General Chemistry

1st year Faculty of Pharmacy

Academic Year: 2020-2021

Tara F. Tahir
Ph.D.Analytical Chemistry (Electrochemistry)
E-Mail: tara.fuad@ishik.edu.iq
tara.fuad@koyauniversity.org
http://sites.koyauniversity.org/tara-fuad/

Why Chemistry is Essential?

- **Chemistry** is the science of the composition, structure, properties and reactions of matter, especially atomic or molecular systems.

- **Chemistry** is life:

* Evolution of life begins from one single organic compound called a nucleotide

* Biochemical processes in the living cell

* Whatever we see, use or consume in everyday life is the gift of research in Chemistry for thousands of years

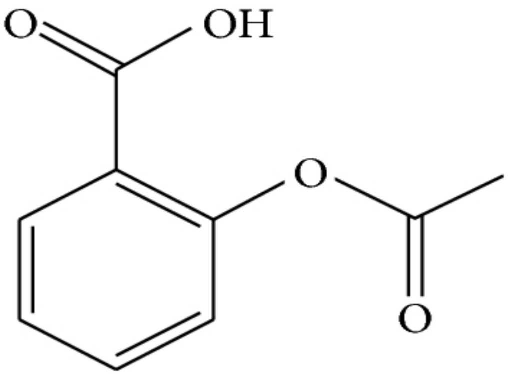
Why Knowledge of Chemistry is essential to be a good pharmacist?

- **Chemistry** has played a major role in pharmaceutical advances and forensic science.
- **Chemistry** plays an important role in understanding diseases and their remedies, i.e. drugs.

* Medicines or drugs that are used for the treatment of various diseases are either organic or inorganic. However, most drugs are organic molecules.

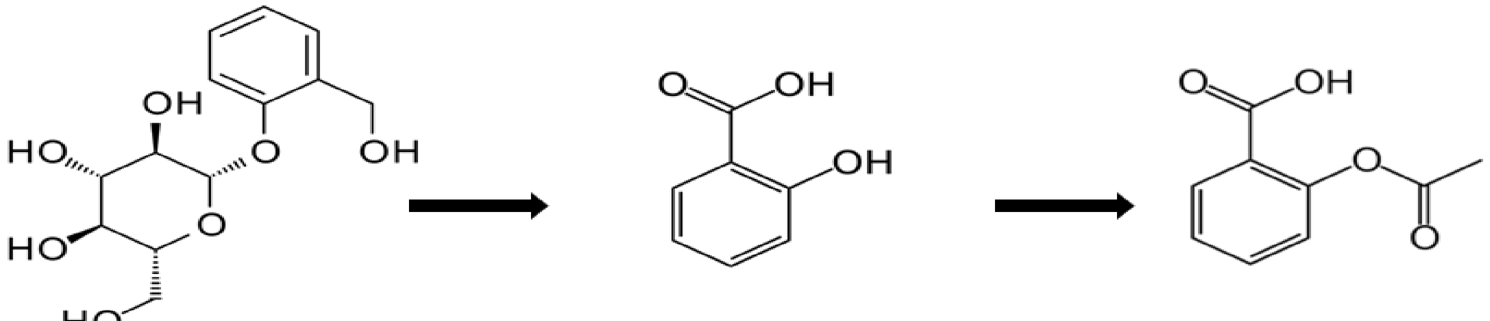
- E.g. the most popular and widely used analgesic drug; **Aspirin**

*** Aspirin is chemically known as acetyl salicylic acid**



Aspirin
Acetyl salicylic acid

Willow tree



- Sources of drug molecules can be natural but whatever the source is, **chemistry** is involved in all processes of drug extraction, isolation, purification, and identification.
- Chemistry is important in the drug development steps, especially in the pre-formulation and formulation studies, the structures and the physical properties, e.g. solubility, pH and storage conditions.

* Drugs having an ester functionality, e.g. aspirin, could be quite unstable in the presence of moisture, and should be kept in a dry and cool place.

Physical Properties of Drug Molecules

1. Physical State:

Physical state of drug molecules is an important factor in the formulation and delivery of drugs.

e.g. amorphous solid, crystalline solid, hygroscopic solid, liquid or gas.

Amorphous solid

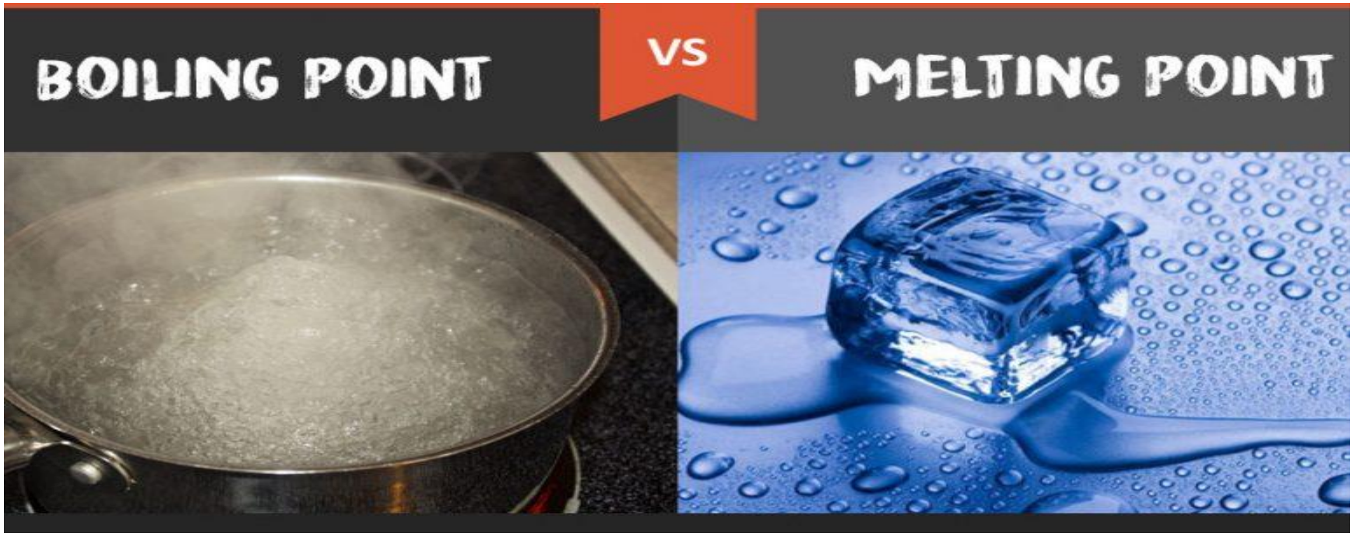
Crystalline solid

2. Melting point and boiling point:

Melting point (m.p.) is the temperature at which a solid becomes a liquid.

Boiling point (b.p.) is the temperature at which the vapour pressure of the liquid is equal to the atmospheric pressure.

e.g. the melting point of water at 1 atmosphere of pressure is 0 °C (32 °F, 273.15 K; this is also known as the ice point) and the boiling point of water is 100 °C.



- Melting point is a characteristic for organic compounds to confirm the purity.

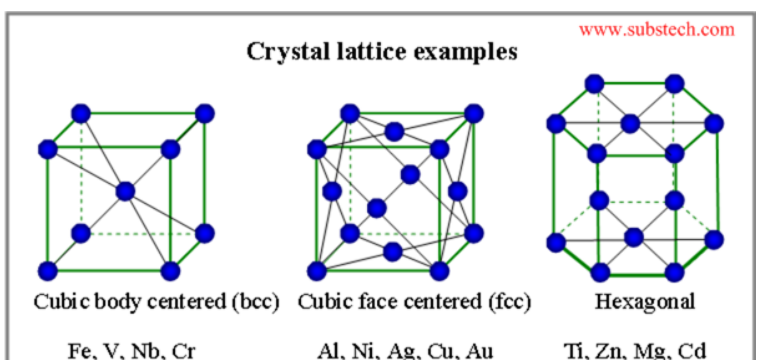
e.g. The melting point of a pure compound is always higher than the melting point of that compound mixed with a small amount of an impurity. The more impurity is present, the lower the melting point.

It can also melt over a wide range of temperatures and is called the "melting point depression."

How impurities affect melting point?

- Most of the solids in the laboratory (organic or inorganic) are crystalline solids. These have molecules or repeating clusters of ions arranged in a regular, tightly packed repeating crystal lattice.

The lattice is held together by various intermolecular forces, which come about because of the chemical nature of the solid. The tighter the crystal lattice, the more energy is required to break it, and eventually melt the compound.

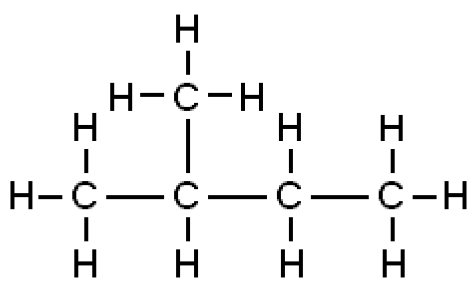


- Examples:

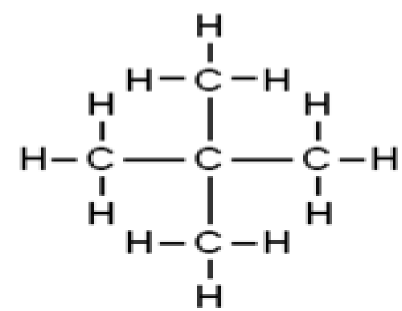
Alkanes with an odd number of carbon atoms pack less tightly, which decreases their melting points. Thus, alkanes with an even number of carbon atoms have higher melting points than the alkanes with an odd number of carbon atoms.

$CH_3CH_2CH_2CH_3$	$CH_3CH_2CH_2CH_2CH_3$
Butan	Pentane
m.p.= -138.4 °C	m.p.= -129.7 °C

- In contrast, between two alkanes having same molecular weights, the more highly branched alkane has a lower boiling point.



Isopentane
b.p. = 27.9 °C



Neopentane
b.p. = 9.5 °C

3. Polarity and Solubility:

- Polarity is a physical property of a compound, which relates other physical properties, e.g. melting and boiling points and solubility. The term bond polarity is used to describe the sharing of electrons between atoms.
- There is a direct correlation between the polarity of a molecule and the number and types of polar or nonpolar covalent bond that are present. In a few cases, a molecule having polar bonds, but in a symmetrical arrangement, may give rise to a nonpolar molecule, e.g. carbon dioxide (CO2).
- In a nonpolar covalent bond, the electrons are shared equally between two atoms. A polar covalent bond is one in which one atom has a greater attraction for the electrons than the other atom. When this relative attraction is strong, the bond is an ionic bond.
- The polarity in a bond arises from the different electronegativity of the two atoms that take part in the bond formation.
- The greater the difference in electronegativity between the bonded atoms, the greater is the polarity of the bond. For example, water is a polar molecule, whereas cyclohexane is nonpolar

- Solubility is the amount of a solute that can be dissolved in a specific solvent under given conditions.
- The process of dissolving is called solvation, or hydration when the solvent is water.
- Like dissolves like i.e., materials with similar polarity are soluble in each other.

Examples:

- A polar solvent, e.g. water, has partial charges that can interact with the partial charges on a polar compound, e.g. sodium chloride (NaCl).

- As nonpolar compounds have no net charge, polar solvents are not attracted to them. Alkanes are nonpolar molecules, and are insoluble in polar solvent, e.g. water, and soluble in nonpolar solvent, e.g. petroleum ether.

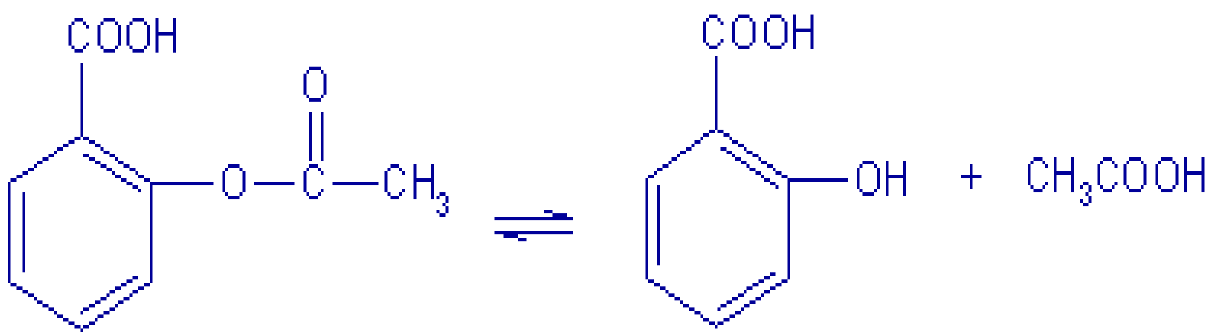
- Saturated solution is a solution in an equilibrium that cannot hold any more solute.
- Solubility is the maximum equilibrium amount of solute that can usually dissolve per amount of solvent.

Factors affecting solubility:

- intermolecular forces between the solvent and solute
- Temperature
- Partial pressure of a solute gas
- Size of the particles
- Stirring
- Amount of solid already dissolved
- pH of the solution

4. Acid-base properties and pH

- Aspirin has an adverse effect which is stomach bleeding and it is partly due to the acidic nature of the drug. In the stomach, aspirin is hydrolysed to salicylic acid. The carboxylic acid group (— COOH) and a phenolic hydroxyl group (—OH) present in salicylic acid make this molecule acidic. Thus, intake of aspirin increases the acidity of the stomach significantly, and if this increased acidic condition remains in the stomach for a long period, it may cause stomach bleeding.
- There are a number of drug molecules that are acidic in nature. Similarly, there are basic and neutral drugs as well.



Aspirin - acetylsalicylic acid   salicylic acid   acetic acid

What is acidic and basic drugs?

- Acid is an electron-deficient species that accepts an electron pair, e.g. hydrochloric acid (HCl)
- Base is a species with electrons to donate, e.g. sodium hydroxide (NaOH).
- A neutral species does not do either of these.

- Most organic reactions are either acid–base reactions or involve catalysis by an acid or base at some point.

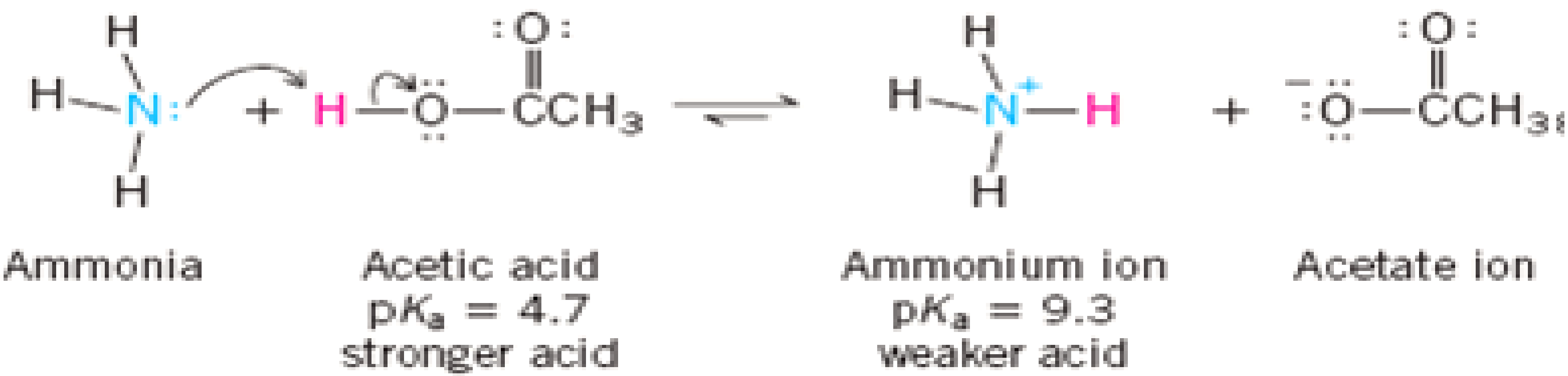
Arrhenius acids and bases

- Acid is a substance that produces hydronium ion (H_3O^+).
- Base produces hydroxide ion (OH^-) in aqueous solution. An acid reacts with a base to produce salt and water.

Brønsted–Lowry acids and bases

- Acid is a proton donor (H^+)
- Base is a proton acceptor (H^+)

- According to the Brønsted–Lowry definitions, any species that contains hydrogen can potentially act as an acid, and any compound that contains a lone pair of electrons can act as a base.
- Therefore, neutral molecules can also act as bases if they contain an oxygen, nitrogen or sulphur atom.
- Acid-base reaction is a reaction where a proton transfer to a base and a base accept a proton.



Lewis theory of acids and bases

- According to Lewis theory, an acid is an electron-pair acceptor, and a base is an electron-pair donor. Thus, a proton is functioned as a Lewis acid. Any molecule or ion may be an acid if it has an empty orbital to accept a pair of electrons. Any molecule or ion with a pair of electrons to donate can be a base.
- Lewis acids are known as aprotic acids, compounds that react with bases by accepting pairs of electrons, not by donating protons.

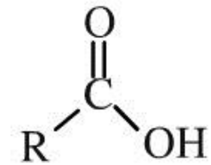
e.g. Borane (BH_3), boron trichloride (BCl_3) and boron trifluoride (BF_3) are known as Lewis acids, because boron has a vacant d orbital that accepts a pair of electrons from a donor species.

$$\underset{\text{Lewis acid}}{BCl_3} + \underset{\text{Lewis base}}{:\ddot{O}(CH_2CH_3)_2} \longrightarrow \underset{\text{Acid-base adduct}}{Cl_3B:\ddot{O}(CH_2CH_3)_2}$$

Organic functional groups and acid-base properties

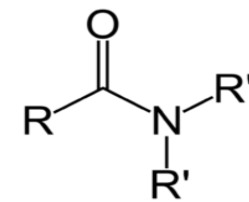
- Drugs are organic molecules that contain different functional groups and those compounds with nonbonding electrons on nitrogen, oxygen, sulphur, or phosphorus can act as Lewis bases or Brønsted bases. They react with Lewis acids or Brønsted acids.
- Lewis acids may be either protic or aprotic acids. Brønsted acids are also called protic acids.

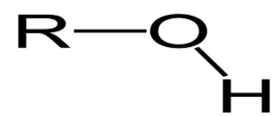
Carboxylic group



$R-NH_2$ Amines group

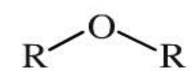
Amide group

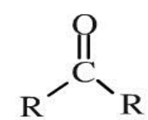




Alcohol group

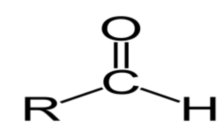
Ether group





Ketone group

Aldehyde group



- Carboxylic acids are moderately strong acid, pK_a = 3-5, e.g. acetic acid pK_a = 4.76. See reaction of page 9.
- Amines are the most important organic bases as well as weak acids. Thus, an amine can behave as an acid and donate a proton, or as a base and accept a proton. See reaction of page 9.
- Alcohols can behave like an acid and donate a proton. However, they are much weaker organic acids. Alcohol may also behave as a base.

e.g., ethanol is protonated by sulphuric acid and gives ethyloxonium ion ($C_2H_5OH_2^+$). See reaction of page 9.

- Ketones can behave as bases. Acetone donates electrons to boron trichloride, a Lewis acid, and forms a complex of acetone and boron trichloride. See reaction of page 10.

- The reaction of an organic compound as an acid depends on how easily it can lose a proton to a base.
- The acidity of the hydrogen atom depends on the electronegativity of the bonded central atom. The more electronegative the bonded central atom, the more acidic are the protons.

$$\text{C-H} < \text{N-H} < \text{O-H}$$

$$\xrightarrow{\hspace{6cm}}$$

Increasing electronegativity

- Weak acids produce strong conjugate bases. Thus, ethane CH_3CH_3 gives a stronger conjugate base than methylamine CH_3NH_2 and methanol CH_3OH.

What is the difference between pK_a and pH?

- pH scale is used to describe the acidity of a solution.
- pK_a is characteristic of a particular compound of how readily the compound gives up a proton.
- The pH of the salt depends on the strengths of the original acids and bases.

$$pH = -\log [H_3O^+]$$

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

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

- Strong acid + strong base= salt pH 7
- Weak acid + strong base= salt > 7
- Strong acid + weak base = salt < 7
- Weak acid + weak base = depends on which one is stronger

- The pH of blood plasma is around 7.4, whereas that of the stomach is around 1.
- Strong acids, e.g. HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_3$ and $HClO_4$, completely ionize in solution, and are always represented in chemical equations in their ionized form.
- Strong bases, e.g. LiOH, NaOH, KOH, RbOH, $Ca(OH)_2$, $Sr(OH)_2$ and $Ba(OH)_2$, completely ionize in solution.
- Hydrolysis is a process where a salt is formed after mixing an acid with a base and the acid releases H^+ ions while the base releases OH^- ions.
- The larger the acid dissociation constant, the stronger is the acid. e.g. Hydrochloric acid has an acid dissociation constant of 10^7, whereas acetic acid has an acid dissociation constant of only 1.74×10^{-5}
- The strength of an acid is generally indicated by its pK_a value rather than its K_a value.

$pK_a = - \log K_a$

Very strong acids pKa < 1
Moderately strong acids pKa = 1-5

Weak acids pKa = 5-15

Extremely weak acids pKa > 15

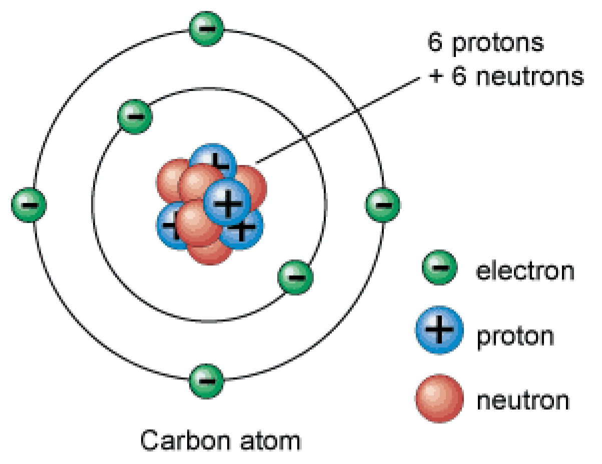
What is buffer?

- Buffer is a solution containing a weak acid and its conjugate base (e.g. CH3COOH and CH3COO-) or a weak base and its conjugate acid (e.g. NH3 and NH_4^+) and contains common ion. It maintains a relatively constant pH even when acidic or basic solutions are added to it.

- Vital example of buffer solution is human blood. Blood can absorb the acids and bases produced by the biological reactions without changing its pH.
- Human cells can only survive in a narrow pH range of 7.4 and any fluctuation in pH would denature most enzymes and hence interfere with the body metabolism. Carbon dioxide from metabolism combines with water in blood plasma to produce carbonic acid (H_2CO_3). The amount of H_2CO_3 depends on the amount of CO_2 present.

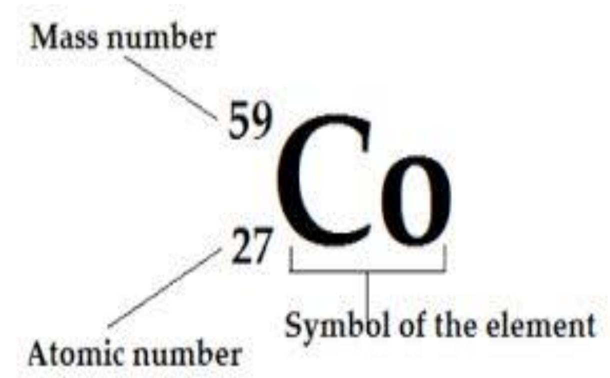
Atoms, elements and compounds

- An atom is the basic building block of the matter that is a collection of subatomic particles of negatively of electrons, positively charged protons and neutral particles of neutrons.
- Electrons move around the nucleus, and are arranged in shells at increasing distances from the nucleus. These shells represent different energy levels and the outer most shell being the highest energy level.
- Both protons and neutrons have mass, whereas the mass of electrons is negligible.

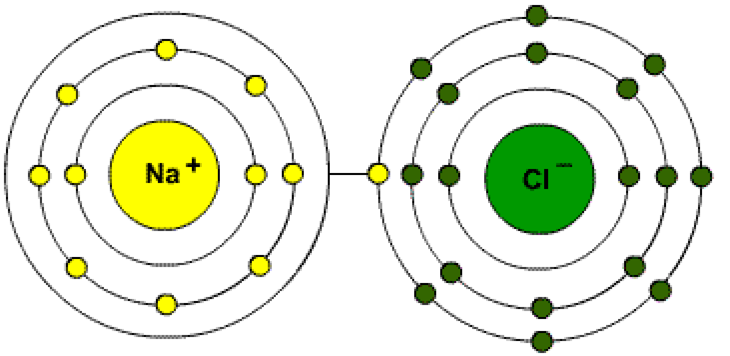


- Atomic number = no. of protons in the nucleus.
- Mass number = Total number of protons + neutrons in the nucleus

e.g., a cobalt atom containing 27 protons and 32 neutrons has a mass number of 59.

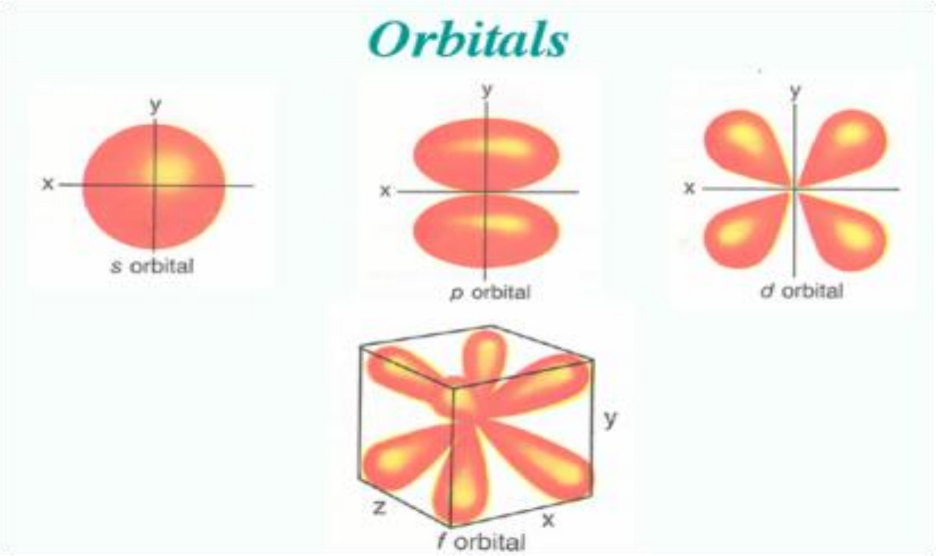


- Elements are substances containing one type of atoms only, e.g. O_2, N_2 and Cl_2.
- Compounds are substances formed when atoms of two or more elements join together, e.g. NaCl, H_2O and HCl.



Orbitals of the atomic structure

- Electrons are involved in the chemical bonding and reactions of an atom and are occupied orbitals in an atom.
- An orbital is a region of space that can hold two electrons.
- Electrons move in region of space called shells and each shell can contain up to $2n^2$ electrons, where n is the number of the shell. Each shell contains subshells called atomic orbitals.



- Electronic configuration is the arrangement of electrons in orbitals, subshells and shells of the atom.

- Valence number is the number of electrons an atom must lose or gain to attain the nearest noble gas or inert gas electronic configuration.

Shells, Subshells, Orbitals

Shell (Corresponds to period)	# of subshells	subshells	# orbitals	# electrons
n = 1	1	1*s*	1	2
n = 2	2	2*s*	1	2
		2*p*	3	6
n = 3	3	3*s*	1	2
		3*p*	3	6
		3*d*	5	10
n = 4	4	4*s*	1	2
		4*p*	3	6
		4*d*	5	10
		4*f*	7	14

- The number of subshells in a shell = shell number
- The first subshell s has 1 orbital. Each successive subshell adds 2 more orbitals (1, 3, 5, 7, etc).
- Each orbital can hold only 2 electrons of opposite spin.
- An atom with $n = 3$ also includes all subshells and orbitals for $n < 3$:
  - 1s, 2s, 2p, 3s, 3p, 3d

Why Chemical Bonds are formed?

- Atoms form bonds in order to obtain a stable electronic configuration, i.e. the electronic configuration of the nearest noble gas.
- Why noble gases are inert?
- because their atoms have a stable electronic configuration in which they have eight electrons in the outer shell except helium (two electrons). Therefore, they cannot donate or gain electrons.
- A filled shell is known as a noble gas configuration.
- Electrons in filled shells are called core electrons. The core electrons do not participate in chemical bonding.
- Electrons in shells that are not completely filled are called valence electrons, also known as outer-shell electrons.

Chemical Bonds Theory: Lewis Theory

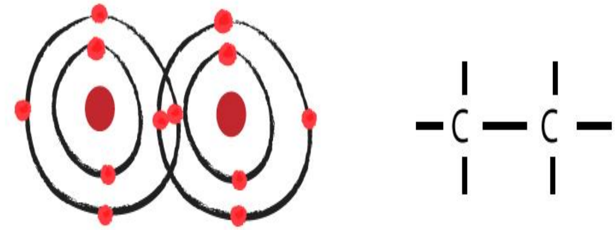
- According to Lewis theory an atom will give up or accept or share electrons in order to achieve a filled outer shell that contains eight electrons and this stable configuration is called an *octet*.

- Lewis structures help us to track the valence electrons and predict the types of bond.

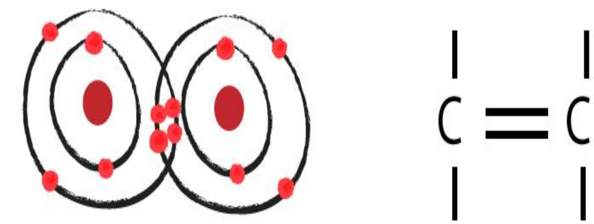
- Elements in organic compounds are joined by covalent bonds, a sharing of electrons, and each element contributes one electron to the bond.

- The number of electrons necessary to complete the octet determines the number of electrons that must be contributed and shared by a different element in a bond. This finally determines the number of bonds that each element may enter into with other elements.

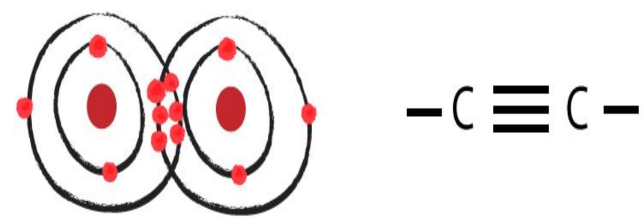
- In a single covalent bond two atoms share one pair of electrons and form a **σ** bond.
- In a double bond they share two pairs of electrons and form a **σ** bond and a π bond.
- In a triple bond two atoms share three pairs of electrons and form a **σ** bond and two π bonds.



SINGLE BOND

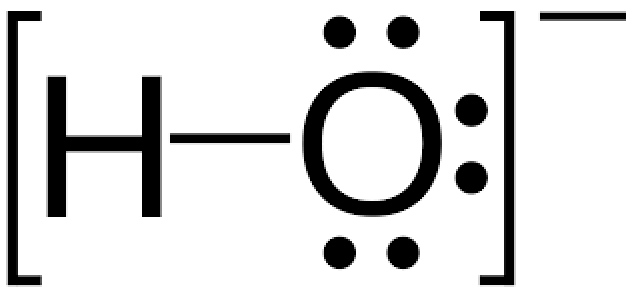


DOUBLE BOND



TRIPLE BOND

- Example; sodium loses a single electron from its 3s orbital to attain a more stable neon gas configuration with no electron in the outer shell.
- The shared electrons are called the bonding electrons and may be represented by a line or lines between two atoms.
- The valence electrons that are not being shared are the nonbonding electrons or lone pair electrons, and they are shown in the Lewis structure by dots around the symbol of the atom.
- A species that has an unpaired electron are called radicals. Usually they are very reactive, and are believed to play significant roles in aging, cancer and many other diseases.



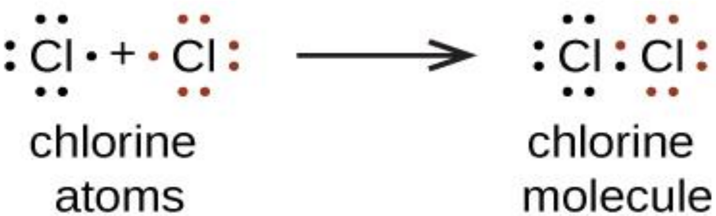
- In neutral organic compounds, C forms four bonds, N forms three bonds (and a lone pair), O forms two bonds (and two lone pairs) and H forms one bond. The number of bonds an atom normally forms is called the valence.

Various types of chemical bonding

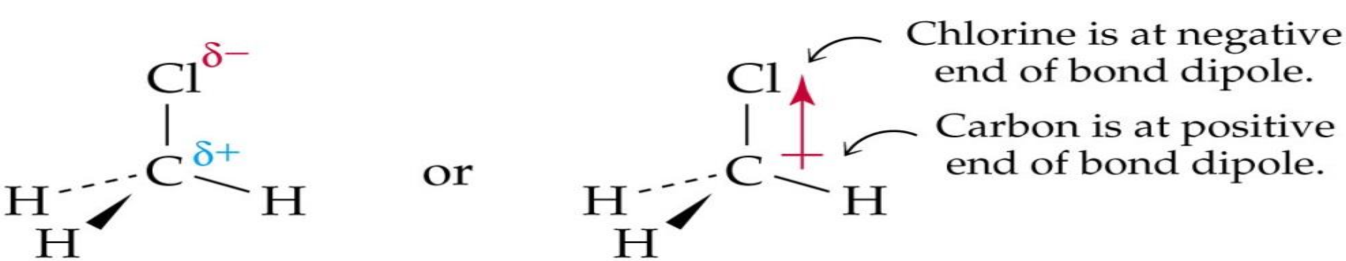
- A chemical bond is the attractive force that holds two atoms together.
- An atom that gains electrons becomes an anion, a negatively charged ion, and an atom that loses electrons becomes a cation, a positively charged ion.
- Metals tend to lose electrons and nonmetals tend to gain electrons, e.g. Fe^{2+} and Cl^{-}
- Cations are smaller than atoms and anions are larger.

- Ionization energy is the energy required for removing an electron from an atom or ion in the gas phase.
- Atoms can have a series of ionization energies, since more than one electron can always be removed, except for hydrogen.
- In general, the first ionization energies increase across a period and decrease down the group.
- Adding more electrons is easier than removing electrons. It requires a vast amount of energy to remove electrons.
- Ionic bonds result from transfer of one or more electrons between atoms. The more electronegative atom gains one or more valence electrons and becomes an anion, while the less electronegative atom loses one or more valence electrons and becomes a cation.
- Ionic bonds are commonly formed between reactive metals, electropositive elements and nonmetals, electronegative elements, e.g. Na (electronegativity 0.9) easily gives up an electron, and Cl (electronegativity 3.0) readily accepts an electron to form an ionic bond.

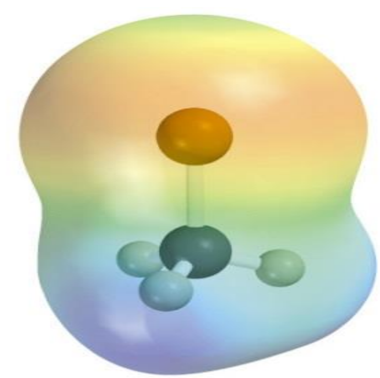
- Covalent bonds result from the sharing of electrons between atoms, e.g. two chlorine atoms can achieve a filled valence shell of 18 electrons by sharing their unpaired valence electrons.



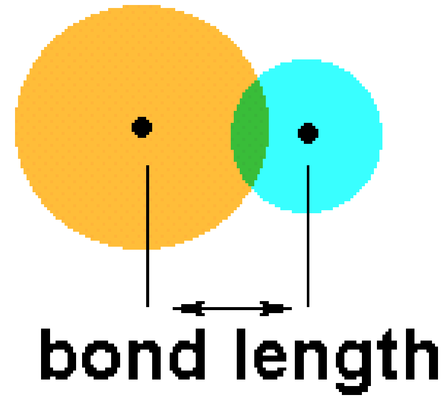
- Nonpolar covalent bond result from sharing electrons equally between two atoms e.g. H_2 and F_2
- Polar covalent bond result from unequally sharing of electrons between two atoms where one atom has a greater attraction for the electrons than the other atom e.g. chloromethane (CH_3Cl)

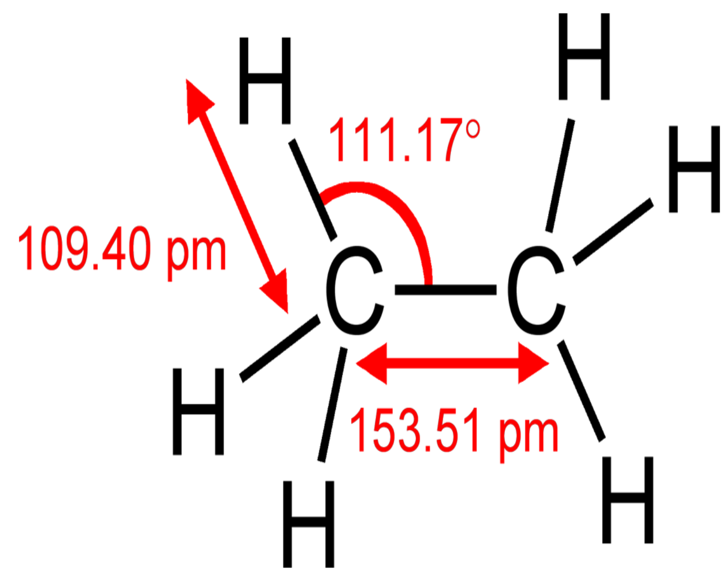


Chloromethane, CH_3Cl



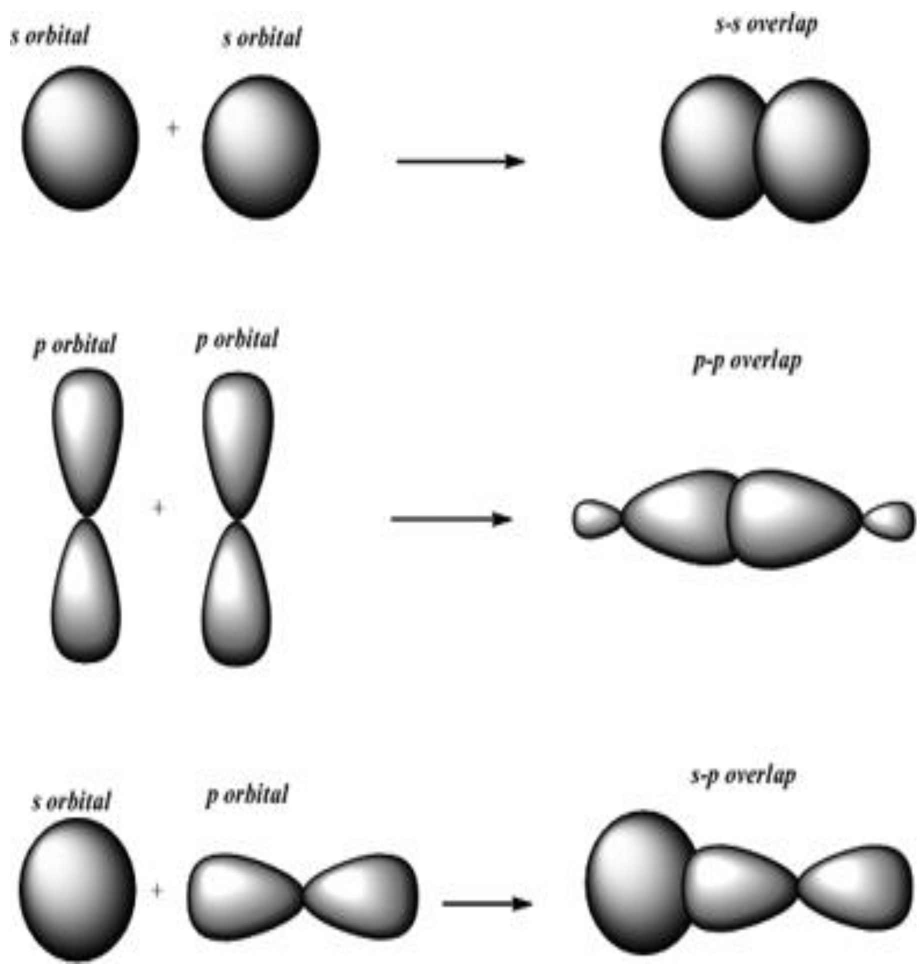
- Dipole moment μ it is a measurement of bond polarity in a unit called debye (D).
- To describe 3D structure of covalent compounds bond lengths and bond angles are used.
- Bond length is the average distance between the nuclei of the atoms that are covalently bonded together.
- A bond angle is the angle formed by the interaction of two covalent bonds at the atom common to both.



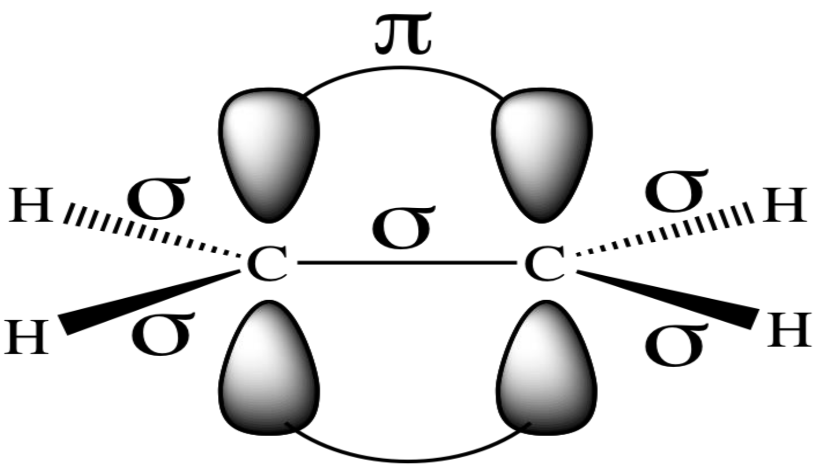


How a covalent bond forms?

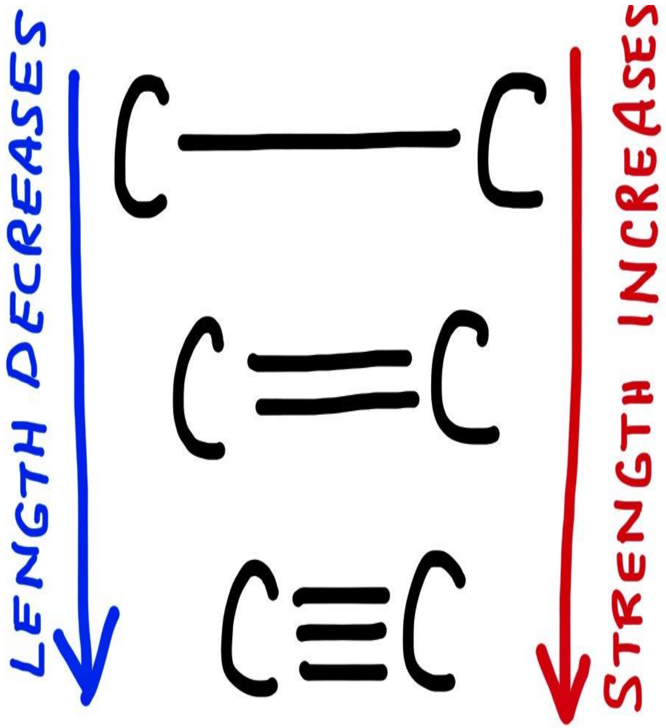
- Hybridization is the overlap of atomic orbitals to form covalent bond and the resulting atomic orbitals are called hybrid orbitals.
- There are two types of orbital overlap, which form sigma (**σ**) and pi (π) bonds. A sigma overlap occurs when there is one bonding interaction that results from the overlap of two s orbitals or an s orbital overlaps a p orbital or two p orbitals overlap head to head.



- A π overlap occurs only when two bonding interactions result from the sideways overlap of two parallel p orbitals.



- σ bond has more length than double bond (π) and triple bond but less strength. Why?



Electronegativity and chemical bonding?

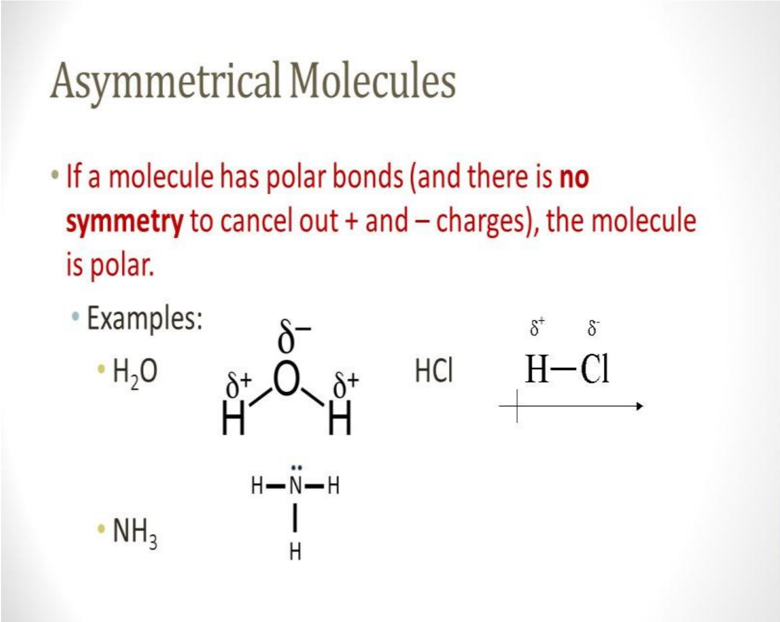
- Electronegativity is the ability of an atom that is bonded to another atom to attract electrons strongly towards it.
- Elements with higher electronegativity values have greater attraction for bonding electron Thus, the electronegativity of an atom is related to bond polarity.
- The difference in electronegativity between two atoms can be used to measure the polarity of the bonding between them. The greater the difference in electronegativity between the bonded atoms, the greater is the polarity of the bond i.e. if the difference is great enough, electrons are transferred from the less electronegative atom to the more electronegative one, hence an ionic bond is formed. Only if the two atoms have exactly the same electronegativity is a nonpolar bond formed.

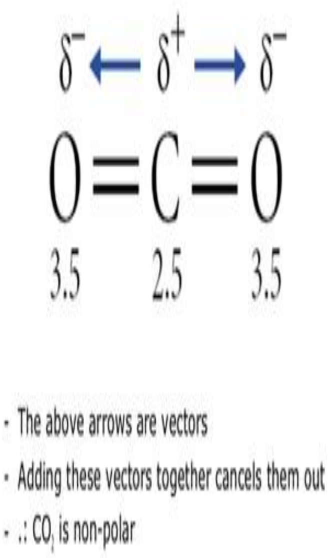
- If the electronegativity difference is:
- 0.5 < Nonpolar covalent bond
- 0.5 - 1.9 Polar covalent bond
- 2.0 or > Ionic bond

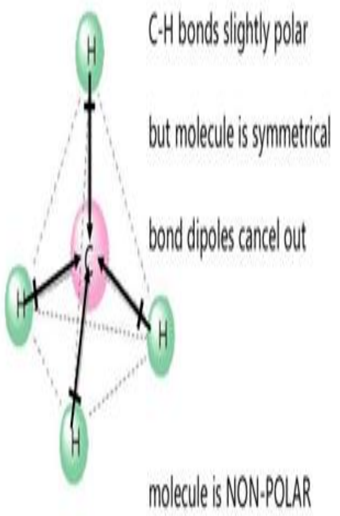
Bond	Difference in electronegativity	Types of bond
C-Cl	3.0 – 2.5 = 0.5	Polar covalent
P-H	2.1 – 2.1 = 0	Nonpolar covalent
C-F	4.0 – 2.5 = 1.5	Polar covalent
S-H	2.5 – 2.1 = 0.4	Nonpolar covalent
O-H	3.5 – 2.1 = 1.4	Polar covalent

- Electrons in a polar covalent bond are unequally shared between the two bonded atoms, which results in partial positive and negative charges. The separation of the partial charges creates a dipole. Therefore, word dipole means two poles, the separated partial positive and negative charges, so a polar molecule results when a molecule contains polar bonds in an unsymmetrical arrangement.

- Nonpolar molecules whose atoms have equal or nearly equal electronegativities have zero or very small dipole moments, as do molecules that have polar bonds but the molecular geometry is symmetrical, allowing the bond dipoles to cancel each other.



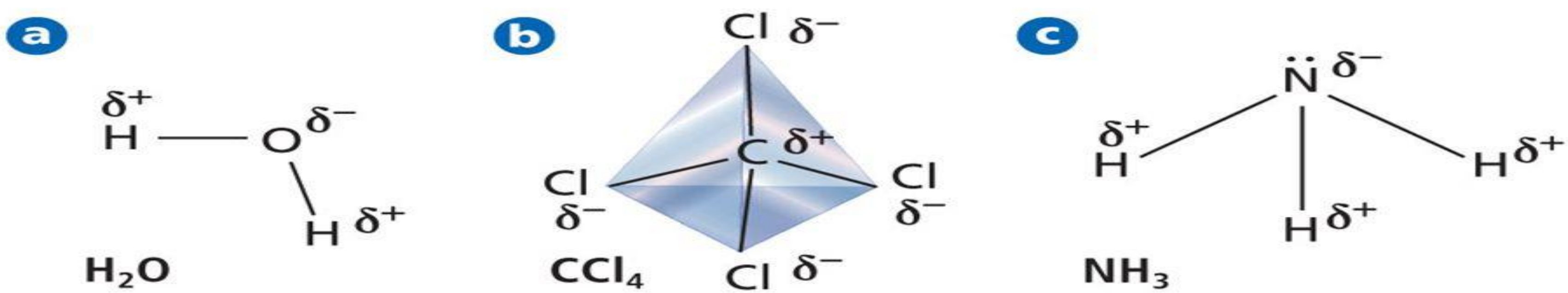




- For more informations watch the vedio:

https://www.youtube.com/watch?v=P3VfhLHrYR4

- Symmetrical molecules have identical atoms distributed on all sides. There are symmetrical linear, trigonal planar, tetrahedral shapes.
- Always asymmetrical are trigonal pyramidal or bent

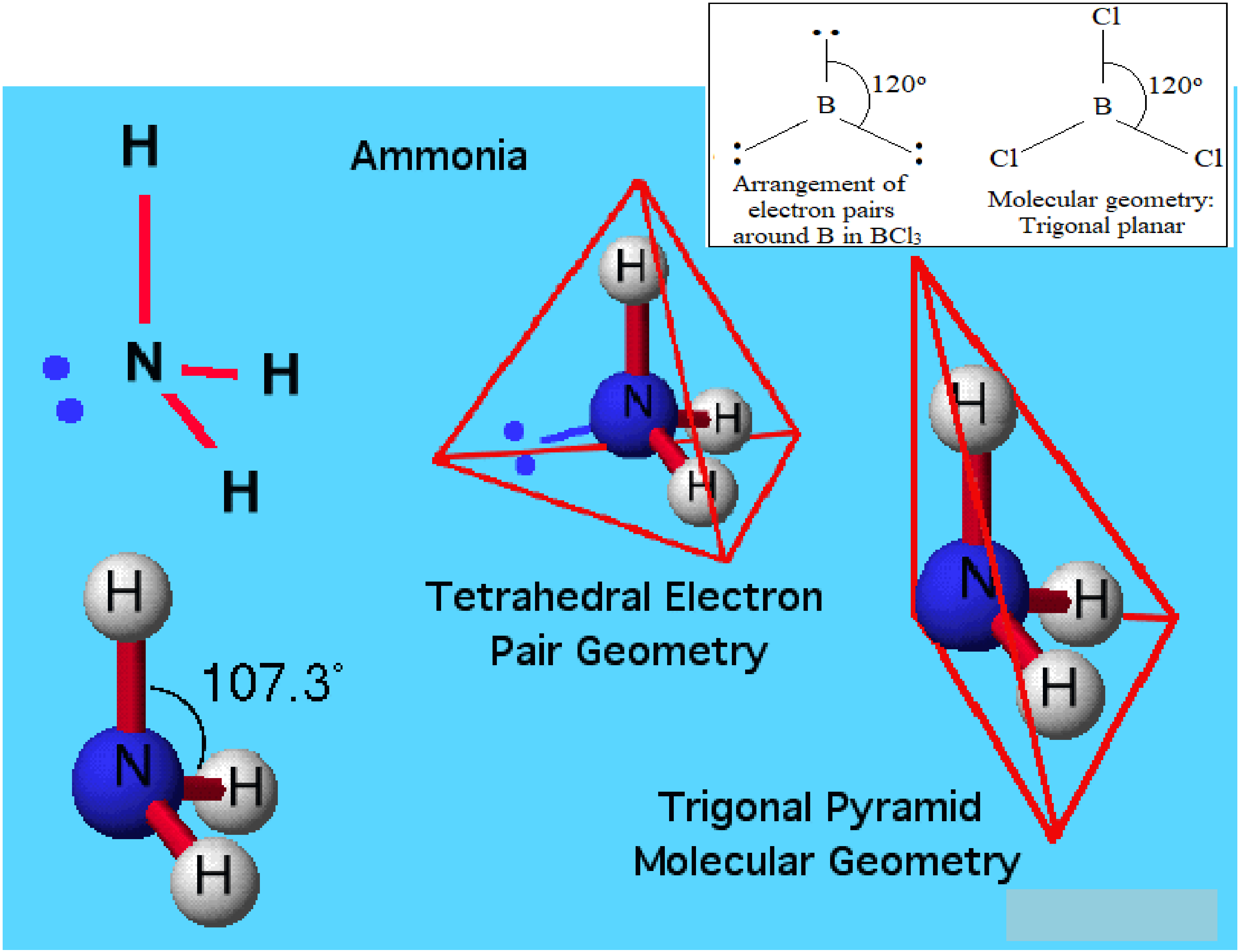


The bent shape of a water molecule makes it polar.

The symmetry of a CCl_4 molecule results in an equal distribution of charge, and the molecule is nonpolar.

The asymmetric shape of an ammonia molecule results in an unequal charge distribution and the molecule is polar.

•*Note: If bonds are polar, asymmetrical molecules are polar and symmetrical molecules are nonpolar.*


Ammonia
H
N
H
H
107.3°
H
N
H
H
Tetrahedral Electron
Pair Geometry
Trigonal Pyramid
Molecular Geometry
B
120°
Arrangement of
electron pairs
around B in BCl3
Cl
B
120°
Cl
Cl
Molecular geometry:
Trigonal planar

Bond polarity and intermolecular forces

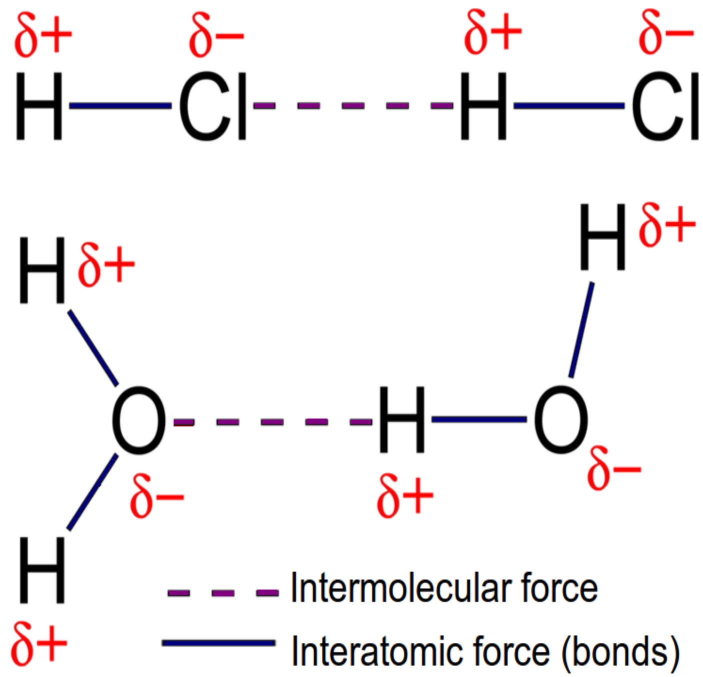
- Bond polarity is referred to unequally sharing of electron pairs between two atoms, e.g.

- Atoms such as nitrogen, oxygen and halogens that are more electronegative than carbon have a tendency to have partial negative charges.

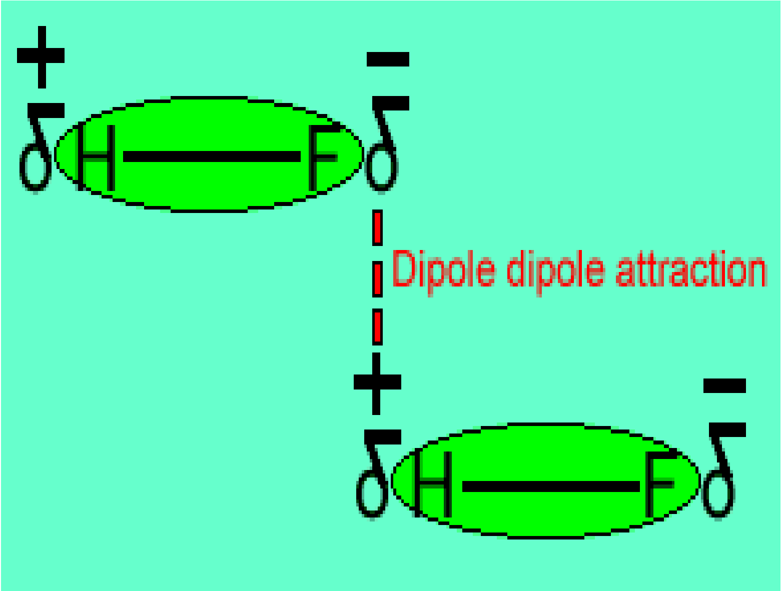
- Atoms such as carbon and hydrogen have a tendency to be more neutral or have partial positive charges.

- Bond polarity arises from the difference in electronegativity of two atoms that participated in the bond formation and also depends on the attraction forces between molecules that are called intermolecular interactions or forces.

- The physical properties, e.g. boiling point, melting points and solubility of the molecules are determined by intermolecular nonbonding interactions.
- There are three types of intermolecular nonbonding interactions: dipole-dipole interactions, van der Waals forces and hydrogen bonding.
- These interactions increase as significantly as the molecular weights increase and also increase with increasing polarity of the molecules.



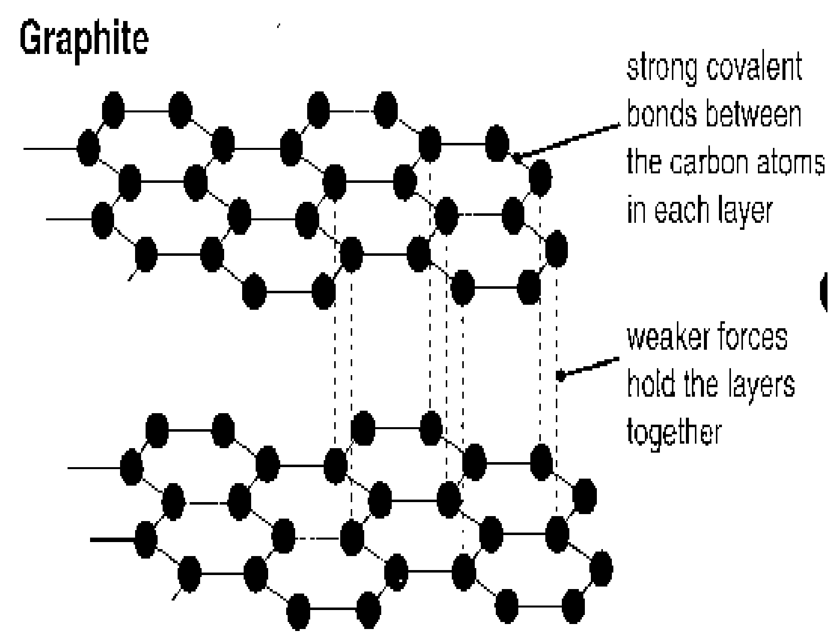
1- Dipole-dipole interactions are the interactions between the positive end of one dipole and the negative end of another dipole and due to this interaction, polar compounds are held together more strongly than nonpolar molecules.



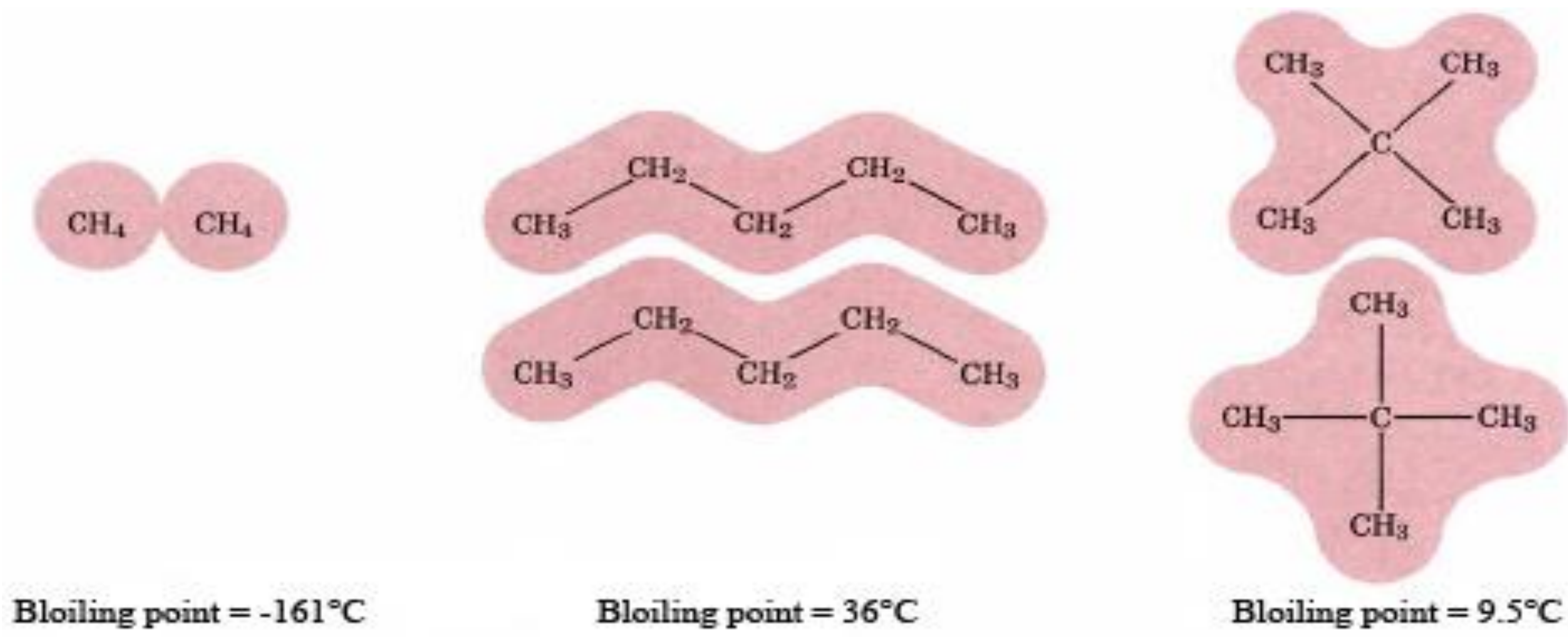
e.g. HF has a dipole moment of 1.98 D because F atom has greater electronegativity than the H atom and it pulls the electrons towards it, leaving partial positive charge on H and partial negative charge on F.

2- Van der Waals forces (London dispersion forces) are relatively weak forces of attraction that exist between nonpolar molecules. These forces are much weaker than the covalent bonds within molecules. Electrons move continuously within bonds and molecules, so at any time one side of the molecule can have more electron density than the other side, which gives rise to a temporary dipole.

- Because the dipoles in the molecules are induced, the interactions between the molecules are also called induced dipole–induced dipole interactions.

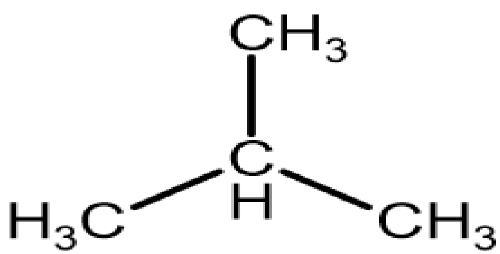


e.g. alkenes are nonpolar molecules because the electronegativity of carbon and hydrogen are similar. Therefore there are no significant partial charges on any of the atoms in an alkane. The size of the van der Waals forces that hold alkane molecules together depends on the area of contact between the molecules. The greater the area of contact, the stronger are the van der Waals forces, and the greater is the amount of energy required to overcome these forces.

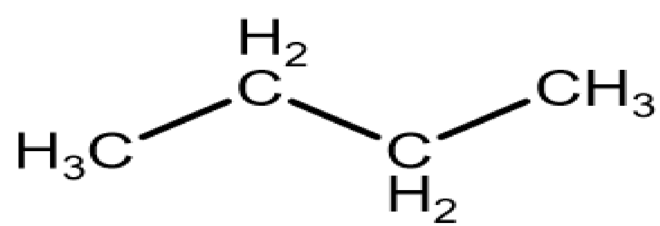


For example, isobutane (b.p.-10.2 °C) and butane (b.p. - 0.6 °C), both with the molecular formula C_4H_{10}, have different boiling points. Isobutane is a more compact molecule than butane. Thus, butane molecules have a greater surface area for interaction with each other than isobutane. The stronger interactions that are possible for n-butane are reflected in its boiling point, which is higher

than the boiling point of isobutane.



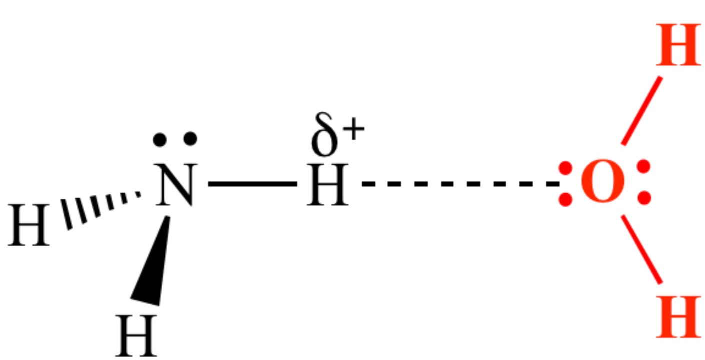
isobutane

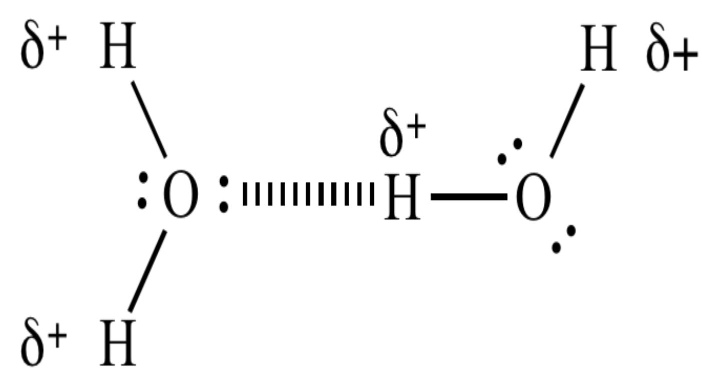


n-butane
butane

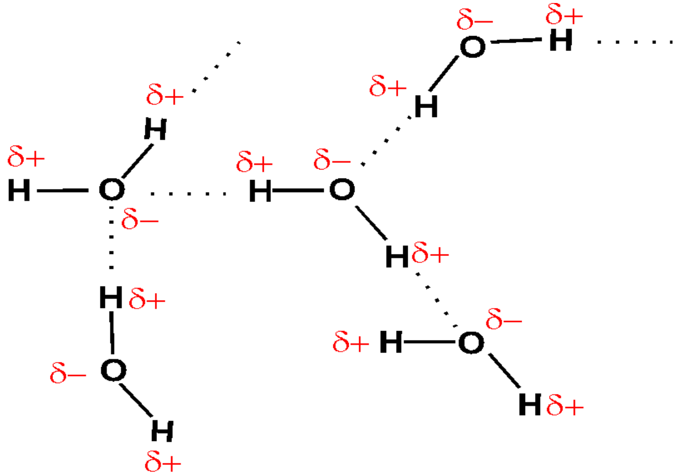
3- Hydrogen bonding is the attractive force between the hydrogen attached to an electronegative atom of one molecule and an electronegative atom of the same (intramolecular) or a different molecule (intermolecular).

- It is an unusually strong force of attraction between highly polar molecules in which hydrogen is covalently bonded to nitrogen, oxygen or fluorine.
- A hydrogen bond is formed whenever a polar covalent bond involving a hydrogen atom is in close proximity to an electronegative atom such as O or N.





- The attractive forces of hydrogen bonding are usually indicated by a dashed line rather than the solid line used for a covalent bond.
- Water is a polar molecule due to the electronegativity difference between hydrogen and oxygen atoms. The polarity of the water molecule with the attraction of the positive and negative partial charges is the basis for the hydrogen bonding.
- Hydrogen bonding is responsible for certain characteristics of water, e.g. surface tension, viscosity and vapour pressure.



- Hydrogen bonding occurs with hydrogen atoms covalently bonded to oxygen, fluorine or nitrogen, but not with chlorine, which has larger atom size. This bonds is the strongest known type of intermolecular interaction.
- The intermolecular hydrogen bonding in water is responsible for the unexpectedly high boiling point of water (b.p. 100 °C).
- Hydrogen bonding is usually stronger than normal dipole forces between molecules, but not as strong as normal ionic or covalent bonds.

Ionic bond > covalent bond > Hydrogen bonding > Dipole forces > van der Waals forces

- Hydrogen bond is called bond of life because The double helix structure of DNA is formed and held together with hydrogen bonds.
- E.g enzymes, hormones, cellulose and hydrogen bonding