number of electrons outside of nucleus.

Atomic number

The atomic number of an element is equal to the number of protons in its nucleus.

Isotopes

Atoms whose nuclei have the same number of protons but different numbers of neutrons are called *isotopes*.A specific isotopes of an element is designated by writing the symbol of the element with the atomic mass number placed as a superscript to the left and atomic number is also added as a subscript to the left .



Isotopes differ in mass because they have different numbers of neutrons in its nuclei.

Table . Naturally Occurring Isotopes of Carbon, Oxygen, and Sulfur

Symbol of Isotope	Number of Protons	Number of Neutrons	Atomic Number	Mass Number
$^{12}_{6}C$	6	6	6	12
$^{13}_{6}C$	6	7	6	13
$^{16}_{8}O$	8	8	8	16
$^{17}_{8}O$	8	9	8	17
$^{18}_{8}O$	8	10	8	18
$^{32}_{16}S$	16	16	16	32
$^{33}_{16}S$	16	17	16	33
$^{34}_{16}S$	16	18	16	34
$^{36}_{16}S$	16	20	16	36

Stable isotopes stable isotopes do not decay .they do not produce radiation .***Radioactive isotopes*** decay, emitting radiation that includes alpha, beta and gamma rays.

Mass number

Mass number of an atom is the sum of the number of protons and number of neutrons in its nucleus.The number of protons within the atom's nucleus is called atomic number and is equal to the number of electrons in the neutral (non-ionized) atom.

Electron arrangement

Electrons are located in shells in the space around the nucleus .The various shells in which electrons are located are giving numbers.The first shell number 1 (called the k shell).This shell is nearest the nucleus .An electron in this shell has the lowest energy of any electron in the atom.Succeeding shells are numbered 2(L shell) ,3(M shell) ,4(N shell) , and so forth.The energy increases in each succeeding shell.Each shell,except the first ,has subshells ,shell 2 has 4 subshells ,3 shell has 9 subshells, 4 shell has 16 subshells.The subshells are regions in space where electron spends most of its time.Such regions are called *atomic orbitals*

Table : . Shells and Their Atomic Orbitals

Shell	Number of Atomic Orbitals	Symbols of Atomic Orbitals
1	1	1s
2	4	2s $2p_x$ $2p_y$ $2p_z$
3	9	3s $3p_x$ $3p_y$ $3p_z$ Five 3d orbitals*
4	16	4s $4p_x$ $4p_y$ $4p_z$ Five 4d and seven 4f orbitals*

* Each of the 3d, 4d, and 4f atomic orbitals also have a subscript. However, we need not concern ourselves with their complete symbol.

$_{8}O:\_1s^2\ 2s^2\ 2p^4$ **or** $_{8}O: 2:6$

$_{17}Cl:\_1s^2 2s^2 2p^6 3s^2 3p^5$ or $_{17}Cl:2:8:7$

$_{53}I:\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$ or $_{53}I:2:8:18:18:7$

$_{53}I:_{36}[Kr]5s^2 4d^{10} 5p^5$

Ionization energy , Electron affinity and Electronegativity

The amount of energy needed to remove electron(or electrons) from the outer shell from an atom called *ionization energy.*

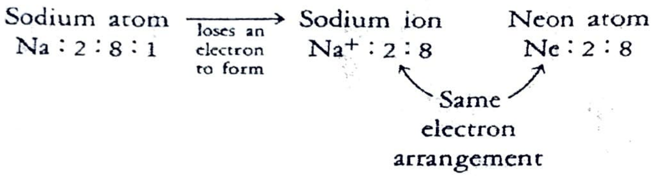
The alkali metals (lithium , sodium , and potassium) lose an electron most easily whearse the noble gases are the most relutant of all the element to lose an electron(noble gases has no ability to lose electron ,noble gases are $_{2}He$,$_{10}Ne$,$_{18}Ar$,$_{36}Kr$,$_{54}Xe$ and $_{86}Rn$).**Electron affinity** is the energy released when atom gains an electron .It is a measure of the ability of an atom to attract an electron. The noble gases are the most relutant of all the element to accept an electron(noble gases has no ability to accept electron).The noble gases have high ionization energy and low electron affinity.That means that atoms have no ability to lose or gain an electron.Sodium atom easily loses an electron to form a positivly charged sodium ion.A sodium ion has 11 proton but only 10 electrons therefore ,it has a net positive charge +1 (Na^+) the sodium loses electron to form the noble gas electron arrangement. Fluorine atom easily accepts an electron to form a negatively charged fluoride ion.A fluoride ion has 9 proton but only 10 electrons therefore ,it has a net negative charge -1 (F^-) the fluoride gains electron to form the noble gas electron arrangement.

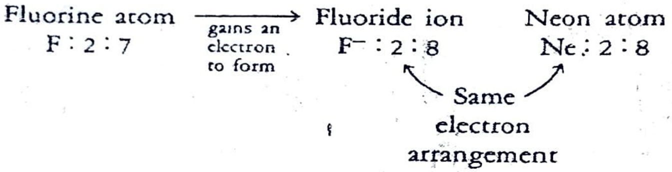
Noble gases unactive because this atoms are stable and has no ability to gain or lose an electron from their outer shells.All atoms may be lose an electron (or electrons) or gain an electron (or electrons) to take noble gases arrangement .

${}_{11}Na$: $1s^2\,2s^2\,2p^6\,3s^1$ (Na^+:$1s^2\,2s^2\,2p^6\,3s^0$)

${}_{9}F$:$1s^2\,2s^2\,2p^5$ (F^-:$1s^2\,2s^2\,2p^6$)

Another example oxygen atom can gain two electrons to form noble gases arrangement (Ne) while Mg can lose two electrons to form noble gases arrangement (Ar).





Electronegativity.

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons to themselves . The Pauling scale is the most commonly used. Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to cesium and francium which are the least electronegative at 0.7.An ion is an atom or molecule that carries an electric charge.Ions are identified by the use of a superscript to show the sign and size of their electric charge.

octet rule

The octet rule is a chemical rule that atoms of main-group elements tend to combine in such a way that each atom has eight electrons in its valence(outer) shell, giving it the same electron configuration as a noble gas. Atoms gain or lose electrons from their outer shell to achieve the noble gas electron arrangement(except helium).

Electrolytes:

Are minerals in your body that have an electric charge. They are in blood, urine and body fluids. Maintaining the right balance of electrolytes helps your body's blood chemistry, muscle action and other processes.

Electrolytes : An element or compound that, when melted or dissolved in water or another solvent, dissociates into ions and is able to conduct an electric current. Levels of electrolytes in our bodies can become too low or too high. That can happen when the amount of water in your body changes, causing dehydration or over-hydration. Causes include some medicines, vomiting, diarrhea, sweating or kidney problems.

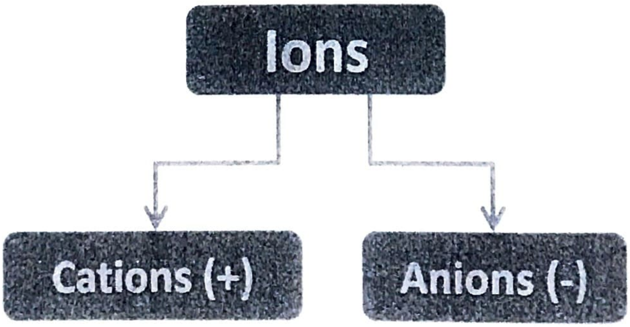
Fluid balance : The amount of water gained each day equals the amount lost. **Electrolyte balance**: The ion gain each day equals the ion loss. **Acid-base balance**: H^+ gain is offset by their loss.

ıntracellular Fluids(ICF): All fluids inside cells.

Extracellular Fluids (ECF) : All fluids outside cells

Anions and Cations

Negatively charged ions are called **anions**, and positively charged ions are called **cations** that is carrying out by forming chemical bond.



Most common cations and anions in the living system are:

Cations	Anions
Na^+	Cl^-
K^+	$H_2PO_4^-$
Mg^{+2}	$HPO4_2^-$
Ca^{+2}	HCO_3^-
Fe^{+2}	--
Fe^{+3}	--

Calcium is the most abundant mineral in the body. Calcium salts are structural components of bones and teeth. Ca^{+2} which are primarily extracellular cations, function in blood clotting, neurotransmitter release, and contraction of muscle. Ca^{+2} level is controlled by parathyroid and calcitrol. • In general, imbalances in calcium concentrations affect the bones, kidney and gastrointestinal tract. Calcium also influences the permeability of cell membranes and thereby regulates neuromuscular activity

$_{20}Ca$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$ or $_{20}Ca$:$_{18}[Ar]\ 4s^2$
Ca^{+2}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0$ or Ca^{+2}:$_{18}[Ar]\ 4s^0$

Sodium affects the osmolality of blood and therefore influences blood volume and pressure and the retention or loss of interstitial fluid.

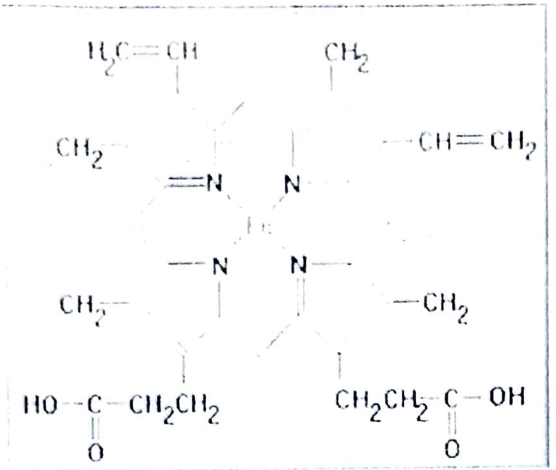
Potassium affects muscular activities, notably those of the heart, intestines and respiratory tract, and also affects neural stimulation of the skeletal muscles. •

Mg^{+2} ions -intracelluar cations that act as cofactors for enzymes.

Phosphate ions are principally intracellular anions and their salts are structural components of bones and teeth, also required for the synthesis of nucleic acids and ATP and participate in buffer reactions. Level is controlled by PTH and calcitrol.

A traces of certain metallic cations are needed to maintain life, like presence of ferrous ion Fe^{+2} in hemoglobin which play important role in transport of oxygen and carbon dioxide:

$Fe^{+2} \leftrightarrow Fe^{+3}$



Pb^{+2} have toxic effect to kidney and also cause nerve damage, while Hg^{+2} cause damage to brain and nerve system.

Free radical

Any element contains a single electron in its outer shell leads to damage for nucleic acids,proteins and lipids in the cell membrane and plasma lipoprotein and can cause cancer.this elements with highly reactive atoms with unpair electrons are called free radicals.Not all free radical is single atom may be compound with atom carry unpair electron.

Pollution

Pollution is the introduction of contaminants into a natural environment that causes instability, disorder, harm or discomfort to the Ecosystem i.e. physical systems or living organisms. • Pollution can take the form of chemical substances or energy, such as noise, heat, or light

Pollutant: A pollutant is a waste material that pollutes air, water or soil. Three factors determine the severity of a pollutant: its chemical nature, the concentration and the persistence

Environment is the land, water and air in which people, animals and plants live. Simply what is around us. The air we breathe, the soil on which we stand and walk and the water we drink are all parts of the environment..

Dump is a place where unwanted waste is taken and left. •

To endanger means to put someone or something in a dangerous or harmful situation.

Different kinds of pollution are found.

- Air Pollution.
- Water Pollution.
- Land Pollution.

Air pollution

contamination of the air by harmful gases and minute particles of solid and liquid matter (particulates) in concentrations that endanger health • Air pollution only occurs outdoors

Air pollution is the accumulation in the atmosphere of substances that, in sufficient concentrations, endanger human health or produce other measured effects on living matter and other materials.

Among the major sources of pollution are power and heat generation, the burning of solid wastes, industrial processes, and, especially, transportation. The six major types of pollutants are carbon monoxide, hydrocarbons, nitrogen oxides, particulates, sulfur dioxide, and photochemical oxidants

ORGANIC AIR POLLUTANTS

Acrylonitrile, Benzene, Butadiene, Carbon disulfide, Carbon monoxide, 1,2-Dichloroethane, Dichloromethane, Formaldehyde, Polycycli aromatic hydrocarbons (PAHs) Polychlorinated biphenyls (PCBs), Polychlorinated dibenzodioxins Styrene, Tetrachloroethylene, Toluene, Trichlorethylene, vinylchloride

INORGANIC AIR POLLUTANTS

Arsenic, Asbestos, Cadmium, Chromium, Fluoride, Hydrogen sulfide, Lead, Manganese, Mercury, Nickel, Platinum, Vanadium

CLASSICAL AIR POLLUTANTS

Nitrogen dioxide, Ozone and other photochemical oxidants, Particulate matter, Sulfur dioxide

What’s in smog

- particulates (especially lead)
- nitrous oxides

- potassium
- Carbon monoxide
- Other toxic chemicals

Smog and haze can reduce the amount of sunlight received by plants to carry out photosynthesis and leads to the production of ozone which damages plants. Ozone pollution can cause respiratory disease, cardiovascular disease, throat inflammation, chest pain, and congestion.

No_2 and O_3 – that are responsible for the buildup of smog.

$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$

NO when release into the atmosphere

$2NO_{(g)} + O_2 \rightarrow 2NO_{2(g)}$ oxidation

$$NO_{2(g)} + hv \xrightarrow{\lambda < 400\ nm} \underset{reactive}{NO_{(g)} + O_{(g)}}$$

$O_{(g)} + O_{2(g)} + M \rightarrow \underset{Ozone}{O_{3(g)}} + M$ M is the inert substance such as N_2

<u>Carbon dioxide and carbon monoxide</u> -1

Both CO_2 & CO are products of combustion. CO_2 is formed in abundant of oxygen. CO & CO_2 is formed in a limited supply of oxygen.

- Carbon dioxide is not a toxic gas, but it does have an asphyxiating effect. In air tight buildings, the concentration of CO_2 can reach as high as 2000 ppm by volume (3ppm outdoor).

Workers exposed to high [CO_2] become fatigued more easily and have difficulty concentrating. Adequate ventilation is the solution to CO_2pollution.

- CO is a colorless and odorless gas and is highly poisonous. The toxicity of CO lies in its unusual ability to bind very strongly to hemoglobin. Both O_2 and CO bind to the Fe II ion in hemoglobin but the affinity of hemoglobin for CO is about 200 in times greater than that for O_2. Carboxyhemoglobin cannot carry the oxygen needed for metabolic processes. CO can cause drowsiness and headache, death may result when half the hemoglobin molecules are complexed with CO. the best first aid response to CO poisoning is to remove the victim immediately to an area with a plentiful oxygen supply.

- *Radon (Rn)*

Rn is a member of group 8A (noble gases), it's an intermediate product of the radioactive decay of Uranium – 238.

All isotopes of radon are radioactive, but radon – 222 is the most hazardous because it has the longest $t_{1/2}$ = 3.8 days. Radon is generated mostly from the phosphate minerals of Uranium.

$${}^{226}_{88}\mathrm{Ra} \rightarrow {}^{222}_{86}\mathrm{Rn} + {}^{4}_{2}\mathrm{He}$$

Radon is found in higher concentrations with Uranium, this is not surprising, since radium is formed as a part of the stepwise decay of uranium.

Radon is radioactive but its radiation is not the major problem. Since it is a gas and chemically inert, it is rapidly exhaled after breathing. Radon decays to Polonium:

$${}^{222}_{86}\mathrm{Rn} \rightarrow {}^{218}_{84}\mathrm{PO} + {}^{4}_{2}\mathrm{He}$$

Polonium isotope is radioactive with long half-life, and is a nonvolatile heavy metal that can attach itself to bronchial or lung tissue and remain for a long time, emitting hazardous radiation.

●

Exhaust Gases of Vehicles

- Pollution from exhaust gases of vehicles is responsible for 60% of all air pollution and in cities up to 80%. There is a large species of harmful chemicals present in these gases, with lead being one of the most dangerous. Motor vehicle emissions are one of the leading causes of air pollution. China, United States, Russia, Mexico and Japan are the world leaders in air pollution emissions.

Combustion of Coal

- The combustion of coal without special precautions can have serious consequences. If winds do not blow away the poisonous gases, they can have fatal effects and may lead to death.

Others Examples of Air Pollution

• Noise Pollution

- Noise pollution or unwanted sounds that are carried by the air, have an irritating and detrimental effect on humans and other animals. Careful planning of streets and buildings in towns and better control over noisy vehicles may add to the control of noise pollution. Noise pollution induces hearing loss, high blood pressure, stress, and sleep disturbance.

Tobacco Smoke

- Tobacco smoke is one of the major forms of pollution in buildings. It is not only the smoker who is infected, but everyone who inhales the polluted air. There is a very strong connection between smoking and lung cancer.

Water Pollution

- Water pollution is the introduction into fresh or ocean waters of chemical, physical, or biological material that degrades the quality of the water and affects the organisms living in it.

Water pollution causes approximately 14,000 deaths per day, mostly due to contamination of drinking water by untreated sewage.

Kinds of Water Pollution

Presence of organic substances degraded by microorganisms .1 (food),so it consume O_2 which reduce O_2 in water that cause a death of many animals living in water.

Microorganism + O_2 H_2S (odour)

From nitrate

In water

2. Presence of Hg in water lead to poisonous of fish human.some microorganism change Hg^+ to CH_3Hg^+ soluble.

3. Suspense particles.

Radioactive isotopes from reactors which contaminate water and underground water

Land Pollution

- Land pollution is the degradation of the Earth's land surface through misuse of the soil by poor agricultural practices, industrial waste dumping, and indiscriminate disposal of urban wastes.

Soil Pollution

- Soil pollution is mainly due to chemicals in herbicides (weed killers) and pesticides (poisons which kill insects). Litter is waste material dumped in public places such as streets, parks, picnic areas, at bus stops and near shops.

Radiation Pollution

Radioactive waste , ionizing radiation from nuclear reactors and nuclear reactions produce radiation such as α ,β ,γ,neutron,this may lead to many reactions that produce free radicals,electrons ,positive species.These are all active and cause to form new species either in air ,solution ,water and soil

Chemistry lab:6 **First stage**

Action strong acid and base on Indicators

Indicators:

There are many substances that show one colour in an acidic medium and another colour in a basic medium. Such substances are called **acid-base indicators**.

An acid/base indicator is a weak organic acid or a weak organic base whose un dissociated form differs in color from its conjugate base or its conjugate acid form.

Litmus is a natural dye found in certain lichens. It was the earliest indicator to be used. It shows red colour in acidic solutions and blue colour in basic solutions. Phenolphthalein and methyl orange are some other indicators.

Procedure:

1- Take about 2 mL solution of acid or base in a test tube.

2- Add a drops of indicator to it and observe the colour. **What to observe?**

3- Tabulate your observations in Table.

Sample solution	**red litmus**	**blue litmus**	**phenolphthalein**	**methyl red**	**methyl orange**

Chemistry lab:9 **First stage**

"Experiment for preparing of ethyne (acetylene) gas in the laboratory"

Chemical properties of alkynes or (acetylenes):

1. Acetylenes are unsaturated hydrocarbons that have carbon atoms in chains linked by one or more triple bond (C≡C).
2. Acetylenes have a formula C_nH_{2n-2}; R-C≡C-H and R-C≡C-R where R=alkyl group, acetylene formula C_2H_2 (H-C≡C-H).
3. Acetylenes have two functional groups (C≡C) and acidic hydrogen atom (R-C≡C-H).

Experiment:

Ethyne gas or (acetylene) is prepared by the reaction of water with calcium carbide (CaC_2). Acetylene is a colorless gas.

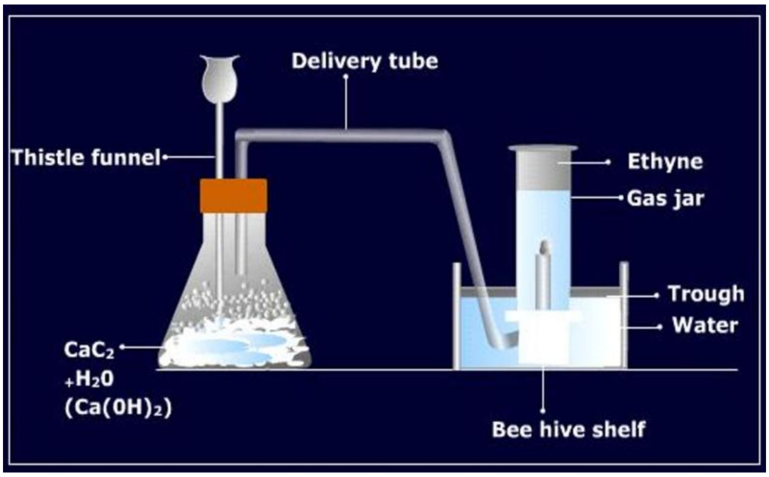
$$CaC_2 + 2H_2O \rightarrow H\text{-}C{\equiv}C\text{-}H + Ca(OH)_2$$

Main Chemicals required:

1. Calcium Carbide (CaC_2) [also known as calcium dicarbide] is a white/gray solid.
2. Water H_2O.

Procedure

1. Set up the apparatus as shown in the diagram.
2. Place few pieces of calcium carbide in a Büchner flash.
3. Add water from the dropping funnel, few drops at a time.
4. Collect the gas produced in the test tube.
5. Discard the first test tube as it a mixture of air and ethyne.



After preparing ethyne now we study some physical properties of ethyne

1-Combustion:

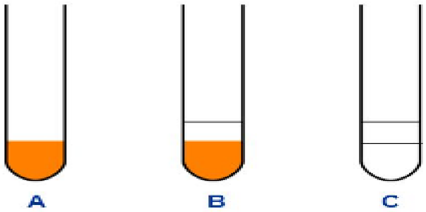
The ethyne gas burns with a **luminous smoky flame** as a result of unburned carbons.

$$2C_2H_2+5O_2 \longrightarrow 4CO_2+2H_2O$$

2-Addition of bromine test:

Addition of bromine water is presence of a triple bond C≡C (unsaturated compound), turns the color from ***yellow/red color*** to ***colorless***. The decolorisation of bromine is the standard test for **unsaturated** in a compound.

A. Place bromine water in the test tube.
B. Add the hydrocarbon to the test tube and shake
C. If the bromine color disappears then the hydrocarbon is unsaturated.



$$H\text{-}C{\equiv}C\text{-}H \quad + \quad 2Br_2 \longrightarrow CHBr_2\text{-}CHBr_2$$

Yellow –red *Colorless*

3-Addition of acidified potassium permanganate ($KMnO_4$) test:

This is also used to test for presence of a triple bond C≡C (unsaturated compound), the color turn from purple to colorless.

A. Place the hydrocarbon to the test tube.

B. Add $KMnO_4$ solution in the test tube and shake.
C. If the $KMnO_4$ color disappears then the hydrocarbon is unsaturated.

$$R\text{-}C\equiv CH \xrightarrow[H_3O^+]{KMnO4} R\text{-}COOH + CO_2$$

$$R\text{-}C\equiv CR' \xrightarrow[H_3O^+]{KMnO4} R\text{-}COOH + R'\text{-}COOH$$

4-Addition of tollen reagent [$Ag(NH_3)_2OH$] test:

The tollen reagent $Ag(NH_3)_2OH$ turns from **colorless** to **white precipitate.**

A. Place the alkyne to the test tube.
B. Add some drops of $Ag(NH_3)_2OH$ solution to the test tube and shake .
C. If the solution colorless turns to white precipitate then the alkyne has acidic hydrogen atom (***terminal Hydrogen atom***)

$$H\text{-}C\equiv C\text{-}H + Ag(NH_3)_2OH \rightarrow H\text{-}C\equiv C\text{-}Ag + 2NH_3 + H_2$$

(Terminal Hydrogen atom)

$$R\text{-}C\equiv C\text{-}R + Ag(NH_3)_2OH \rightarrow N.R$$

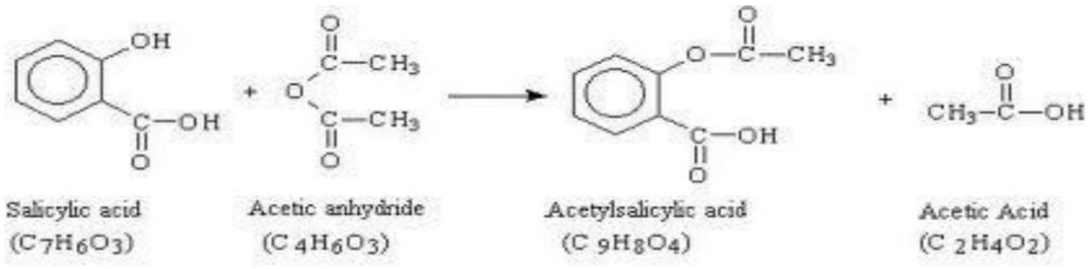
(No terminal Hydrogen atom)

Chemistry lab:10 **First stage**

Aspirin synthesis

Principle

Aspirin is formed by reaction of salicylic acid with acetic anhydride in the presence of sulfuric acid and heat according to the equation:



Acetic anhydride serves as both a reactant and a solvent.

Reagents:

Salicylic acid, acetic anhydride, sulfuric acid.

Procedures:

1-Weight 2 gm of salicylic acid (2- hydroxyl benzoic acid). Place the solid in a test tube and take it to the hood.

2-Carefully add 3 ml of acetic anhydride and then 5 drops of sulfuric acid to the solid, and shake well for 5 minutes. (sometimes product forms at this stage).

3-Place the test tube in a water bath (60 C) and heat, with intermetten shaking, for 10 minutes.

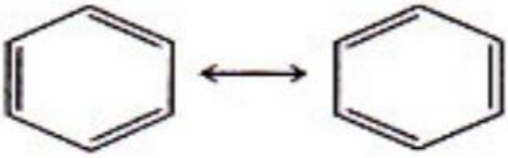
4-Place the tube in an ice water bath. Add 5 ml for distilled water and stir for 5 minutes (keep the tube in the bath while stirring)

6- Leave the crystals to dry overnight

7-Determine the weight of the crystals.

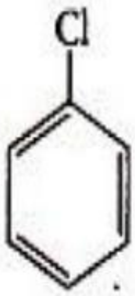
Aromatic Compounds

Benzene has a high degree of unsaturation but is unreactive.It does not undergo the usual addition reactions of alkenes and alkynes. It does react under more vigorous conditions(heat or catalysts) to form products of substitution. This difference in chemical reactivity is due to the structure of benzene.

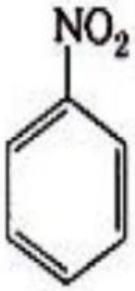


Naming Aromatic Compounds

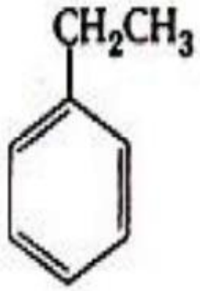
When one group is attached to benzene, the compound is named by placing the name of the group as prefix to the benzene.



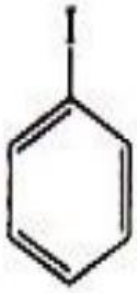
Chlorobenzene



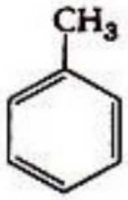
Nitrobenzene

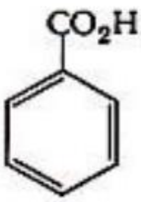


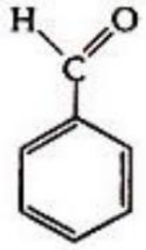
Ethylbenzene

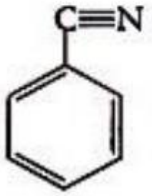


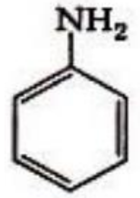
Iodobenzene

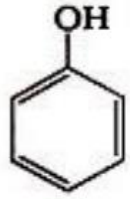

Toluene

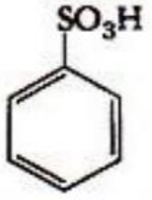

Benzoic acid

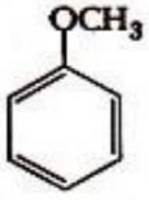

Benzaldehyde

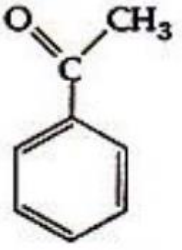

Benzonitrile


Aniline

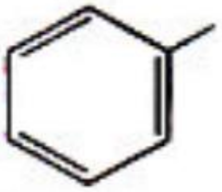

Phenol

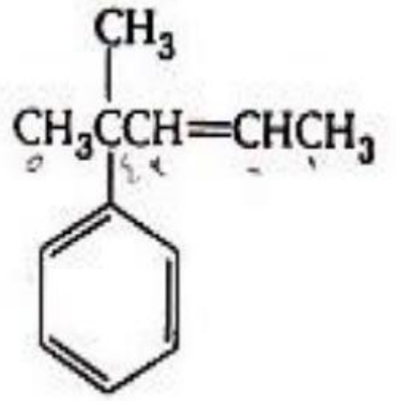

Benzenesulfonic acid

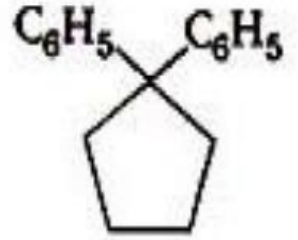

Anisole

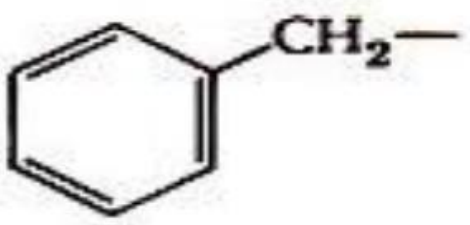

Acetophenone

C_6H_5— or


Phenyl group

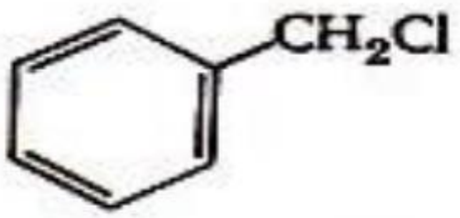

4-Methyl-4-phenyl-2-pentene


1,1-Diphenylcyclopentane

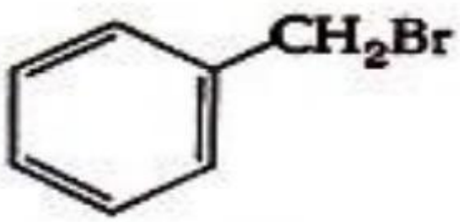


Benzyl group

For example:

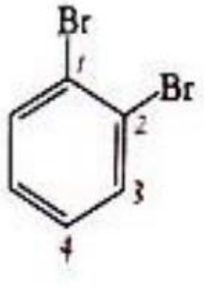


Benzyl chloride

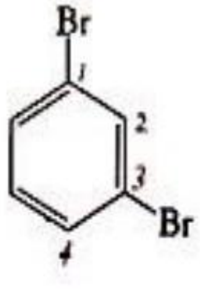


Benzyl bromide

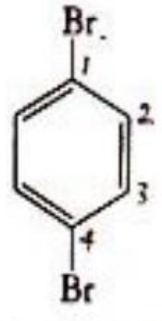
If two groups are attached to the benzene ring, the name must not only tell what groups are present, but also where they are located. We can differentiate the three possible isomers of a disubstituted benzene in two ways.



o-Dibromobenzene
1,2-Dibromobenzene

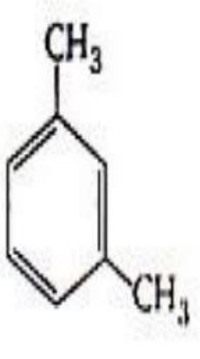


m-Dibromobenzene
1,3-Dibromobenzene

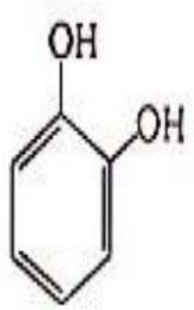


p-Dibromobenzene
1,4-Dibromobenzene

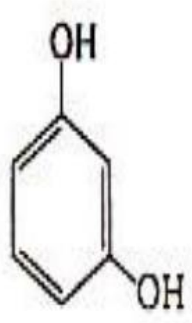
Several disubstituted benzenes have been given names that, give no indication of the kind of groups attached to the ring.



m-Xylene



Catechol



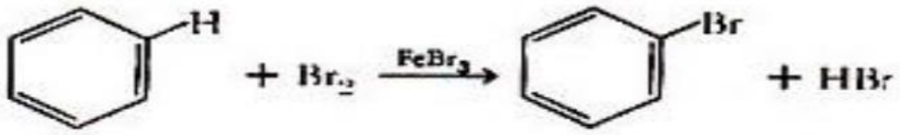
Resorcinol

Hydroquinone

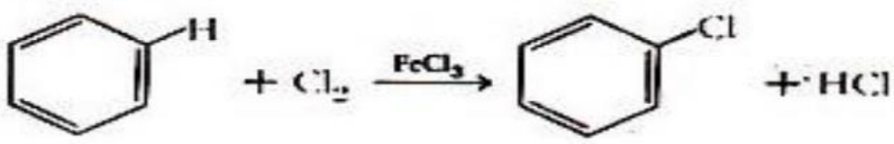
p-Cresol

Substitution Reactions Of Aromatic Compounds

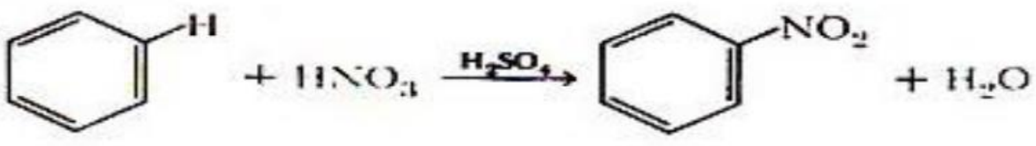
Bromination

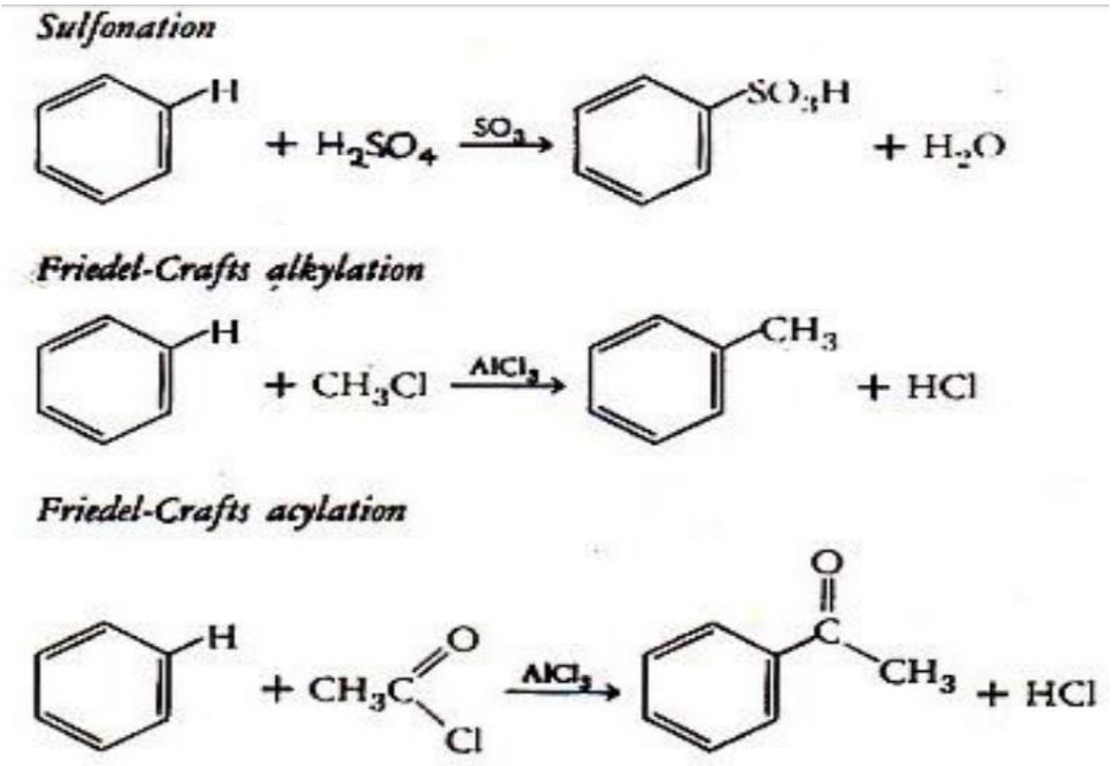


Chlorination



Nitration





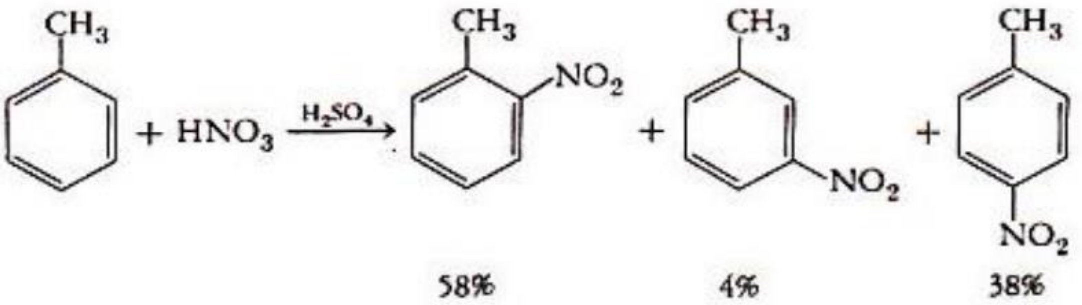
Sulfonation

Friedel-Crafts alkylation

Friedel-Crafts acylation

Toluene,nitrobenzene,chlorobenzene and many other aromatic compounds also undergo substitution reactions. These compounds differ from benzene in that they already have a substituent on the benzene ring. The remaining five hydrogens are not equivalent. When these hydrogens undergo substitution reactions, isomers can be formed depending on which hydrogen is replaced.

Examples:



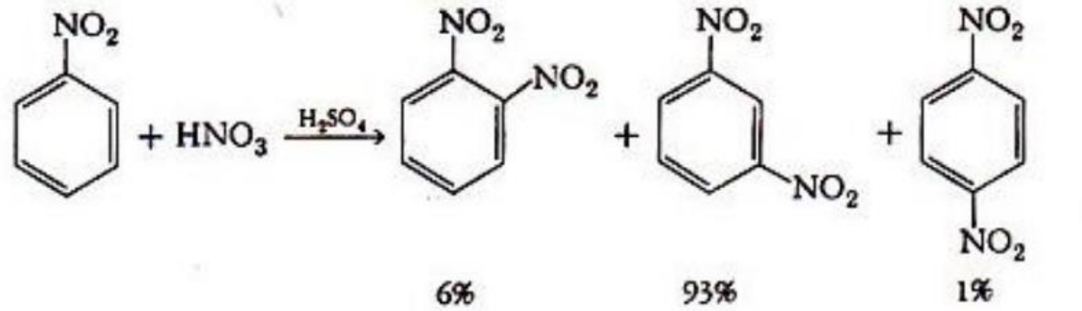
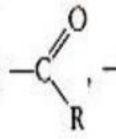
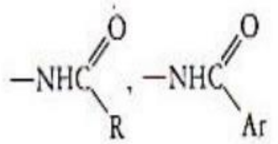


Table 13-1. Directing Effect of Substituents[a]

Predominant Directing	
o, p	m
$-NH_2$, $-NHR$, $-NR_2$	$-NO_2$
$-NH_2$, $-NHAr$, $-N(Ar)_2$	$-CN$
$-OH$, $-OR$, $-OAr$	$-CO_2H$, $-CHO$,  $-C(=O)R$, $-CO_2R$, $-CO_2Ar$
 $-NHC(=O)R$, $-NHC(=O)Ar$	$-SO_3H$
Alkyl groups	$-\overset{+}{N}H_3$, $-\overset{+}{N}H_2R$, $-\overset{+}{N}HR_2$, $-\overset{+}{N}R_3$
Halogens	$-\overset{+}{N}H_3$, $-\overset{+}{N}H_2Ar$, $-\overset{+}{N}H(Ar)_2$, $-\overset{+}{N}(Ar)_3$

[a] R = alkyl groups; Ar = aryl groups.

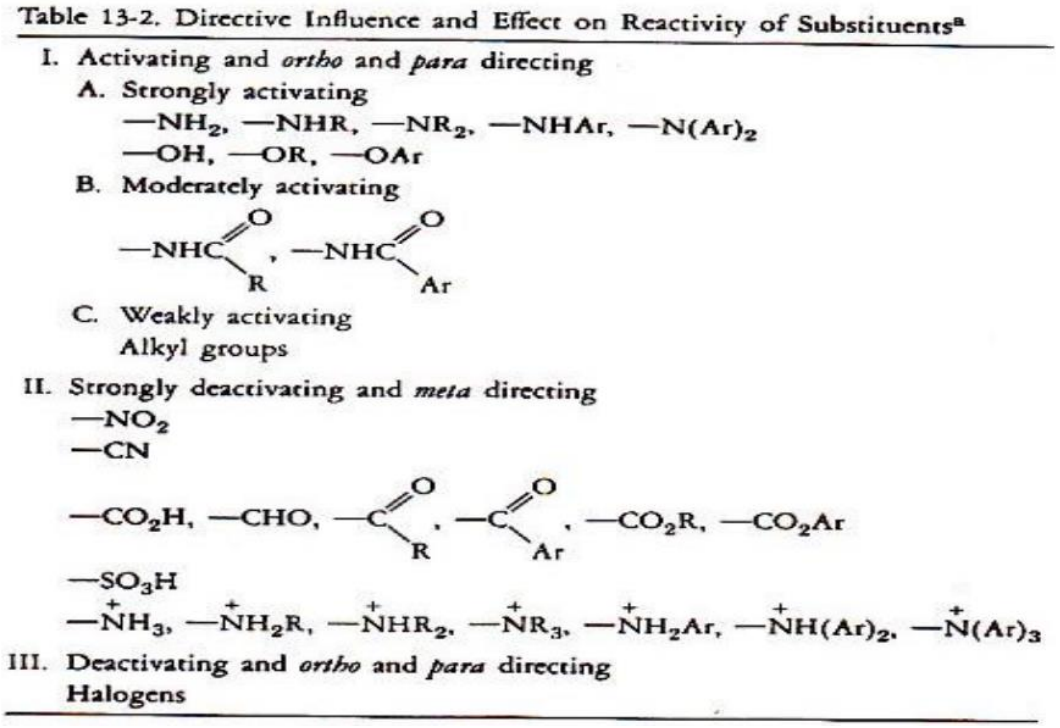


Table 13-2. Directive Influence and Effect on Reactivity of Substituents[a]

I. Activating and *ortho* and *para* directing

- A. Strongly activating
  - $-NH_2$, $-NHR$, $-NR_2$, $-NHAr$, $-N(Ar)_2$
  - $-OH$, $-OR$, $-OAr$
- B. Moderately activating
  - $-NHC(=O)R$, $-NHC(=O)Ar$
- C. Weakly activating
  - Alkyl groups

II. Strongly deactivating and *meta* directing

$-NO_2$

$-CN$

$-CO_2H$, $-CHO$, $-C(=O)R$, $-C(=O)Ar$, $-CO_2R$, $-CO_2Ar$

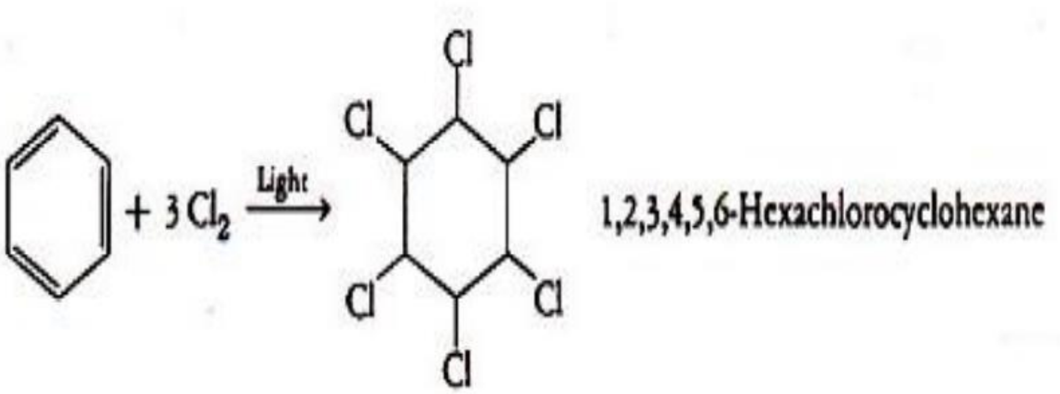
$-SO_3H$

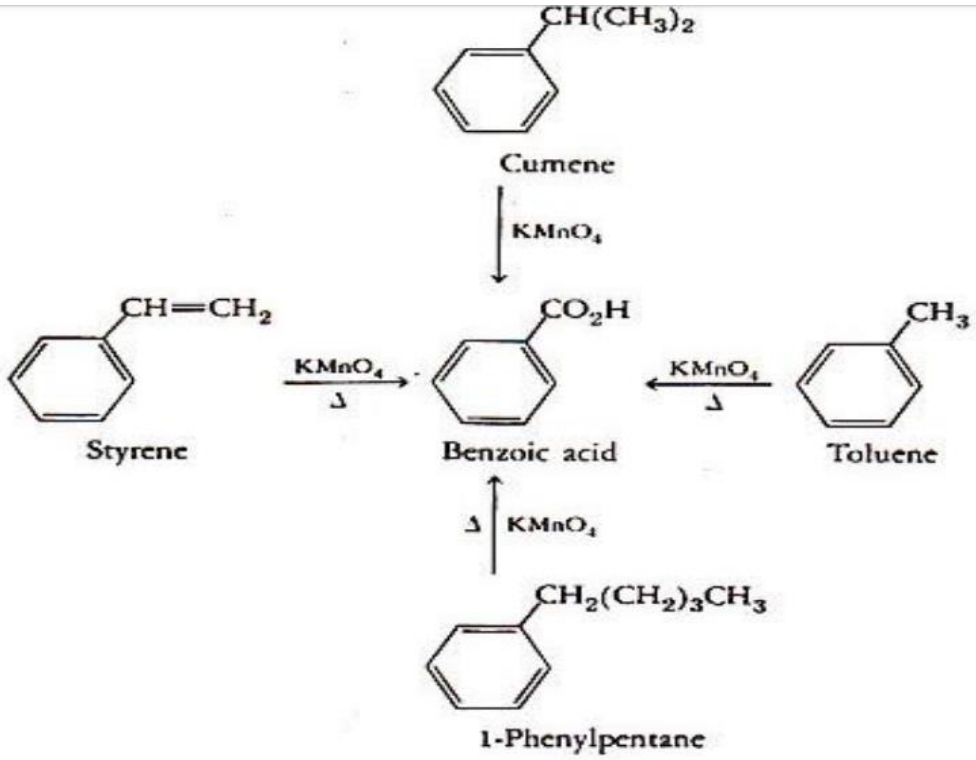
$-\overset{+}{N}H_3$, $-\overset{+}{N}H_2R$, $-\overset{+}{N}HR_2$, $-\overset{+}{N}R_3$, $-\overset{+}{N}H_2Ar$, $-\overset{+}{N}H(Ar)_2$, $-\overset{+}{N}(Ar)_3$

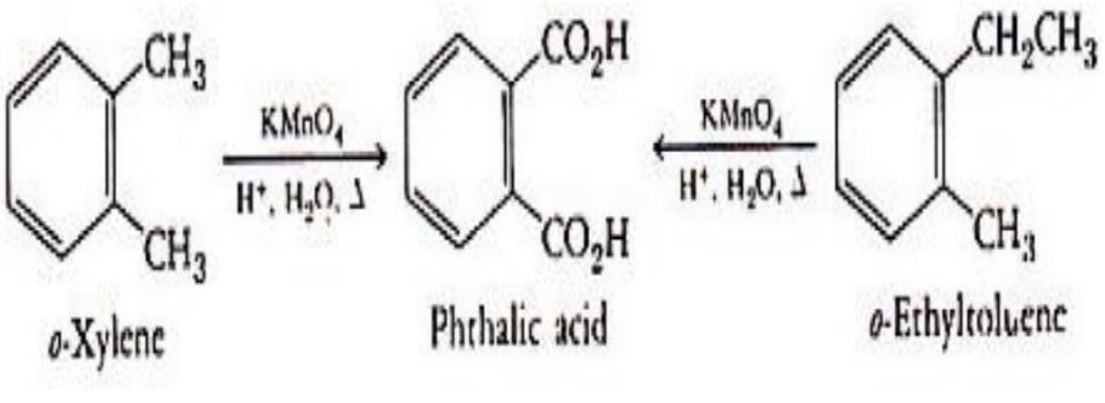
III. Deactivating and *ortho* and *para* directing

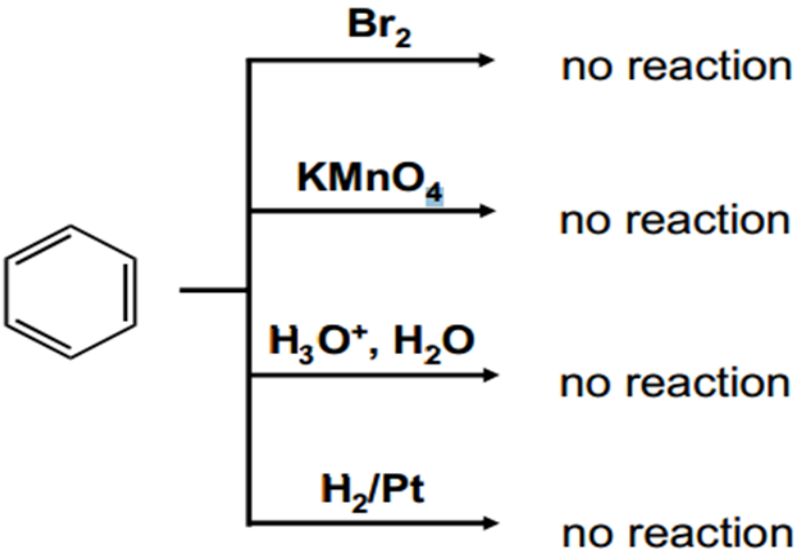
Halogens

Other Reactions Of Aromatic Compounds

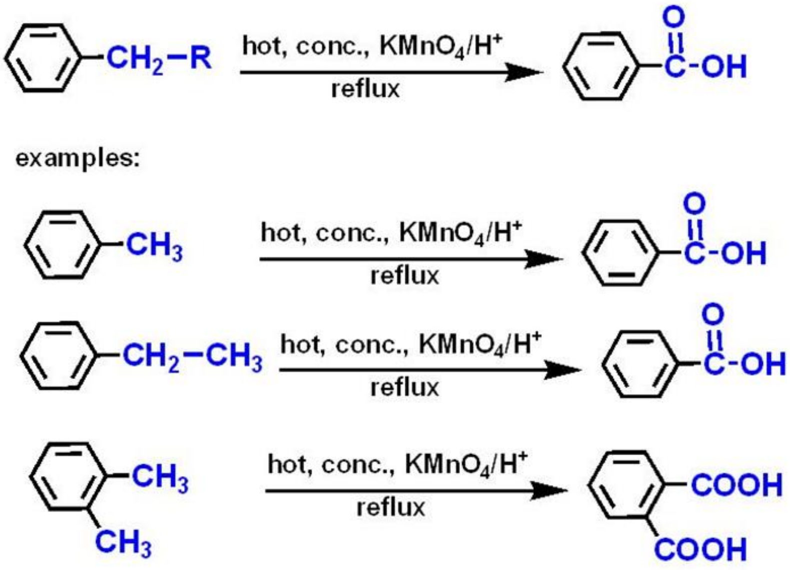



CH(CH3)2
Cumene
KMnO4
CH=CH2
KMnO4
Δ
CO2H
KMnO4
Δ
CH3
Styrene
Benzoic acid
Toluene
Δ
KMnO4
CH2(CH2)3CH3
1-Phenylpentane


CH3
KMnO4
H+, H2O, Δ
CO2H
KMnO4
H+, H2O, Δ
CH2CH3
CH3
CO2H
CH3
o-Xylene
Phthalic acid
o-Ethyltoluene



OXIDATION REACTION OF ALKYLBENZENE



Complete the following reactions

1-Benzonitrile + Br_2 / $FeBr_3$
2-Acetophenone + HNO_3 /H_2SO_4
3-Anisole + CH_3Cl /$AlCl_3$
4-Phenol + H_2SO_4/ SO_3

Aromatic compounds in natural

1-benzen is toxic compound

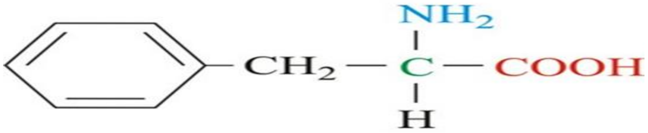
2-It cause severe liver damage.

3-It is possible carcinogen(cause cancer).

4-It is not found in natural .

Many compounds that have a benzene ring in their basic structure such compound are called derivatives of benzene

1-Phenylalanine(essential amino acid) a compound essential for life contains a benzene a ring and is derivative of benzene



This compound cannot be made in the human body and must be obtained from food we eat .However ,plants can make phenylalanine (**called essential amino acids**).Phenylalanine can be used to prepare other important compounds such as hormones epinephrine and Dopa.

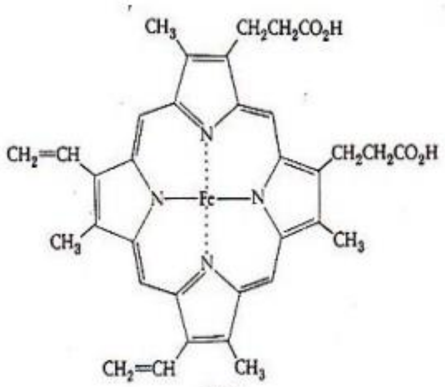
Phenylalanine is transformation into tyrosine(isnot essential amino acid)

Aromatic Heterocyclic compounds

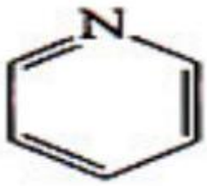
Resemble benzene in they are highly unsaturated yet undergo electrophilic substitution ,They contain at least one element other than carbon or hydrogen usually nitrogen.They contain 6 electrons in ring and heteroatom.

1-Four pyrrole rings joined together by CH bridge to form porphyrins that their ability to bind metal ions in space between the four nitrogen atoms

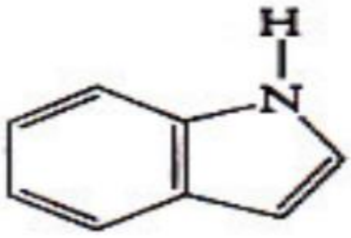
Hemoglobin is composed of two parts .One parts protein portion called globin.The other part is porphyrin part called heme .Heme contains an iron atom in ferrous oxidation state in space between the nitrogen atoms.



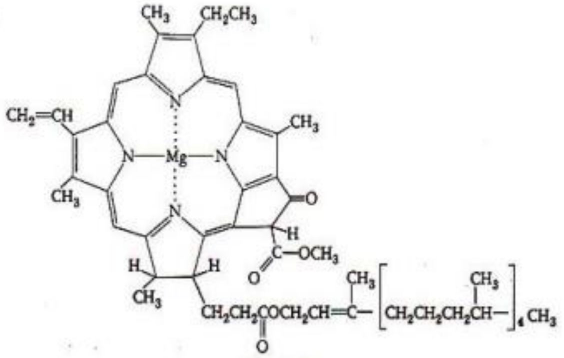
Heme



Pyridine



Indole



Chlorophyll a

2-pyridine

Pyridine can be regarded as benzene in which one carbon has been replaced by a nitrogen atom .

Some exampleNicotinamide ,Nicotine ,Nicotinic acid and pyridoxine (vitamin B_6)

3-Indole

Indole is an aromatic heterocyclic compound containing a pyrrole ring and benzene ring with one side in common .The single most important derivative of indole in living system is the amino acid tryptophane it is essential amino acid for mammals must get it from foods.Many important compounds are prepared from tryptophane in living systems .One of them is serotonin,an important hormone that affects blood pressure and is a factor in the functioning of brain.

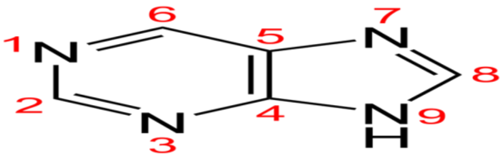
4-Imidazole is a compound that contains tow nitrogen atoms in five membrane ring

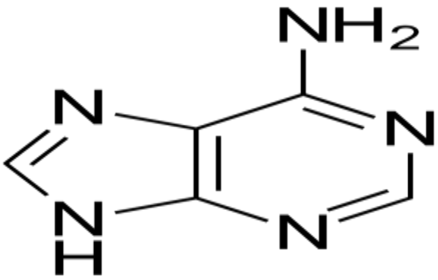
The most important derivative of imidazole is the amino acid histidine .

In living systems enzyme transform histidine to the pharmacologically active histamine

Histamine which causes dilation of capillaries and stimulation of muscle

5-Purine is a compound that contains two heteroaromatic rings that have one side in common .A derivative of purine with an $-NH_3$ (amine)substitutuent in the 6 position is called adenine .It is part of adenine triphosphate (ATP).





pH of solutions

The pH of solution is defined as following

$$pH = -\log_{10}[H^+] \qquad [H^+] = 10^{-pH} \text{ (in mol/L)}$$

$$pOH = -\log_{10}[OH^-] \qquad [OH^-] = 10^{-pOH} \text{ (in mol/L)}$$

$$pH + pOH = 14 \qquad [H^+] \times [OH^-] = 10^{-14} \text{ (in mol}^2\text{/L}^2\text{)}$$

$$K_w = [H^+] \times [OH^-] \qquad \text{At } 25^{o}C,\ K_w = 10^{-14}\ mol^2/L^2$$

The pH of some common Body Fluids

The various body fluids differ both in their,acidity and in their range of acidities.Stomach acid is the most acid is the most aciditic ,and it has a narrow pH rang(1 to 3).Blood plasma is slightly basic,and it has a very narrow pH range .If the pH of blood changes to a value outside this range,the ability of the blood to transport oxygen is reduced.Therefore,maintaining the pH of blood plasma within a narrow range is important to human life .Urine has a wide pH range .It can be acidic ,basic or neutral.This wide range of the pH of urine is due to many acidic and basic substances are removed from the body through the urine to help maintain the pH of blood plasma.A list of the pH values of body fluids given in the following table.

Fluid	pH range
Blood	7.35-7.45
Gastric juices	1.6-1.8
Bile	7.8-8.6
Urine	5.5-7.0
Saliva	6.2-7.4

Buffer solution

A buffer solution is an mixture of either a weak acid plus a salt of this weak acid or weak base plus a salt of this weak base .Such a mixture reacts with both acids or strong base cause little change in its pH.For example: acetic acid and sodium acetate dissolved in water is an example of solution. Weak acid molecules ($C_2 H_3 O_2 H$) and conjugate base of the the acetic ion of sodium acetate ($C_2H_3O_2^-$).Now if small amount of strong acid added to the buffer solution it will react with the conjugate base as follows:

$$H_3O^+ + C_2H_3O_2^- \underline{\hspace{3cm}} C_2H_3O_2H + H_2O$$

If hydroxide ions added to the buffer solution they will react with molecules of acetic acid to form acetate ions and water as following :

$$\mathbf{OH^- + C_2H_3O_2H \longrightarrow C_2H_3O_2^- + H_2O}$$

The pH of a buffer solution is determined by the pKa of weak acid and log of the ratio of the concentration of the conjucate base to the concentration of the acid

Henderson-Hasselblach equation

$$PH= PKa + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$$

A small change in the concentration of either acetate or acetic acid will hardly change the PH.

A buffer solution has a limited ability to react with acids and bases without changing its PH.

A solution act as a buffer because its contains both members of a conjucate acid-base pair. Remove one of these two by either a chemical or physical process destroy the buffer action of the solution.

Biological Buffer Significance

1- Ability to prevent large change in pH is an important property of most intact biological organism.

2- Blood plasma is highly effective buffer solution and designed to keep the range of pH of blood within (0.2) pH unit from (7.35-7.45) outside this range are compatable with life.

3- Many of metabolites constantly produced and utilized in the cell are weak acids.

***BLOOB BUFFER*:**

The blood retains its fairly constant pH because of the presence of buffer .These buffer present both in blood plasma and in red blood cells
In plasma are sodium buffer,in red blood cells potassium buffers
The blood buffer consist the following:

1- ***BicarbonateBuffers***

The bicarbonate buffer system in the red blood cell consists of carbonic acid (H_2CO_3),and potassium bicarbonate ($KHCO_3$).The bicarbonate buffer system in the blood plasma consists of carbonic acid and sodium bicarbonate ($NaHCO_3$).If we assume that a strong acid (such as HCl) is added to a sample of blood,it will react with the salt part of the buffer and undergo the following reactions

$$HCl + KHCO_3 \xrightarrow{\hspace{2cm}} H_2CO_3 + KCl \text{ (in blood cells)}$$

$$HCl + NaHCO_3 \xrightarrow{\hspace{2cm}} H_2CO_3 + NaCl \text{ (in blood plasm}$$

The carbonic acid produced (H_2CO_3) is part of the original buffer .Note that the strong acid (HCl).has been replaced by very weak one (H_2CO_3).The other products (KCl and NaCl) are neutral salts and will not affect the pH of systems:

$$KOH + H_2CO_3 \xrightarrow{\hspace{2cm}} KHCO_3 + H_2O \text{ (in the blood cells)}$$

$$NaOH + H_2CO_3 \xrightarrow{\hspace{2cm}} NaHCO_3 + H_2O \text{ (in the blood plasma)}$$

The salts ($KHCO_3$ and $NaHCO_3$) are part of the original buffer system and the water produced is neutral. So that the pH again is unaffected in both cases (reaction with a strong acid or a strong base) more of the buffer is produced plus a neutral compound.

The bicarbonate buffers and the blood protein buffers play a major part in the control of the pH; the phosphate buffers have an important role inside the cell and in the urine.

2- Phosphate Buffers

The phosphate buffers consist of mixtures of K_2HPO_4 and KH_2PO_4 (also Na_2HPO_4 and NaH_2PO_4). Which function similarly to the bicarbonate buffers in neutralizing excess acid and base.

$$HCl + K_2HPO_4 \longrightarrow KH_2PO_4 + KCl$$

$$KOH + KH_2PO_4 \longrightarrow K_2HPO_4 + H_2O$$

3- Hemoglobin Buffers

The hemoglobin buffers account for more than one half of the total buffering action in the blood. There are hemoglobin buffers and oxyhemoglobin buffers.

Hemoglobin Buffer	**Oxyhemoglobin Buffer**
HHb	$HHbO_2$
KHb	$KHbO_2$

Acid-Base balance in blood

Carbonic acid-bicarbonate ion (conjugate acid-base pair) acts as a buffer in the control of blood .

Carbonic acid is formed by dissolving carbon dioxide in aqueous body fluids.It a weak acid that ionizes to bicarbonate ion.The equation for these two equilibrium reactions is as follows:

$$CO_2 + H_2O \longleftrightarrow H_2CO_3 \longleftrightarrow HCO_3^- + H^+$$

Normally, in body fluids such as blood ,there is 24 m Eq / L of bicarbonate ion to 1.2 m Eq / L of carbonic acid .

$$\frac{[HCO_3^-]}{[H_2CO_3]} = \frac{24 \text{ mEq/L}}{1.2 \text{ mEq/L}} = 20$$

1-The pH of blood is within normal its normal range of 7.35-7.45 when this ratio, 20 parts carbonate ion to 1 part carbonic acid.
2-The pH of blood becomes more acidic when the ratio
$[HCO_3^-]$ / $[H_2CO_3]$ become less than 20 /1 or 12 /1.Say 16/1 or 12/1
condition of blood signified by a pH less than 7.35 is called acidemia
3-The pH of blood becomes more basic when the ratio
$[HCO_3^-]/[H_2CO_3]$ become more than 20/1.say 25/1 or 30/1.
The alkaline condition of the blood signified by a pH greater than 7.45 is called alkalemia.

amount of carbon dioxide (CO_2),deeper and faster breathing called hyperventilation occurs.

This causes a decrease in the acidity of the blood because the carbon dioxide formed is lost through the lungs.The kidneys can help by releasing more bicarbonate ion into the blood and removing hydrogen ions .In these ways ,the body returns the $[HCO_3^-]/[H_2CO_3]$ ratio to its normal value of 20 and maintain the acid-base balance in the blood.

Death occure if the pH of blood is more acidic than 6.8 or more basic than 7.8.

Buffers in the body differ in one important respect from those in the laboratory.The body can supply components of the buffer solution as they are used up or can remove from the body any excess component.To explain this,let us consider how the body uses the carbonic acid-bicarbonate ion buffer system to overcome the increase in either the acid or the base concentration in the blood.

Acidosis

Consider a patient who has an illness that causes an increase in the concentration of acidic products in the blood.The physiologic processes causing academia are call acidosis.The acidic products react with bicarbonate ions to produce carbonic acid.

$$H_3O^+ + HCO_3^- \underline{\qquad\qquad\qquad} H_2CO_3 + H_2O$$

This causes a dcrease in the ratio $[HCO_3^-]$ /$[H_2CO_3]$,so acidosis will occur.One of the functions of both the lungs and the kidneys is to maintain the pH of the blood by supplying the buffer components that are used up or removing any excess components from the body.

The circulation of air into and out of the lungs,called ventilation .An increase in the amount of carbonic acid in the blood cause increase ~~in the~~ ~~Blood causes increase~~ in amount of carbon dioxide formed from decomposition of carbonic acid .To loss this excess of carbon dioxide ,deeper and faster breathing ,called hyperventilation occurs.This causes a decrease in the acidity of the blood because the carbon dioxide formed is lost through the lungs.The kidneys can help by releasing more bicarbonate ion into blood and removing hydrogen ion .in this way ,the body retains the $[HCO_3^-]/[H_2CO_3]$ ratio to its normal value of 20 and maintain the acid-base balance in the blood

Alkalosis

Consider another patient who has an illness causing an increase in concentration of basic products in the blood.The physiologic process causing alkalemia.

The lungs and the kidneys will both involved in this case. These basic products react with carbonic acid to form bicarbonate ions.This time the ratio $[HCO_3^-]/[H_2CO_3]$ increases.

$$\underset{\text{Basic product}}{OH^-} + H_2CO_3 \underline{\qquad\qquad\qquad} HCO_3^- + H_2O$$

The simplest way to prevent this ratio from increasing is to keep the carbon dioxide in the blood and use it to produce carbonic acid .To do this loss of carbon dioxide through the lungs is minimized by slower and shallow breathing ,called hypoventilation.

As before,the kidneys can help by removing bicarbonate ions and adding hydrogen ions to the blood.Thus ,the lungs and kidneys can function to maintain the pH of the blood within its normal range of 7.35to 7.45

●

Measuring pH

The pH of a solution can be calculated if the concentration of either H_3O^+ or OH^- is known. Alternatively, measurement of pH allows the calculation of H_3O^+ or OH^- concentration. The pH of aqueous solutions may be approximated by using indicating paper (pH paper) that develops a color related to the solution pH. Alternatively, a pH meter can give us a much more exact pH measurement. A sensor measures an electrical property of a solution that is proportional to pH.

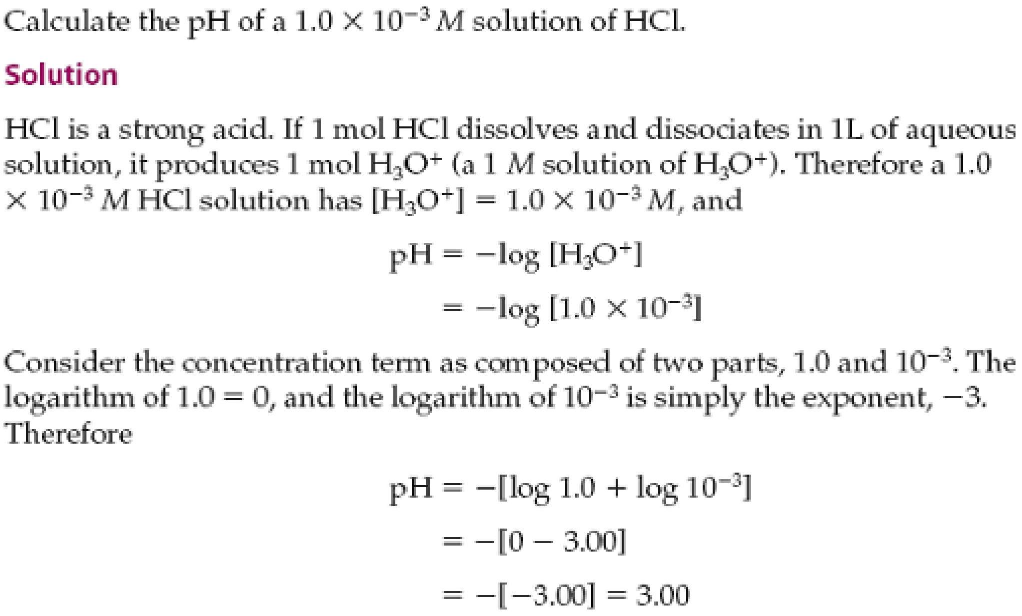
Calculating pH

One of our objectives in this chapter is to calculate the pH of a solution when the hydronium or hydroxide ion concentration is known, and to calculate $[H_3O^+]$ or $[OH^-]$ from the pH. We will develop this skill using two different approaches, one requiring a working knowledge of logarithms, the other using decimal logic.

The pH of a solution is defined as the negative logarithm of the molar concentration of the hydronium ion:

$$\mathbf{pH = -\log [H_3O^+]}$$

Example 1



Calculate the pH of a 1.0×10^{-3} *M* solution of HCl.

Solution

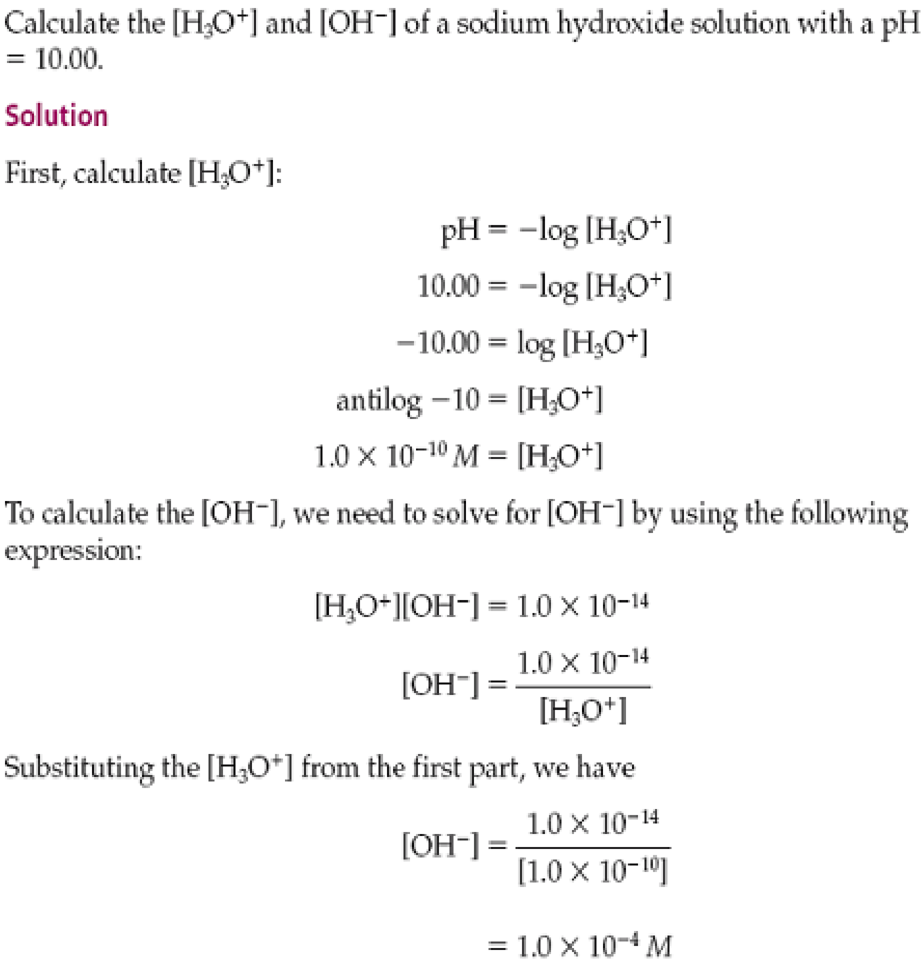
HCl is a strong acid. If 1 mol HCl dissolves and dissociates in 1L of aqueous solution, it produces 1 mol H_3O^+ (a 1 *M* solution of H_3O^+). Therefore a 1.0×10^{-3} *M* HCl solution has $[H_3O^+] = 1.0 \times 10^{-3}$ *M*, and

$$\begin{aligned} pH &= -\log [H_3O^+] \\ &= -\log [1.0 \times 10^{-3}] \end{aligned}$$

Consider the concentration term as composed of two parts, 1.0 and 10^{-3}. The logarithm of 1.0 = 0, and the logarithm of 10^{-3} is simply the exponent, −3. Therefore

$$\begin{aligned} pH &= -[\log 1.0 + \log 10^{-3}] \\ &= -[0 - 3.00] \\ &= -[-3.00] = 3.00 \end{aligned}$$

Example 2



Calculate the $[H_3O^+]$ and $[OH^-]$ of a sodium hydroxide solution with a pH = 10.00.

Solution

First, calculate $[H_3O^+]$:

$$\begin{aligned} \text{pH} &= -\log [H_3O^+] \\ 10.00 &= -\log [H_3O^+] \\ -10.00 &= \log [H_3O^+] \\ \text{antilog } -10 &= [H_3O^+] \\ 1.0 \times 10^{-10}\,M &= [H_3O^+] \end{aligned}$$

To calculate the $[OH^-]$, we need to solve for $[OH^-]$ by using the following expression:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$

Substituting the $[H_3O^+]$ from the first part, we have

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[1.0 \times 10^{-10}]}$$

$$= 1.0 \times 10^{-4}\,M$$

Chemistry lab:7 First stage

"Experiment for preparing of methane gas in the laboratory"

Chemical properties of alkanes:

1. Alkanes have a formula C_nH_{2n+2}, n=1,2,3 …etc.
2. Alkane doesn't have functional group because it's formed from saturated hydrocarbon (C-C and C-H).
3. Alkane structure (R-R or R-H) R=alkyl group.
4. Alkane is unaffected by many common chemical reagents but reacts with chlorine or fluorine in the presence of light.
5. The alkanes can exist as gases, liquids, or solids at room temperature. The unbranched alkanes (methane CH_4, ethane C_2H_6, propane C_3H_8, and butane C_4H_{10}) are gases; while from (pentane C_5H_{12} to hexadecane $C_{17}H_{36}$) are liquids; and the larger than hexadecane are solid.

Procedure:

Methane can be prepared in the laboratory by heating a mixture of sodium acetate with sodium hydroxide according to the equation:

$$CH_3COONa + NaOH \xrightarrow{\Delta} CH_4 + Na_2CO_3$$

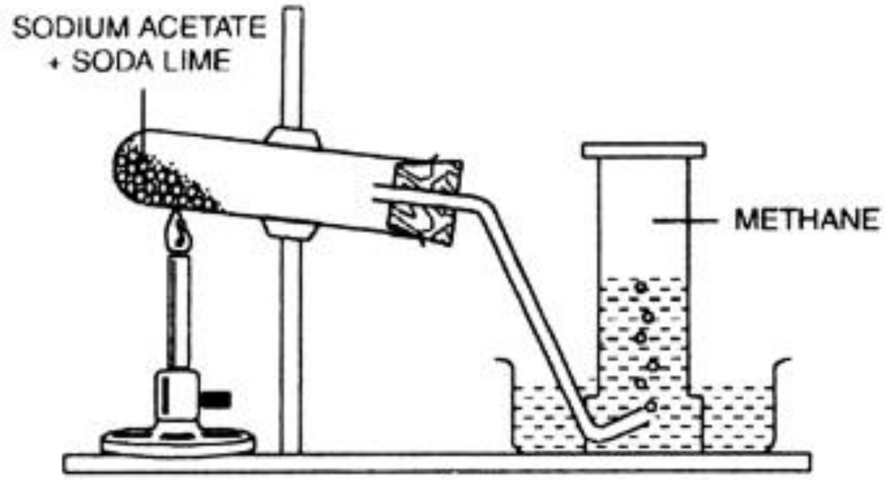
1- Firstly measure out 5 gm of sodium acetate or (sodium ethanoate).

2- Then add 2.5 gm of pulverized sodium hydroxide (or soda lime) and mix these two compounds well.

3- Transfer the mixture into a hard glass test tube / boiling tube and heat strongly.

4- (Sodium hydroxide NaOH and calcium oxide CaO) are called **soda lime**.

5- Methane can be collected by tube of the gas will replace the water.



After prepare methane we can study physical properties of methane or (alkanes):

1- Combustion: in combustion of methane gas is produce carbon dioxide, water vapor and heat.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + Heat\ \Delta$$

$$Alkane + O_2 \rightarrow CO_2 + H_2O + Heat\ \Delta$$

2- It is a colorless and odorless gas.

3- It is lighter than water and insoluble in water but soluble in non-polar solvents (organic solvents) like; alcohol (R-OH), carbon tetrachloride (CCl_4) … etc.

NOW:

1. Can you prepare hexane, butane and ethane?
2. Define sodium lime?
3. Write chemical equation for combustion of propane?

4- Ethane can be solved in ………. and ……… but not in …………..

Tikrit University **College of Dentistry**

Chemistry lab:4 **First stage**

Preparation and standardization of 0.1 M(HCl) hydrochloric acid solution

Hydrochloric acid is present in the digestive juices of the human stomach. Excessive secretion of the acid causes gastric ulcer, while a marked deficiency of it impairs the digestive process and is sometimes the primary cause of deficiency anemia. Concentrated hydrochloric acid causes burns and inflammation of the skin.

Theory:- Hydrochloric acid is produced in solutions up to 38% HCl (concentrated grade). Higher concentrations up to just over 40% are chemically possible, but the evaporation rate is then so high that storage and handling need extra precautions, such as pressure and low temperature. Laboratory grade hydrochloric acid is not sufficiently pure to be used as a primary standard, because it evaporates easily. In this experiment, a standard solution of sodium carbonate is used to determine the exact concentration of a hydrochloric acid solution. The neutralization reaction that occurs is as follows:

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

Methyl orange indicator solution is used. At the end-point – when neutralization just occurs – the indicator changes color from yellow to peach-pink.

Procedure:-

1. **Preparing (50 ml) 0.1 M HCl Solution**:38 % HCL shows density 1.19 g/mL and we can find M by next : -

$$M = \frac{sp.gr * \% * 1000}{M.wt}$$

Calculate the volume of HCl (conc.):- We must dilute it to preparing 0.1 M HCl in 50 ml from next: $(M^* V)_{conc.} = (M * V)$

dilute

$$M * V_{ml} = 0.1 * 50_{ml}$$

Transfer V ml by cylinder to clean and dry beaker containing 30 ml D.W, transfer the solution to volumetric flask capacity 50 ml, and complete the volume to the mark by D.W.

2. **Preparing (50 ml) 0.1 M Na_2CO_3 Solution**:-calculate amount from sodium carbonate for prepare 0.1 M in 50 ml -

$$M = \frac{Wt.(gm)}{M.Wt.} * \frac{1000}{V(ml)}$$

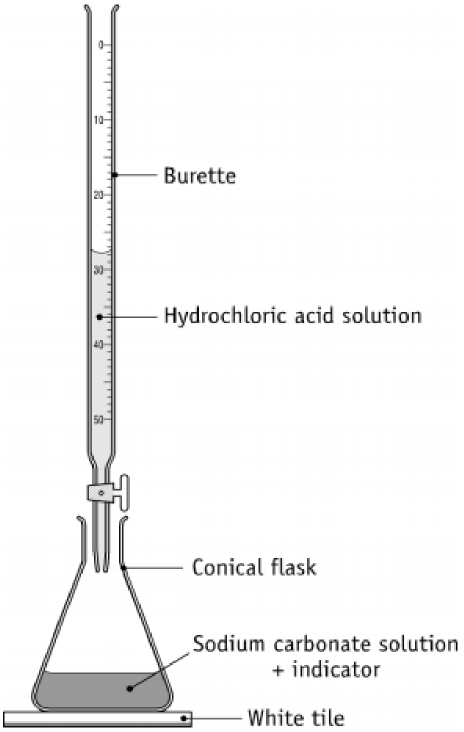
$$0.1 = \frac{Wt.(gm)}{106} * \frac{1000}{50}$$

$$Wt. = 0.53\ gm$$

Weigh 0.53 gm. from Na_2CO_3 in clean and dry beaker and dilute in 30 ml D.W, transfer solution to volumetric flask capacity 50 ml and complete the volume to the mark by D.W.

3. Transfer known volume of 5 ml the sodium carbonate solution, with a pipette, to a conical flask then add one or two drops of methyl orange to this solution.

4. Add the acid unknown solution from the burette gradually with continuous swirling of the solution in the conical flask and near the end point, the acid is added drop by drop. Continue the addition of the acid until the color of the solution passes from yellow to faint red.



5. Repeat the experiment three times and tabulate your results then take the mean of the three readings.

Calculations: Calculate the molarity of HCl:-

$$\text{m mol HCl} = \text{m mol } Na_2CO_3$$

$$(M * V)\,HCl = (M * V)\,Na_2CO_3 * \frac{1}{2}$$

$$(M * V_{\text{burette}}) = (0.1 * 5) * \frac{1}{2}$$

Discussion:-

1. What the difference between primary and secondary standard substances?

2. Calculate the volume of conc. HCl required for preparing 250 ml 0.1 M?

Chemistry lab: 8 First stage

"Experiment for preparing of ethene(ethylene) gas in the laboratory"

Alkenes:

1-Alkenes are unsaturated hydrocarbons have (C=C) that consider the functional group.

2-Alkenes formula C_nH_{2n} where n= 2, 3, 4 ………

3-Alkene reacts with brom water (Br_2), when the red brom water is adding into alkene solution its color will disappear.

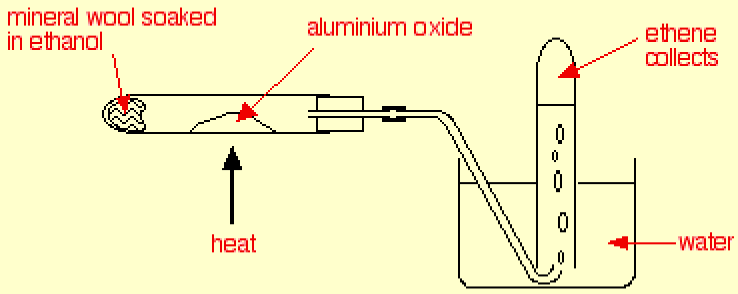
Procedure:

The dehydration of ethanol to give ethene (ethylene):

1-This is a simple way of making gaseous alkenes like ethene. If ethanol vapor is passed over heated aluminium oxide powder, the ethanol is essentially broken to give ethene and water vapor.

$$CH_3\text{-}CH_2\text{-}OH \xrightarrow{Al_2O_3} CH_2{=}CH_2 + H_2O$$

To make a few test tubes of ethene, you can use this apparatus:



It wouldn't be too difficult to imagine scaling this up by boiling some ethanol in a flask and passing the vapor over aluminium oxide heated in a long tube.

2-Ethanol is heated with an excess of concentrated sulphuric acid at a temperature of 170°C. The gases produced are passed through sodium hydroxide solution to remove the carbon dioxide and sulphur dioxide produced from side reactions.

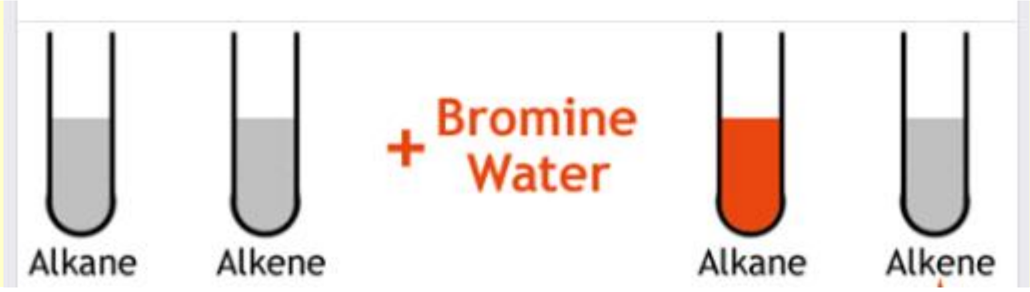
$$CH_3\text{-}CH_2\text{-}OH \xrightarrow{\text{conc } H_2SO_4} CH_2{=}CH_2 + H_2O$$

Do the following

1-Distinguish between ethane and ethene.

2-By the same way prepare propene.

3-see the diagram then explain why alkane has red color.



4- Dehydration reagents are ………. and ………. .

5- See the equation then complete

$CH_3\text{-}CH{=}CH\text{-}CH_3 + Br_2 \rightarrow$

$CH_3\text{-}CH_2 \text{–} CH_2\text{-}CH_3 + Br_2 \rightarrow$