Intermolecular Forces (IMF's)

Molecules are independent, neutral particles. As we have seen, they have a particular shape and consist of a definite composition of atoms that are covalently bonded together.

Molecular substances are found in all three phases as gases, liquids, or solids. However, in order to form into the liquid or solid phases, the molecules must have some kind of attractive force that allows them to stick together.

Inter- = between Intra- = within

Intermolecular forces are forces that act BETWEEN molecules. These are the forces that stick molecules to each other or to other particles (such as ions.)

Three Predominant IMFs

1. **Dipole-Dipole Interactions**
2. **Hydrogen Bonding**
3. **Dispersion Forces**

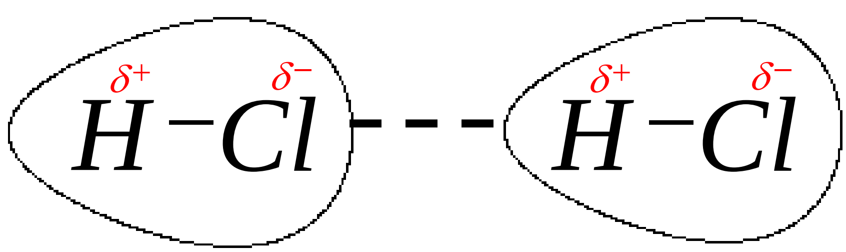
Dipole-Dipole interactions, are caused by permanent dipoles in molecules. Molecules are polar due to two main reasons:

- having polar covalent bonds
- having certain molecular shapes

Polarity due to polar bonds

When one atom is covalently bonded to another with a significantly different electronegativity, the electronegative atom draws the electrons in the bond nearer to itself, becoming slightly negative. Conversely, the other atom becomes slightly positive. Electrostatic forces are generated between the opposing charges and the molecules align themselves to increase the attraction (reducing potential energy).

An example of dipole-dipole interactions due to bond type can be seen in hydrochloric acid (HCl):



dipole molecule **dipole molecule**

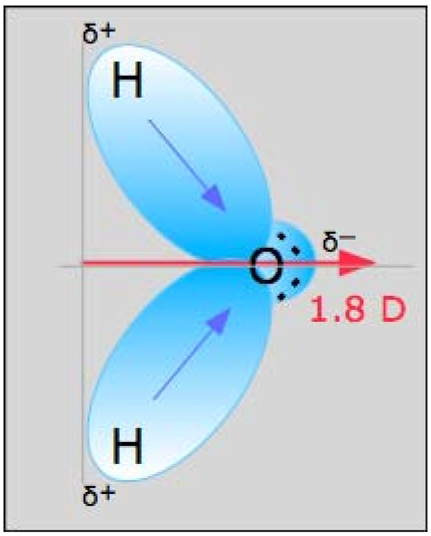
The H-Cl bond has an EN difference of 0.96 relating to a % ionic character of about 21%. This bond therefore is polar covalent.

Dipoles behave like tiny bar magnets.
(+)(-) attract, but (+)(+) and (-)(-) repel.

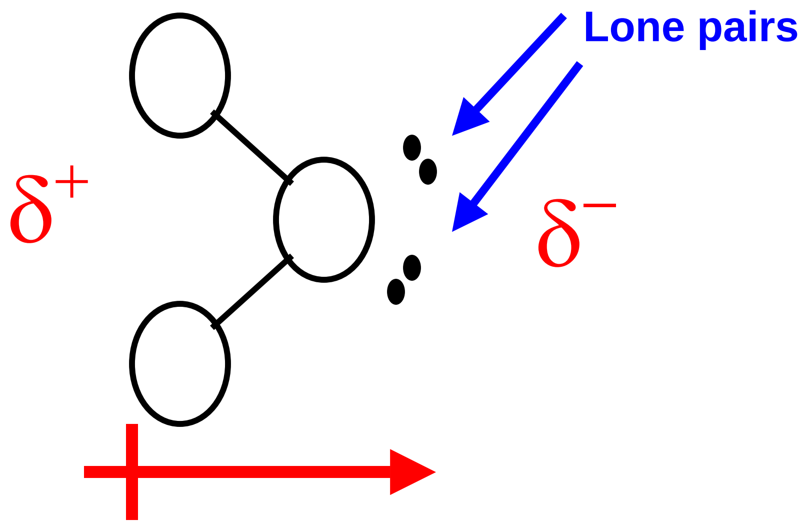
Polarity due to molecular shape

When the shape of a molecule causes the charge density to be un-evenly distributed, a polar situation is created.

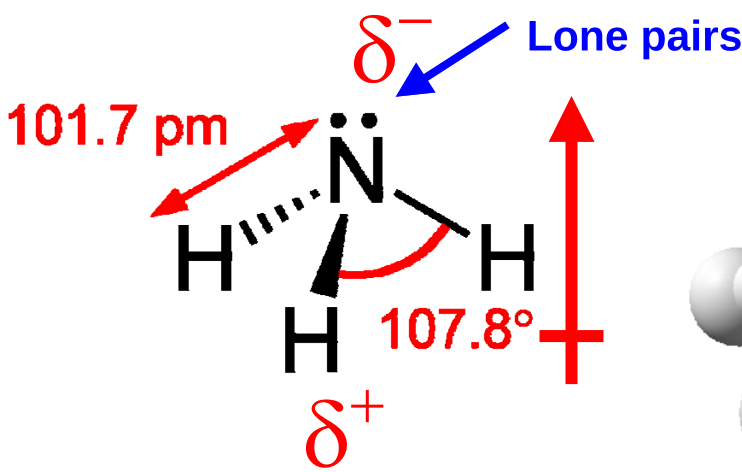
An example of dipole-dipole interactions due molecular shape can be seen in water (H_2O):

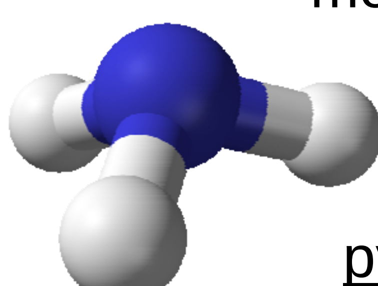


The water molecule is polar because it has a <u>bent shape</u>.



Another example of dipole-dipole interactions due molecular shape can be seen in ammonia (NH_3):





The ammonia molecule is polar because of it's <u>pyramidal shape</u>.

Hydrogen Bonding (H-Bonding)

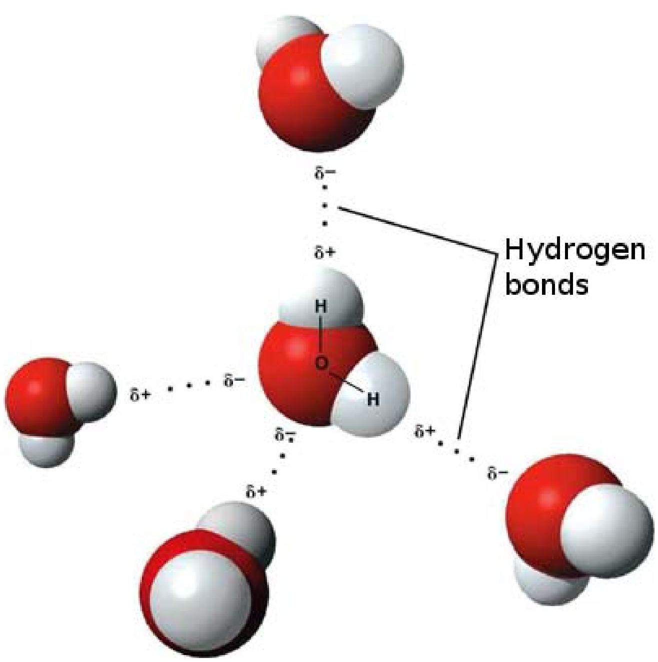
Hydrogen bonds are caused by highly electronegative atoms. They only occur between hydrogen and oxygen, fluorine or nitrogen, and are the strongest intermolecular force.

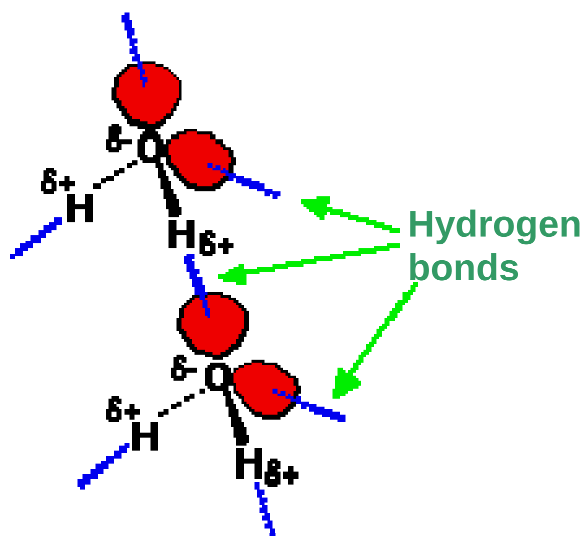
The high electronegativities of F, O and N create highly polar bonds with hydrogen, which leads to strong bonding between hydrogen atoms on one molecule and the lone pairs of F, O or N atoms on adjacent molecules.

Therefore to have H-bonding you must have both:

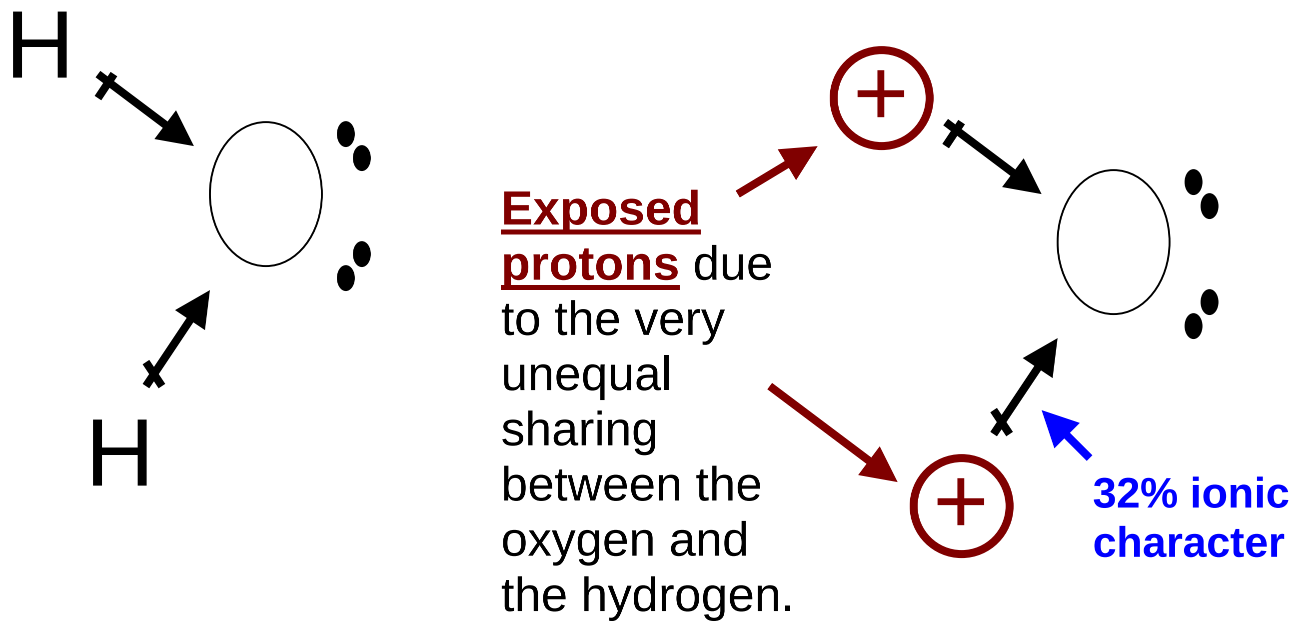
- H – bonded to either F, O, or N
- Lone pairs of electrons available on the F, O, or N

An example of hydrogen bonding between water molecules is:

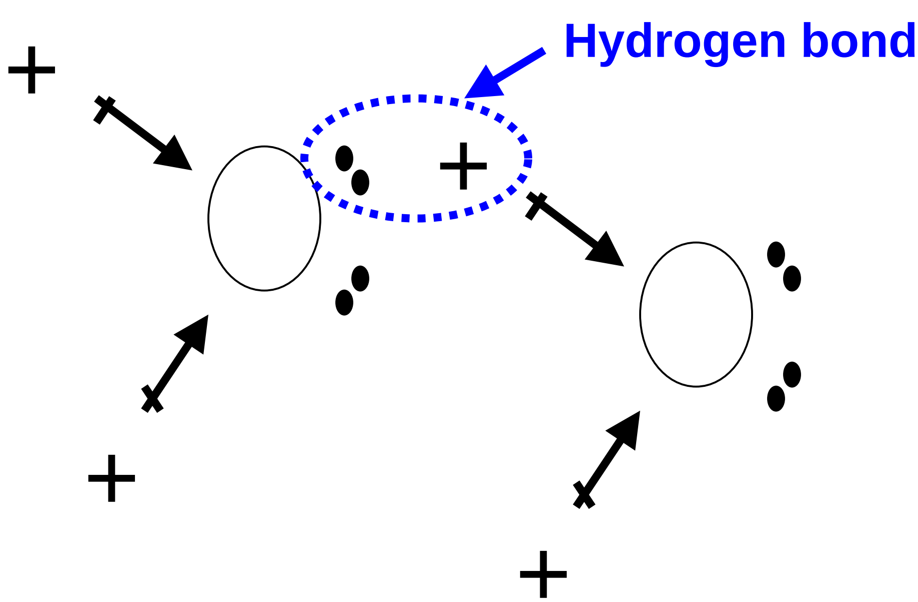




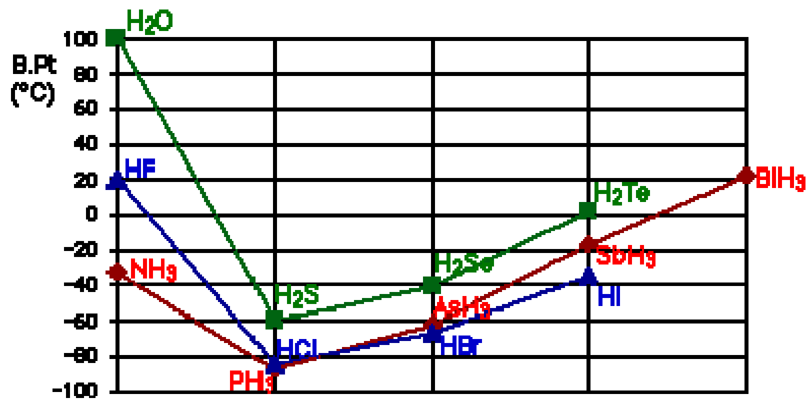
How does H-bonding work?



The hydrogen atom is very much exposed (naked) due to the unequal sharing with the oxygen. The naked proton is attracted to the lone pairs of electrons on other molecules. A temporary association of these two parts of adjacent molecules produces the hydrogen bond.



Evidence for hydrogen bonding:

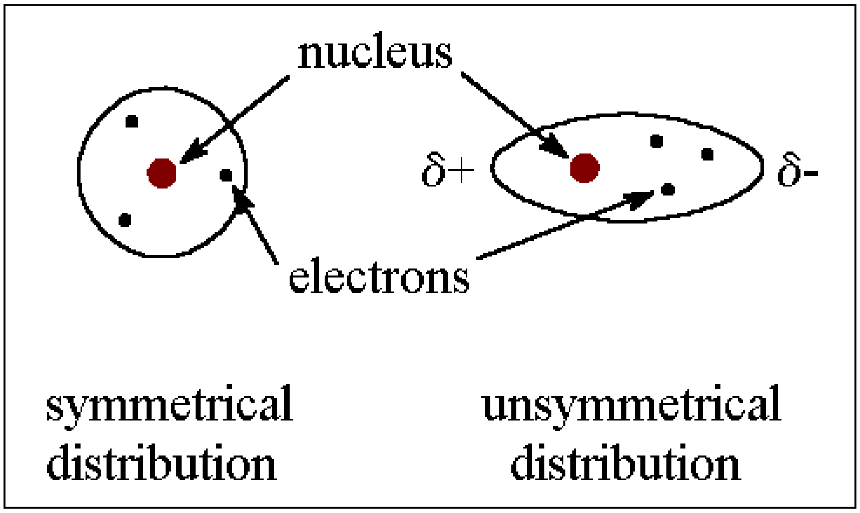


The high boiling points of water, hydrogen fluoride (HF) and ammonia (NH_3) is an effect of the extensive hydrogen bonding between the molecules.

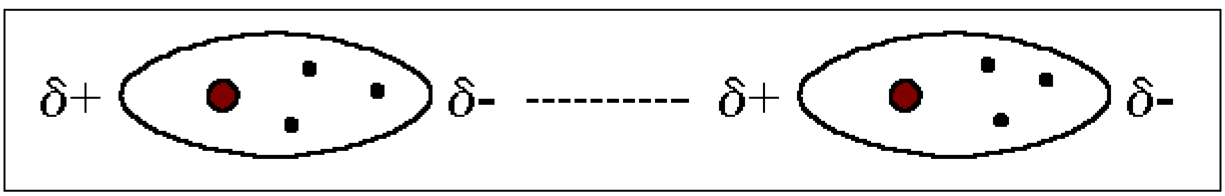
London Dispersion Forces (LDFs)

The London dispersion force is caused by random and temporary changes in the polarity of atoms, caused by the location of the electrons in the atoms' orbitals. The location of an electron in an orbital is random. When an electron is on one side of the nucleus, this side becomes slightly negative; this in turn repels electrons in neighboring atoms, making these regions slightly positive. This <u>induced dipole</u> causes a brief electrostatic attraction between the two molecules (or atoms). When the electron eventually moves to another location, the electrostatic attraction is broken and the dipole disappears.

London forces are the attractive forces that cause nonpolar substances to condense to liquids and to freeze into solids when the temperature is lowered sufficiently.



Because of the constant motion of the electrons, an atom or molecule can develop a temporary (instantaneous) dipole when its electrons are distributed unsymmetrically about the nucleus.



While LDFs are generally the weakest of the IMFs, in very large molecules they can be comparatively the strongest.

The main strength factor is the number of electrons.

More electrons = stronger LDF

(more polarizable)

This phenomenon is the only attractive intermolecular force at large distances present between neutral atoms (e.g., a noble gas), and is the major attractive force between non-polar molecules, (e.g., nitrogen or methane). Without London forces, there would be no attractive force between noble gas atoms, and they wouldn't exist in liquid form.

London forces become stronger as the atom or molecule in question becomes larger. This is due to the increased polarizability of molecules with larger, more dispersed electron clouds. This trend is exemplified by the halogens (from smallest to largest: F_2, Cl_2, Br_2, I_2).

Halogen element	Number of electrons	Phase at 25°C
F	9	Gas
Cl	17	Gas
Br	35	Liquid
I	53	solid

The London forces also become stronger with larger amounts of surface contact. Greater surface area means closer interaction between different molecules.

The first explanation of the attraction between noble gas atoms was given by Fritz London in 1930.

van der Waal's Forces

Johannes van der Waals a Dutch physicist first proposed that molecules have attractive forces between them in the late 1800s. Since he was one of the first to postulate an intermolecular force, such a force is now sometimes called a **van der Waals force**.

It is also sometimes used loosely as a synonym for the totality of intermolecular forces.

Comparing the Relative Strength of Intermolecular Forces

Bond type	Dissociation energy (kJ)
Covalent	**1675**
Hydrogen bonds	**50-67**
Dipole-dipole	**2 - 8**
London Dispersion Forces	**< 4**

Note: this comparison is only approximate AND only for relatively small molecules of approx. equal size- the actual relative strengths will vary depending on the molecules involved.