INTERMOLECULAR FORCES

Intermolecular forces (in order of decreasing strength) are: ion-ion, metallic, dipole-dipole and London dispersion (or induced dipole) forces. (Strictly speaking, covalent bonding, present in covalent network solids, is not an inter-molecular force since the solid in this case is a single giant molecule). 'Hydrogen bonding' is a special case of a dipole-dipole force, where an extra large dipole exists between the hydrogen covalently bonded to a small electronegative atom, such as **N**, **O** or **F**. Hydrogen bonding is an inter-molecular force between the hydrogen of one molecule and the lone pair of electrons on the nitrogen, oxygen or fluorine of a neighboring molecule. For molecules with a net dipole moment (or large individual bond dipole), the dominant interaction will be dipole-dipole interactions (such liquids are said to be polar). If the molecules have only weak dipoles (e.g., **C-H** bonds) then London dispersion (induced dipole) forces become important. If the molecules have no dipole moment, (e.g., $\mathbf{H_2}$, noble gases etc.) then the only interaction between them will be the weak London dispersion (induced dipole) force. Large atoms (or non-polar molecules) have larger London dispersion forces as there larger electron clouds are farther away from the nuclei and are therefore more polarizable. For liquids, stronger intermolecular forces result in higher viscosity, surface tension, boiling point and melting point and lower vapour pressure (see table below).

General Affects of Intermolecular Forces on Physical Properties of Liquids

PROPERTY	VOLATILE LIQUIDS	NON-VOLATILE LIQUIDS
	(weak intermolecular forces)	(strong intermolecular forces)
Viscosity	Low	High
Surface tension	Low	High
Boiling point (b.p.)	Low	High
Melting point (m.p.)	Low	High
Vapour pressure (P^o)	High	Low
Rate of evaporation	High	Low
Heat of vapourization (ΔH^o_{vap})	Low	High
Specific heat	Low	High

Sample Problem: Place the following compounds in order of increasing boiling point.

$\mathbf{H_2S}$, $\mathbf{H_2O}$, $\mathbf{CH_4}$, $\mathbf{H_2}$, **KBr**

Solution: **KBr** is an ionic solid (metal and non-metal) so it has the highest boiling point of the listed compounds. Water exhibits hydrogen bonding between adjacent molecules and boils at the next highest temperature. $\mathbf{H_2S}$ is polar, but not H-bonded and has the third highest boiling point. The molecules of methane and hydrogen are both non-polar, and will be held together by London dispersion forces. Since methane is larger than hydrogen, it's electron cloud is more diffuse and more readily distorted (i.e., more polarizable). As the strength of London forces increases with the degree of polarizability, $\mathbf{CH_4}$ will have a higher boiling point than $\mathbf{H_2}$.

So the boiling point order is; $\mathbf{H_2 < CH_4 < H_2S < H_2O < KBr}$