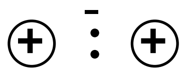
Bonding

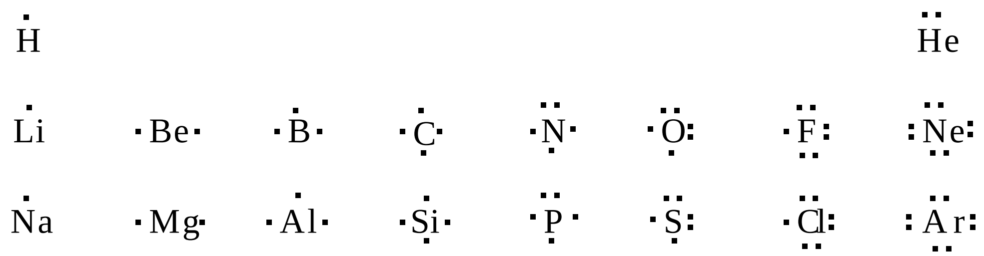
Bonding Between Atoms

Atoms in a **covalent** bond are held together by electrostatic forces of attraction between positively charged nuclei and negatively charged electrons.

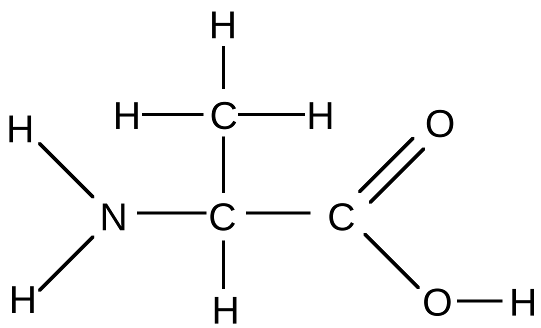


Usually one electron is supplied by each atom so the number of pair bonds (A:A or A-A) an atom can form depends on its valency (number of unpaired electrons in the outer energy level).

Valency 1 2 3 4 3 2 1 0



e.g. Alanine (an amino acid)



	No. of bonds
H	1
O	2
N	3
C	4

ELECTRONEGATIVITY

An atom's **electronegativity** is its power to attract the bonding pair of electrons.

Electronegativity increases across a period in the periodic table (L->R) because the nuclear charge increases.

Electronegativity **decreases** down a group (T->B). As the atoms get bigger the electrons get further away from the nucleus and are therefore less tightly held.

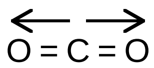
BOND POLARITY

The electron pair is distorted slightly towards the more electronegative atom making it slightly negative (δ^-) and the other slightly positive (δ^+) i.e. the bond is **polar**

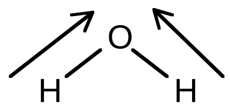
e.g.

$\overset{\delta+}{C} = \overset{\delta-}{O}$; $\overset{\delta+}{C} - \overset{\delta-}{Br}$; $\overset{\delta+}{H} - \overset{\delta-}{O}$; $\overset{\delta+}{H} - \overset{\delta-}{Cl}$

The overall polarity of the molecule depends on its shape :

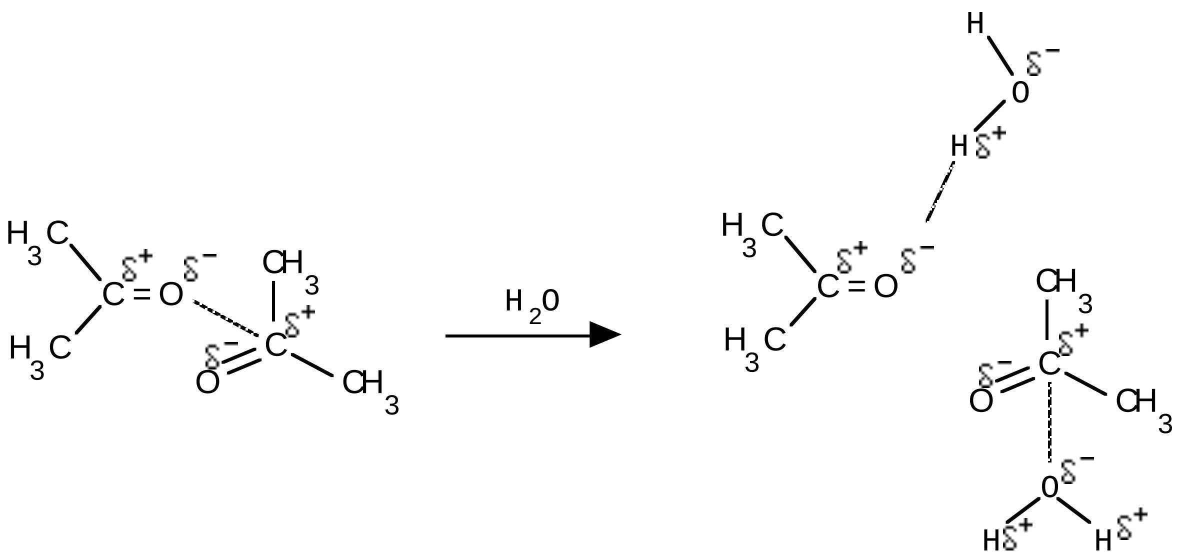


is non-polar (the two polar bonds oppose, cancelling the effect)

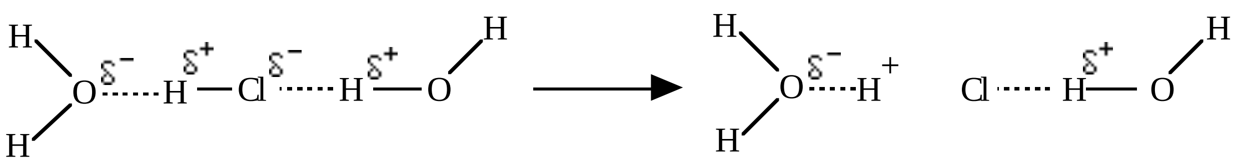


is polar (the two polar bonds do not cancel the effect)

In general, polar substances dissolve in polar solvents like Water which are able to pull the molecules apart e.g. Propanone dissolves in Water :

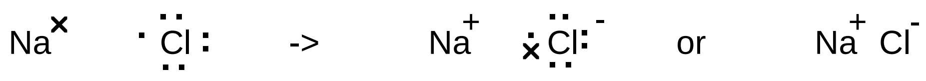


Some molecules with polar bonds ionise in aqueous solution. The attractive forces exerted by the water molecules are sufficient to break the polar covalent bonds in the solute e.g.



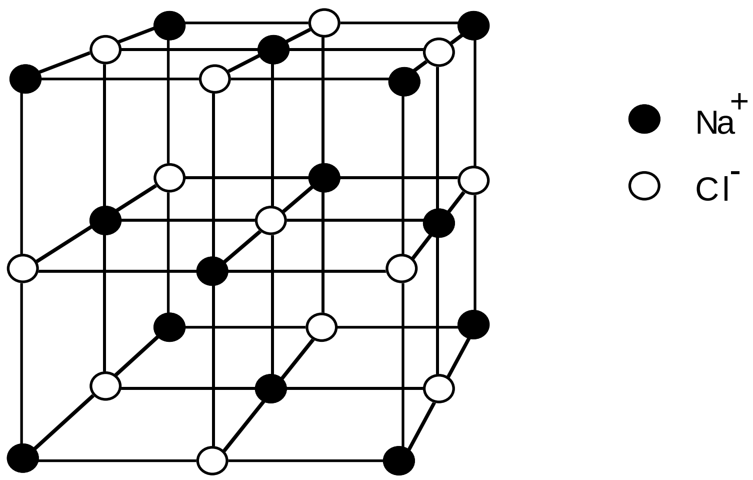
or HCl $\xrightarrow{H_2O}$ H^+ Cl^-

When the two atoms are far apart in the periodic table, one atom is <u>much more</u> electronegative than the other. The more electronegative atom (e.g. Chlorine) pulls the electron pair almost completely off the other atom (e.g. Sodium) forming ions e.g.



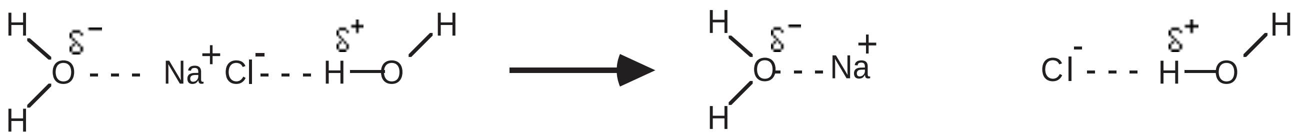
The attraction between positive and negative charges holds the ions together forming an **electrovalent bond**. Such bonds are very strong and are broken only at high temperatures e.g. Na^+Cl^- M.P. 801 ^{0}C ; B.P. 1439 ^{0}C.

The ionic structure is a lattice held together by the attraction of the oppositely charged ions :



Electrovalent substances dissolve only in highly polar solvents which are able to pull the ions apart

e.g. Sodium chloride dissolves in Water :



Electrovalent substances conduct electricity when molten, not in the solid state. When molten the ions can move and 'carry the current'; in the solid the ions cannot move.

Hydrogen's electronegativity is higher than it should be from its position in the table. This is because it is such a small atom that its electrons can get very close to the nucleus. Hydrogen's electronegativity is equivalent to that of Carbon.

Where the two atoms have equal electronegativities (e.g. Carbon and Hydrogen) the electron pair is equally shared and the bond is **non-polar** e.g. C-C ; C-H etc.

Non-polar substances dissolve in non-polar solvents e.g. Methane dissolves in Hexane.

Bonds Between Molecules in Covalent Liquids and Solids

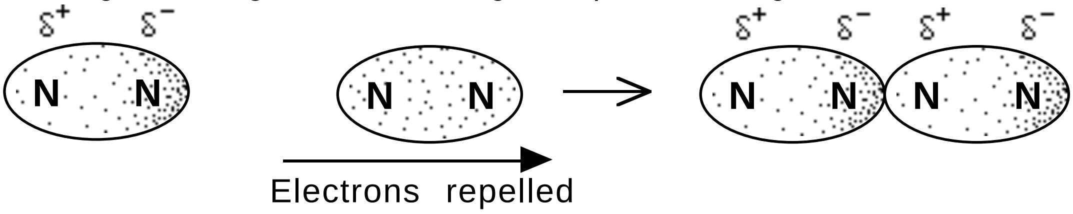
When covalent substances melt or boil the bonds **between** the molecules (intermolecular bonds) are broken. These bonds are weak so covalent substances have low M.P.'s and B.P.'s

e.g. Propane MP - 188 ^{0}C BP - 42.2 ^{0}C.

van der Waals bonds

These occur between **all** molecules in liquids and solids.

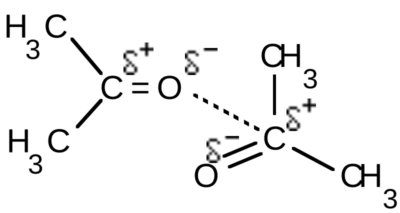
Since electrons move about over the surface of an atom or molecule, at any instant of time, the distribution will probably be distorted and a small polarity will exist which can induce a small polarity in a neighbouring molecule e.g. liquid Nitrogen :

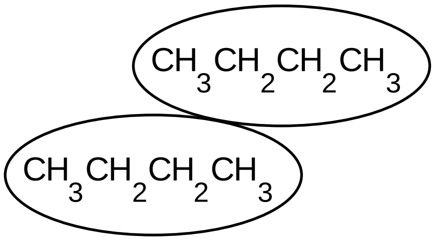


The larger the molecule the greater the surface area and the stronger the van der Waals bond. Hence M.P.'s and B.P.'s in covalent substances increase with molecular size e.g. going down the alkanes series.

Molecules with a permanent polarity e.g. Propanone have, of course, a stronger attraction

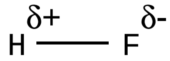
e.g. Compare the boiling points of Propanone (56.2 ^{0}C) and Butane (-0.5 ^{0}C) both with the same molecular mass (58) :





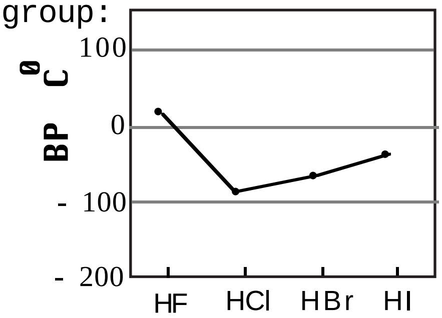
Hydrogen Bonds

When Hydrogen is bonded to one of the three most electronegative elements (N, O or F) the electron pair is greatly distorted towards the electronegative atom exposing the Hydrogen nucleus. This creates a highly polar bond.

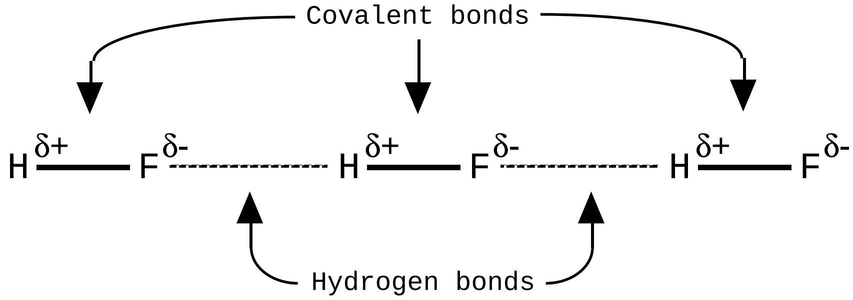


Stronger attraction between molecules called **HYDROGEN BONDING** is a major consequence e.g.

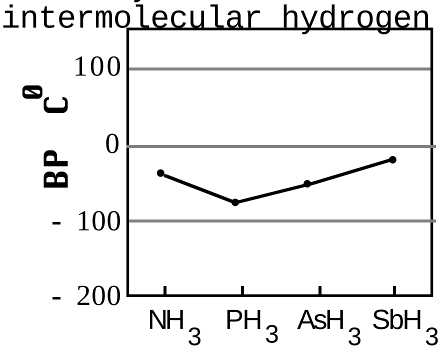
1. Hydrogen fluoride HF, though a small molecule, has an abnormally high boiling point (19 ⁰C) compared to the other hydrides in its group:

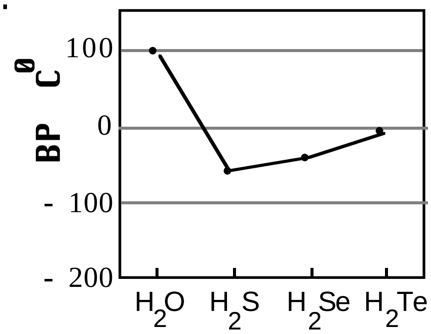


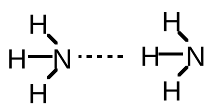
This is due to intermolecular hydrogen bonding increasing the attraction between neighbouring molecules in the liquid state:

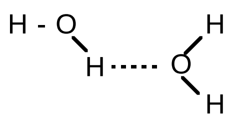


2. Similarly, Ammonia NH_3 and Water H_2O have abnormally high boiling points (- 33 and 100 ⁰C respectively) compared to the other hydrides in their groups. This is also due to intermolecular hydrogen bonding:



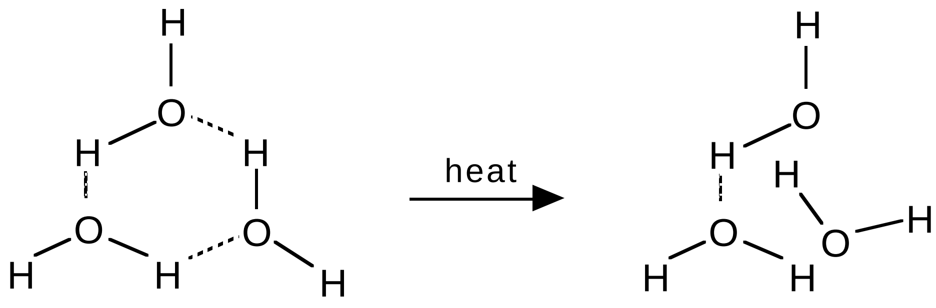






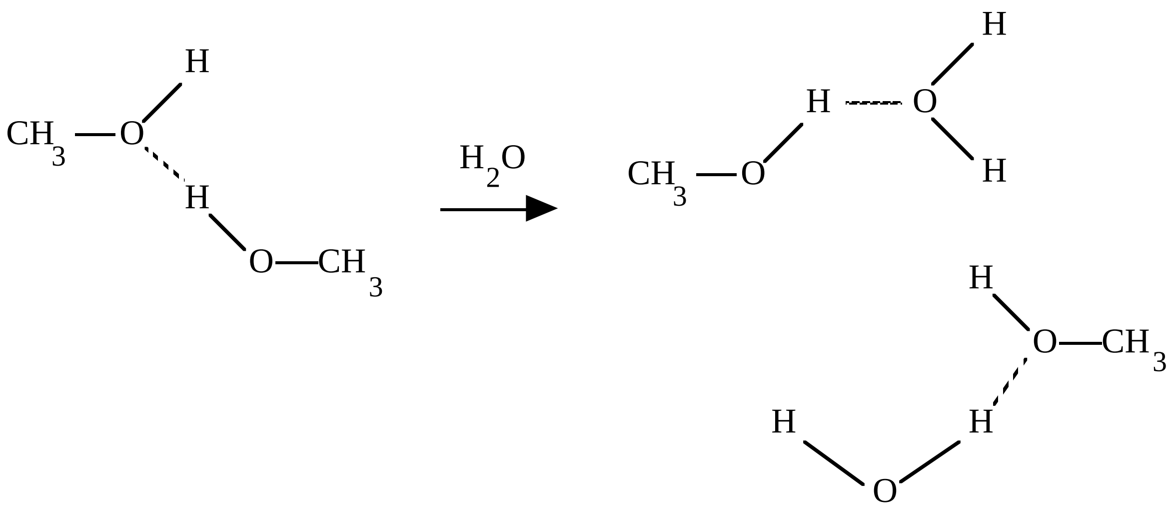
------- Hydrogen bond

3. The molecules of ice are held out in a lattice by hydrogen bonds. At 0 ^{0}C some of the hydrogen bonds start to break. The lattice collapses. The ice melts. The molecules are then more closely packed so the density increases.



Above 4 ^{0}C the increased movement of the molecules results in the density decreasing in the usual way.

4. Many substances dissolve in water because they can hydrogen bond to the water molecules e.g. Methanol:



5. Many substances are very viscous (sticky) because of hydrogen bonding between their molecules restricting the movement of the molecules e.g. Propan-1,2,3-triol: