METALLIC BONDING

What is a metallic bond?

Metallic bonding in sodium

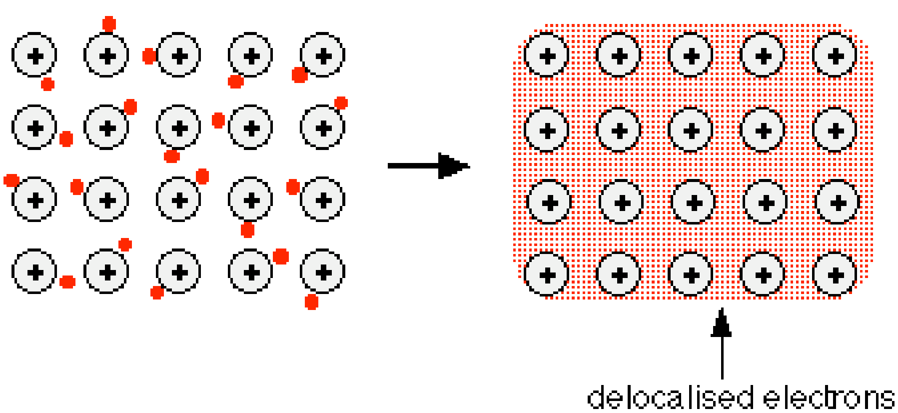
Metals tend to have high melting points and boiling points suggesting strong bonds between the atoms. Even a metal like sodium (melting point 97.8°C) melts at a considerably higher temperature than the element (neon) which precedes it in the Periodic Table.

Sodium has the electronic structure $1s^2 2s^2 2p^6 3s^1$. When sodium atoms come together, the electron in the 3s atomic orbital of one sodium atom shares space with the corresponding electron on a neighbouring atom to form a molecular orbital - in much the same sort of way that a covalent bond is formed.

The difference, however, is that each sodium atom is being touched by eight other sodium atoms - and the sharing occurs between the central atom and the 3s orbitals on all of the eight other atoms. And each of these eight is in turn being touched by eight sodium atoms, which in turn are touched by eight atoms - and so on and so on, until you have taken in all the atoms in that lump of sodium.

All of the 3s orbitals on all of the atoms overlap to give a vast number of molecular orbitals which extend over the whole piece of metal. There have to be huge numbers of molecular orbitals, of course, because any orbital can only hold two electrons.

The electrons can move freely within these molecular orbitals, and so each electron becomes detached from its parent atom. The electrons are said to be ***delocalised.*** The metal is held together by the strong forces of attraction between the positive nuclei and the delocalised electrons.



This is sometimes described as "an array of positive ions in a sea of electrons".

If you are going to use this view, beware! Is a metal made up of atoms or ions? It is made of ***atoms.***

Each positive centre in the diagram represents all the rest of the atom apart from the outer electron, but that electron hasn't been lost - it may no longer have an attachment to a particular atom, but it's still there in the structure. Sodium metal is therefore written as Na - *not* Na^+.

Metallic bonding in magnesium

If you work through the same argument with magnesium, you end up with stronger bonds and so a higher melting point.

Magnesium has the outer electronic structure $3s^2$. Both of these electrons become delocalised, so the "sea" has twice the electron density as it does in sodium. The remaining "ions" also have twice the charge (if you are going to use this particular view of the metal bond) and so there will be more attraction between "ions" and "sea".

More realistically, each magnesium atom has one more proton in the nucleus than a sodium atom has, and so not only will there be a greater number of delocalised electrons, but there will also be a greater attraction for them.

Magnesium atoms have a slightly smaller radius than sodium atoms, and so the delocalised electrons are closer to the nuclei. Each magnesium atom also has twelve near neighbours rather than sodium's eight. Both of these factors increase the strength of the bond still further.

Metallic bonding in transition elements

Transition metals tend to have particularly high melting points and boiling points. The reason is that they can involve the 3d electrons in the delocalisation as well as the 4s. The more electrons you can involve, the stronger the attractions tend to be.

The metallic bond in molten metals

In a molten metal, the metallic bond is still present, although the ordered structure has been broken down. The metallic bond isn't fully broken until the metal boils. That means that boiling point is actually a better guide to the strength of the metallic bond than melting point is. On melting, the bond is loosened, not broken.

METALLIC STRUCTURES

The structure of metals

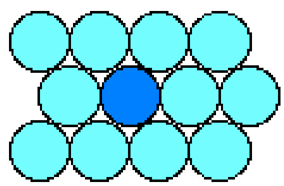
The arrangement of the atoms

Metals are giant structures of atoms held together by metallic bonds. "Giant" implies that large but variable numbers of atoms are involved - depending on the size of the bit of metal.

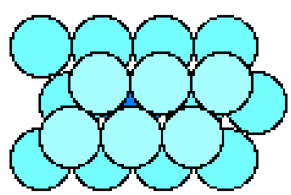
12-co-ordination

Most metals are *close packed* - that is, they fit as many atoms as possible into the available volume. Each atom in the structure has 12 touching neighbours. Such a metal is described as 12-co-ordinated.

Each atom has 6 other atoms touching it in each layer.



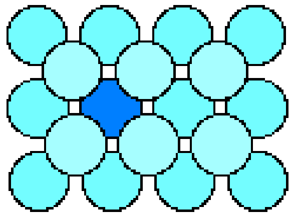
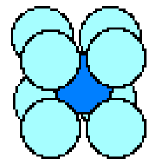
There are also 3 atoms touching any particular atom in the layer above and another 3 in the layer underneath.



This second diagram shows the layer immediately above the first layer. There will be a corresponding layer underneath. (There are actually two different ways of placing the third layer in a close packed structure, but that goes beyond the requirements of current A'level syllabuses.)

8-co-ordination

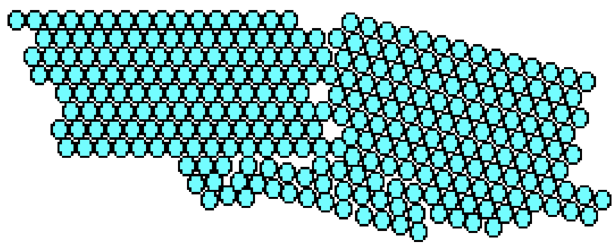
Some metals (notably those in Group 1 of the Periodic Table) are packed less efficiently, having only 8 touching neighbours. These are 8-co-ordinated.

The left hand diagram shows that no atoms are touching each other within a particular layer . They are only touched by the atoms in the layers above and below. The right hand diagram shows the 8 atoms (4 above and 4 below) touching the darker coloured one.

Crystal grains

It would be misleading to suppose that all the atoms in a piece of metal are arranged in a regular way. Any piece of metal is made up of a large number of "crystal grains", which are regions of regularity. At the grain boundaries atoms have become misaligned.



The physical properties of metals

Melting points and boiling points

Metals tend to have high melting and boiling points because of the strength of the metallic bond. The strength of the bond varies from metal to metal and depends on the number of electrons which each atom delocalises into the sea of electrons, and on the packing.

Group 1 metals like sodium and potassium have relatively low melting and boiling points mainly because each atom only has one electron to contribute to the bond - but there are other problems as well:

- Group 1 elements are also inefficiently packed (8-co-ordinated), so that they aren't forming as many bonds as most metals.

- They have relatively large atoms (meaning that the nuclei are some distance from the delocalised electrons) which also weakens the bond.

Electrical conductivity

Metals conduct electricity. The delocalised electrons are free to move throughout the structure in 3-dimensions. They can cross grain boundaries. Even though the pattern may be disrupted at the boundary, as long as atoms are touching each other, the metallic bond is still present.

Liquid metals also conduct electricity, showing that although the metal atoms may be free to move, the delocalisation remains in force until the metal boils.

Thermal conductivity

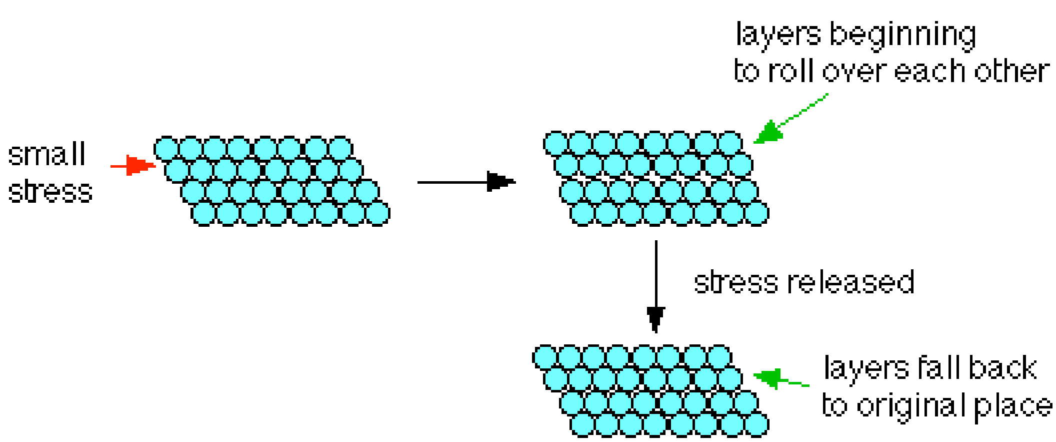
Metals are good conductors of heat. Heat energy is picked up by the electrons as additional kinetic energy (it makes them move faster). The energy is transferred throughout the rest of the metal by the moving electrons.

Strength and workability

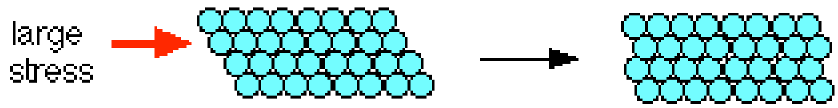
Malleability and ductility

Metals are described as ***malleable*** (can be beaten into sheets) and ***ductile*** (can be pulled out into wires). This is because of the ability of the atoms to roll over each other into new positions without breaking the metallic bond.

If a small stress is put onto the metal, the layers of atoms will start to roll over each other. If the stress is released again, they will fall back to their original positions. Under these circumstances, the metal is said to be ***elastic.***



If a larger stress is put on, the atoms roll over each other into a new position, and the metal is permanently changed.



The hardness of metals

This rolling of layers of atoms over each other is hindered by grain boundaries because the rows of atoms don't line up properly. It follows that the more grain boundaries there are (the smaller the individual crystal grains), the harder the metal becomes.

Offsetting this, because the grain boundaries are areas where the atoms aren't in such good contact with each other, metals tend to fracture at grain boundaries. Increasing the number of grain boundaries not only makes the metal harder, but also makes it more brittle.

Controlling the size of the crystal grains

If you have a pure piece of metal, you can control the size of the grains by ***heat treatment*** or by ***working the metal.***

Heating a metal tends to shake the atoms into a more regular arrangement - decreasing the number of grain boundaries, and so making the metal softer. Banging the metal around when it is cold tends to produce lots of small grains. Cold working therefore makes a metal harder. To restore its workability, you would need to reheat it.

You can also break up the regular arrangement of the atoms by inserting atoms of a slightly different size into the structure. ***Alloys*** such as brass (a mixture of copper and zinc) are harder than the original metals because the irregularity in the structure helps to stop rows of atoms from slipping over each other.

`http://www.chemguide.co.uk/atoms/bonding/metallic.html`

`http://www.chemguide.co.uk/atoms/structures/metals.html#top`