SEE VIDEO 5

First Ionisation Energy

The first ionisation energy is the energy required to remove one electron from each atom in one mole of gaseous atoms to form one mole of gaseous 1+ ions.

In equation form for oxygen:

$$O(g) \longrightarrow O^+(g) + 1e^- \qquad +1314\ KJ\ mol^{-1}$$

1 mol gaseous atoms (not molecules)

1 mol gaseous 1+ ions

There are 3 factors that affect ionisation energy:

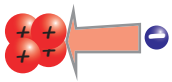
1. Distance from the nucleus or atomic radius
2. Nuclear charge
3. Electron shielding

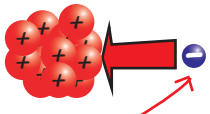
Distance from the nucleus

This electron will experience a greater attractive force from the positive nucleus because it is closer

The further an electron is away from the nucleus, the weaker the attractive force from the nucleus and the easier it will be to remove and so the ionisation energy will be lower

Nuclear charge





This electron will experience a greater attractive force from the nucleus because there are more protons (greater nuclear charge)

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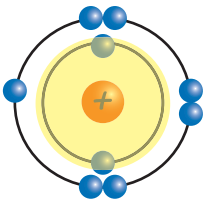
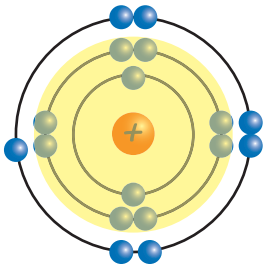
The greater the number of protons, the greater the nuclear charge and the stronger the force of attraction between the nucleus and the electrons and so the ionisation energy increases.

Electron Shielding

Electron shielding is the repulsion of outer electrons from electrons in inner shells that reduces the attractive force from the nucleus.



Flourine

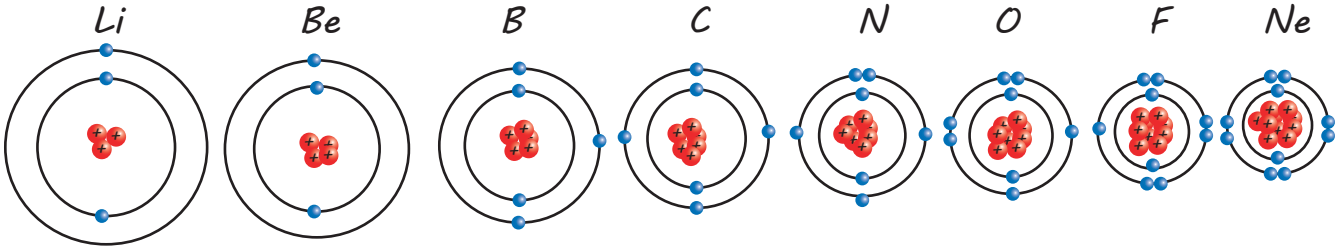


Chlorine

Greater number of electrons in inner shells causes greater repulsion of the outer electrons - Electron shielding is greater

The outer electrons of fluorine experience less electron shielding than the outer electrons of chlorine due to the fact there are fewer electrons in inner shells to repel the outer electrons. Increased electron shielding leads to a lowering of first ionisation energy since the attractive force between the nucleus and the outer electrons is reduced.

Trends in 1st Ionisation Energy Going Across a Period



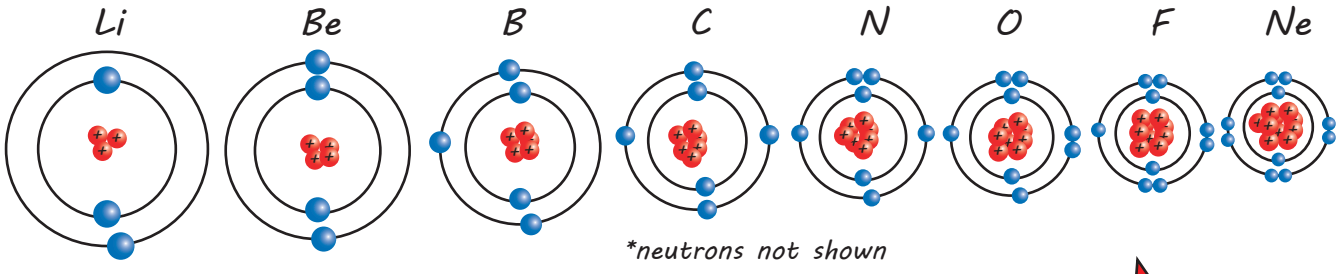
*neutrons not shown

ATOMIC RADIUS DECREASES GOING ACROSS A PERIOD

This is because the number of protons increases which causes the electrons to be pulled closer to the nucleus

A decrease in atomic radius therefore causes the 1st ionisation energy to increase going across a period.

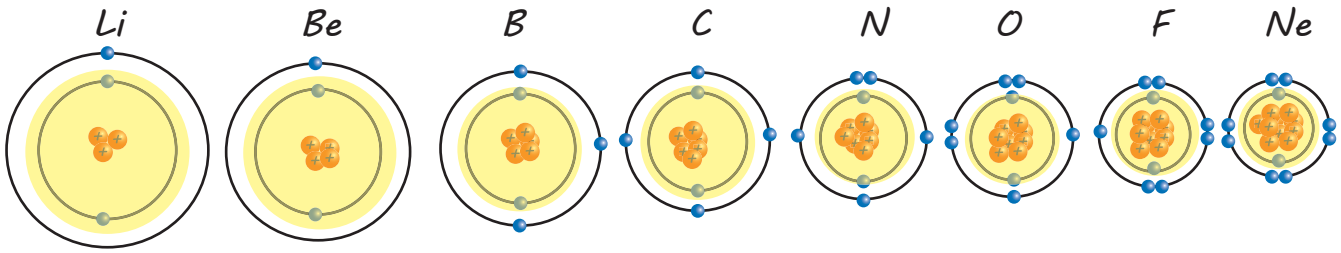
SEE VIDEO 5



*neutrons not shown

NUCLEAR CHARGE (No. OF PROTONS) INCREASES

The nuclear charge increases going across a period which causes the outer electrons to be more strongly attracted to the nucleus causing the first ionisation energy to increase going across a period.



*neutrons not shown

ELECTRON SHEILDING STAYS THE SAME

Since the number of electrons in inner shells remains the same, electron shielding does not change going across a period and so electron shielding has no effect on the 1st ionisation energy.

To summarise:

A decrease in atomic radius going across a period causes the 1st ionisation energy to increase.

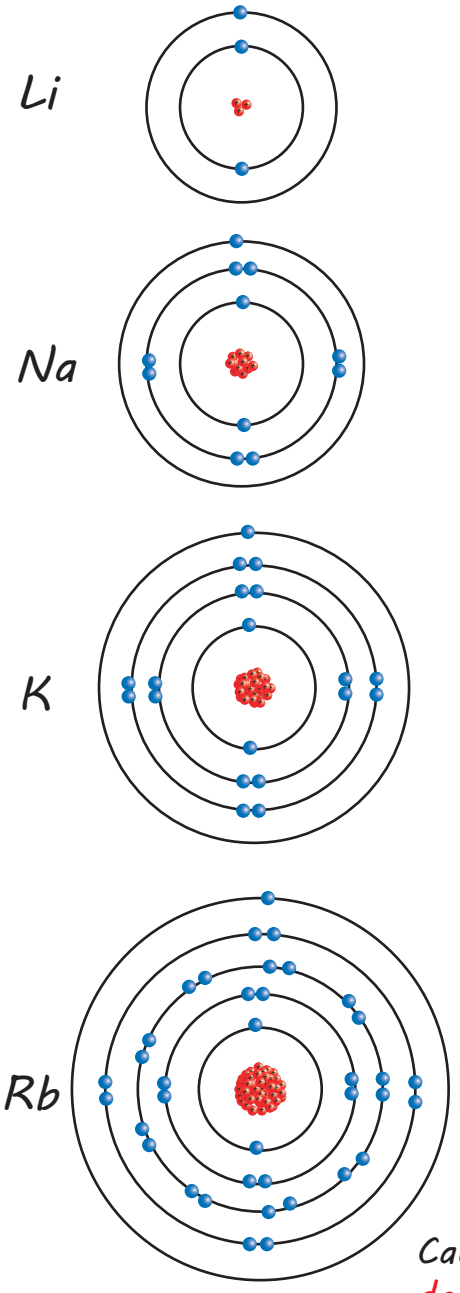
An increase in the nuclear charge going across a period causes 1st ionisation energy to increase.

Electron shielding stays the same going across a period which has no effect on the 1st ionisation energy.

Overall - The first ionisation energy increases going across a period

SEE VIDEO 5

Trends in 1st Ionisation Energy Going Down a Group



*neutrons not shown

ATOMIC RADIUS INCREASES

Causes a decrease in 1st ionisation energy

NUCLEAR CHARGE INCREASES

Causes an increase in 1st ionisation energy

ELECTRON SHIELDING INCREASES

Causes a decrease in 1st ionisation energy

Overall the effects from the atomic radius increasing and the electron shielding increasing predominate over the effect from nuclear charge causing a decrease in 1st ionisation energy as you go down a group.

SEE VIDEO 6

Successive Ionisation Energies

Successive ionisation energies are the energies to remove each electron in turn.

Let's look at what this means:

First we have one mole of gaseous atoms and we remove a single electron from each atom to form one mole of gaseous 1+ ions. This is the first ionisation energy. For example:

$$N(g) \longrightarrow N^{+}(g) + 1e^{-} \qquad 1402 \text{ kJ mol-1}$$

The second ionisation energy is the energy required to remove an electron from a 1+ ion to form a 2+ ion.

$$N^{+}(g) \longrightarrow N^{2+}(g) + 1e^{-} \qquad 2856 \text{ kJ mol-1}$$

The third ionisation energy is the energy required to remove an electron from a 2+ ion to give a 3+ ion. The important thing to remember is that a single electron is removed each time and the energy quoted is the energy for that particular electron being removed.

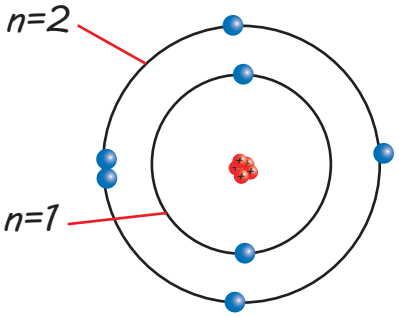
$$N^{2+}(g) \longrightarrow N^{3+}(g) + 1e^{-} \qquad 4578 \text{ kJ mol-1}$$

Evidence for shells

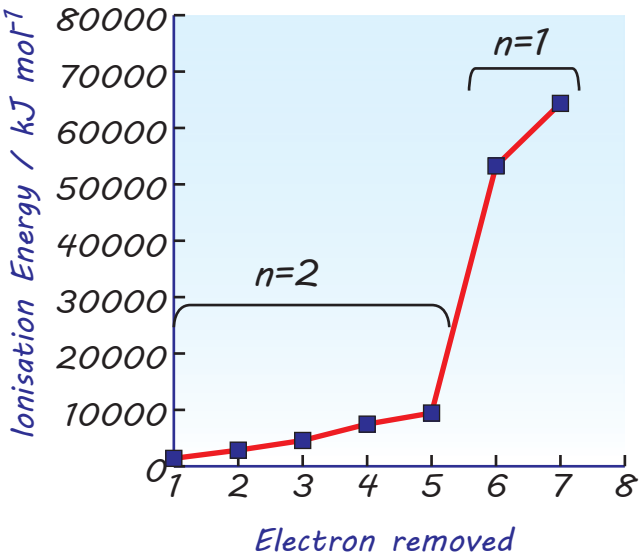
Successive ionisation energies give evidence of shells. Let's look at nitrogen and plot a graph for the ionisation energy as we remove each electron.

SEE VIDEO 6

Successive Ionisation Energies for Nitrogen



*neutrons not shown



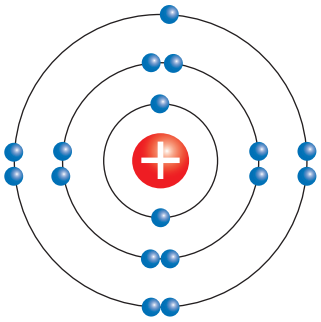
We see the first 5 ionisation energies, corresponding to the 5 outermost electrons in the n=2 shell show a steady increase. Then when we go to removing the next electron, that is in the innermost n=1 shell, there is a sharp rise in the energy required. This is because this 6th electron is closer to the nucleus and so experiences a greater nuclear attraction. Also, the electron shielding is lower for this electron.

By inspecting ionisation energies and identifying where the sudden jumps occur, it is possible to work out the number of electrons in each shell.

SEE VIDEO 7

Orbitals, Subshells and Shells

For many purposes, it is helpful to think of an atom as consisting of a central positive nucleus with electrons orbiting around it in a circular motion.



This is now considered to be an over simplification. We now believe that electrons occupy orbitals.

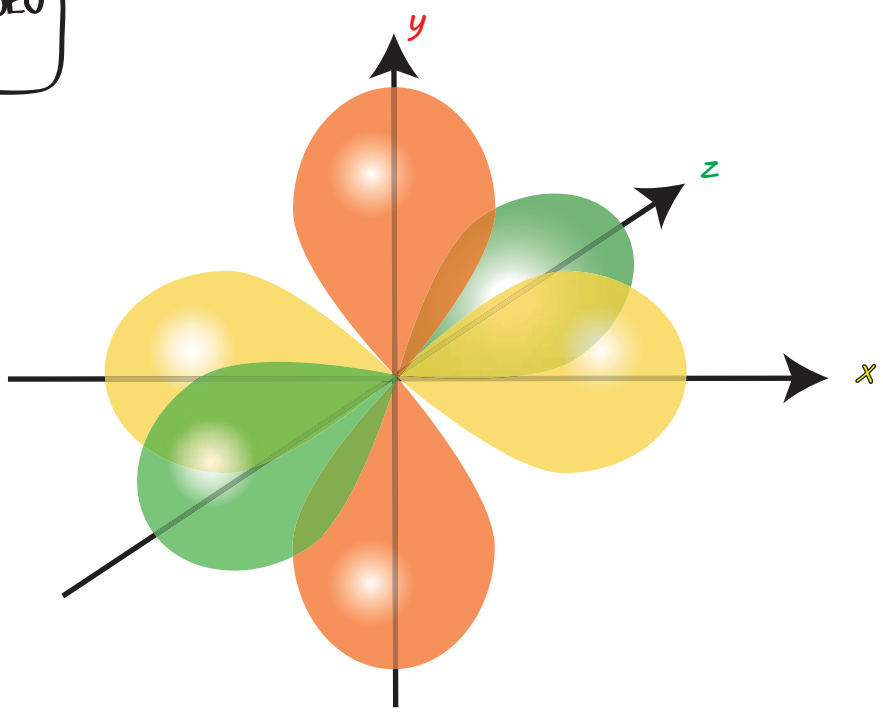
There are different types of orbital which have different shapes. For AS chemistry, you only need to know about two types of orbital; the s orbitals and p orbitals.

The s orbital is spherical in shape and NOT circular.

Each s orbital can accommodate 2 electrons with opposite spins. At any instance in time an s electron is probably located somewhere within the spherical region.

p orbitals have a 3 dimensional dumbell shape with the centre at the nucleus. There are three p orbitals.

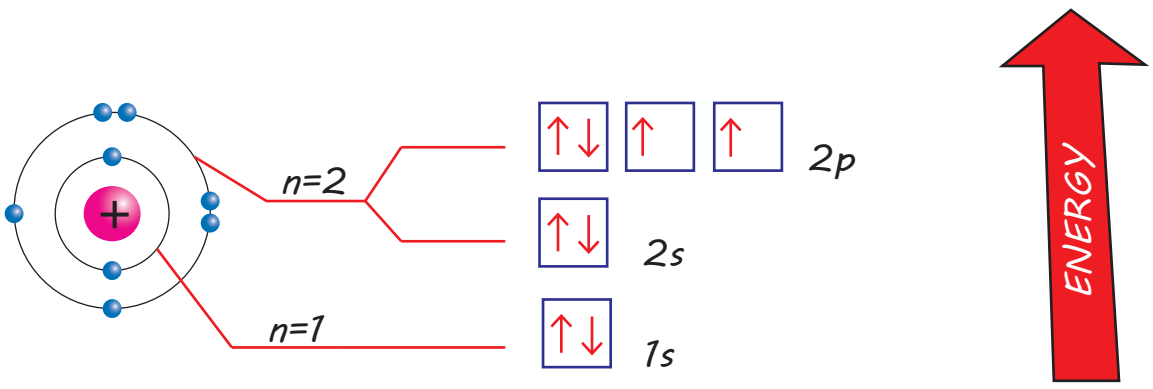
SEE VIDEO 7



One of the p orbitals lies across the x axis, another runs at right angles to this along the y axis, the third p orbital runs at right angles to these across the z axis. Each p orbital can accommodate two electrons with opposite spins. There are therefore a maximum of six p electrons in a p subshell.

An atomic orbital is a region that can occupy up to a maximum of 2 electrons providing the spins are opposite.

It is convenient to represent orbitals using boxes and arrows to represent electrons. Let's look how this can be done for oxygen that has eight electrons; 2 electrons in shell n=1, and 6 electrons in shell n=2



Oxygen

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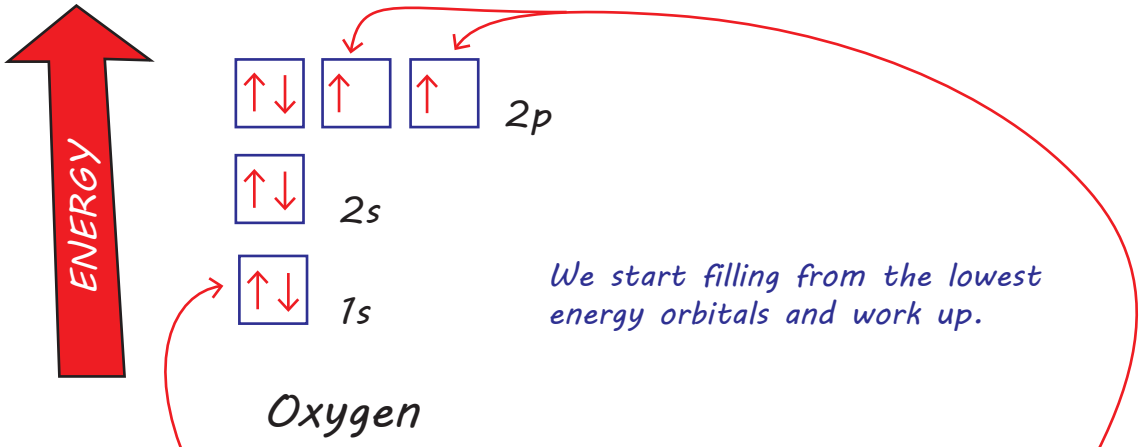
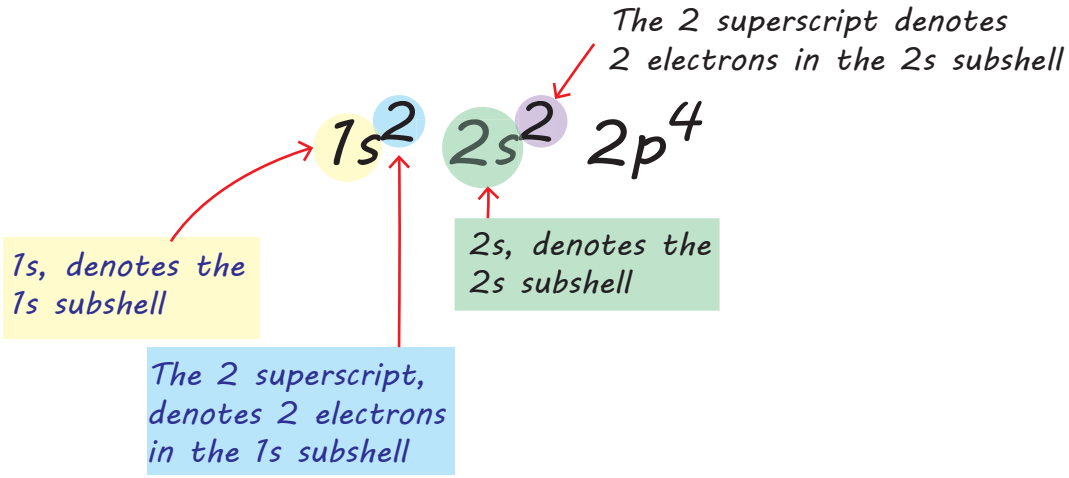
Filling rules

For oxygen, we have 8 electrons to distribute



We would write the full electronic configuration of oxygen as:

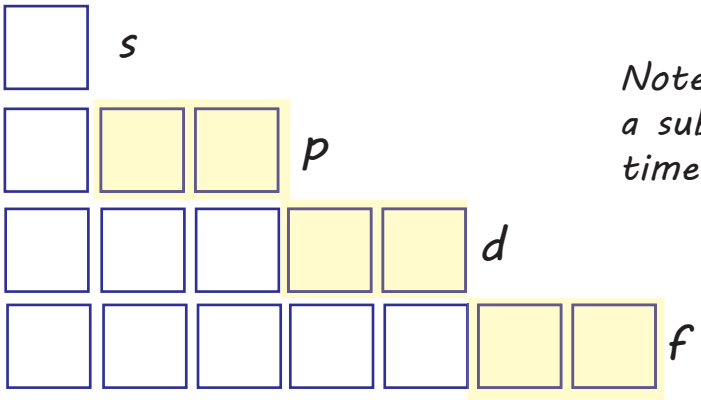


How do we know how many sub-shells there are for each shell?

For shell n=1, there is only one subshell 1s
For shell n=2, there are two subshells 2s and 2p
For shell n=3, there are three subshells 3s, 3p and 3d
For shell n=4, there are four subshells 4s, 4p, 4d and 4f
and so it continues.

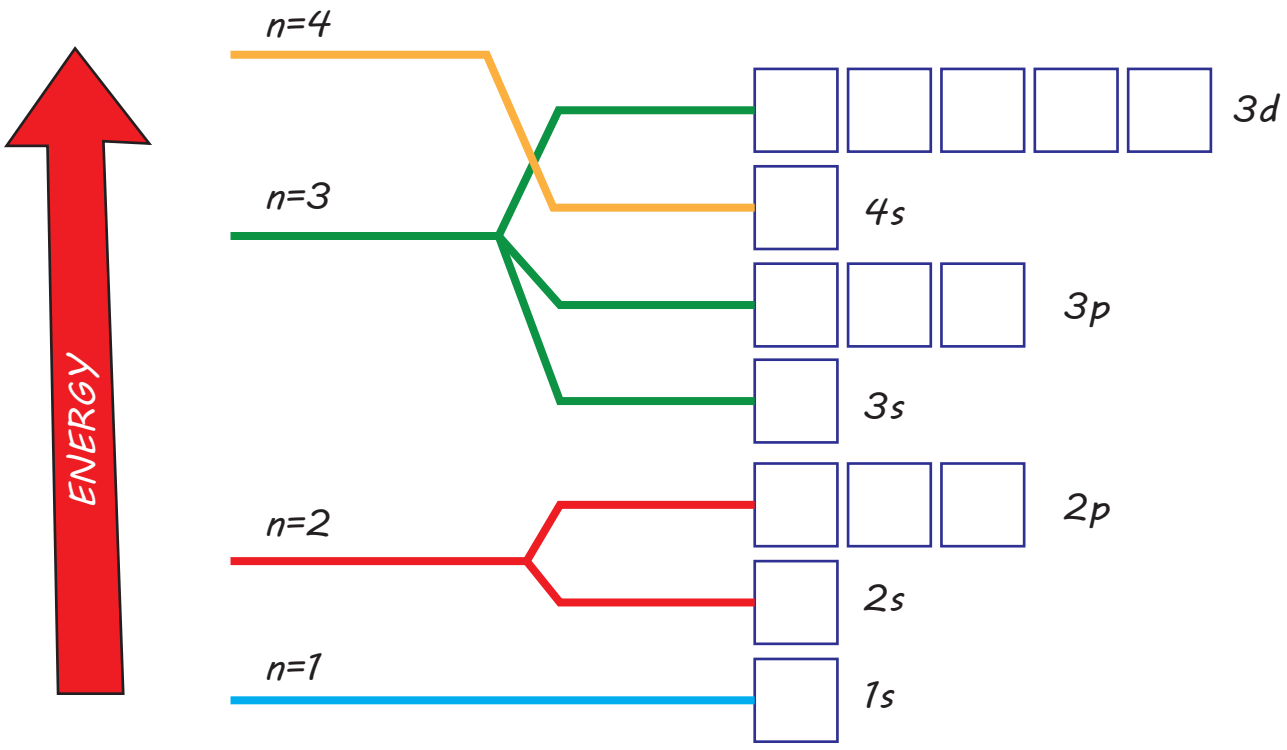
SEE VIDEO 7

How do you know how many orbitals are in a given subshell?



Note: The number of orbitals in a subshell increases by two each time.

When filling up orbitals with electrons, you need to be careful since there is overlap between some of the shells. For example:



We can see that the 4s subshell is lower in energy than the 3d sub-shell and so the 4s subshell fills before the 3d.

Let's look at an example working out the electronic configuration for titanium which has an atomic number of 22. This means we have 22 electrons to distribute.

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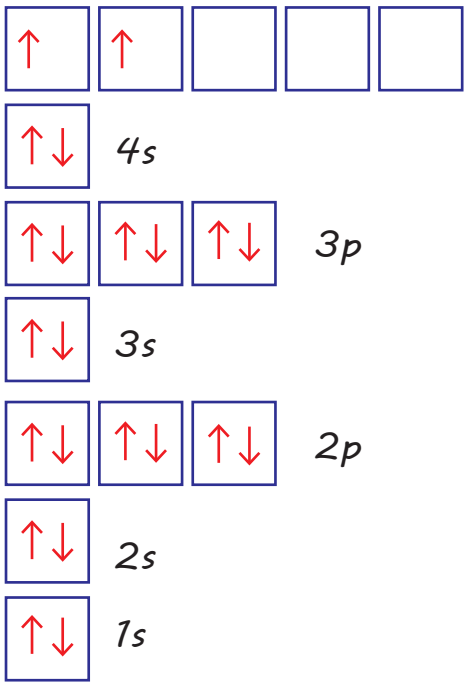
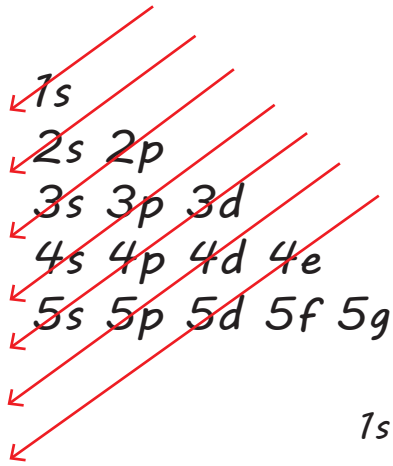


The electronic configuration of titanium is therefore:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

The filling order of the subshells can be worked out by writing each subshell out in a list, starting a new line for each shell:



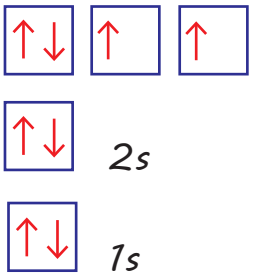
You then fire arrows in diagonally and the order in which they strike the subshell gives the filling order:

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p etc

Electronic configuration of ions

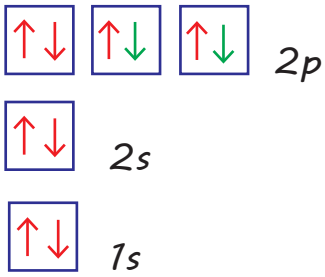
With ions, you either add the required number of electrons to the base atom, in the case of negative ions or remove them, in the case of positive ions.

For example, O^{2-}



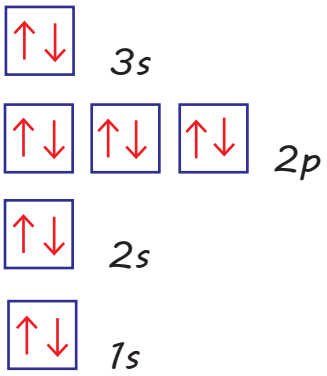
Oxygen base atom
$1s^2\ 2s^2\ 2p^4$

The charge is 2-
so 2 electrons
need to be added



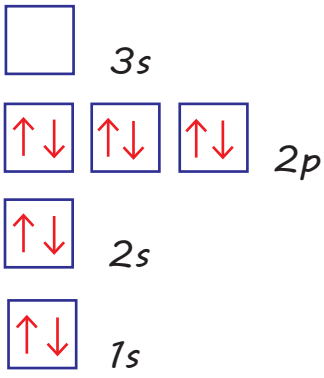
O^{2-}
$1s^2\ 2s^2\ 2p^6$

Now let's look at a positively charged ion, Mg^{2+}:



Magnesium base atom
$1s^2\ 2s^2\ 2p^6\ 3s^2$

We need to remove 2
electrons from the
highest energy subshell



Mg^{2+}
$1s^2\ 2s^2\ 2p^6$

Be careful when removing electrons from the atoms with a 3d subshell since the 4s electrons are removed before the 3d electrons.

REMEMBER: The 4s subshell fills before the 3d but the 4s electrons are also the first to leave.

For example, let's consider Ti^{2+}

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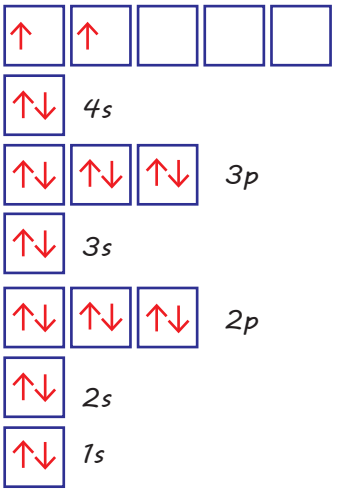
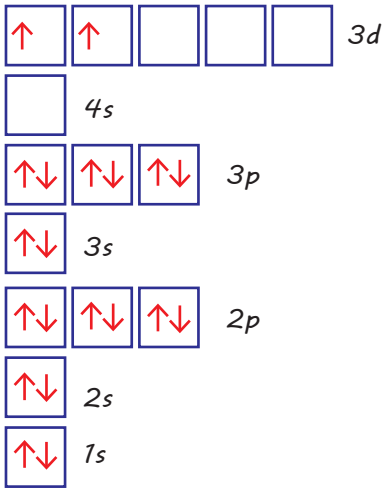


We need to remove 2 electrons but these need to come from the 4s subshell rather than the 3d*



This is the electronic configuration of titanium

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

Ti^{2+}

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$

*The reason for this is because when the the 4s and 3d subshell are occupied with electrons, they swap over and 4s becomes higher in energy than 3d.

Shortened electronic configurations

You sometimes see electronic configurations written like this:

For example, for potassium, K:

Full electronic configuration

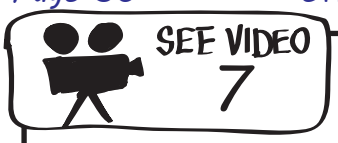
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

Shortened electronic configuration

$[Ar]\ 4s^1$

This is the electronic configuration of the nearest noble gas to potassium, going backwards, and so rather than write out $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ you write the symbol for Ar in square brackets.

Similarly, sodium would be, $[Ne]3s^1$



Principal quantum number, n, is simply the shell number however for the exams you do need to learn the proper definition:

Principal quantum number, n, is number representing the overall relative energy of each of a group of orbitals - this increases with increased distance from the nucleus.

A shell is a group of atomic orbitals with the same principal quantum number.

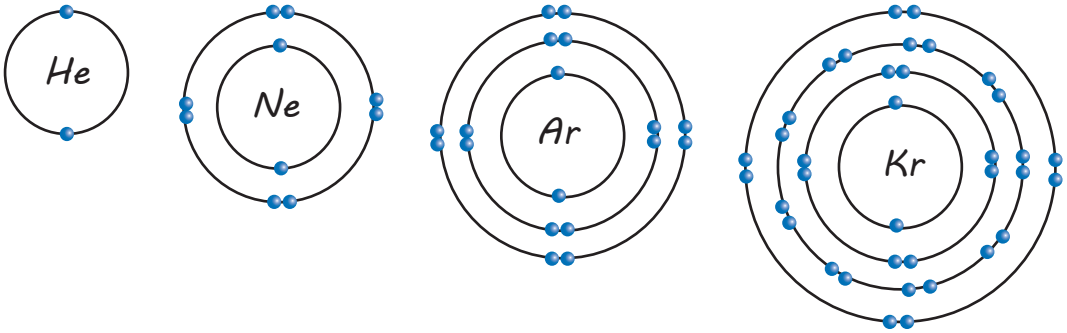
SEE VIDEO 8

Chemical Bonding

Noble gases

4.0	He	Helium 2
20.2	Ne	Neon 10
39.9	Ar	Argon 18
83.8	Kr	Krypton 36
131.3	Xe	Xenon 54
(222)	Rn	Radon 86

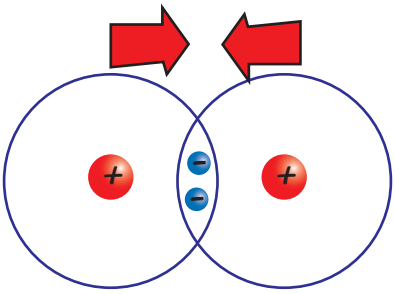
The noble gases are unusual in that they exist on their own as single atoms. This is to do with their stable electronic configurations:



Helium has a full outer shell of two electrons and the other noble gases have eight electrons is their outer shell. Other elements try to achieve the same electron configurations as the noble gases by chemical bonding. When elements achieve 8 electrons in their outer shell through bonding we say they have obeyed the octet rule.

Covalent Bonding

Covalent bonding generally occurs between non metallic elements and involves atoms sharing pairs of electrons:



The two positive nuclei are attracted to the shared pair of electrons - bonding the two atoms together to make a molecule

A covalent bond involves two atoms sharing a pair of electons.

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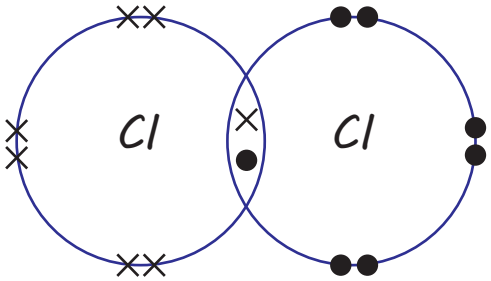
Covalent bonding can be represented using dot and cross diagrams.

For example, the chlorine molecule, Cl_2

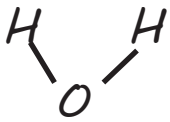


In dot and cross diagrams, you only need to draw the electrons in the outer shell

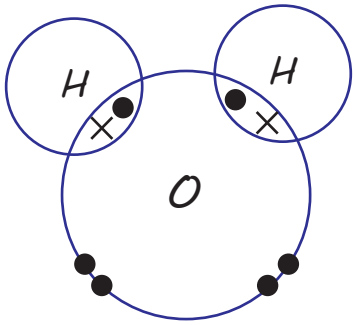
We can simplify this by drawing a straight line to represent a shared pair of electrons:

Cl — Cl

In a similar way, water, H_2O, can be represented like this:



Or in a dot and cross diagram like this:



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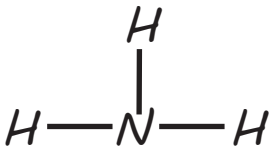
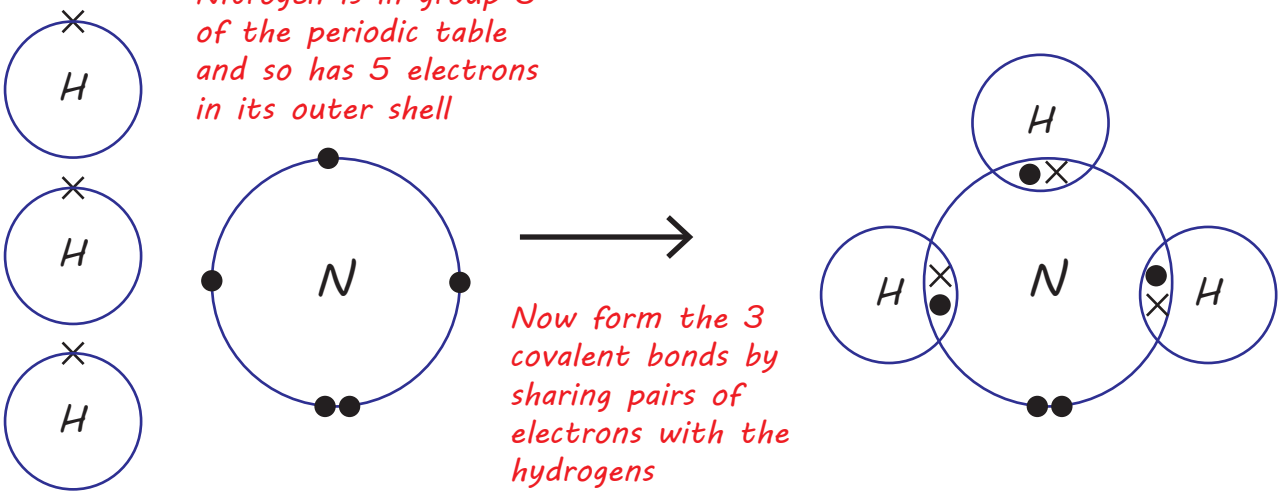
Similarly ammonia, NH_3, can be represented like this:



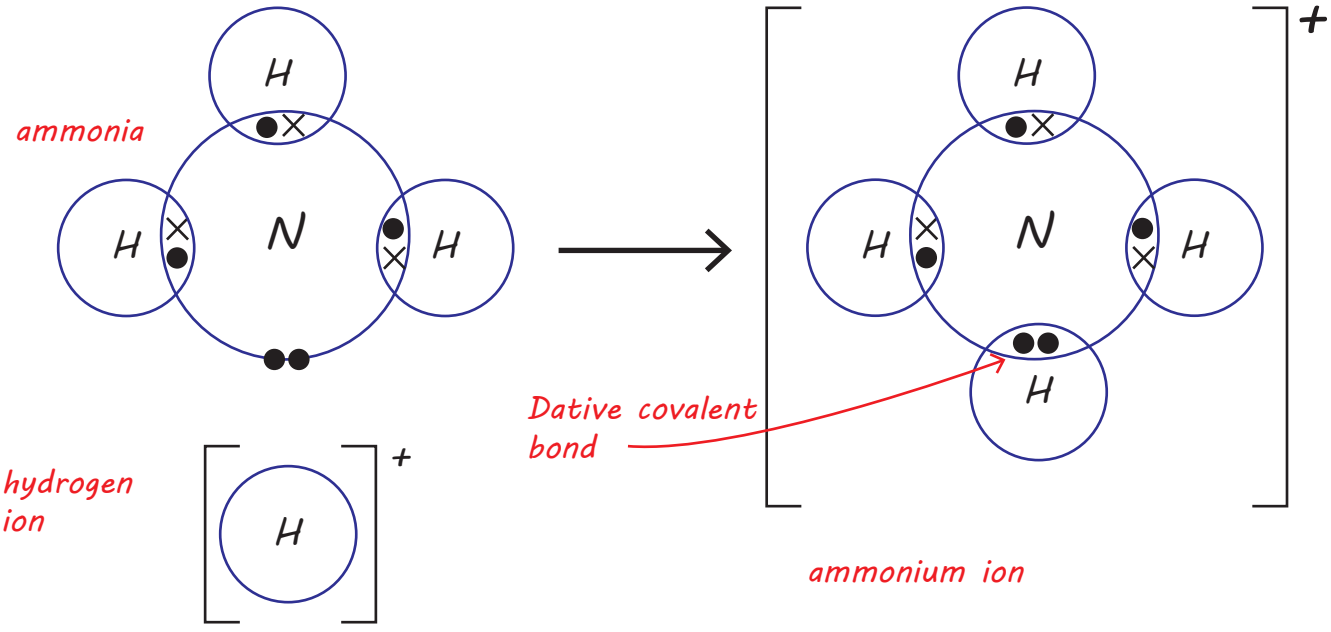
Or, as a dot and cross diagram:

Start with the base atoms:

Nitrogen is in group 5 of the periodic table and so has 5 electrons in its outer shell



The ammonium ion, NH_4^+, contains a special type of covalent bond known as a dative covalent bond. This is where two electrons that form the bond are donated from one atom; in this case nitrogen.



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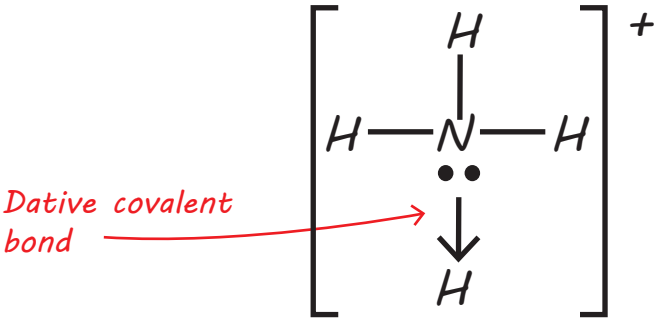
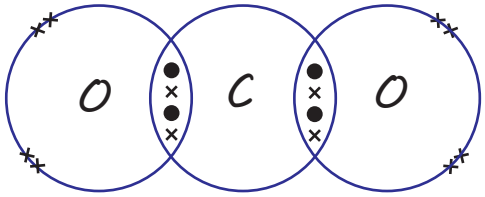
We can draw the ammonium ion in short form as follows:



Carbon dioxide contains two carbon to oxygen double bonds which are represented in short form like this:

$$O{=}C{=}O$$

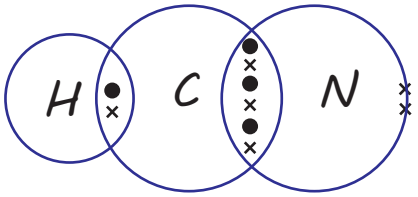
And in a dot and cross diagram like this:



The highly toxic molecule, hydrogen cyanide, HCN, can be represented like this:

$$H{-}C{\equiv}N$$

And:



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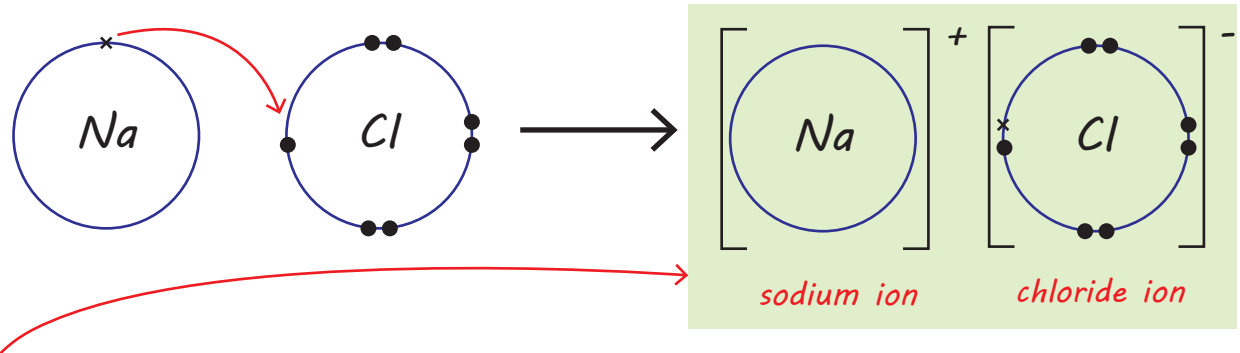
Ionic Bonding

Ionic bonds tend to occur between metals and non-metals.

An ionic bond is and electrostatic attraction between oppositely charged ions.

Ionic bonding involves the transfer of electrons. In sodium chloride, the single outer electron from sodium is transferred to chlorine.

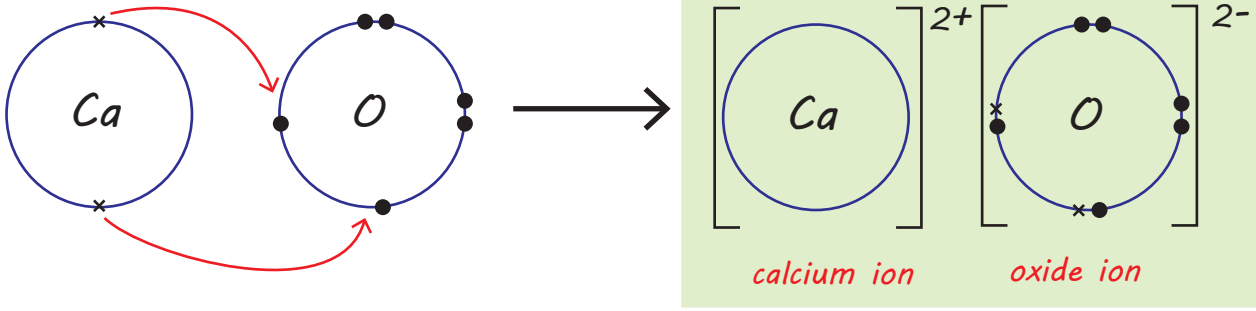


Both the sodium and the chlorine end up with an octet in their outer shell. The sodium, because it has lost an electron, ends up with a single positive charge. The chlorine, having gained an electron ends up with a single negative charge.

When asked to draw a dot and cross diagram in an exam, you only have to draw the final diagram of the charged ions. For the metal ion, it is perfectly acceptable to show just the empty outer shell.

Now let's look at calcium oxide, CaO:

Calcium, being a group 2 element, has 2 electrons in its outer shell and so will need to lose these in order to obtain an octet. Oxygen being a group 6 element, has 6 electrons in its outer shell and needs to gain 2 electrons to obtain its octet.



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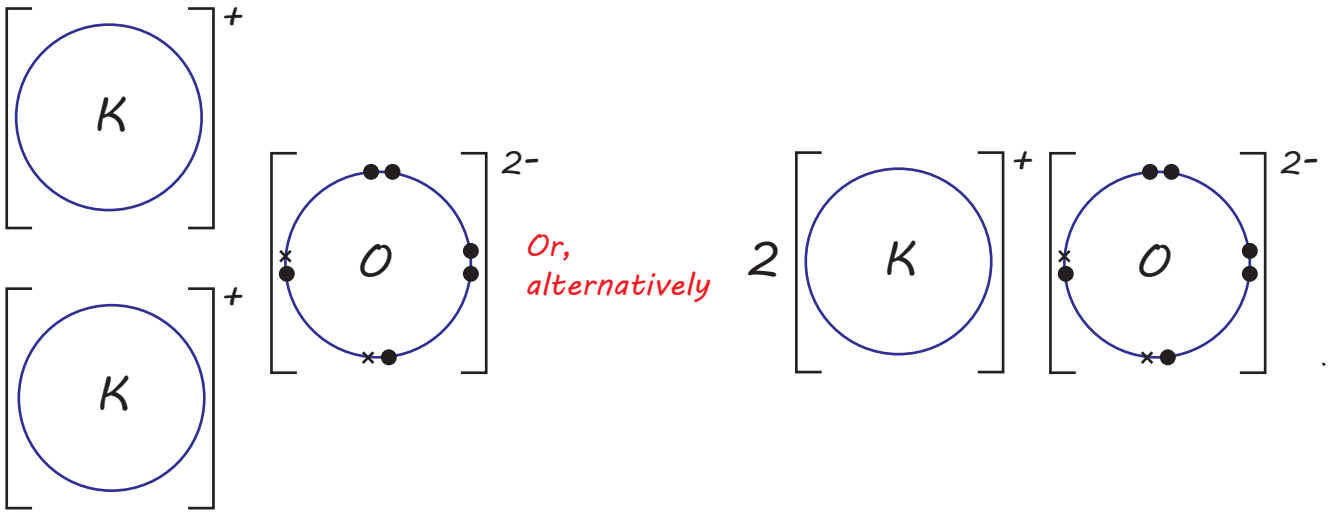
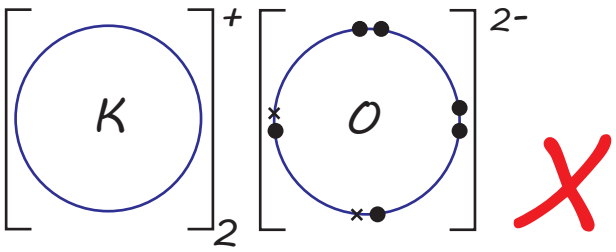
Now let's look at potassium oxide, K_2O:

This can be represented in two ways:



However, this way is incorrect:

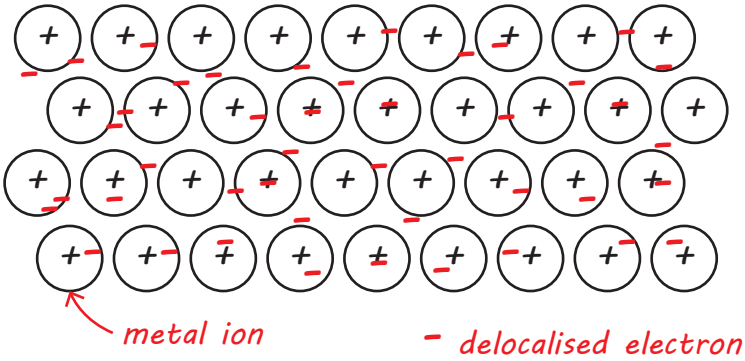


Metallic Bonding

Metallic bonding occurs in metals.

Metallic bonding is the electrostatic attraction between positive metal ions and delocalised electrons.

Sodium has a single electron in its outer shell and this is delocalised, meaning it is not attached to any single atom but is free to move around. The metallic bonds in sodium are therefore formed from singly charged positively charged sodium ions and delocalised electrons.



When asked to draw a diagram like this in an exam, make sure you draw at least 3 rows of metal ions and you label the ions and the delocalised electrons

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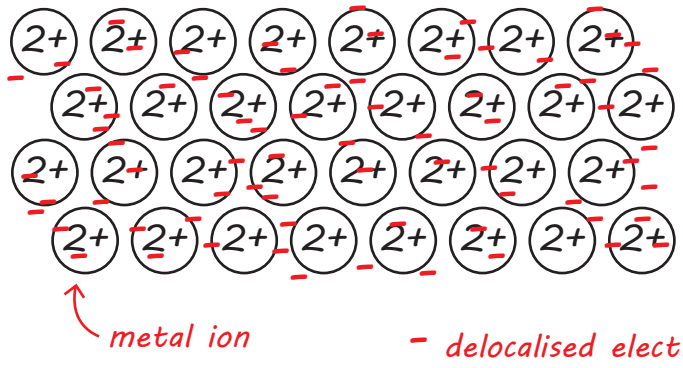
Magnesium, being a group 2 element, has two electrons in the outer shell and these 2 electrons are delocalised which leaves a metallic ion with a 2+ charge.



The increased charge on the magnesium metal ion and the greater number of delocalised electrons makes a stronger metallic bond than sodium and consequently explains why magnesium has a higher melting point than sodium since more energy is required to break the metallic bonds.

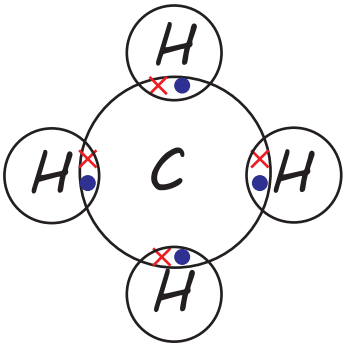
Why do metals conduct electricity?

The delocalised outer electrons in metals account for why metals conduct electricity since these delocalised electrons are free to move and carry the charge.

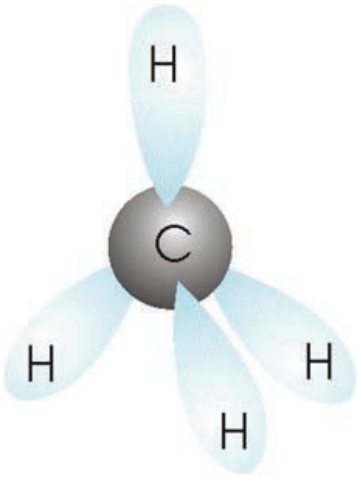
SEE VIDEO 9

Shapes of Molecules

With methane, CH_4, carbon has 4 electrons in its outer shell, hydrogen has just one and so 4 single covalent bonds are formed.

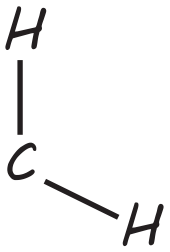


There are four pairs of bonding electrons. These bonding pairs of electrons will repel each other equally and the bonds will space themselves out as far as possible.

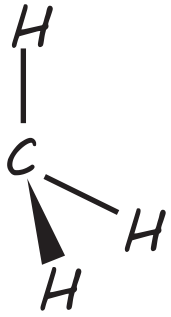


When the four pairs of bonding electrons are separated as far as possible, the molecule ends up in a tetrahedral arrangement.

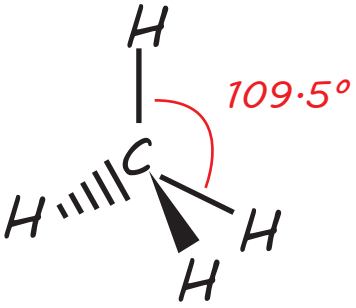
When drawing a tetrahedral molecule in 2D, we first draw the carbon and then two of the bonds to the hydrogen which are in the same plane as the paper.



We then need to draw a bond coming out of the paper towards us and we do that with a solid wedge.



That leaves the final bond that goes away from us to the rear and we draw that with a dashed wedge. In methane all the bond angles are 109·5°



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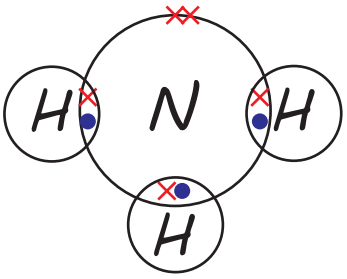
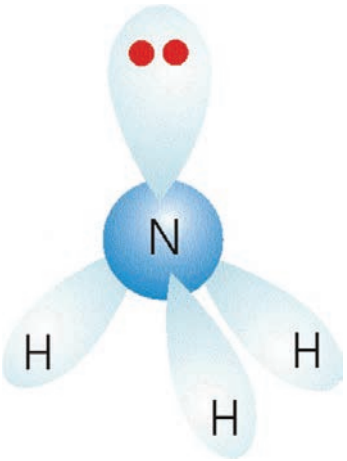
Now let's look at ammonia, NH_3. Nitrogen has 5 electrons in its outer shell and will form 3 covalent bonds with 3 hydrogen atoms.

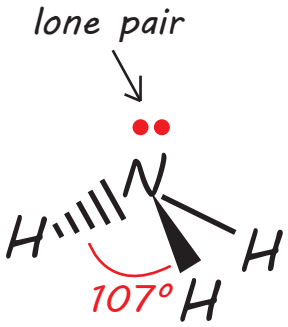


The arrangement here is different from the previous example with methane in that there are 3 pairs of bonding electrons and a lone pair of non-bonding electrons. Non-bonding electrons do however influence the shape since these repel the bonding electrons and as a result the bonding and non-bonding electrons arrange themselves like this:



However, when we assign a shape to the molecule we ignore the non-bonding electrons and just consider the arrangement of the atoms. The shape is pyramidal.

An important point is the nonbonding electrons have a greater repulsion on the bonding electrons and so the bond angles in ammonia are smaller than methane.



The effect of the lone pair leads to a reduction in the bond angle by 2·5° and so the bond angle in ammonia is 107°

Remember it is the electron pairs that are repelling, not the atoms.

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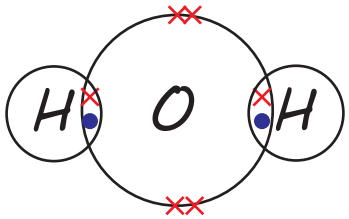
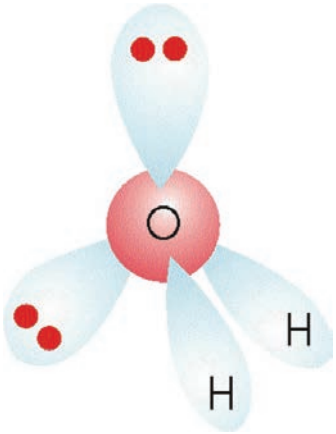
Now let's look at water, H_2O. Oxygen has 6 electrons in the outer shell and forms 2 single covalent bonds with 2 hydrogen atoms and this results in 2 bonding pairs of electrons and 2 lone pairs of electrons from the oxygen that need to distribute themselves to minimise the repulsion of like charges.

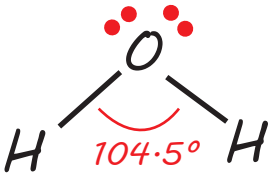


The bonding electrons and lone pairs of electrons arrange themselves like this:

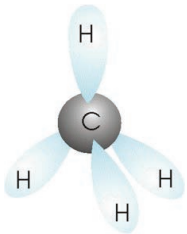


The shape is non-linear or bent.

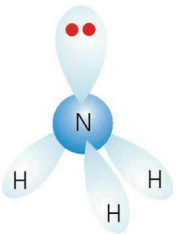
Since there are now two lone pairs, the repulsion of the lone pairs is greater still and so the bond angle is reduced to 104.5°



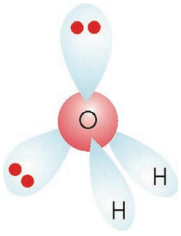
For every pair of non-bonding electrons present, the bond angle is reduced by 2.5°



No lone pairs present so bond angles are 109.5°



One lone pair present so bond angle is reduced by 2.5° to 107°



Two lone pairs so bond angle is reduced by 5° to 104.5°

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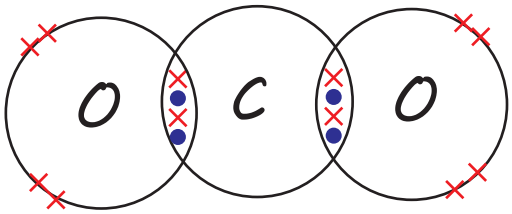
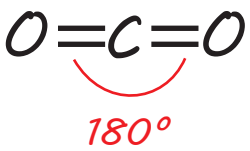
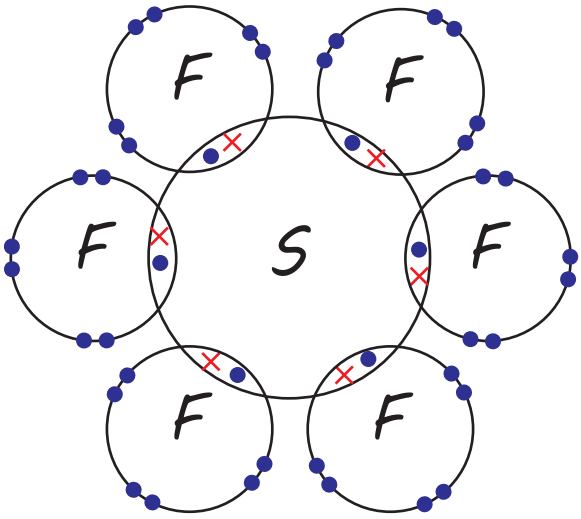
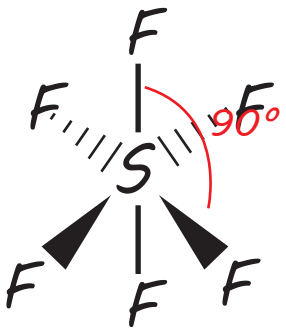
Carbon dioxide, CO_2, has two carbon to oxygen double bonds. There are no non-bonding pairs on the carbon and so the electrons that form the double bonds will space themselves out as far as possible to give a linear shape with a bond angle of 180°





With sulfur hexafluoride, SF_6, sulfur has six electrons in its outer shell (group 6 of periodic table) and forms 6 single covalent bonds with fluorine. There are no non-bonding electron pairs so the bonding electrons arrange themselves to minimise repulsion of each other.





The shape of SF_6 is octahedral and the bond angles are 90° throughout

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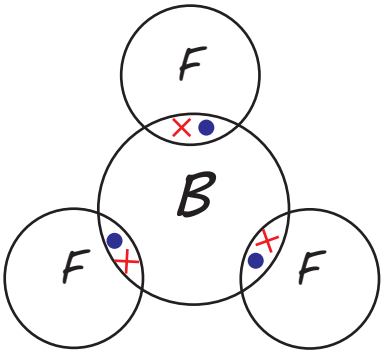
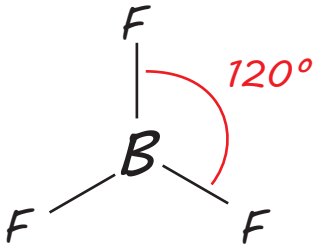
With boron trifluoride, BF_3, there are 3 bonding pairs of electrons and no lone pairs.



This leads to a trigonal planar arrangement of atoms similar to a 3 bladed propellor with bond angles of 120°



Note the spelling of trigonal planar, it is trigonal and not triagonal.

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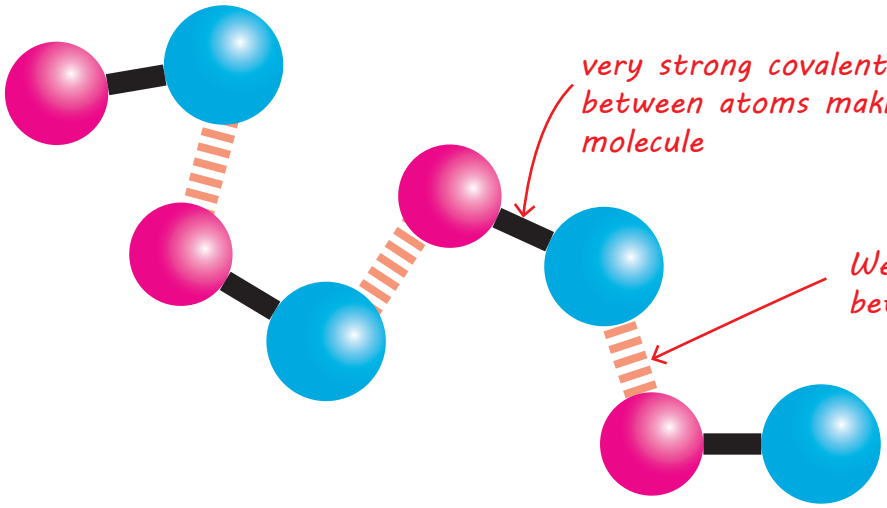
Intermolecular Forces

Covalently bonded molecules experience forces of attraction between each other:



These intermolecular forces of attraction hold material together when it is solid and prevent the material becoming a gas when it is a liquid.

There are 3 types of intermolecular force:

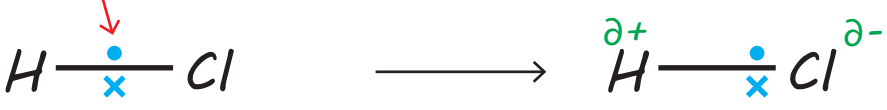
1. Permanent dipole-dipole
2. Hydrogen bonding
3. van der Waals' forces

1. Permanent dipole-dipole

Permanent dipole-dipole forces occur when one of the bonded elements has a greater attraction for the bonded pair of electrons. Let's look at hydrogen chloride, HCl:

shaired pair of electrons making up covalent bond

shaired electrons are pulled towards chlorine making the chlorine slightly negative (∂-) and the hydrogen slightly positive (∂+)



chlorine has a greater attraction for this shaired pair of electrons

this difference in polarity across the molecule is called a permanent dipole

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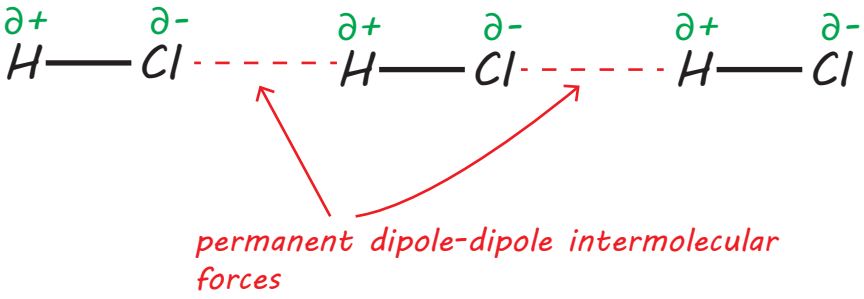
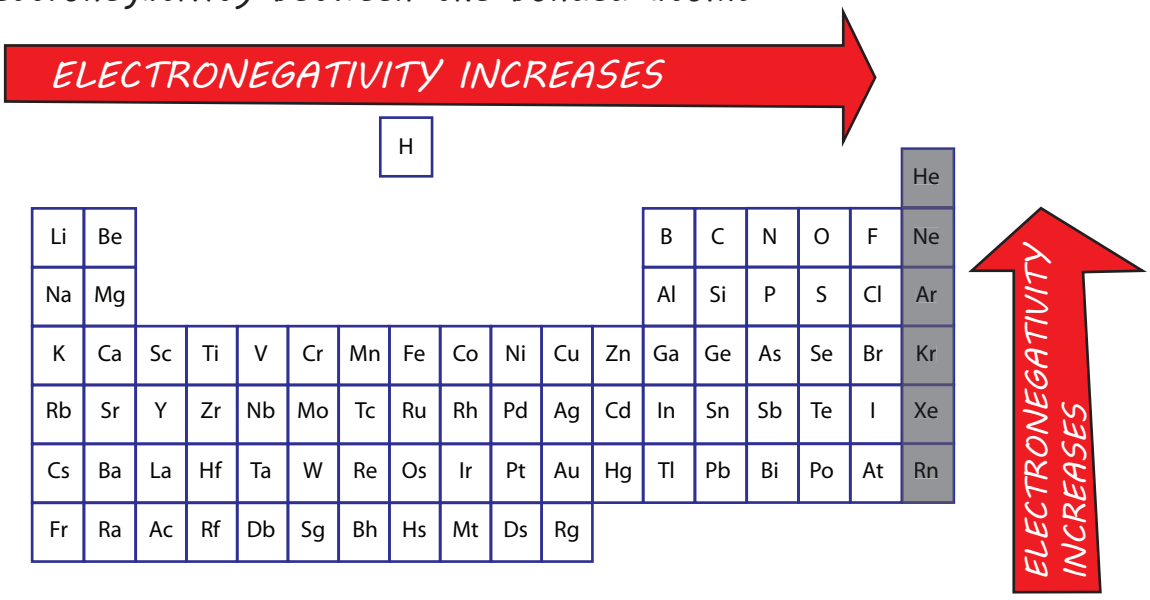
Since chlorine has a greater attraction for electrons than hydrogen, we say it has a greater electronegativity.

Electronegativity is the measure of attraction a bonded atom has for a bonded pair of electrons.



The ∂+ on the hydrogen of one molecule has a weak attractive permanent diplole-dipole force with the ∂- of a chlorine from another molecule. Permanent dipoles occur when there is a difference in electronegativity between the bonded atoms.



Electronegativity increases going from left to right across the periodic table (ignoring the noble gases) and going up the group. Fluorine is therefore the most electronegative element.

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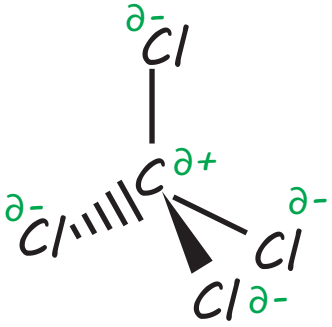
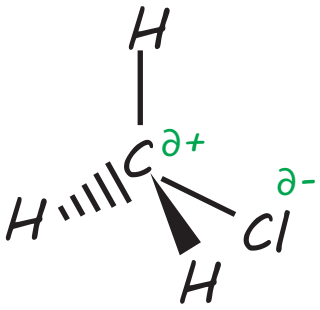
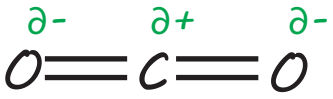


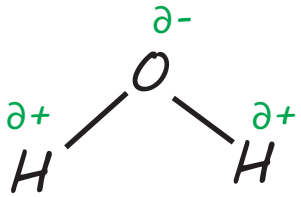
With a symmetrical molecule like the tetrahedral tetrachloromethane, CCl_4, the dipoles cancel which leaves the molecule nonpolar.



However, with chloromethane, this is not symmetrical and so it does have a permanent dipole making this solvent polar in nature.



Carbon dioxide is symmetrical and so the dipoles cancel making the molecule non-polar.



Water however, is not symmetrical and so is polar.

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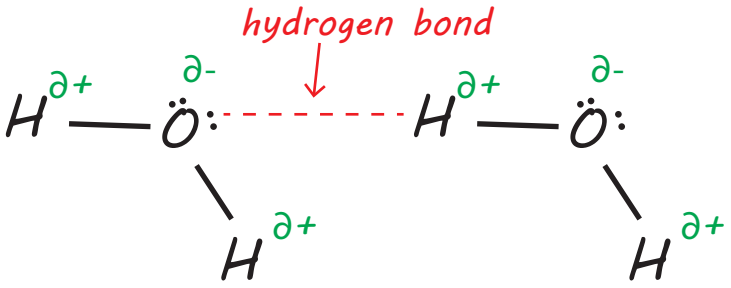
2. Hydrogen Bonding

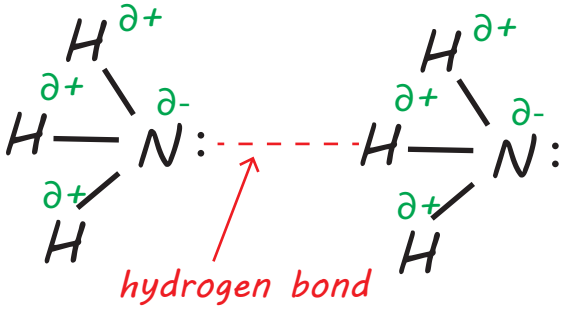
There is a special type of permanent dipole-dipole force when hydrogen is bonded to either nitrogen or oxygen. In such cases, the permanent dipole is a strong one. A hydrogen bond forms between the ∂+ of the hydrogen and the lone pair of electrons from the nitrogen or oxygen in another molecule.

Hydrogen bonding in water:



When drawing a hydrogen bond, be sure to include the dipoles and then draw the hydrogen bond, as a dotted line going from the lone pair of the nitrogen or oxygen to the ∂+ on the hydrogen from another molecule. Label the hydrogen bond.

Hydrogen bonding in ammonia:



Although called a hydrogen bond, the hydrogen bond is not a chemical bond but an intermolecular force.

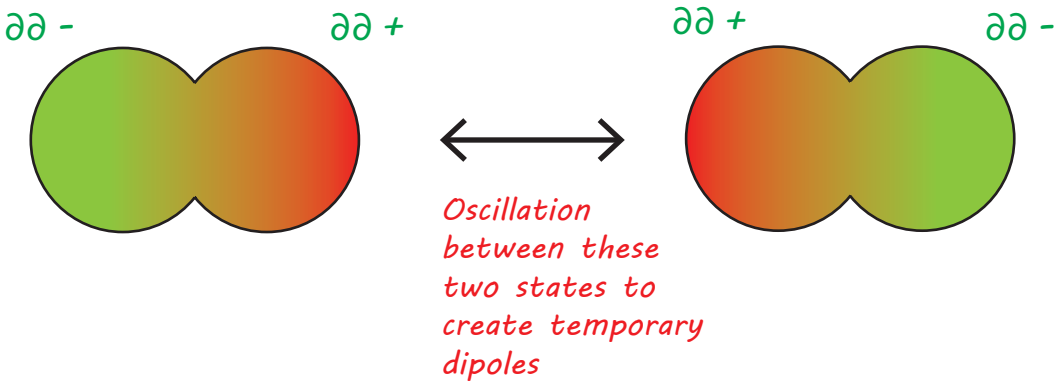
Hydrogen bonding explains why water has a relatively high boiling point and freezing point a high surface tension and why as a solid, ice is less dense than water - due to the open lattice structure.

SEE VIDEO 10

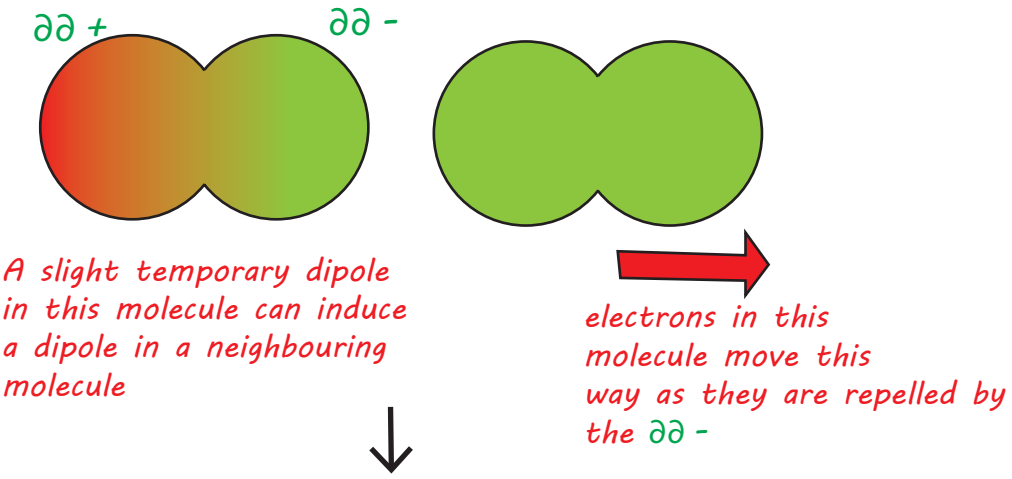
3. van der Waals' Forces

In situations where the electronegativity between bonded atoms are the same, for example in covalently bonded elements like Cl_2 and O_2, there are no permanent dipoles since each atom has an equal attraction for the bonded pair of electrons.

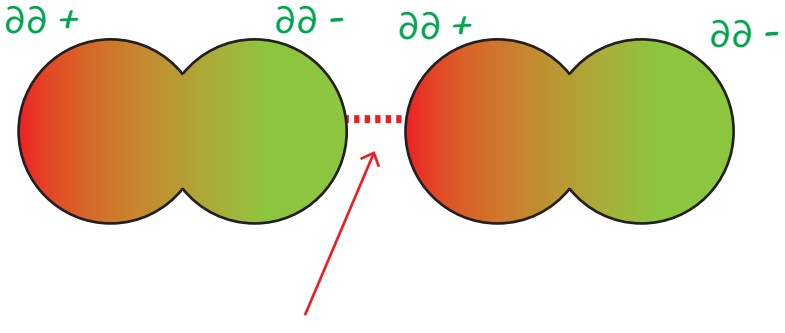
However, the electrons do oscillate from one side of the molecule to the other and therby create weak temporary dipoles:



A temporary dipole in one molecule can induce a dipole in another molecule:



This results in a weak intermolecular force called a van der Waals' Force:



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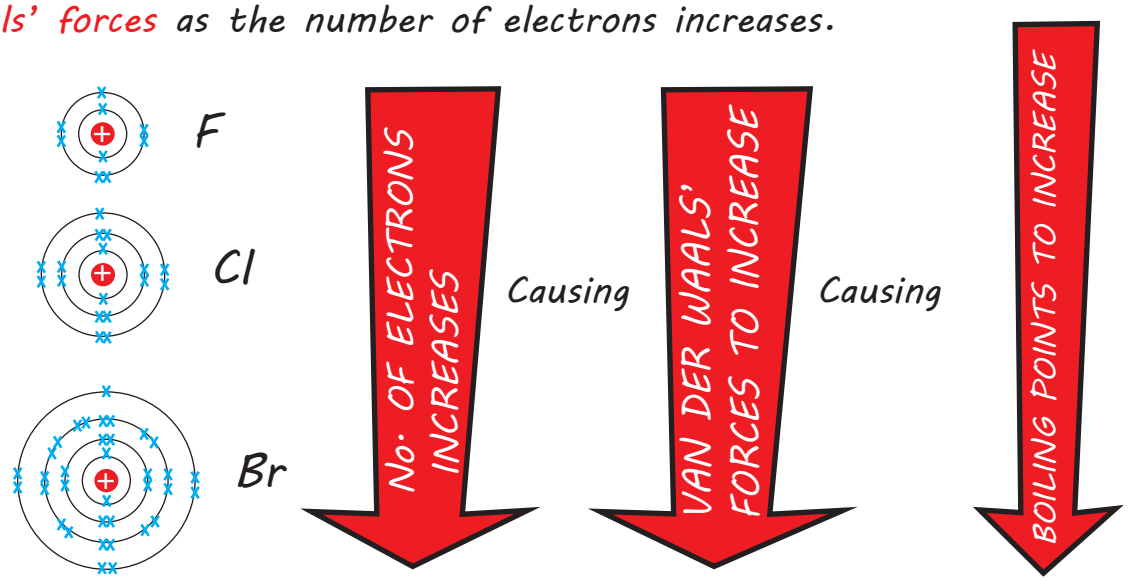
If you are asked in an exam to explain how van der Waals' forces occur, you may write the following:

Temporary dipoles can occur due to oscillations in the movement of electrons. These temporary dipoles can induce dipoles in neighbouring molecules giving rise to a weak intermolecular force called van der Waals' forces.

van der Waals' forces increase with increased number of electrons and so the boiling points of group 7 elements increase going down the group since more energy is required to overcome the stronger van der Waals' forces as the number of electrons increases.



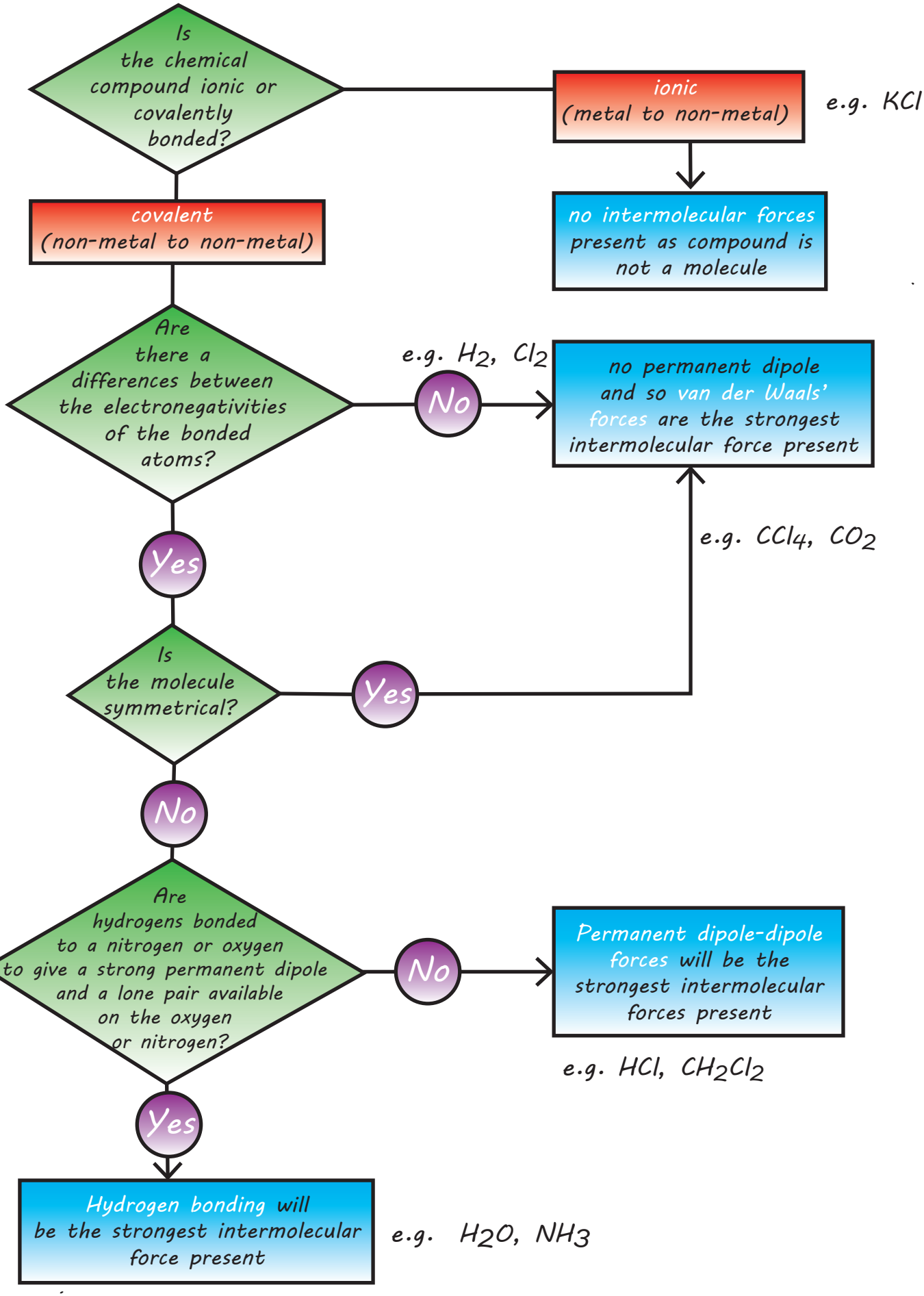
It is important to realise that van der Waals' forces occur in all molecules although only in nonpolar molecules is this the strongest intermolecular force present. For example, in HCl, van der Waals' forces are present but the permanent dipole-dipole interaction is the strongest intermolecular force.

The order of strength of intermolecular forces are:

Hydrogen bonding > Permanent dipole-dipole > van der Waals' forces

SEE VIDEO 10

How can we tell which are the strongest intermolecular forces present?



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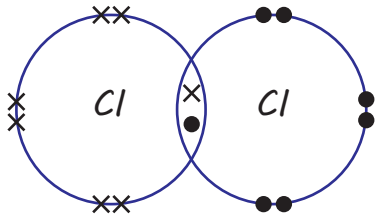
Bonding and Structure

Covalently bonded molecules

When non-metals bond they tend to form covalent bonds by sharing pairs of electrons:

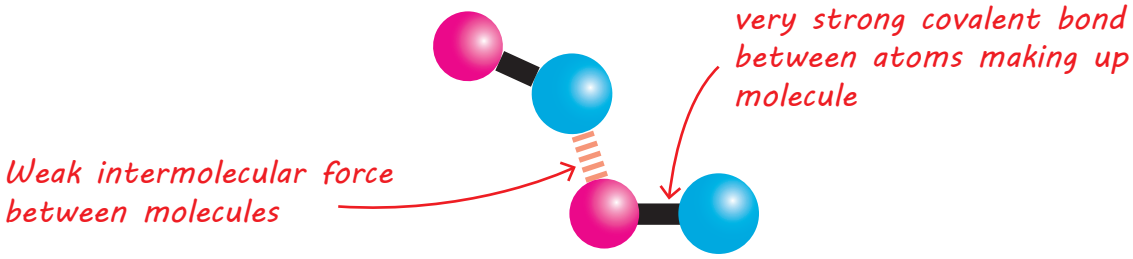


or $Cl—Cl$

Covalently bonded compounds can have the following structures:

1. Simple molecular structure

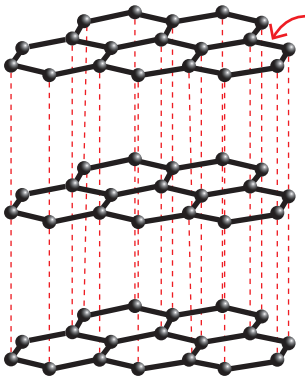
   For example, O_2, Cl_2 CH_4



Compounds with simple molecular structures tend to have low melting and boiling points since it is the weak intermolecular forces that need to be overcome to melt and boil them rather than the strong covalent bonds.

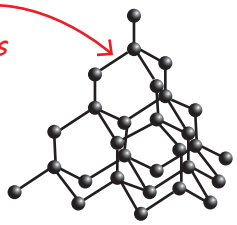
2. Giant covalent lattice structure

   In these compounds, strong covalent bonds extend throughout the entire lattice. For example, in the carbon allotropes of graphite and diamond:



graphite

strong covalent bonds extend throughout the lattices



diamond

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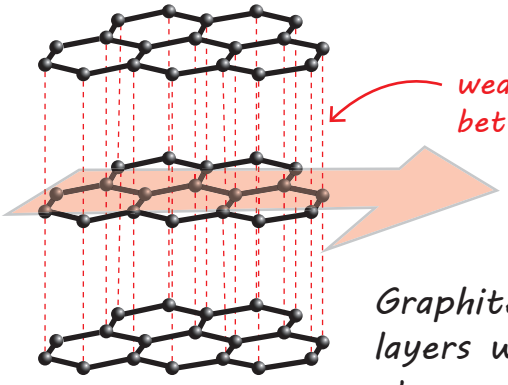
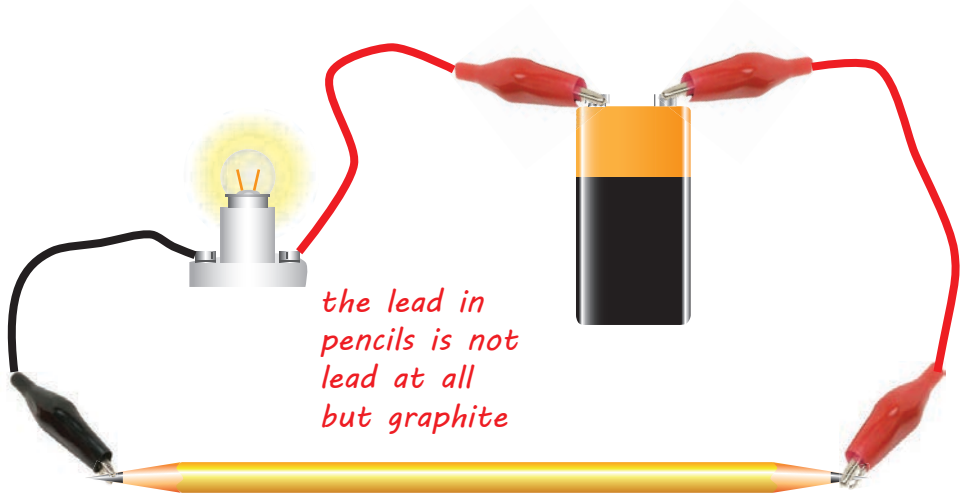
Compounds with giant covalent lattice structures have high melting and boiling points because these compounds have strong covalent bonds throughout the entire lattice and it is these strong covalent bonds that need to be overcome in order to melt and boil them.



graphite

Graphite is unusual in that it forms hexagonal layers which consist of covalently bonded carbon atoms. There are weak van der Waals' forces between the layers which allows the layers to slide over one another and this explains why graphite is soft and can be used as a lubricant.

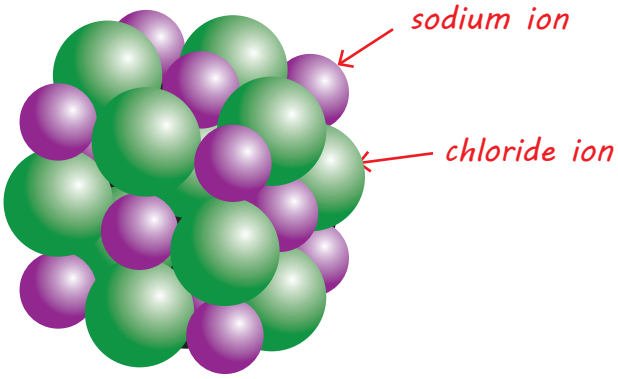


Only 3 of the 4 outer electrons in carbon are used for bonding in graphite. The other electron is delocalised and this explains why graphite conducts electricity.

SEE VIDEO 11

Ionically bonded compounds

When non-metals and metals bond, they tend to form ionic bonds. Ionic compounds have a giant ionic lattice structure. For example, sodium chloride's structure is made up of positively charged sodium ions which are surrounded by negatively charged chloride ions in a continuous lattice.



The ionic bonds within the giant ionic lattice extend throughout the entire lattice. Each sodium ion is surrounded by chloride ions and is bonded equally to all of them. Similarly each chloride ion is surrounded by sodium ions and is ionically bonded to them.

All ionic compounds have giant ionic lattice structures and as a result, they all have high melting and high boiling points since it is the strong ionic bonds that need to be overcome to melt and boil them.

Ionic compounds and conduction of electricity

In a solid state, ionic compounds do not conduct electricity as the ions are locked in place with strong ionic bonds.

However, in a molten state, or when dissolved, the ions become free to move and carry the charge and so do conduct electricity. Note with ionic compounds, it the ions that carry the charge and **NOT** delocalised electons.

SEE VIDEO 11

Metallic bonding

Metals form metallic bonds which are formed from positive metal ions and the outer electrons which are delocalised. Metals form giant metallic lattices, which like all giant structures, have high melting and high boiling points since the strong metallic bonds need to be overcome to melt them.

Metals conduct electricity as the delocalised electrons are free to move and carry the charge.

Summary

There are 3 types of chemical bond and 4 different structures:

Covalent
1. Simple molecular structures
2. Giant covalent lattice structures

Ionic
Giant ionic lattice structures

Metallic
Giant metallic lattice structures

All giant structures have high melting and boiling points because strong chemical bonds need to be overcome to melt and boil them

Have low melting and boiling points because weak intermolecular forces need to be overcome to melt and boil them

Generally do not conduct electricity. Graphite conducts electricity because it has delocalised electrons to carry the charge

ionic compounds only conduct electricity when molten or in aqueous solution as only then are the ions mobile and free to carry the charge

Metals conduct electricity because the electrons in their outer shells are delocalised and these can carry the charge

Three types of intermolecular force:

1. Hydrogen bonds - formed when hydrogen is bonded to nitrogen or oxygen to give a strong permanent dipole and a lone pair of electrons are available from an oxygen or nitrogen to make a hydrogen bond between the lone pair and the $\partial+$ on the hydrogen

2. Permanent dipole-dipole forces which occur due to a difference in electronegativity between bonded atoms (providing molecule is not symmetrical)

3. van der Waals' forces that occur in non-polar molecules due to temporary dipoles that induce dipoles in neighbouring molecules giving rise to a weak attractive force. van der Waals' forces increase with increased number of electrons