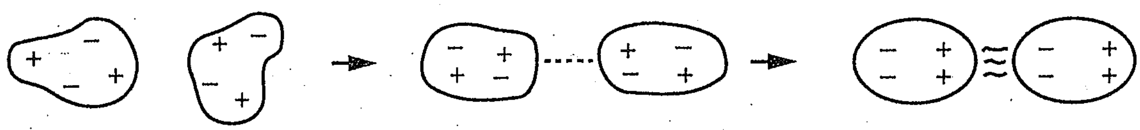
Student Handout 2 of 3: Intermolecular Forces

Here's a look at Intermolecular Forces <-- The forces neighboring molecules exert upon one another.

- There are a couple of generalities you can depend on
- The higher the melting point and boiling point the stronger the intermolecular forces
- Like dissolves like—things held together by electrostatic attraction (ions, polar molecules) will be soluble in polar solvents

A. Dispersion Forces (sometimes called London Forces)

All substances have these weak interactions. They happen because you can induce dipoles in molecules. When two molecules get close to one another they can influence each other in a local way. The molecules may be neutral or nonpolar but locally if there is some + charge hangin' around on a spot it will attract some - charge on the other molecule. The dipoles that are induced are temporary.



How strong dispersion forces are depends upon how easily polarized the electrons are in each molecule. Usually the bigger the atoms in the nucleus, the farther the electrons are from the nucleus and the easier they are to influence.

large molecule --> large molecular mass --> very polarizable --> stronger dispersion forces.

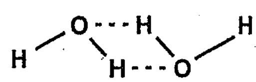
B. Dipole Forces

These are straight-up electrostatic interactions (by the way, electrostatic interactions just means the attraction between opposite charges) between polar molecules. These are weaker than ionic forces because dipoles are only partial charges whereas ions have full charges.

C. Hydrogen Bonding

This is a special case of dipole forces in which the molecule is very polar (large electronegativity difference between bonded atoms in the molecule) and the molecules are constructed so they can get close to each other.

The classic example of hydrogen bonding is H_2O. We know water is a very polar molecule and it also fits together with itself very well.



Because of water's polarity, all the H's and O's have strong electrostatic attractions. On top of that, the molecule fits together with itself very well.

Only a few molecules can hydrogen bond, you can tell from their properties because they will have abnormally high boiling points and melting points. --> You sort of expect these things (B.P and M.P.) to increase with molecular weight, so as you go down a group (V, VI and VII) and look at the B.P. and M.P. of the hydrides you expect them to increase. But right at the beginning of the groups, NH_3, H_2O, HF the properties are way-high because of hydrogen bonding. CH_4 doesn't hydrogen bond so the properties of the Group IVA hydrides simply increase with molecular size.

Also use these rules to determine if hydrogen bonding will occur:

H has to be bonded to F, O, N (big electronegativity differences induce molecular polarity) and there must also be a lone pair on the cental atom (F, O_2, or N)in the molecule under consideration (the positive and negative poles on the molecule have to fit together).