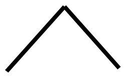
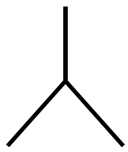
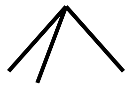
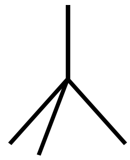
Small Molecules

Small molecules are formed either of non-metallic elements or of compounds of non-metallic elements.

The shapes of small molecules can be predicted by VSEPR (valence shell electron pair repulsion) theory.

				
linear	V-shaped	trigonal planar	trigonal pyramidal	tetrahedral

Method for drawing molecules which have one central atom:

1. Draw an electron-dot diagram first
   A. Write the central atom (the element there is only one atom of in the molecule)
   B. Write the other atoms around it.
   C. The central atom will share enough electrons to make all the other atoms stable. Draw these electrons around the central atom.
   In a negative ion, the central atom may share more than it originally had in its valence shell.
   D. Each outer atom will share the same amount as is being shared with it. Draw these electrons between the central atom and each other atom.
   E. Draw any remaining electrons (non-bonding electrons) so all the atoms are stable.
   For most atoms this will mean 8 electrons (except hydrogen which needs 2). Sometimes the central atom will have a total of more than 8, this is called 'expanding the octet' and is not a problem.

2. Draw the stick-bond diagram
   A. Arrange the shape so that regions of negative charge are far away from each other.
   Electron pairs within double and triple bonds do not repel one another and can be treated as a single 'cloud'.
   B. If the molecule is a polyatomic ion, draw positive charge on the central atom or negative charge evenly spread around the outer atoms.
   C. The number of sticks between any two atoms is half the electrons drawn between them (since they are electron pairs), *unless* the outer atom has a negative charge, in which case only a single bond is drawn.
   D. Draw any central-atom non-bonding electrons.

Note: Draw electrons in pairs, where possible.

Polar and Non-polar Bonding

The polarity of a covalent bond is related to the difference in electronegativity between the two atoms involved in the bond.

Non-polar covalent: There is little or no difference between the electronegativity of bonding atoms

Polar covalent: Significant difference between the electronegativities of bonding atoms

Note: if the difference is significant enough that the valence electron(s) move from one atom to the other and are no longer shared, the bond is *ionic*.

In a polar covalent bond the electron density is greater around the more electronegative atom so that end of the bond will have a partial negative charge (δ^-) and the other will have a partial positive charge (δ^+). This is known as a *dipole* (arrangement with two poles).