Intermolecular forces

MOLECULAR POLARITY

Whether a molecule is polar, or not, depends both on the relative electronegativities of the atoms in the molecule and on its shape. If the individual bonds are polar then it does not necessarily follow that the molecule will be polar as the resultant dipole may cancel out all the individual dipoles.

non-polar (resultant dipole zero) — polar — polar — non-polar (resultant dipole zero)

Van der Waals' forces

Even in non-polar molecules the electrons can at any one moment be unevenly spread. This produces temporary instantaneous dipoles. An instantaneous dipole can induce another dipole in a neighbouring particle resulting in a weak attraction between the two particles. Van der Waals' forces increase with increasing mass.

increasing van der Waals' forces →

	F_2	Cl_2	Br_2	I_2
M_r	38.0	70.9	160	254
b. pt / °C	–188	–34.0	58.0	183

increasing van der Waals' forces →

	CH_4	C_2H_6	C_3H_8	C_4H_{10}
M_r	16.0	30.0	44.0	58.0
b. pt / °C	–162	–88.6	–42.2	–0.5

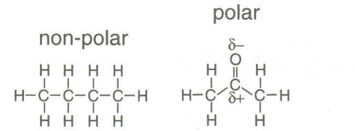
INTERMOLECULAR FORCES

The covalent bonds between the atoms *within* a molecule are very strong. The forces of attraction *between* the molecules are much weaker. These intermolecular forces depend on the polarity of the molecules.

Dipole:dipole forces

Polar molecules are attracted to each other by electrostatic forces. Although still relatively weak the attraction is stronger than van der Waals' forces.

non-polar — polar



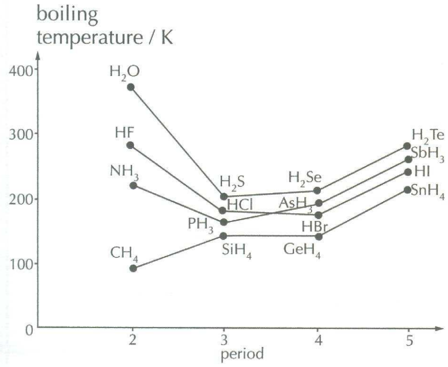
butane M_r = 58
b. pt –0.5 °C

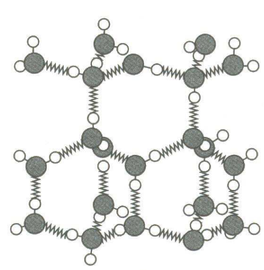
propanone M_r = 58
b. pt 56.2 °C

identical masses
(different intermolecular forces)

Hydrogen bonding

Hydrogen bonding occurs when hydrogen is bonded directly to a small highly electronegative element, such as fluorine, oxygen, or nitrogen. As the electron pair is drawn away from the hydrogen atom by the electronegative element all that remains is the proton in the nucleus as there are no inner electrons. The proton attracts a non-bonding pair of electrons from the F, N, or O resulting in a much stronger dipole:dipole attraction. Water has a much higher boiling point than the other group 6 hydrides as the hydrogen bonding between water molecules is much stronger than the dipole:dipole bonding in the remaining hydrides. A similar trend is seen in the hydrides of group 5 and group 7. Hydrogen bonds between the molecules in ice result in a very open structure. When ice melts the molecules can move closer to each other so that water has its maximum density at 4 °C.





the ice lattice

= hydrogen bond