Covalent Bonding part 2

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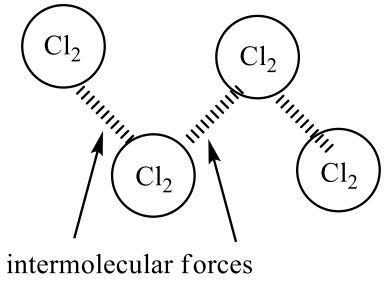
Covalent molecules can be split into two types: **simple** molecular (or simple covalent) structures and **giant** molecular structures.

Simple molecular

Almost *all* covalent molecules that you know are **simple covalent** molecules e.g. O_2, N_2, H_2, Cl_2, CH_4, NH_3, CO_2 etc.

These molecules have **low** boiling and melting points.

✓ remember that covalent bonds are **not** broken when boiling one of these molecules. It is only **weak** intermolecular forces that are broken.



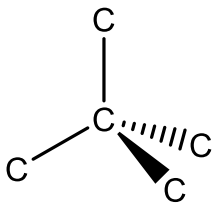
For example, as in Cl_2 above, we have lots of Cl_2 molecules that are held together by these weak intermolecular forces. These forces are broken upon boiling. The Cl-Cl covalent bond is **not broken**. To break covalent bonds requires a huge amount of energy.

Giant molecular structures

Diamond, graphite and silicon dioxide are the classic examples that they tend to talk about. These structures are giant lattices and have **high melting points** due to covalent bonds being broken.

Diamond

Diamond is made up from carbon atoms that form a **tetrahedral** structure (each carbon has 4 other carbons bonded to it), which can be described as a crystal lattice.

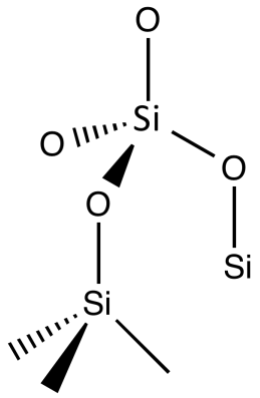


each carbon bonds to 4 carbons
contains a massive number of
these units

There isn't too much more to say about diamond other than it is very hard, has a very high melting point (~3800 K), it doesn't conduct electricity and it is completely insoluble.

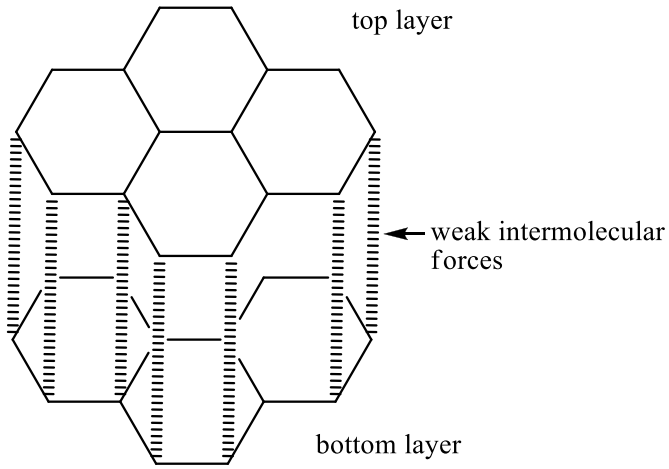
✓ SiO_2 (silicon dioxide) also has a tetrahedral structure but it is not a carbon structure. Each silicon is surrounded by 4 oxygen atoms, but on average, there will be two oxygens *per* Si.

A section of this structure is shown below. It will just keep extending in all directions in a giant lattice type structure:



Graphite

Graphite is a bit different. Like diamond, it consists of only carbon but it is arranged in a **layer** structure with **delocalised** electrons, both of which give it different properties.



Graphite only makes **3** bonds per carbon atom, which means there is one d**elocalised** (free) electron per carbon atom.

These delocalised electrons are free to move throughout the layer and can carry electric current. Therefore, graphite **conducts electricity**.

The weak intermolecular forces enables the layers to slide over the top of each other, which gives graphite uses such as in pencils and lubricants i.e. it is **fluid**.

The layers are also quite far apart which makes graphite a lot **less dense** than diamond.

Similarly to diamond, graphite has a very high melting point (~3900 K) and is insoluble.

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