

1.1 Metallic Elements

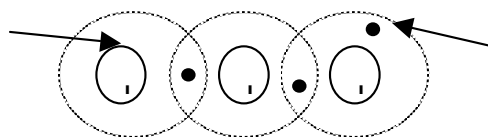
A the end of this unit you should be able to

1.

Metallic Lattice

In a solid the atoms are arranged neatly in a simple pattern known as a lattice.

Core of an atom is the nucleus and inner shells of electrons



The electrons are delocalised which means that they are free to move from one atom to the next

ost electrons are free to move between atoms. The atoms are held together by the attraction of the positively charged cores and the negative free moving electrons.

Physical properties

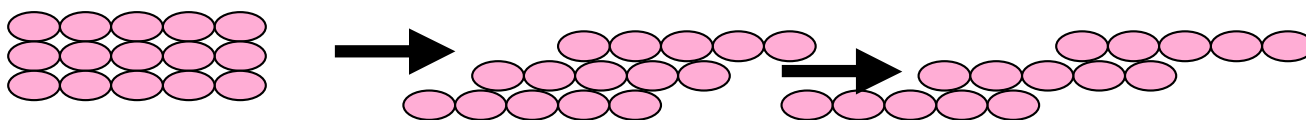
The physical properties of a substance are those which do not involve the substance in chemical reactions

e.g. colour, state, melting point, conductivity, malleability and ductility.

Malleable means that the substance can be rolled into sheets or hammered into shape.

Ductile means that the substance can be pulled into wires or simply stretched.

Metals are malleable and ductile. Good conductors of heat and electricity



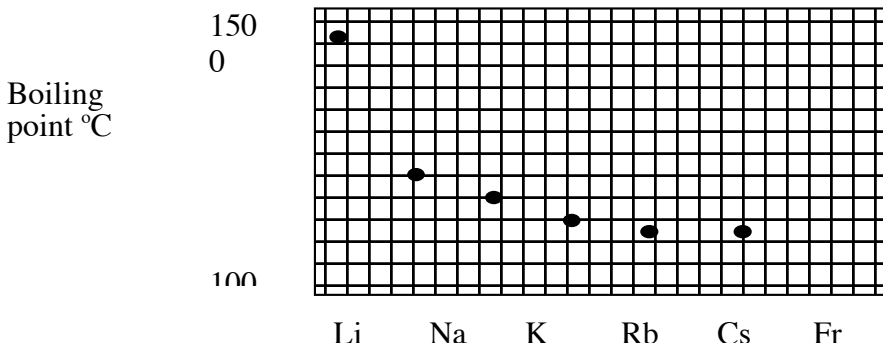
The layers of metal atoms can slip over each other without breaking the metallic bonds; it is this property which allows them to be both malleable and ductile.

The free moving electrons allow the metal to be a good conductor of electricity.

As the atoms are close packed vibrations can quickly pass from one atom to the next; it is this which makes metals good conductors of heat.

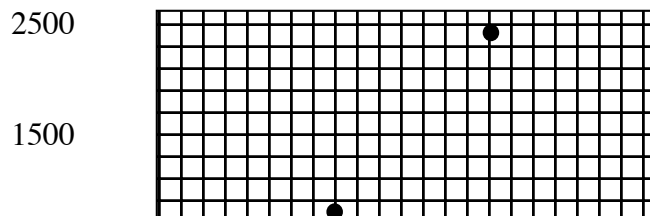
Boiling Points

When a substance is boiled all the forces of attraction between its atoms must be overcome to allow it to move freely as a gas. The boiling points give an indication of the strength of the metallic bond.



The Boiling and melting points of the alkali metals decrease as we come down the group. This tells us that the bonding between the atoms decrease in strength as we come down the group.

The decrease can be explained as the attractions between the positive nuclei and negative electrons is decreased by number of electron shells in the core of the atoms.

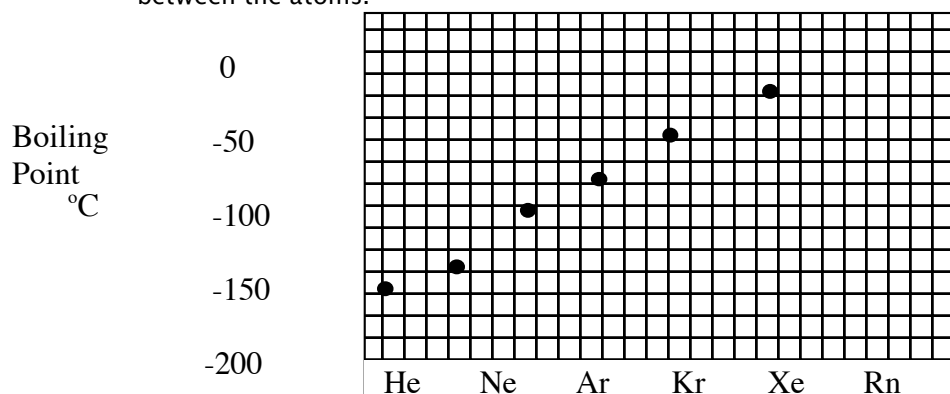


Boiling point °C

As the number of delocalised electrons (metals) increase so the boiling points increase.

1.2 Monatomic Elements

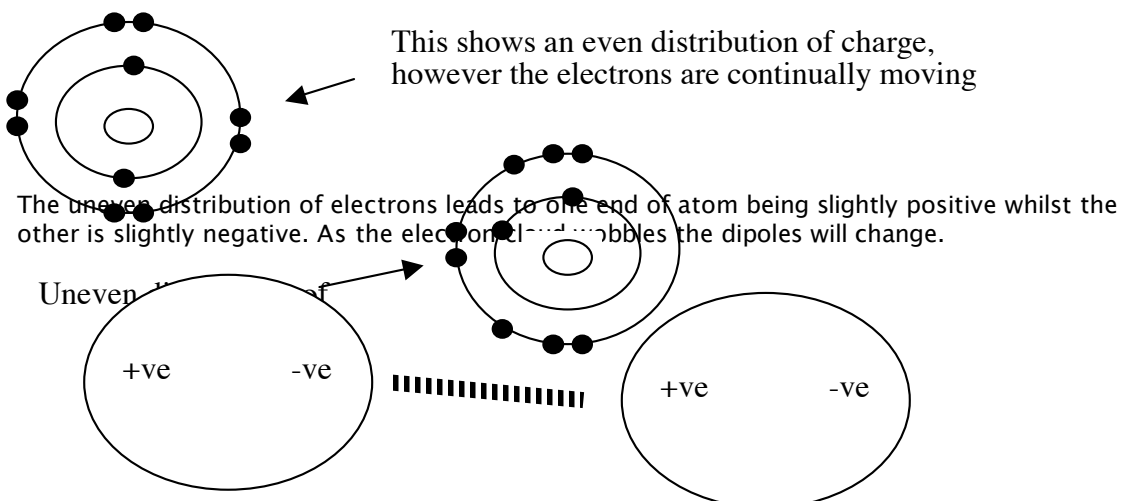
The Noble Gases are the only elements to exist as single atoms. There are very weak forces between the atoms.



As we descend Group 8 of the Periodic Table the boiling point of the Noble Gases increases. This tells us that the forces holding the atoms together must be getting stronger.

Van der Waals forces.

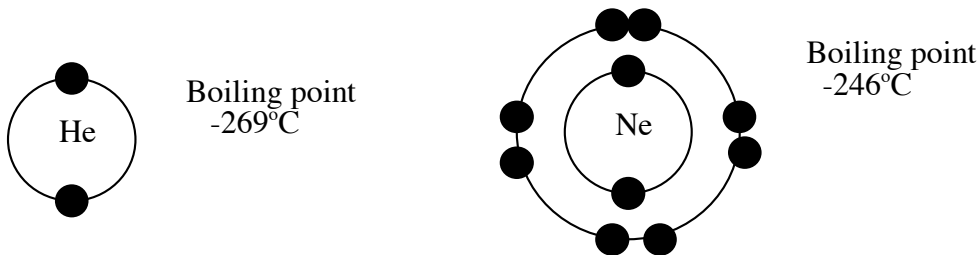
When we examine a neutral atom e.g. Neon



It is the attraction between the temporary dipole in one atom with the temporary dipole caused in an adjacent atom, which represents the Van der Waals force.

In addition when two atoms approach each other the electron clouds will repel each other causing dipoles in nearby atoms. These in turn will cause further dipoles.

The Van der Waals force is the electrostatic attraction between a temporary dipole and a temporary induced dipole.
 Larger molecules will have more wobbles in their outer electrons than smaller ones. The farther from the positive nucleus the electrons the less hold the nucleus will exert on it.
 This leads to more wobble and greater number of dipoles. This in turns increases the Van der Waals forces.



In general the greater the number of electrons in an atom or molecule the greater the Van der Waals forces.

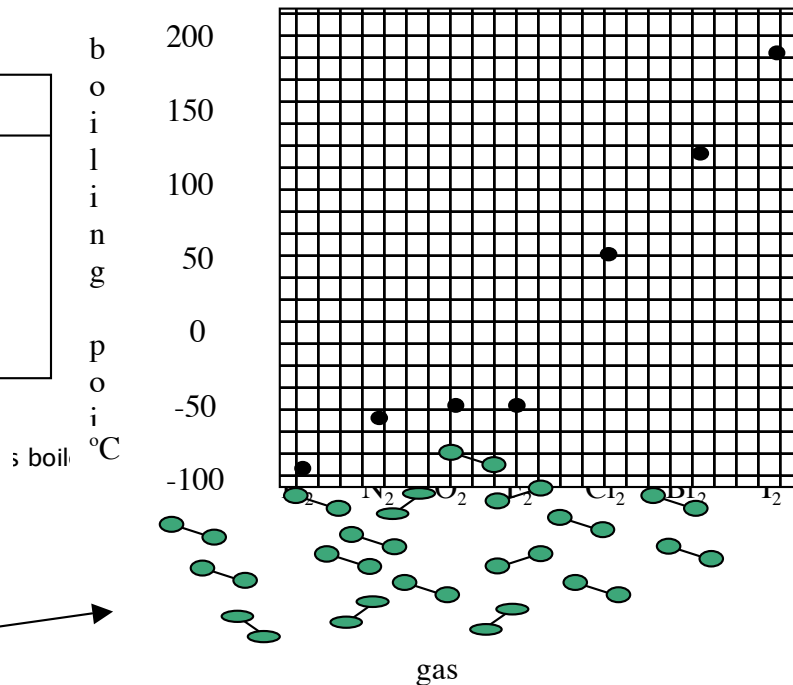
1.3 Molecular Elements

Elements, which are not metallic or monatomic, are covalent. The majority of the covalent elements are also molecular. Some are gases, some liquid and others solid at room temperature. In this section we learn how to explain why and even predict.

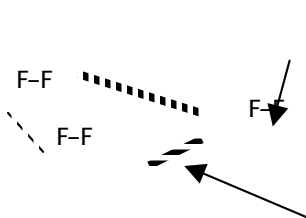
1. Diatomic elements

Diatomic element		Boiling point °C
Hydrogen	H ₂	-253
Nitrogen	N ₂	-196
Oxygen	O ₂	-183
Fluorine	F ₂	-188
Chlorine	Cl ₂	-35
Bromine	Br ₂	59
Iodine	I ₂	184

The boiling points and hence the Van der Waals forces increase as the number of electrons increase.

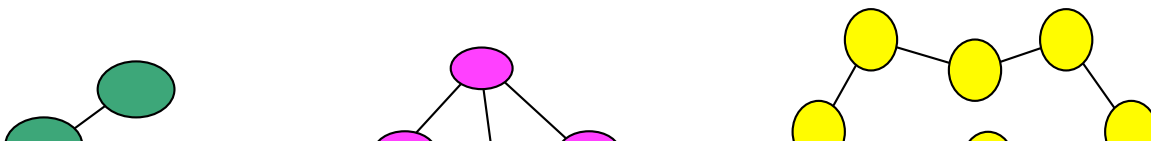


The weak Van der Waals forces are broken but the strong covalent bonds are not.



Strong covalent bonds are not broken when substance is boiled

2. Larger covalent molecule: Weak vander waals forces are broken when substance is boiled



Chlorine

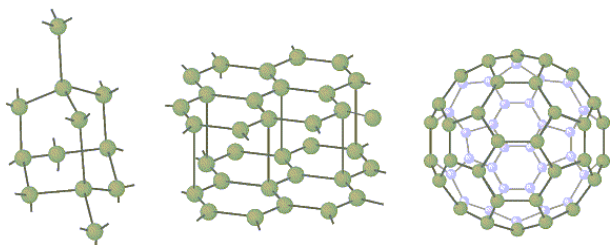
Phosphorus

Sulphur

Element	Formula	Boiling Point (°C)
chlorine		
phosphorus		
sulphur		

As expected the boiling points increase as the number of atoms, and the total number of electron, in the molecule increase.

3. Molecular Carbon



Diamond

Graphite

Buckminsterfullerene

© Microsoft Corporation. All Rights Reserved.

Buckminsterfullerene is a large molecule contains 60 carbon atoms (C_{60}) and its Melting and boiling points would be reasonably high. It is however hollow and not close packed like diamond and graphite and will have a low density. Note graphite and diamond are covalent network (see 1.4)

1.4 Covalent network elements

1. Diamond structure

In diamond each carbon is tetrahedrally surrounded by 4 identical carbon atoms. It can be described as a lattice as it has a regular repeating pattern. To bend it the strong covalent bond would need to be broken, this makes diamond one of the hardest known mineral (Mohr scale 10) It is so hard it is used to cut glass and as tips in drills

2. Graphite structure

The distance between the atoms in the layer are smaller (covalent bonds) than the distance between the layers (Van der Waals forces) In the layer there are only 3 electrons used in bonding, the other electron is free to move and hence graphite will conduct electricity. Graphite can be used as a solid lubricant as the

Property	Diamond	Graphite
Appearance	colourless transparent solid	black/grey shiny solid
Conduction	no	yes
Feel	smooth not	slippery
Hardness	scratches glass	rubs off on paper

3. Networks and melting points

Only three elements exist as covalent networks boron, carbon and graphite

They are all solids as to melt them the strong covalent bonds must be broken
Silicon has the same bonding structure as it like carbon has 4 electrons in its outer shell

It takes a lot of heat energy to break strong covalent bonds this is only available at very high temperatures.

Boron would be hard as to break it the strong covalent bonds must be broken

Atomic Size 1.5

- Atomic radii
e.g. solid iodine

The covalent radius of an element is $1/2$ the distance between two bonded atoms ($1/2 \times 266\text{pm}$)
The Van der Waals radius is $1/2$ the distance between two non bonded atoms ($1/2 \times 430$)

- Atomic size and the Periodic Table

- a) The atomic radius increases down a Group of the periodic table as the number of electron shells increase.
- b) The atomic size decreases across a period as the attraction for the shells of electrons increase because the number of protons in the nucleus increases.
Note The number of shells remains constant across a period.

Density of elements

Density is mass per unit volume. So the greater the mass of an atom the higher its density will be. The larger the size of the atom the smaller its density will be.

For the metallic elements the density increases from left to right across a period.

As we move across a period from left to right the electrons in each atom are more tightly held by the nucleus due to the increasing number of protons in the nucleus. This causes a decrease in the atomic size and so more atoms can be packed together in the same volume.

For the metallic elements the density decreases from top to bottom in a group. This is due to an increase in the number of occupied electron energy levels giving larger atoms, which cannot pack as closely together.

The alkali metals have lower densities than other metals as they have the greatest atomic volume.

Density increases across a group as the mass increases and the volume decreases.

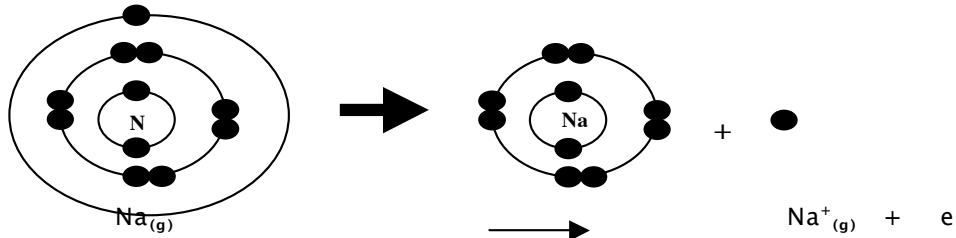
Density increases going down a group as the increase in mass has a greater affect than the increase in volume.

1.6 Ionisation Energies

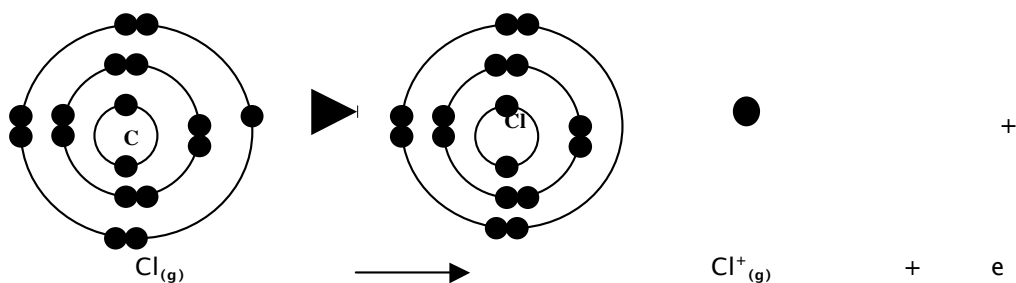
1. 1st Ionisation Energy

This is the energy required to remove one electron from each atom in a mole of gaseous (free) atoms.

The easiest electron to remove is the one with the highest energy i.e. one from the outer electron shell.

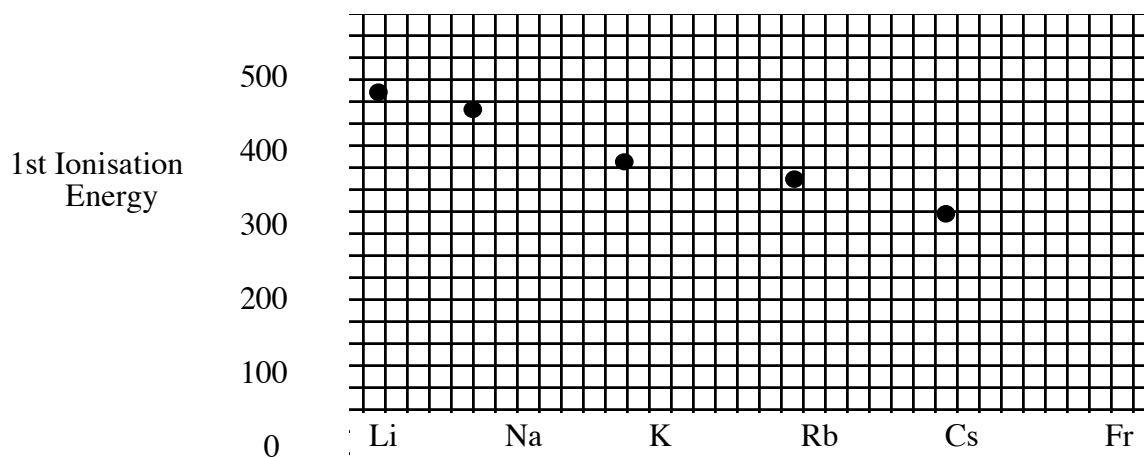


The energy needed to remove one mole of electrons from one mole of gaseous atoms is + 502 kJ mol⁻¹



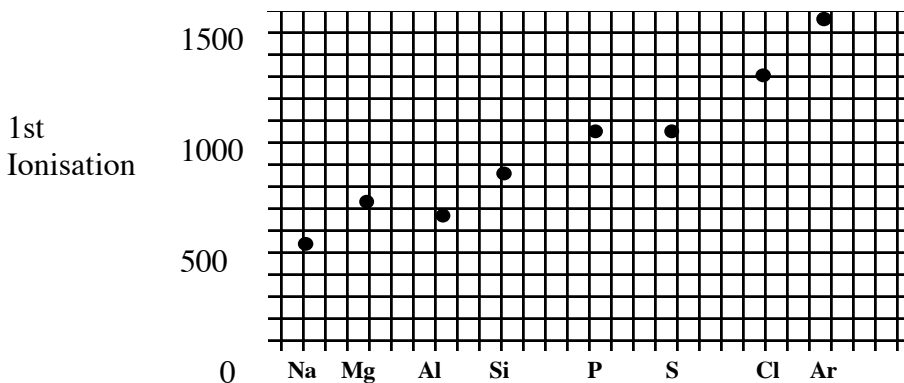
The energy needed to remove one mole of electrons from one mole of gaseous atoms is +1260 kJ mol⁻¹

2. 1st Ionisation energy across a group



As we move across a period the increased attraction (caused by increased nuclear charge) holds the outermost electron more tightly, it takes more energy to remove it.

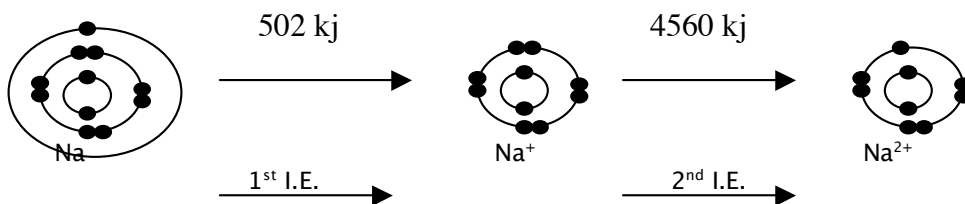
3. 1st Ionisation energy down a group



- As we descend down a group the outermost electron is further from the nucleus this decreases the energy needed to remove an electron.
- The outermost electron is shielded (screened) from the nucleus by inner shells of electrons, this increases down the group, making it easier to remove an electron.
- As we descend down a group the nuclear charge increases making it harder to remove an outer electron.

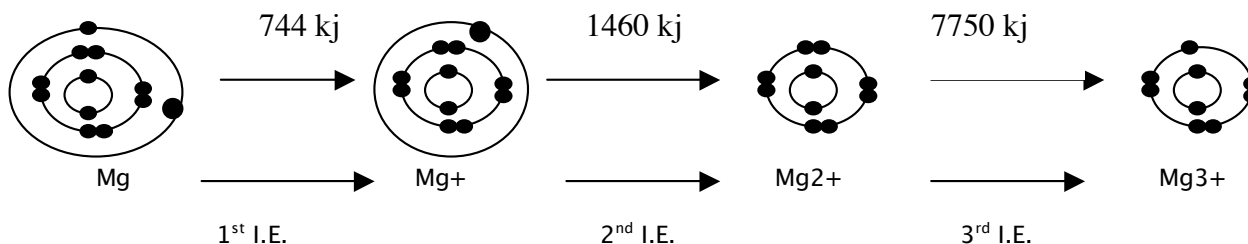
Since the 1st Ionisation Energy decreases down the group a) and b) above must have a greater effect than c).

4. Successive ionisation energies



There is a big increase between 1st and 2nd ionisation energies for sodium as the second electron must be removed from a new electron shell which is

- Nearer to the nucleus
- Has lost screening effect of inner electron shell
- Removing an electron from positive particle rather than a neutral one.



Similarly for magnesium there is a large rise between ionisation energies 2 and 3.