**Key Con**c**epts**

Three types of force can operate between covalent molecules:

* Van der Walls Forces  
  also known as London Dispersion Forces (named after Fritz London who first described these forces theoretically 1930)
* Dipole-dipole interactions
* Hydrogen bonds

Relative strength of Intermolecular Forces:

* Intermolecular forces (Van der Walls forces, dipole-dipole interactions and hydrogen bonds) are much weaker than intramolecular forces (covalent bonds, ionic bonds or metallic bonds)
* Van der Walls forces are the weakest intermolecular force (one hundredth-one thousandth the strength of a covalent bond), hydrogen bonds are the strongest intermolecular force (about one-tenth the strength of a covalent bond).
* Van der Walls forces < dipole-dipole interactions < hydrogen bonds

Van der Walls Forces

* are very weak forces of attraction between molecules resulting from:

momentary dipoles occurring due to uneven electron distributions in neighbouring molecules as they approach one another

* The more electrons that are present in the molecule, the stronger the Van der Walls forces will be.
* Van der Walls forces are the only type of intermolecular force operating between [non-polar](http://www.ausetute.com.au/molpolar.html) molecules, for example, Van der Walls forces operate between hydrogen (H2) molecules, chlorine (Cl2) molecules, carbon dioxide (CO2) molecules, dinitrogen tetroxide (N2O4) molecules and methane (CH4) molecules.

Dipole-dipole Interactions

* are stronger intermolecular forces than Van der Walls forces
* occur between molecules that have permanent net dipoles ([polar](http://www.ausetute.com.au/molpolar.html) molecules), for example, dipole-dipole interactions occur between SCl2 molecules, PCl3 molecules and CH3Cl molecules.   
  The partial positive charge on one molecule is electrostatically attracted to the partial negative charge on a neighbouring molecule.
* *If the permanent net dipole within the polar molecules results from a covalent bond between a hydrogen atom and either fluorine, oxygen or nitrogen, the resulting intermolecular force is referred to as a hydrogen bond (see below).*

Hydrogen bonds

* occur between molecules that have a permanent net [dipole](http://www.ausetute.com.au/molpolar.html) resulting from hydrogen being covalently bonded to either fluorine, oxygen or nitrogen. For example, hydrogen bonds operate between water (H2O) molecules, ammonia (NH3) molecules, hydrogen fluoride (HF) molecules, hydrogen peroxide (H2O2) molecules, alkanols (alcohols) such as methanol (CH3OH) molecules, and between alkanoic (caboxylic) acids such as ethanoic (acetic) acid (CH3COOH) and between organic amines such as methanamine (methyl amine, CH3NH2).
* are a stronger intermolecular force than either Van der Walls forces or dipole-dipole interactions since the hydrogen nucleus is extremely small and positively charged and fluorine, oxygen and nitrogen being very [electronegative](http://www.ausetute.com.au/bondpola.html) so that the electron on the hydrogen atom is strongly attracted to the fluorine, oxygen or nitrogen atom, leaving a highly localised positive charge on the hydrogen atom and highly negative localised charge on the fluorine, oxygen or nitrogen atom. This means the electrostatic attraction between these molecules will be greater than for the polar molecules that do not have hydrogen covalently bonded to either fluorine, oxygen or nitrogen.

**Effect of Intermolecular forces on melting and boiling points of molecular covalent substances:**

Since melting or boiling result from a progressive weakening of the attractive forces between the covalent molecules, the stronger the intermolecular force is, the more energy is required to melt the solid or boil the liquid.

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| If only Van der Walls forces are present, then the more electrons the molecule has (and consequently the more [mass](http://www.ausetute.com.au/mmcalcul.html) it has) the stronger the Van der Walls forces will be, so the higher the melting and boiling points will be.  Consider the hydrides of Group IV, all of which are non-polar molecules, so only Van der Walls forces act between the molecules.  CH4 (molecular mass ~ 16), SiH4 (molecular mass ~ 32), GeH4 (molecular mass ~ 77) and SnH4(molecular mass ~ 123) can all be considered non-polar covalent molecules.  As the mass of the molecules increases, so does the strength of the Van der Walls force acting between the molecules, so more energy is required to weaken the attraction between the molecules resulting in higher boiling points. | http://www.ausetute.com.au/images/grabpg4h.gif *Boiling Points of Group IV Hydrides* |
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| If a covalent molecule has a permanent net dipole then the force of attraction between these molecules will be stronger than if only Van der Walls forces were present between the molecules. As a consequence, this substance will have a higher melting or boiling point than similar molecules that are non-polar in nature.  Consider the boiling points of the hydrides of Group VII elements.  All of the molecules HF (molecular mass ~ 20), HCl (molecular mass ~ 37), HBr (molecular mass ~ 81) and HI (molecular mass ~ 128) are polar, the hydrogen atom having a partial positive charge (Hδ+) and the halogen atom having a partial negative charge (Fδ-, Clδ-, Brδ-, Iδ-).  As a consequence, the stronger dipole-interactions acting between the hydride molecules of Group VII elements results in higher boiling points than for the hydrides of Group IV elements as seen above.   With the exception of HF, as the [molecular mass](http://www.ausetute.com.au/mmcalcul.html) increases, the boiling point of the hydrides increase.  HF is an exception because of the stronger force of attraction between HF molecules resulting from hydrogen bonds acting between the HF molecules. Weaker dipole-dipole interactions act between the molecules of HCl, HBr and HI. So HF has a higher boiling point than the other molecules in this series. | http://www.ausetute.com.au/images/grabpg7h.gif*Boiling Points of Group VII hydrides* |

**Effect of Intermolecular Forces on Solubility**

In general like dissolves like:

* non-polar solutes dissolve in non-polar solvents   
  Paraffin wax (C30H62) is a non-polar solute that will dissolve in non-polar solvents like oil, hexane (C6H14) or carbon tetrachloride (CCl4).   
  Paraffin wax will NOT dissolve in polar solvents such as water (H2O) or ethanol (ethyl alcohol, C2H5OH).
* polar solutes such as glucose (C6H12O6) will dissolve in polar solvents such as water (H2O) or ethanol (ethyl alcohol, C2H5OH) as the partially positively charged atom of the solute molecule is attracted to the partially negatively charged atom of the solvent molecule, and the partially negatively charged atom of the solute molecule is attracted to the partially positively charged atom of the solvent molecule.

Glucose will NOT dissolve in non-polar solvents such as oil, hexane (C6H14) or carbon tetrachloride (CCl4).

* Ionic solutes such as sodium chloride (NaCl) will generally dissolve in polar solvents but not in non-polar solvents, since the positive ion is attracted the partially negatively charged atom in the polar solvent molecule, and the negative ion of the solute is attracted to the partially positively charged atom on the solvent molecule.

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| **Introduction:** The dissolving process involves a consideration of the relative strength of three intermolecular attractive forces. The type of forces between **solute-solute molecules** and **solvent-solvent molecules** must be considered. These intermolecular attractions must be broken before new **solute-solvent** attractive forces can become effective. Perhaps the bond breaking and bond forming processes take place simultaneously. A solute will dissolve in a solvent if the solute-solvent forces of attraction are great enough to overcome the solute-solute and solvent-solvent forces of attraction. A solute will not dissolve if the solute-solvent forces of attraction are weaker than individual solute and solvent intermolecular attractions. Generally, if all three of the intermolecular forces of attraction are roughly equal, the substances will be soluble in each other. |
| **Solubility Rule and Summary**  **LIKES DISSOLVE LIKES.**  This means that ionic or polar solutes dissolve in polar solvents. Non-polar solutes dissolve in non-polar solvents.  Polar and ionic solutes DO NOT dissolve in non-polar solvents and vice versa.  Remember that when applying the solubility rule: "Likes Dissolve Likes", that there are no absolutes and there are exceptions with a small amount of solubility possible. The rule is most useful when making comparisons between a series of compounds. |
| **IONIC SOLUTE - POLAR SOLVENT:**When an ionic crystal such as NaCl is placed in water, a dissolving reaction will occur. Initially, the positive and negative ion are only attracted to each other. The water molecules are hydrogen bonded to each other. If the crystal is to dissolve, these bonds must be broken.  Negative chloride ions on the surface are attracted by neighboring positive sodium ions and by the partially positive hydrogen atom in the polar water molecule (See the graphic on the right).  Similarly, the positive sodium ions are attracted by both chloride ions and the partially negative oxygen atom in the polar water molecule.  A "tug-of-war" occurs for the positive and negative ions between the other ions in the crystal and the water molecules. Whether the crystal dissolves is determined by which attractive force is stronger. If the internal ionic forces in the crystal are the strongest, the crystal does not dissolve. This is the situation in reactions where precipitates form. If the attractions for the ions by the polar water molecules are the strongest, the crystal will dissolve. This is the situation in sodium chloride. |
| **POLAR SOLUTE - POLAR SOLVENT:**  Polar ammonia molecules dissolve in polar water molecules. These molecules mix readily because both types of molecules engage in hydrogen bonding. Since the intermolecular attractions are roughly equal, the molecules can break away from each other and form new solute (NH3), solvent (H2O) hydrogen bonds.  A wide variety of solutions are in this category such as sugar in water, alcohol in water, acetic and hydrochloric acids. |
| **NON-POLAR SOLUTE - NON-POLAR SOLVENT:**  In all types of non-polar compounds, the only intermolecular attractions are the very weak Van der Walls. The weak attractive forces formed by the solute-solvent molecules compensate for breaking those weak bonds in the two pure non-polar substances. An example is solid iodine (I2) dissolved in liquid bromine (Br2).  In addition to diatomic molecules with identical atoms, the **most common type of non-polar compounds are the hydrocarbons**. Many C-C and C-H non-polar bonds are present. Hydrocarbons are present in oils, grease, fats, dry cleaning solvents, turpentine, gasoline, etc. |
| **NON-POLAR SOLUTE - POLAR SOLVENT:**  Non-polar Iodine is not very soluble in water. An intermolecular bond between an induced dipole (I2) and a polar bond in water is not very strong compared to the hydrogen bonds in water. The water molecules would rather remain hydrogen bonded to each other, then to allow an iodine molecule come between them. The water molecules effectively "squeeze" out the non-polar iodine. The intermolecular forces are not roughly equal, therefore, the "unlike" substances are not soluble in each other.  Various gases such as O2, N2, H2, CO2 are not very soluble because the gases are essentially non-polar. Of course you may say that oxygen must be dissolved in water to sustain fish life -- true, but the solubility is very low. Carbon dioxide is soluble in water such as carbonated beverages -- again this is true but why does it fizz when opened or lose the bubbles on standing? Carbon dioxide is not very soluble in water |