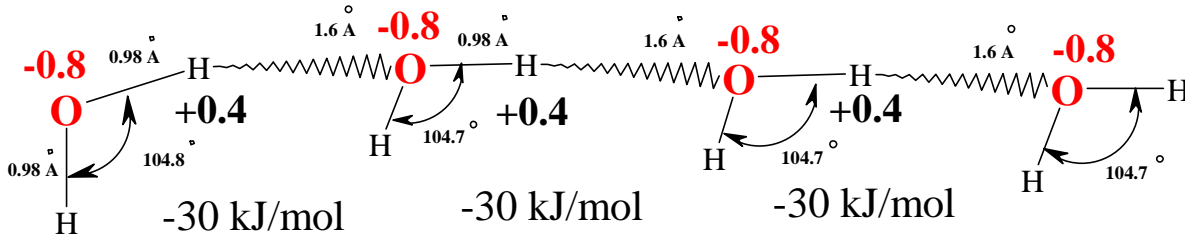


INTERMOLECULAR BONDS

Hydrogen bond polar water molecule attraction force

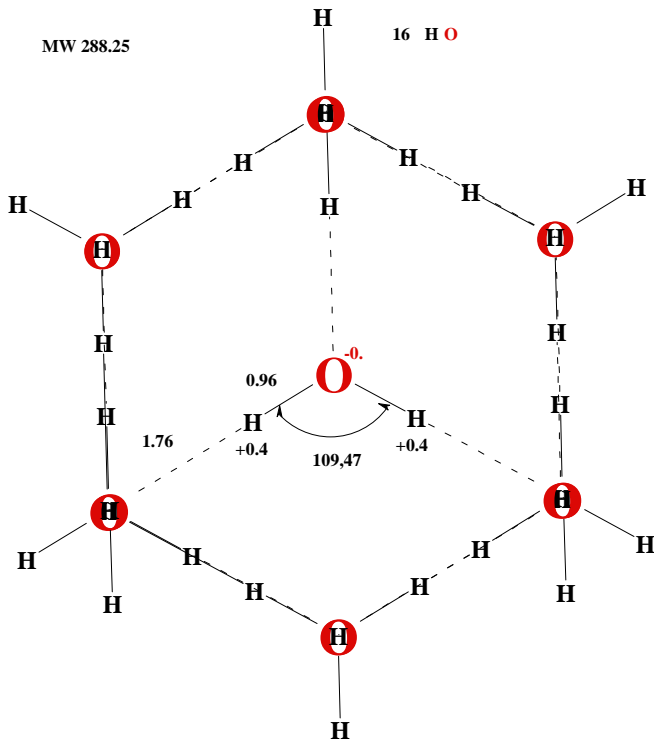
Liquid water $4 \text{H}_2\text{O} \Rightarrow (\text{H}_2\text{O})_4$ tetramer structure all water molecules are hydrogen bonding clustered no one free water molecule in liquid state:



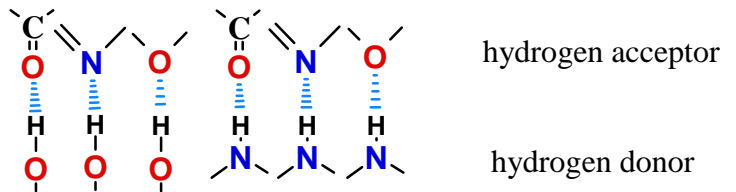
Three hydrogen bonds as polar water interaction resulting in tetramer $(\text{H}_2\text{O})_4$ three hydrogen bonds energies sum $\Delta G = -90 \text{ kJ/mol}$.

Ice has clustered the water molecules by hydrogen bonding lattice in hexagonal structure.

Pure hydrogen bonding **crystalline** lattice hexagonal structure.



1. 1. Hydrogen bonds of **O** and **N** atoms between hydrogen atom **H** acceptor and other side **H** donor:



2. Salt bridge $\text{AspCOO}^- \dots ^+\text{H}_3\text{Nlys}$

3. **Hydrophobic bonds** . water tetramer press together non polar \diamond atomic groups $(\text{H}_2\text{O})_4 \rightarrow \diamond \leftarrow (\text{H}_2\text{O})_4$ residues of amino acids or base pairing tablets in DNA and RNA molecules

4. **Coordinative bonds**

coordinative donor acceptor bond

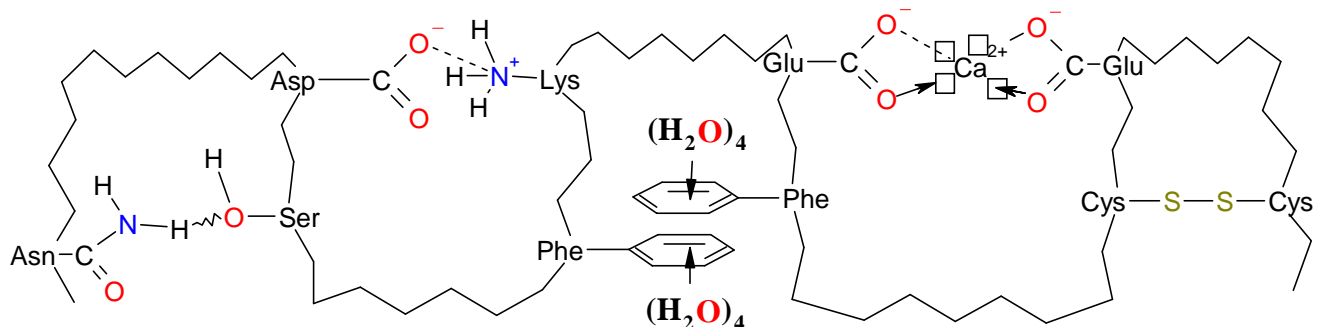
calcium ion with carboxyl groups

$-\text{COO}^- \rightarrow \square \text{Ca}^{2+} \leftarrow \text{:OO}^-$ or **iron(II)** ion

on center of hem $\rightarrow \square \text{Fe}^{2+} \leftarrow$

5. **Disulphide bonds** Cys-S-S-Cys

5 Intermolecular forces (intermolecular bonds) for proteins and other bio polymer molecules



HYDROGEN BOND

This far we have been discussing the bonds inside one molecule. However, bonds and electrostatic interactions can exist also between two or several molecules. Such bonds are called *intermolecular bonds*.

The Physiologic most important kind of intermolecular bonds is *hydrogen bond*.

Hydrogen bonds are established between molecules in compounds, in which a hydrogen atom is linked to a strongly electro negative (EN) element **F, O, N, Cl**, where physiological significant are **O** and **N**.

In these cases the common electron pair, which carries out the bond between **H** and the other atom, is shifted very much towards the very electronegative (EN) atom (**O, N**). For this reason the **1s** orbital of **H** atom can be considered as “half-empty” □ **H**. If the orbital of **H** atom is half-empty □ **H**, it (as well as any empty orbital) can be used for donor-accepter bond formation. In other words, an unshared electron pair of another particle can be inserted into the empty orbital of □ **H** and □ **H** atom can act as an accepter of unshared electron pair.

Let us take **H₂O** as an example. The nucleus of oxygen atom is surrounded by four unshared electron pairs ::::. Two of these electron pairs is used for formation of bond with **H** atom. Now one of the remaining two electron pairs can be placed into the half-empty orbital of the **H** atom of **another H₂O** molecule:

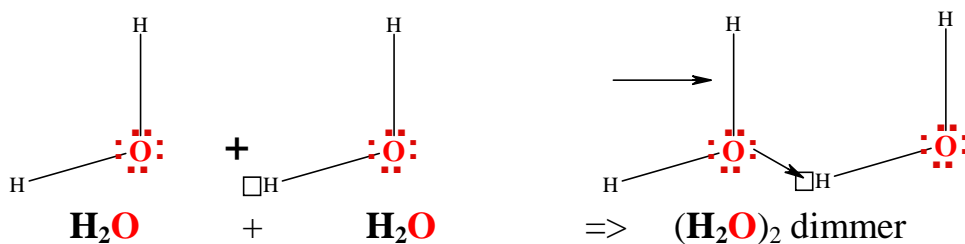


Fig.11.15. Formation of hydrogen bonding between two molecules of **H₂O**.

(the electron pairs of oxygen are shown as pairs of dots ::, ::, ::, ::)

In such a way, *hydrogen bond is a kind of donor-accepter bond*. As, in fact, the **1s** orbital of **H** is only partly empty and not completely empty, the strength of hydrogen bond is smaller, than the one of usual chemical bonds. The bond energy of hydrogen bond in different cases **EN** (of **F, O, N, Cl**) reaches **8-40** kJ/mol while strengths of a real chemical bond inside one molecule usually is **150-400** kJ/mol. For water molecule **30** kJ/mol.

One must note that *hydrogen is the only element*, which is able to form this kind of bonds (and it is the reason why this type of bond has been named *hydrogen bond*).

Let us explain, why no other element can form a bond similar to hydrogen bond. A hydrogen atom has just one electron on its only **1s**-orbital. If this electron is shifted towards an atom of a strongly electronegative element, there practically is just a nucleus of hydrogen atom instead of the atom itself. An atomic nucleus is always positively charged, hence it **attracts** an unshared electron pair, offered by another atom. The unshared electron pair of the other atom can therefore easily be inserted into the partly empty **1s** orbital of hydrogen atom forming a hydrogen bond.

All the elements other than hydrogen have more than one electron shell surrounding their nuclei. For instance, an atom of **Li** has two electron shells, but as well as an atom of **H** has just one electron in its outer shell. The difference is, that **Li** has this electron on its **2s** orbital and if this only electron of the outer shell is displaced towards a more electronegative element and the **2s** orbital of **Li** becomes partly empty, two electrons still remain at **1s** orbital of **Li** atom. If another element tries to insert its unshared electron pair into the partly empty orbital of **Li**, the remaining two electrons from **1s** orbital of **Li** will **repel** this unshared pair and no hydrogen-type bond will be established.

The Physiologic role of **H** bond is very important for life maintenance. Let us see some examples.

1. Due to hydrogen bond water is dimer (H_2O)₂, trimer, tetramer and no monomer H_2O molecule form.

(H_2O)₂ in its solutions dissociates in two stages, 1st, cleaving one proton: $(\text{H}_2\text{O})_2 \Rightarrow \text{H}^+ + \text{H}_2\text{O} \cdot \text{HO}^-$

The hydrogen bond still keeps together a molecular ion $\text{H}_2\text{O} \cdot \text{HO}^-$. $\text{H}_2\text{O} \cdot \text{HO}^- \Rightarrow \text{H}^+ + \text{HO}^- + \text{HO}^-$
 At the second dissociation stage one more hydrogen ion cleaved and now two individual, negatively charged hydroxide ions are formed, as there is no more hydrogen bond to keep them together.

What is even more important, due to the hydrogen bonding water is a very weak acid, that could seem surprising, if one takes into account **OH** group the great acid strength of others H_2SO_4 and HNO_3 .

2. Hydrogen bonds are responsible for the structure of water both in the state of ice and in the liquid state.

In fact, if one measures the molecular mass of liquid water, the result is about **72** instead of **18**, that can be calculated from the formula of H_2O and is observed in gas phase. The reason for this is, that water molecules in liquid water are associated in tetrahedric particles, consisting of **4** molecules of H_2O , bonded by **H** bonds. The high melting point of water **0°C** is also due to the presence of hydrogen bonds (compare to the melting point **-86°C** of H_2S). Thus, if **H** bonds didn't exist, water on the Earth would exist in the gas phase only.

3. Hydrogen bonds are responsible for formation of dimmers in liquid alcohols and organic acids. In both of these compounds the hydrogen atom of the **OH** group is linked to a very electronegative element. As well as in the example of water, the electron, which originally belonged to hydrogen is very much shifted towards oxygen and the **1s** orbital of **H** atom is partly empty. Oxygen atom from the neighboring alcohol (organic acid) molecule inserts its unshared electron pair into this partly empty orbital, a hydrogen bond is established and the two neighboring molecules further exist in a form of dimer, see an example in fig.11.16.

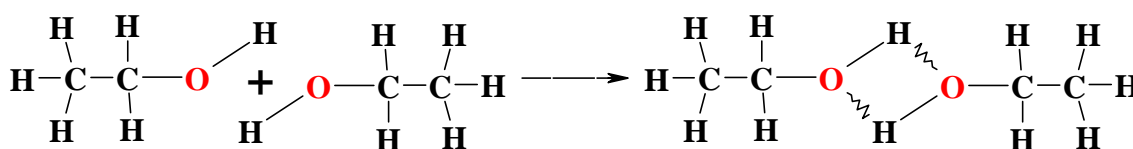


Fig.11.16. Dimerization of ethyl alcohol.

Hydrogen bonds between hydrogen and oxygen atoms of neighboring molecules are shown as wave lines

4. Finally, hydrogen bonds are responsible for both the helix-type α -structure and the belt-type β -structure as the possible kinds of protein secondary structures. As well, hydrogen bonds are responsible for a great part of distortions and folding of protein secondary structures in known as tertiary structures of proteins. This point deserves a longer discussion, that can be found in these studies of Medical Chemistry.

POLARITY OF COVALENT BONDS.- POLARITY OF MOLECULES

When a covalent bond is formed between two atoms, having equal **EN**, the bond is non-polar. As soon as the **EN** of both atoms is different, the bond between the two atoms becomes a *dipole*.

A system, consisting of two electric charges equal by value but having opposite signs, that are placed in distance l from each other, is called a dipole.

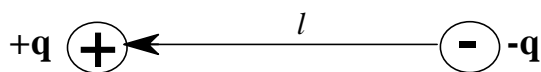


Fig.11.13. A dipole $\mu = q \cdot l$ moment for **H₂O** $\mu = 0.8 \text{ el} \cdot 2.61 \text{ \AA} = 2.088 \text{ el} \cdot \text{\AA}$.

Polarity of the bond is characterized by the dipole momentum

The polarity of bond depends both on the value of the two opposite charges and on the distance l between them. An increase of l causes a greater polarity as well as an increase of q .

If there are several bonds in the molecule, the polarity of the overall molecule can be found, following certain rules (note, that the molecule can be non-polar, although there are polar bonds in it!):

1) If there is just one polar bond in the molecule, the molecule is **polar**. For instance in **HCl** the bond is polar and the molecule is polar, as well. In **CHCl₃** the **C-H** bonds are considered to be non-polar (the difference of **EN** between **C** and **H** is ignorable), but the **C-Cl** bond is polar, therefore the overall molecule is polar.

2) If there are several polar bonds in the molecule, polarity of molecule depends on the symmetry of molecule:

a) if the polar bonds are situated symmetrically, the molecule is **non-polar**. For instance, in the molecule of **CO₂** (it is linear, see above) both **O=C=O** bonds are oriented symmetrically and their dipole moments compensate each other, and the overall molecular $\mu = 0$. Carbonate anion of **CO₃²⁻**, having 3 polar bonds, is also non-polar for the same reasons - the three **sp²** symmetric orbitals of carbon atom are symmetrically positioned in one plane but anion charge -2 is delocalised even around three oxygen atoms as **CO₃²⁻**.

b) if the polar bonds are situated non-symmetrically, their dipole moments don't compensate each other, therefore the molecular $\mu \neq 0$ and the molecule is polar. For instance, **CHCl₃** is a polar molecule, **H₂O** is a polar molecule and **NH₃** is a polar. Molecular dipole moment μ .

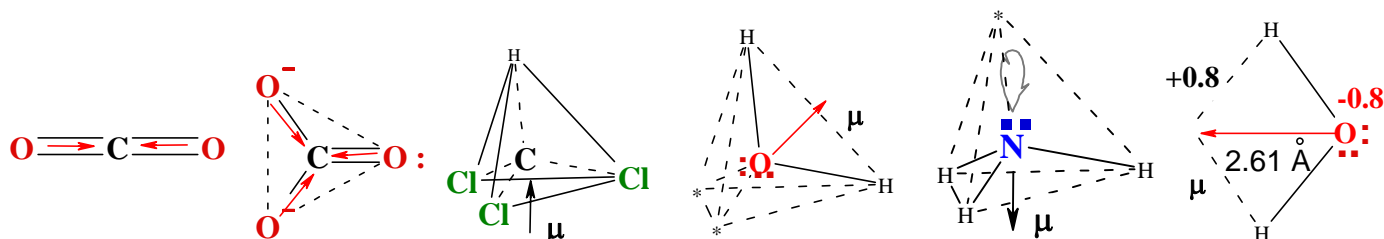


Fig.11.14. Symmetric and symmetric absence bonded atoms creates polar bonds in the molecules and ions.

POLARIZATION OF MOLECULES IN OUTER ELECTRICAL FIELD

When a molecule is subjected to an outer electrical field, the dipole momentum of the molecule is changed - non-polar molecules become polar and polar molecules become more polar. When a molecule is subjected to electric field, its dipole momentum becomes:

$$\mu = \mu^0 + \mu_{\text{ind}},$$

where

μ^0 is the dipole momentum of the molecule itself ($\mu^0 = 0$ for a non-polar molecule or $\mu^0 > 0$ for a polar one) and

μ_{ind} is the dipole momentum, induced by the outer field.

The value of the induced dipole momentum can be found as:

$$\mu_{\text{ind}} = E \cdot \alpha,$$

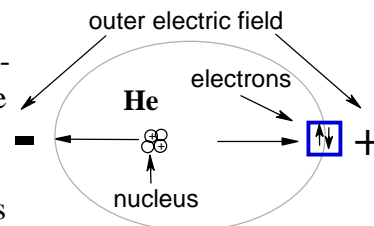
where

E is the intensity of the electrical field (volts/m) and

α is the *polarizability* of the molecule.

Polarization of the molecule in the outer electrical field originates in such a way, that the negative pole of outer field attracts the positively charged atomic nuclei, but the positive pole of the outer field attracts the negatively charged electrons. In such a way, the positive and the negative charges are dragged to opposite directions, the distance l between the centers of positive and negative charges becomes greater and the dipole momentum is increased.

Electron mass is 7294,3 times smaller, than He atom nucleus, the polarization is practically carried out as a distortion of the "electron cloud" around the molecule - it is shifted towards the positive pole of the outer electric field. For this reason the polarizability of molecule depends on the size of its "electron cloud". (In other words, the nuclei are too heavy to be displaced under the action of electric field and the action of outer electric field practically causes a displacement of electrons only). Frequency of green $\lambda = 555$ nm photons $f = 540 \times 10^{12}$ Hz;



POLARIZATION IN CHEMISTRY

POLARIZATION UNDER ACTION OF WATER

When chemical compounds are dissolved in a polar solvent, the electric field of the solvent (e.g. water) molecules provides an extra polarization to the molecules of solute. For instance, **HCl**, which has just a polar covalent bond in the gas phase, is additionally polarized by the electric field of water molecules in solution, therefore the dipole momentum of **HCl** molecules becomes great enough to become able to dissociate into ions.

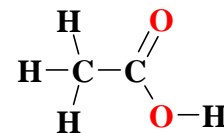
*Due to the additional polarization under action of the electric field of water molecules, **HI** in water solution is a stronger acid than **HBr**, which, on its turn, is a stronger acid, than **HCl**. This phenomenon looks strange at the first sight: as the electro negativity of **Cl** is greater, than the one of **Br** and **I**, a molecule of **HCl** in gas phase is more polar, than the ones of **HBr** or **HI**.*

When these three compounds are dissolved in water, their molecules are additionally polarized by the electric field of water molecules. The value of the induced dipole momentum depends on the size of the "electron cloud" around the molecule - the greater it is, the easier electrons are shifted towards the positive pole of the outer field in the molecules of hydrogen halogenides. The size of the molecular electron cloud is mainly determined by the one of the appropriate halogen atom, because the **1s** orbital of hydrogen is rather small. As the size of an **I** atom is greater, than the size of **Br** and **Cl** atoms, the molecules of **HI** can be polarized much easier, than the molecules of **HBr** and **HCl**.

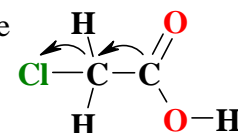
Thus, **HI** happens to be the strongest of these three acids in their water solution due to its greater polarizability, although its μ^\ominus was smaller than the one of **HBr** or **HCl**. For the same reason **H₂S** is more acidic than **H₂O** and **PH₃** is less basic (more acidic) than **NH₃** etc.

INTRAMOLECULAR POLARIZATION

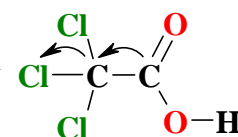
In many cases, especially in organic chemistry, introduction of a substituent atom (**O**, **N**, **Cl**), having a great electro negativity **EN**, causes a polarization of all bonds in the molecule.



For instance, when a hydrogen atom in acetic acid is replaced by a more electronegative atom (of **O**, **N**, **Cl**), which is highly electronegative (**EN**), attracts electrons and therefore shifts the electron density of the entire molecule:



As a result, the **O-H** bond in the carboxyl group becomes more polar and therefore chloroacetic acid appears to be a much stronger acid than acetic acid, compare the dissociation constants $K = 1.75 \cdot 10^{-5}$ for acetic and $K = 1.4 \cdot 10^{-3}$ for chloroacetic acid. If all the three hydrogen atoms in the methyl group are substituted by **Cl** atoms, forming trichloroacetic acid one gets an acid, the strength of which can be compared to hydrochloric acid.

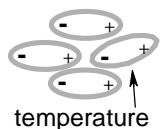


INTERMOLECULAR (VAN DER WAALS'S - London) FORCES.

Weakest interactions between the molecules are called Van der Waals's or London forces. Sometimes called also Van der Waals's bonds. Molecular coordinates viewers have Van der Waals's electron orbitals.

By the most, they act in solid and liquid states (composite lipid's formed cell membranes of phospholipids and cholesterol or other sterids), as in the gas phase the distances between the molecules are great and, as these forces are not very strong, they are not able to keep the molecules together. Three main types of Van der Waals's or London forces are known, called *orientation, induction and dispersion forces*.

ORIENTATION FORCES.



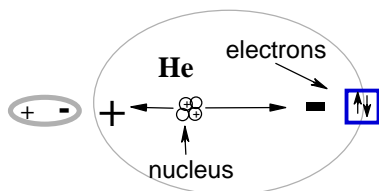
Orientation forces can be established between two polar molecules (for water dipoles). When two polar molecules come close to each other, they are oriented so, that the negative pole of one molecule can interact electrostatic with the positive pole of the other one.

As the result, the molecules can stick together for a short period of time. An increase of temperature weakens the orientation forces, as the thermal motion of molecules becomes more intensive at higher temperatures.

As both of the molecules are polar, both dipole moments are used in this kind of intermolecular interaction, therefore the orientation forces are the strongest ones out of all the 3 kinds of Van der Waals's interactions.

INDUCTION FORCES.

Induction forces exist between a polar and a non-polar molecules.



When a polar molecule comes close to a non-polar one, a dipole is induced in the non-polar molecule. If the polar molecule comes to the non-polar with its positive pole forward, the electron cloud of the non-polar molecule is shifted towards the positive pole of the polar molecule. If the polar molecule comes with its negative pole forward, the electron cloud of the non-polar molecule is shifted away from the negative end of the polar molecule. In both cases a dipole is induced in the formerly non-polar molecule and now this induced dipole can interact electrostatic with the dipole of the polar molecule.

As in this case only the polar molecule "works" - it induces a dipole in the other molecule and after that interacts with the induced dipole, one can easily understand, that the efficiency of the interaction must be weaker, than in the case of an orientation type interaction.

DISPERSION FORCES.

Dispersion forces act between two non-polar molecules. Any non-polar molecule looks non-polar only, if it is observed in a time interval longer as 10^{-15} femto seconds. In fact, due to the motion of electrons, a non-polar molecule exists in at least three different states:



(the common pair of electrons is shown as a pair of dots)

As one can see, part of time the common pair of electrons can be found behind atom **A**, as in the first state (then atom **A** obtains a partly negative charge and atom **B** - a partly positive one), a part of time the electron pair can be found between the atoms and the molecule is really non-polar and a part of time - when the common electron pair is behind the atom **B** the polarity is opposite to the one in the first case.

Thus, although the molecule in average is non-polar, at long time intervals (these time intervals can be about 10^{-14} seconds long). If the two non-polar molecules happen to meet each other at a moment, when their polarities are opposite, they can interact electrostatic by attraction with dispersion London forces.

In such a case the short-living positive pole of one molecule attracts the short-living negative pole of the other molecule and the molecules are kept together for some time. Of course, this kind of interaction is the weakest of the discussed ones, but it cannot be ignored. For instance, the existence of induction forces is the only possible explanation for adsorption of non-polar molecules on non-polar surfaces, that is very well known and can have measurable values.