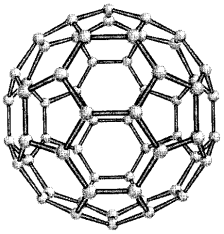
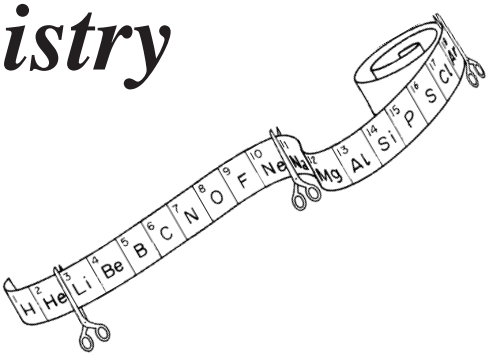
Higher Chemistry





Unit 1:

Periodicity, Polarity & Properties

Student:

Lesson	*Activities*			*Done*	*Checked*
1.1 ***Metallic Elements***	1.	Metallic Lattice			
	2.	Physical Properties			
	3.	Boiling Points			
1.2 ***Monatomic Elements***	1.	Very Weak Attractions			
	2.	Noble Gas Series			
		Check Test 1a	Score: / ?		
		Home Practice 1a	Score: / 20		
1.3 ***Molecular Elements***	1.	Diatomic Elements			
	2.	Larger CovalentMolecules			
1.4 ***Covalent Network Elements***	1.	Diamond Structure			
	2.	Other Networks			
	3.	Graphite Structure			
		Check Test 1b	Score: / ?		
		Home Practice 1b	Score: / 20		
1.5 ***Atomic Size***	1.	Atomic Radius			
	2.	Atomic Size and the Periodic Table			
	3.	Atomic Size and Density			
		Check Test 1c i	Score: / ?		
1.6 ***Ionisation Energies***	1.	First Ionisation Energy			
	2.	Across aPeriod and Down a Group			
	3.	Second Ionisation Energy			
		Check Test 1c ii	Score: / ?		
		Home Practice 1c	Score: / 20		
1.7 ***Electronegativity***	1.	How Strong an Attraction ?			
		Check Test 1d	Score: / ?		
1.8 ***Molecular Compounds***	1.	Bonding and Properties			
	2.	Effect of Increasing Atomic Size			
	3.	Effect of Increasing the Number of Atoms			
		Check Test 1e	Score: / ?		
		Home Practice 1e	Score: / 20		
Consolidation Work		Consolidation A	Score: / 10		
		Consolidation B	Score: / 10		
		Consolidation C	Score: / 10		
		Consolidation D	Score: / 10		
End-of-Section Assessment	Score: %		Grade:		

Lesson	Activities			Done	Checked
1.9 ***Polar Covalent Bonds***	1.	Unequal Sharing			
	2.				
	3.				
1.10 ***Polar - Polar Attractions***	1.	Shape & Polarity			
	2.	Solubilities			
	3.	Comparing Boiling Points			
	4.	Polar Molecules			
		Check Test 1f	Score: / ?		
		Home Practice 1f	Score: / 17		
1.11 ***Hydrogen Bonding***	1.	Hydrogen Bonds			
	2.	Boiling Points			
	3.	Viscosity & Bonding			
		Check Test 1g i	Score: / ?		
1.12 ***Water***	1.	Boiling Point			
	2.	Structure of Ice			
		Check Test 1g ii	Score: / ?		
		Home Practice 1g	Score: / 20		
1.13 ***Ionic Compounds***	1.	Atomic Ions			
	2.	Simple Ionic Structures			
		Check Test 1h	Score: / ?		
		Home Practice 1h	Score: / 10		
1.14 ***Dissolving***	1.	Dissolving In Water			
1.15 ***Covalent Network Compounds***	1.	Silicon Carbide			
	2.	Silicon Dioxide			
	3.	Properties & Uses			
		Check Test 1i	Score: / ?		
Consolidation Work		Consolidation A	Score: / 10		
		Consolidation B	Score: / 10		
		Consolidation C	Score: / 10		
		Consolidation D	Score: / 10		

End-of-Section Assessment Score: % Grade:

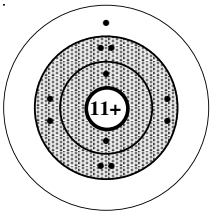
1.17 ***Redox Agents***	1.	Definitions			
	2.	Oxidising & Reducing Agents			
	3.	Donating & Accepting Electrons			
	4.	Electrochemical Series			
		Check Test 1j i	Score: / ?		
1.18 ***Redox Equations***	1.	Combining Simple Ion-Electron Equations			
	2.	More Complex Examples			
		Check Test 1j ii	Score: / ?		
1.19 ***Redox Titration***	1.	The Redox Reaction			
	2.	The Titration			
	3.	The Calculation			
		Check Test 1j iii	Score: / ?		
		Home Practice 1j	Score: / 30		
Consolidation Work		Consolidation A	Score: / 10		
		Consolidation B	Score: / 10		
		Consolidation C	Score: / 10		
		Consolidation D	Score: / 10		

End-of-Section Assessment Score: % Grade:

1.1 Metallic Elements

*This first topic looks at **bonding** in metals and how this helps explain some of the typical **properties** of a metal.*

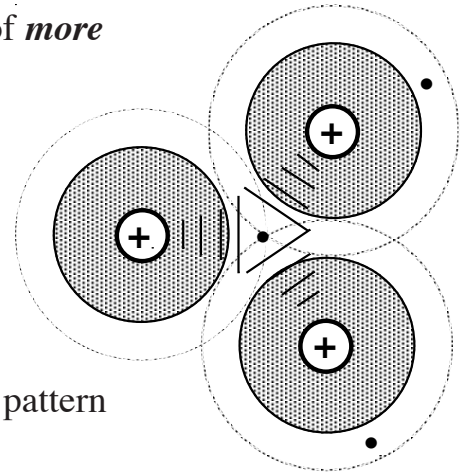
Metallic Bond



The *inner **elec*** of a metal atom are tightly held by the ***nuc***. They are not able to escape from the atom. They are said to be ***loc***.

The *outermost **elec*** are also strongly attracted by the ***nuc***, but they are ***free to move*** from atom to atom. They are said to be ***deloc***.

Each delocalised electron can be ***attr*** to the ***nuc*** of ***more than one*** neighbouring atom.



The ***metallic bond*** is the electrostatic ***attr*** between the ***neg*** charged outermost electrons and the ***pos*** charged nuclei of neighbouring atoms.

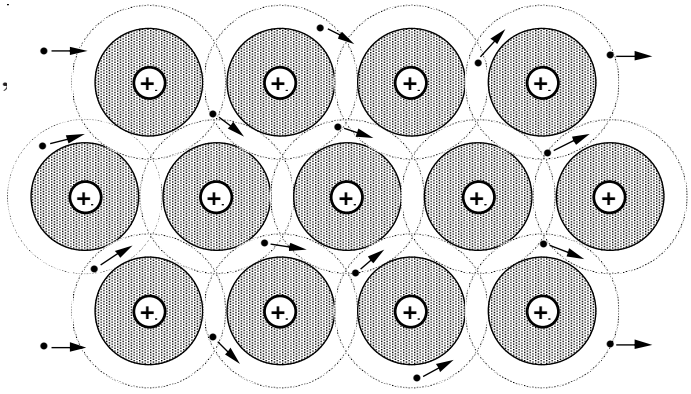
The ***metallic bond*** is usually ***strong*** and ***acts in all directions*** around each atom.

In a solid metal, the atoms are ***arra*** neatly in a simple pattern known as a ***lattice***

Physical Properties

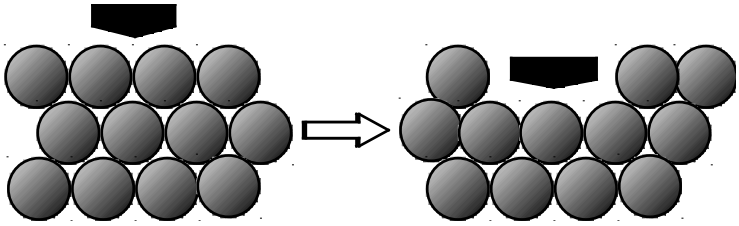
The Physical properties of a substance are those properties which do not involve chemical reactions.

Metals make ***good cond***.



Since the outermost electrons are ***deloc*** it is easy to make them all move in the ***same*** direction.

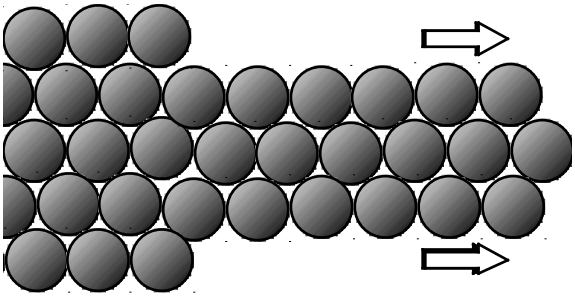
Once a ***volt*** is applied across the metal, the ***elec*** will move away from the ***neg*** electrode towards the ***pos*** electrode.



*The shapes of metals can be altered by beating them (horseshoes) or by rolling into sheets (tin foil). This property is called **malleability**.*

Met bonding is very 'fluid'. This makes it easy for metal ***ato*** to be slid around each other and yet remain strongly ***bon*** together. Metals are described as ***malleable***.

Metals can be drawn into thin wires by stretching. This property is called ***ductility****.*

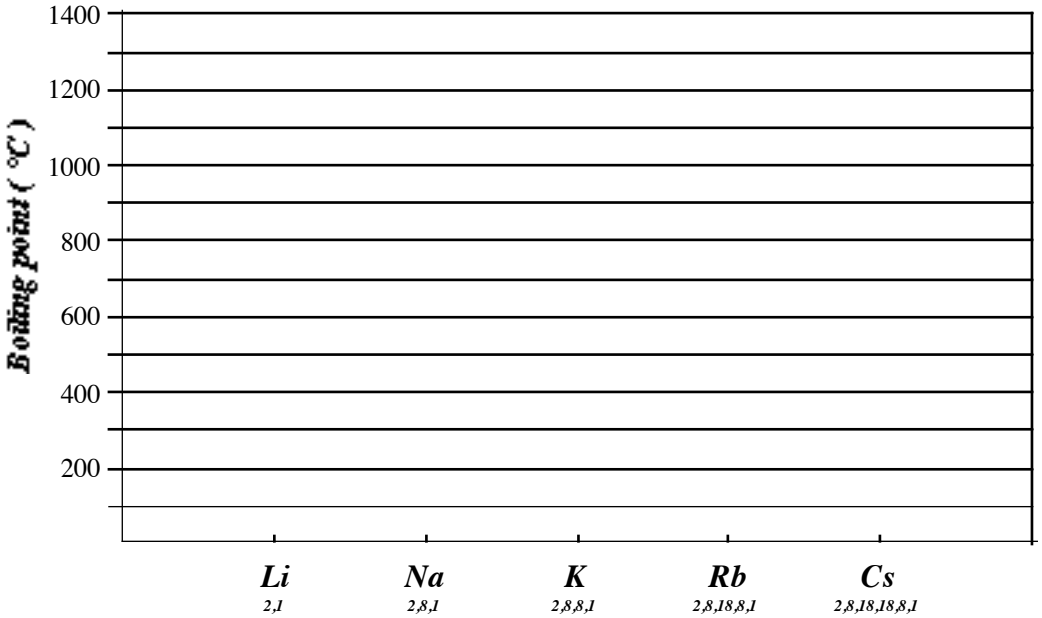


Again, it is because it is easy for metal atoms to ***sli*** around each other while staying strongly ***bon*** to each other. Metals are described as ***ductile***.

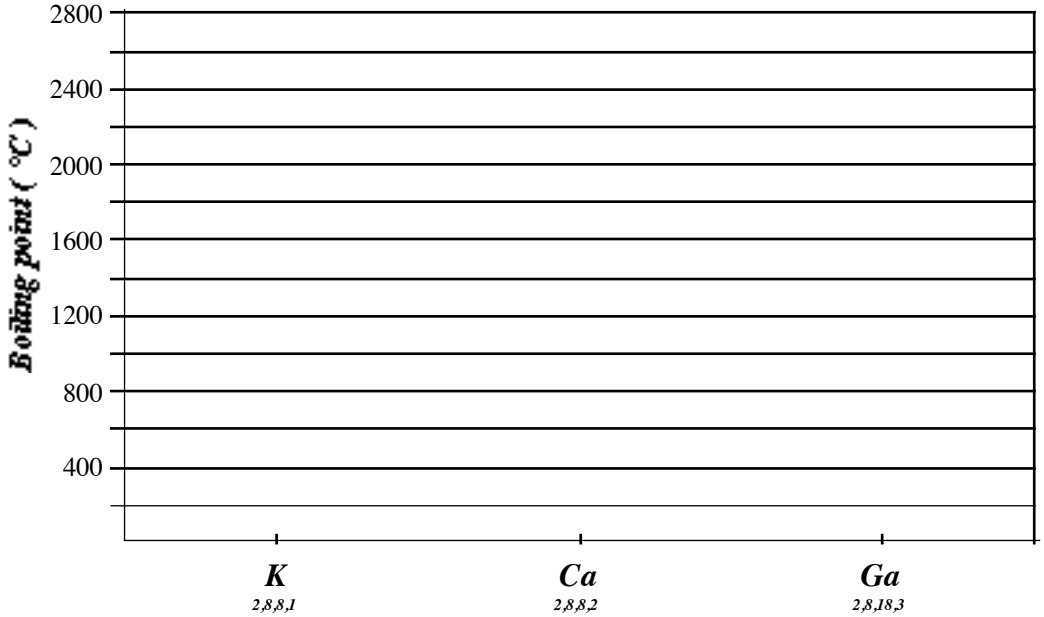
Boiling Points

To boil a metal, and allow its atoms to move around freely as a gas, the metallic bond must be overcome.

The Boiling Points of metals are a guide to the relative strengths of the metallic bonds holding the atoms together.



As you go ***do*** Group 1, the outermost ***elec*** is further and further away from the ***nuc***. The strength of ***attr*** between electrons and nuclei will, therefore, ***decr***. This makes it easier for the metal atoms to move further apart as shown by the ***decr*** in Boiling points.



As you go ***across*** the fourth period, the outermost ***elec*** are always to be found in the ***fou*** shell, at a similar ***dist*** from the ***nuc***. The ***number of electrons*** is, however, ***incr***.

Since the Boiling points are , the ***bond strength*** must ***inc*** as the ***num*** of outermost electrons in a metal ***incr***.

The strength of the Metallic bond

decreases as the outermost electrons become further from the nucleus

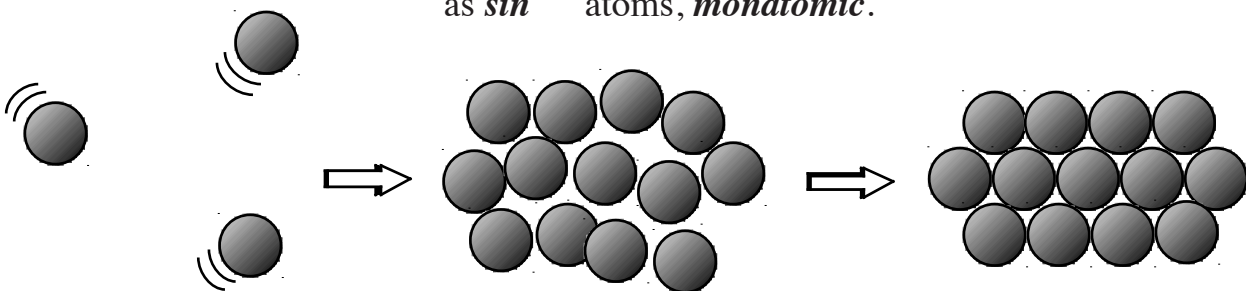
increases as the number of outermost electrons increase

1.2 Monatomic Elements

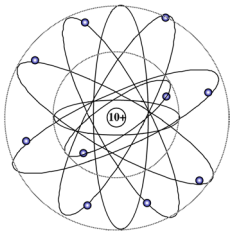
This topic is about the monatomic elements and the very weak type of bonding force that exists between atoms of these elements.

Very Weak Attractions

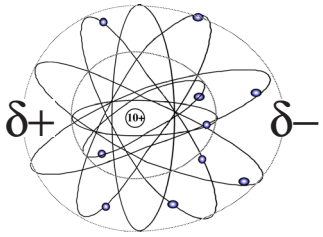
The ***No*** gases are the only elements normally found as ***sin*** atoms, ***monatomic***.



At ***extremely low temp*** elements like Neon become ***liq*** or even ***sol*** ,like every other substance. At these low temperatures, there must be some ***weak for*** holding these atoms together. Very little ***ene*** is needed to overcome these weak forces, as the low Melting and Boiling Points for the Noble gases demonstrate. *These very weak bonding forces are called* ***London Dispersion Forces***.

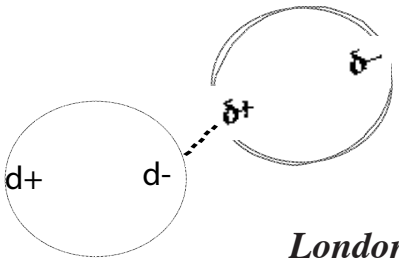


Normally, the ***elec*** are *spread evenly* within the set of electron clouds (***orb***) that make up an electron ***sh***. The centre of ***neg*** charge and the centre of ***pos*** charge will coincide at the centre of the atom.



Uneven distribution of the ***elec*** causes the electron cloud to 'wobble' and the centre of ***neg*** charge no longer coincides with the ***pos*** charge. One side of the atom is temporarily slightly negative(δ–), while the other side becomes temporarily positive (δ+). *We call this a temporary* ***dipole***.

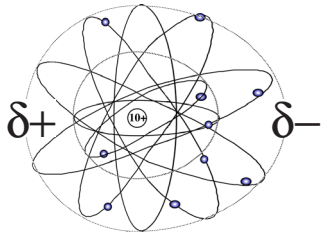
The ***dip*** is only ***temp*** as electron repulsions quickly act to ensure a more *even distribution*. However, the brief ***wob*** on one atom can affect the ***elec*** clouds on any nearby atoms.



This causes, or ***induces***, a temporary ***dip*** in the nearby atom. Short lived ***attr*** will then exist between the slightly ***neg*** (δ–) side of one atom and the slightly ***pos*** (δ+) side of the other atom.

London Dispersion Forces *are the shortlived attractions that exist between a slight charge on one atom (**temporary dipole**) and the opposite charge set up on a nearby atom (**induced dipole**).*

Noble Gas Series

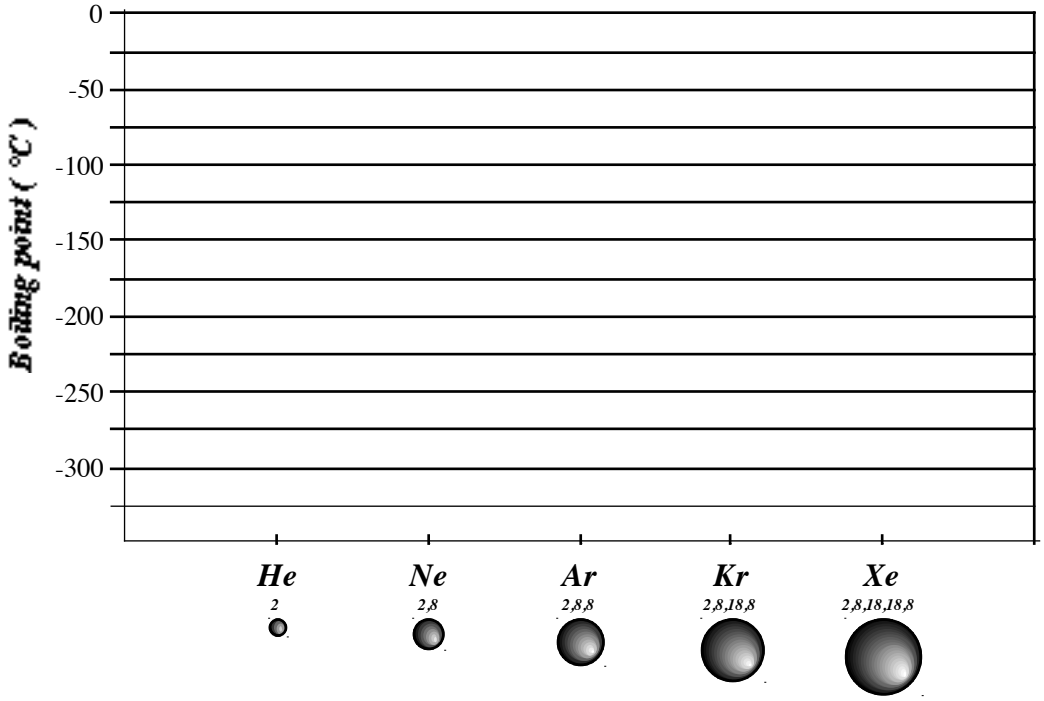


Every atom has a ***nuc*** and clouds of ***elec*** around it. So all atoms have *London Dispersion forces* acting between them (as long as they are close enough to each other).

They are such ***sm*** forces that they can usually be ignored, especially if other, much ***stro*** , bonding forces are present, e.g. ***met*** bonds.

*However, for the noble gases, London Dispersion forces are the **only** type of bonding force present.*

So the melting points and boiling points of the noble gases depend entirely on the strength of the London Dispersion forces between their atoms.



The ***lar*** an atom is, the ***fur*** out its outermost ***elec*** , the more likely it is that the electron clouds will be able to 'wobble' and set up a ***temp*** ***dipole***.

The ***lar*** the dipole set up, the ***stro*** the ***London Dispersion*** forces that will exist. As the ***attr*** between atoms increase, the ***ene*** needed to move the atoms further apart will ***inc*** as shown by the ***inc*** in Boiling points.

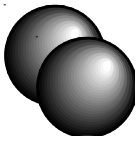
London Dispersion forces are very small, but they are larger between bigger atoms.

1.3 Molecular Elements

This topic is about the elements which exist normally as discrete molecules.

Diatomic elements

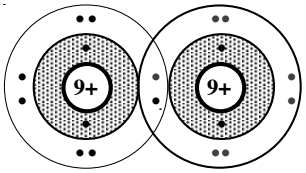
H_2 N_2 O_2 F_2 Cl_2 Br_2 I_2



Molecular elements are elements made up of individual molecules i.e. distinct groups of atoms joined together by ***covalent*** *bonds.*

The simplest molecular elements are ***diatomic*** *i.e. each molecule is made of just* ***two atoms****.*

Substances which are molecular have ***more than one type*** *of bonding present.*

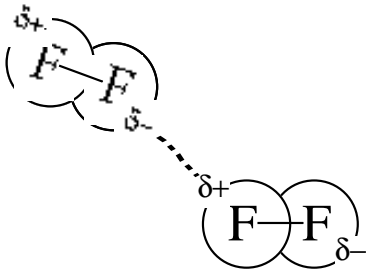


F — F

Within each molecule the atoms are ***strongly*** held together by a ***cov*** bond.

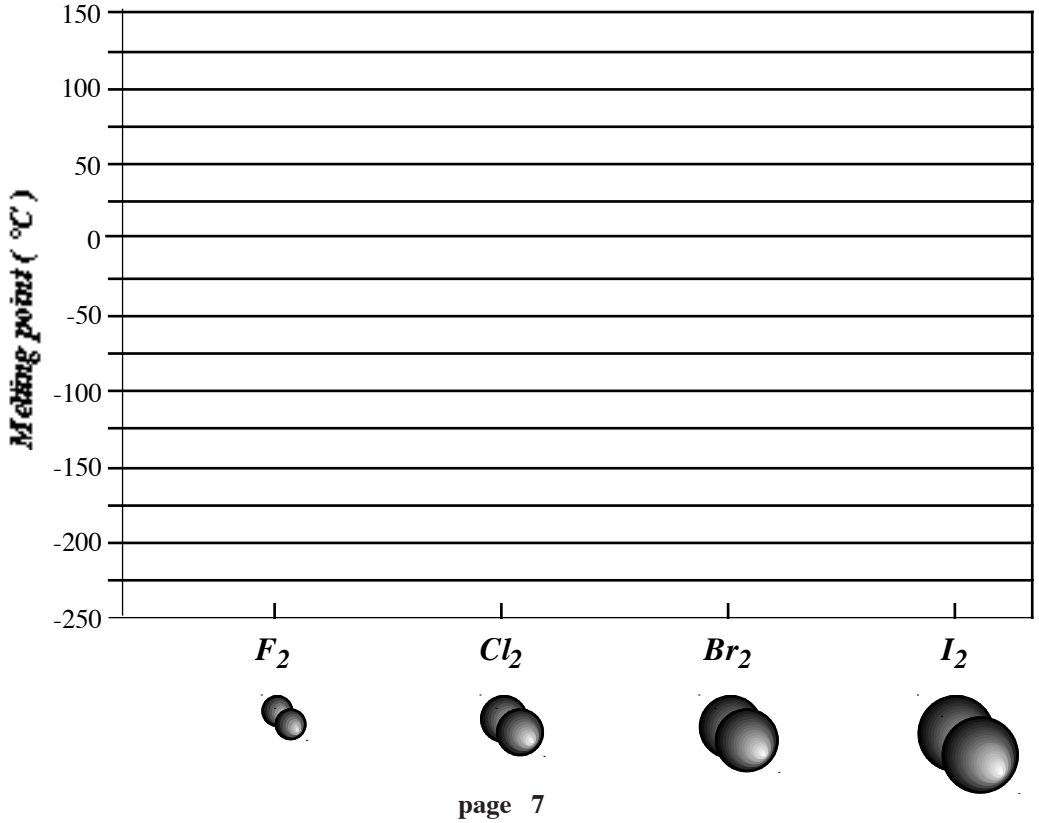
A ***covalent bond*** *is the electrostatic attraction between a* ***shared pair of electrons*** *and the nuclei of the two sharing atoms.*

Usually, the ***cov*** ***bond*** is represented by a line.



Between the ***mol*** of fluorine there are the same ***we*** forces that exist between atoms - ***London Dispersion forces***.

Though very weak, these ***intermolecular forces*** can still determine ***phy*** properties such as ***melt*** and ***boil*** points.



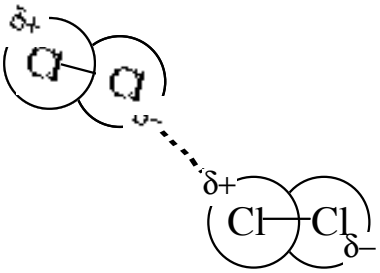
The ***lar*** a molecule is, the ***fur*** out the outermost ***elec*** , the more likely it is that the electron clouds will be able to 'wobble' and set up a ***temp*** ***dipole***.

The ***lar*** the dipole set up, the ***stro*** the ***London Dispersion*** forces that will exist.

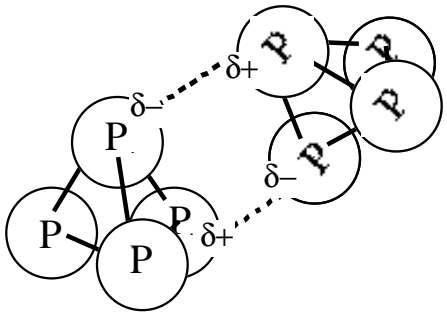
As the ***attr*** between molecules increase, the ***ene*** needed to move the molecules further apart will ***inc*** as shown by the ***inc*** in Melting points.

Larger Covalent Molecules

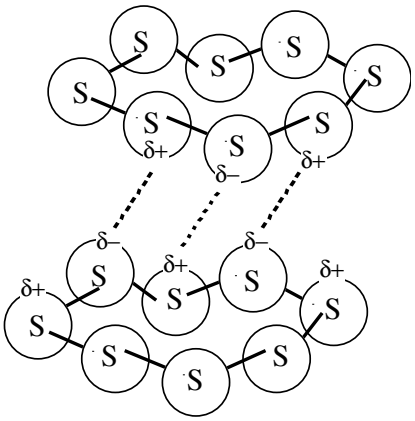
*The **size** of atoms has an effect on the strength of London Dispersion attractions between molecules. The **number** of atoms can also have an effect.*



Chlorine molecules have the formula =
and a melting point = °C.



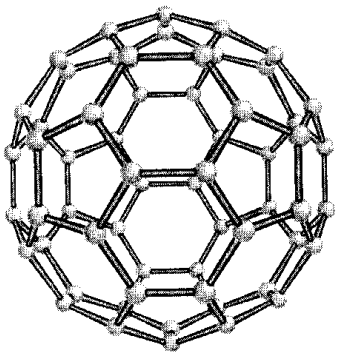
Phosphorus molecules have formula =
and a melting point = °C.



Sulphur molecules have formula = S_8
and a melting point = °C.

Chlorine, phosphorus and sulphur atoms are of a ***sim*** ***size*** so are likely to produce ***sim*** ***dipoles***, and ***sim*** ***London Dispersion attractions***. However, with ***more atoms*** in a molecule, a ***lar*** number of London Dispersion forces are possible.

In all these examples, it is important to remember that there is no need to break the very strong ***cov*** bonds ***within*** these molecules.



Larger molecules like ***full*** will have even ***high*** melting points, though, they are still relatively low compared to other substances which have ***stro*** bonds than London Dispersion forces to be broken.

*The **strength** of London Dispersion forces*

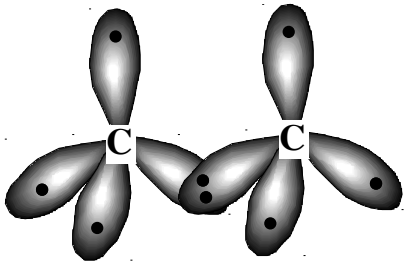
increases as the size of the atom increases

increases as the number of atoms in the molecule increases

1.4 Covalent Network Elements

This topic is about elements where the atoms are joined together by strong covalent bonds to form vast networks of atoms.

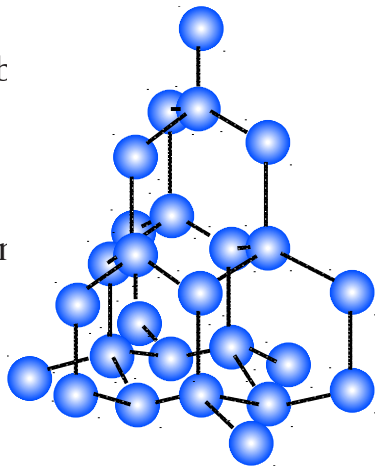
Diamond Structure



Car atoms have 4 ***unpaired elec*** which allows them to form a total of ***4 bonds***.

Each ***car*** can bond with another 3 carbons each, and so on until a huge ***net*** of ***cova*** bonded atoms has been formed.

The arrangement of carbon atoms can be described as ***tetrahedral*** because of the position of the 4 carbon atoms around a central carbon atom.



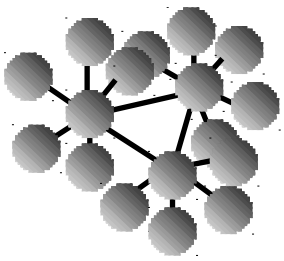
Rings of 6 carbon atoms, ***hexa*** , can also be seen.

Dia are not malleable - the carbon atoms can not slide around each other without breaking ***strong cov*** bonds.

Diamonds are strong and hard since many ***strong cov*** bonds would need to be overcome if a diamond is to be cut or broken.

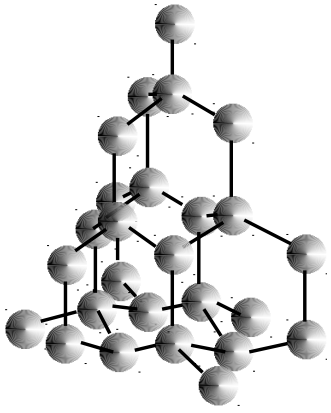
Other Networks

The only other ***cov*** ***net*** in the first 20 elements are ***boron*** and ***silicon*** (which has a diamond structure like carbon).



Not surprisingly they are all ***high mel*** ***point solids***.

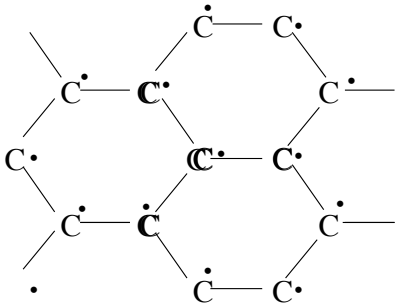
A lot of ***strong cov*** bonds have to be broken to melt these elements.

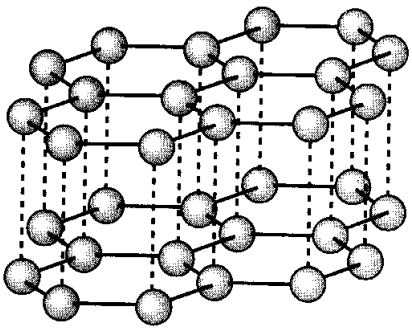


Melting point of ***boron*** = °C
Melting point of ***silicon*** = °C
Melting point of ***carbon*** = °C

Boron and silicon are also very ***hard*** substances, though not as hard as diamond, of course.

Graphite Structure





Unusually, ***car*** has two ***net*** structures - ***dia*** and ***gra***. Their ***properties*** are very different, and so must their ***structures*** be.

In ***gra***, each carbon atom is only bonded to 3 other carbon atoms by ***strong covalent*** bonds.

This forms flat ***sheets*** of 6 membered rings. The fourth outermost electrons are free to move (***delocalised***) between all the carbon atoms in a sheet or layer.

Between the layers are the much weaker ***London Dispersion forces*** that exist between all atoms.

The distance between the layers is ***gre*** than the the distance between carbon atoms within the same layer

We can use the ***structure*** of graphite to explain its ***properties***.

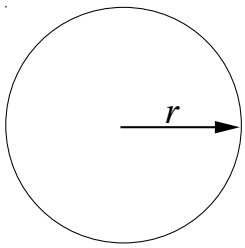
Conductivity ***gra*** is able to ***con*** *along layers* because it has some ***deloc*** electrons. (Conductivity between layers is not possble).

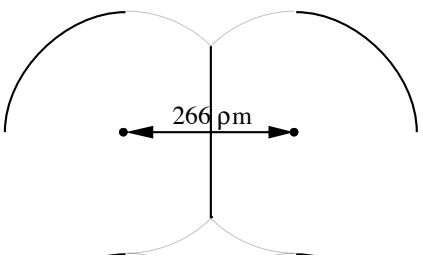
Lubrication powdered graphite makes an excellent solid lubricant because the ***weak London Dispersion forces*** allow the layers to ***slide*** over each other.

1.5 Atomic Size

This lesson looks more closely at the size of atoms and the way they vary within the Periodic Table.

Atomic Radius



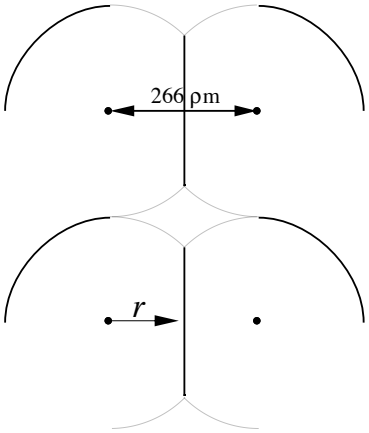


We are used to thinking of atoms as ***sphe***, and of drawing them as ***cir***, so the size of an atom could be defines by its ***rad*** the distance from the ***cen*** of the atom to its ***edge***.

Unfortunately, outside the ***nuc*** is mainly ***emp*** ***space*** and there is no obvious edge.

What can be detected (by, for example, X-rays) are the ***nuc*** of atoms. This allows the ***distance between the nuclei*** of neighbouring atoms to be measured, and this can be used to 'calculate' the ***atomic radius*** of individual atoms.

The radius can be calculated by assuming it is ***half the distance between the nuclei of two atoms***. However, there are two ways that this can be done.



Covalent Radius

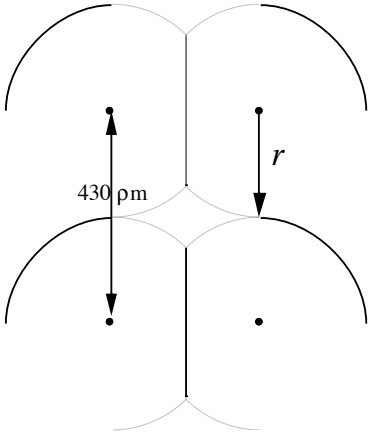
When two atoms form a ***cov*** ***bond*** they ***overlap***.

Half the distance between the two nuclei would give a value slightly ***sma*** than the true size of the atom.

In the case of iodine, I_2,

covalent radius = / 2 = ρm

(where ρm = picometre = 1 x 10^{-12} m)



Van der Waals Radius

Between some atoms and many ***mol*** , there exist weak *London Dispersion* forces which do not involve any overlap.

Half the distance between the two nuclei of atoms in neighbouring molecules would give a value ***closer*** to the true size of the atom.

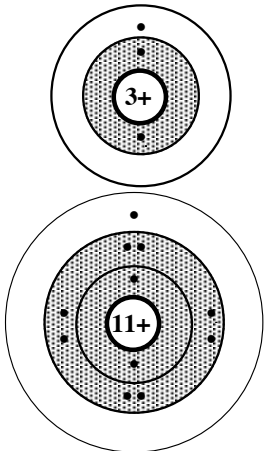
In the case of iodine, I_2,

van der waals radius = / 2 = ρm

(where ρm = picometre = 1 x 10^{-12} m)

Unfortunately, there are very few atoms whose size could be determined using elements, in the solid state, held together by just London Dispersion forces. Therefore, ***Covalent Radii*** are the values most often quoted in Data Books and these are the values you will use most often.

Atomic Size & the Periodic Table



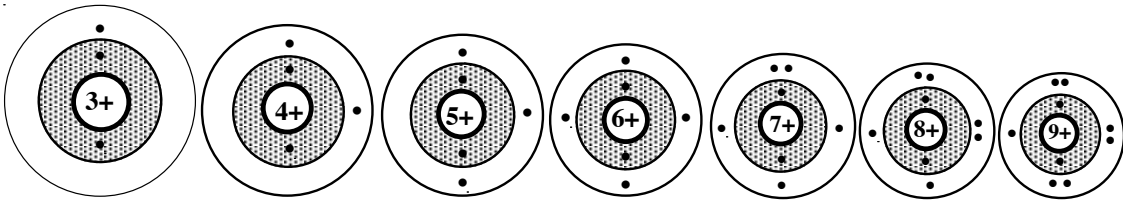
As you go do a Group (column) there is an inc in the size of the atoms,(covalent radii).

Li	2,1	covalent radius =	ρm
Na	2,8,1	covalent radius =	ρm
K	2,8,8,1	covalent radius =	ρm

This is not surprising, as each atom has an ***extra energy level***, (***sh***) further out from the ***nuc*** .

More surprisingly is that:-

as you go acr a Period (row) there is a dec in the size of atoms, (covalent radii).



The ***num*** of energy levels is the ***same***, but the increasing size of the ***nuclear charge*** causes increased ***attr*** between the outermost ***ele*** and the ***nuc***. This causes the outer shell to be ***pulled clo*** and closer to the nucleus, and the size of the atoms decrease.

Atomic Size & Density

*Density is **mass per volume**.*

So, mass of atom ↑ Density ↑
and, size of atom ↑ Density ↓

As you go do a Group (column) there is an inc in the density

The ***mass*** of the atoms will be , and the ***size*** of the atoms will be , but the ***inc*** in mass must be the larger effect.

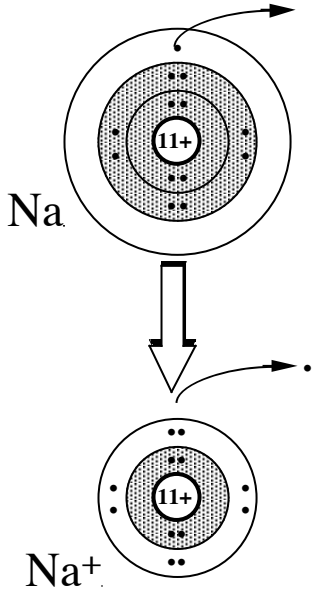
As you go acr a Period (row) there is an inc in the density

The ***mass*** of the atoms will be , and the ***size*** of the atoms will be , so both factors will lead to an ***inc*** in density. (Changes in structure (***net →mol***) and state (***so →ga***) explain the low ***den*** of many of the elements on the ***ri*** of the Periodic Table.)

1.6 Ionisation Energies

*This lesson is about **ionisation energy** and how it is related to an element's position in the periodic table.*

First Ionisation Energy



*The **first ionisation energy** is the **energy required** to remove **one electron** from every atom in **a mole** of **free** atoms.*

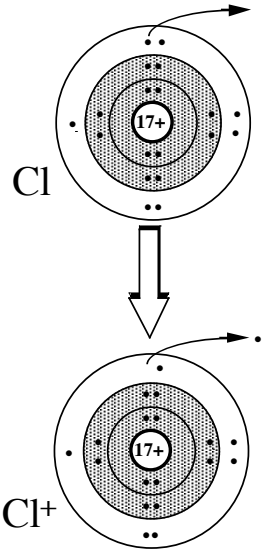
*It is measured in kJ mol^{-1}, i.e. **kilojoules per mole***

$$Na_{(g)} \rightarrow Na^{+}_{(g)} + e^{-} \quad 502 \text{ kJ mol}^{-1}$$

The sodium atoms have to be ***fr***, i.e. <u>not</u> ***bon*** to any other atom. This can only be done by ***hea*** the sodium until it is a ***g***.

The ***elec*** lost appears as a ***prod***, on the ***ri*** hand side of the ***equ***.

All atoms, including non-metal atoms, can be forced to lose an ***elec*** . This results in the formation of a ***pos*** ***ion***, wheras in normal reactions, non-metals will usually ***ga*** ***electrons*** and form ***neg*** ***ions***.

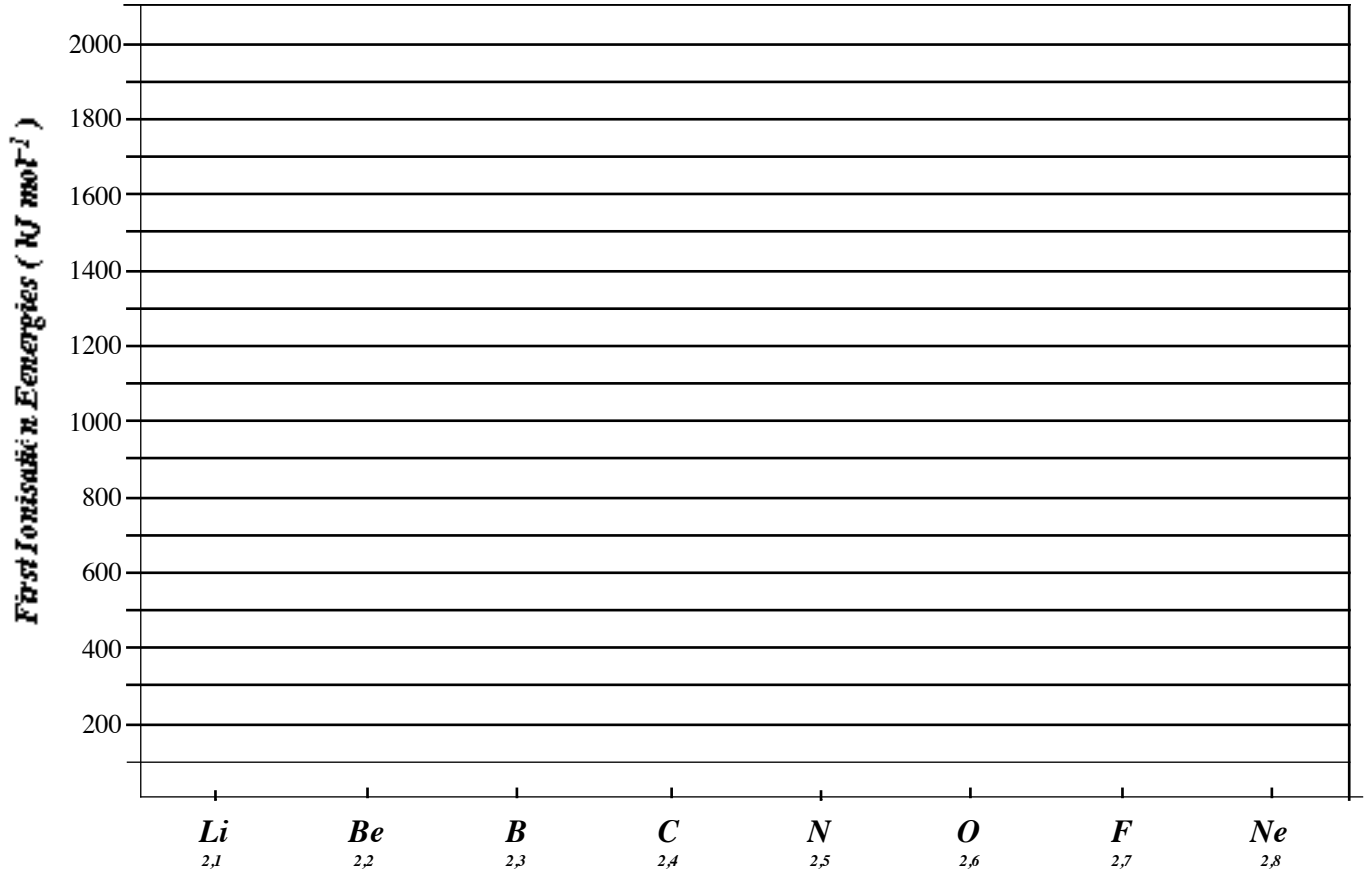


$$Cl_{(g)} \rightarrow Cl^{+}_{(g)} + e^{-} \quad 1260 \text{ kJ mol}^{-1}$$

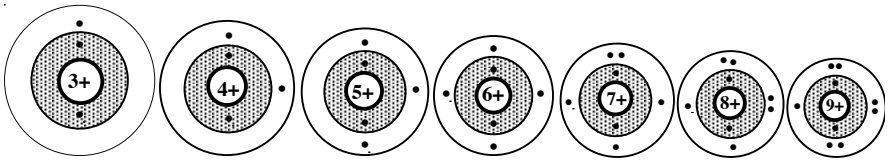
The chlorine atoms have to be ***fr*** , i.e. <u>not</u> ***bon*** to any other atom. This can only be done by ***brea*** the chlorine ***mol*** apart.

The ***elec*** lost appears as a ***pro*** , on the ***ri*** hand side of the ***equ*** .

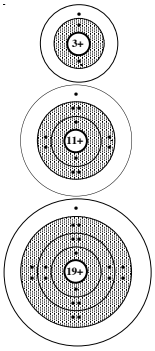
Across a Period & Down a Group



As you go across a per ***the general trend is for the first ionisation energy to inc*** . *(The fluctuations from the general trend require a more advanced explanation that goes beyond this level of chemistry)*



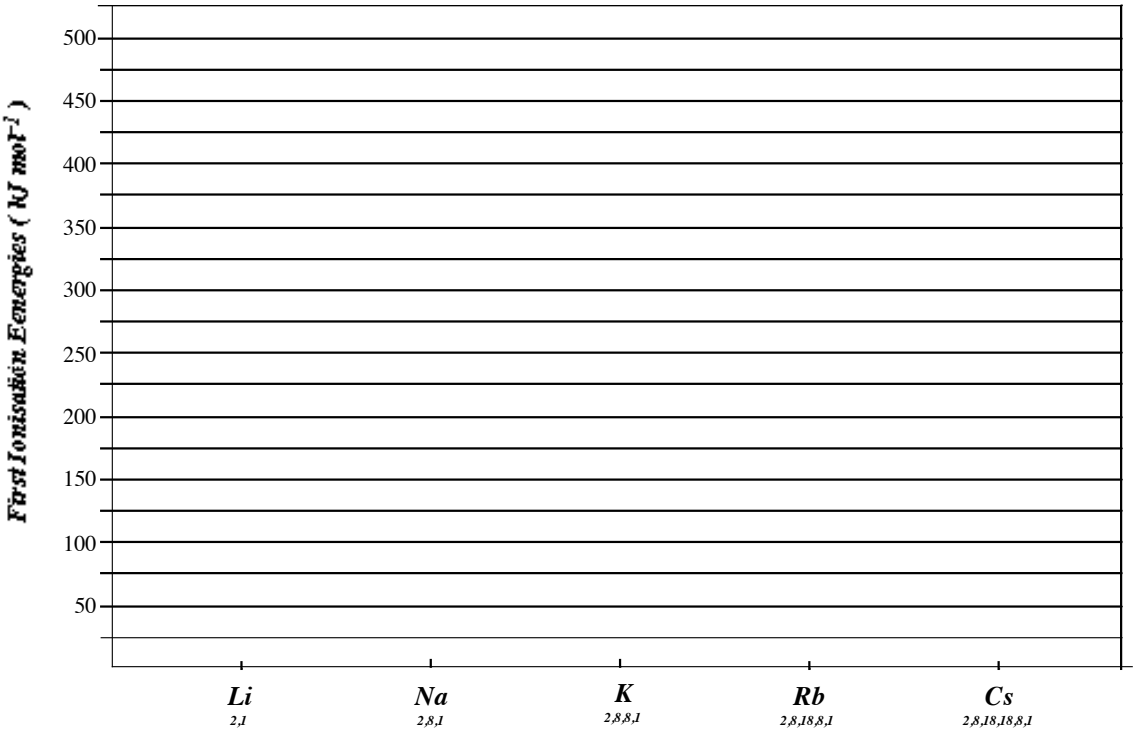
As mentioned earlier, (p11), the ***num*** of energy levels is the ***same***, but the increasing size of the ***nuc*** ***charge*** causes increased ***attr*** between the outermost ***ele*** and the ***nuc*** . This makes it harder to ***rem*** an electron from the ***out*** shell.



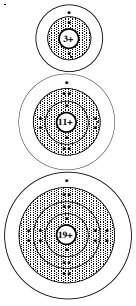
*On going **down a group**, a new energy level (shell) is added each time. Each outer layer is further from the nucleus. This could make it **easier** to remove an electron.*

*The outer shell is screened by more and more shells closer to the nucleus. This could make it **easier** to remove an electron.*

*On the other hand, the nucleur charge is also increasing. This could make it **harder** to remove an electron.*



As you go down a gr the general trend is for the first ionisation energy to de-

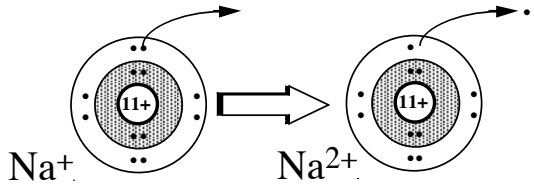


On going ***down a group***, an ***extra shell*** is added each time. Each shell is ***fur*** from the ***nuc***. This makes it ***eas*** to remove an ***elec***.

The outer shell is ***screened*** by more and more shells ***clo*** to the ***nuc***. This makes it ***easier*** to remove an ***elec***.

Even though the nucleur charge is increasing, the other two factors mean that it still gets ***easier to remove an electron as you go down a group.***

Second Ionisation Energy



*The **second ionisation energy** is the **energy required** to remove **a second electron** from every 'atom' in **a mole** of **free** 'atoms'.*

$(Na_{(g)} \quad Na^{+}_{(g)} + e^{-} \quad 502 \text{ kJ mol}^{-1})$

$$Na^{+}_{(g)} \rightarrow Na^{2+}_{(g)} + e^{-} \quad 4560 \text{ kJ mol}^{-1}$$

Each ionisation will require ***mo energy*** than the previous one; a ***dramatic inc*** occurs whenever an electron has to be removed from the ***next shell*** closer to the nucleus.

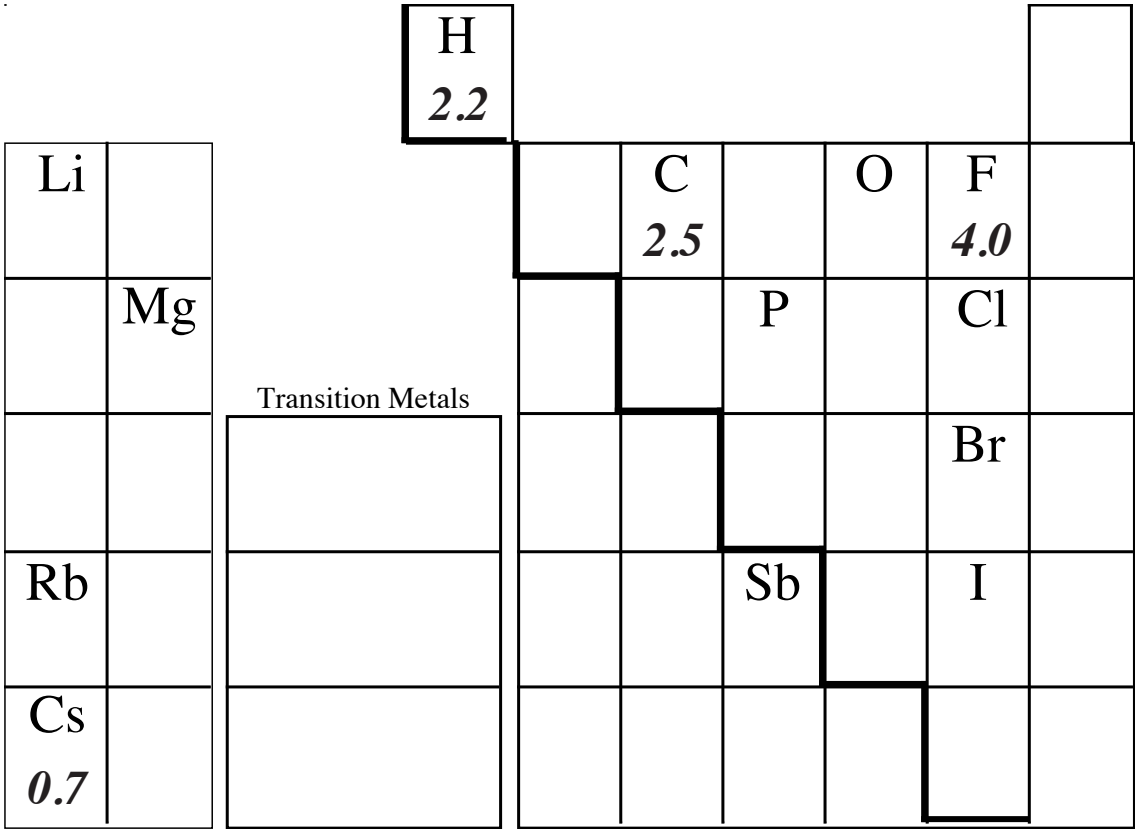
1.7 Electronegativity

This lesson is about ***the unequal sharing of electrons in covalent bonds***

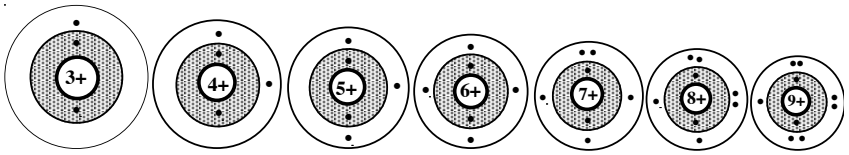
How Strong An Attraction ?

This activity is about how different atoms have different strengths of attraction for the shared electrons in a covalent bond.

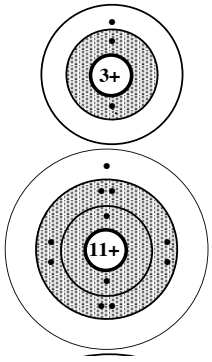
Electronegativity *is a measure of the attraction an atom has for the shared electrons in a covalent bond.*



As you go across a per ***the trend is for the electronegativity to inc*** .



The atoms are ***'shri*** ***'*** in size bringing any ***sha*** ***electrons clo*** ***and clo*** to the ***nuc*** .

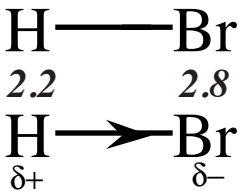


As you go do ***a Group there is a dec*** ***in the electronegativity***

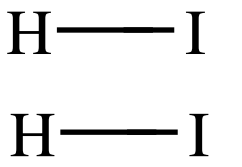
This is not surprising, as each atom has an ***extra energy level***, (***sh***) further out from the ***nuc*** .

The ***fur*** from the ***nuc*** the ***sha*** ***pair*** of ***elec*** are, the ***wea*** the ***attr*** .

Electronegativity values can be used to determine the ***polarity*** of a ***cov bond***. The atom with the ***hi electronegativity value*** will have a ***str attraction*** for the ***sha pair of elec***. This atom will end up ***sli neg***, while the other atom is ***sli pos***.



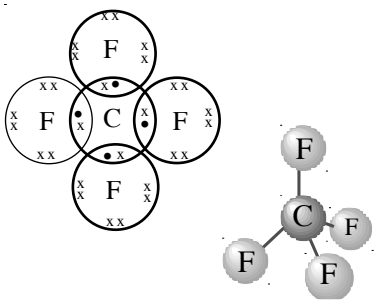
The ***rela polarity*** of two bonds can also be compared using *electronegativity values*. Subtracting one value from the other provides a rough indication. The value for the H—Br bond is ***0.6*** (2.8 - 2.2). The value for the H—I bond is (-), showing that the bond is ***more polar***



1.8 Molecular Compounds

This topic is about molecular compounds and some of the bonding forces that are present in their structures.

Bonding and Properties

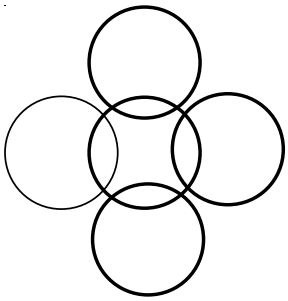
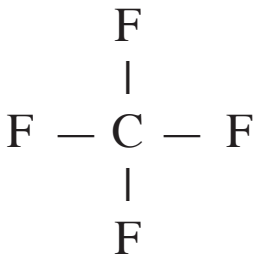
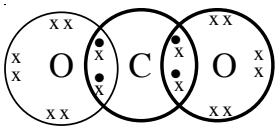
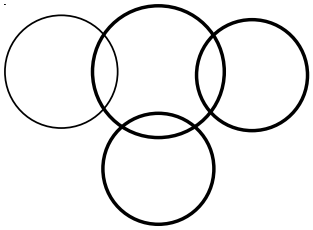
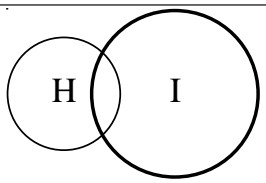
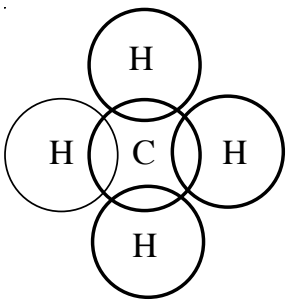


This activity considers the bonding present in molecular compounds and how this affects properties.

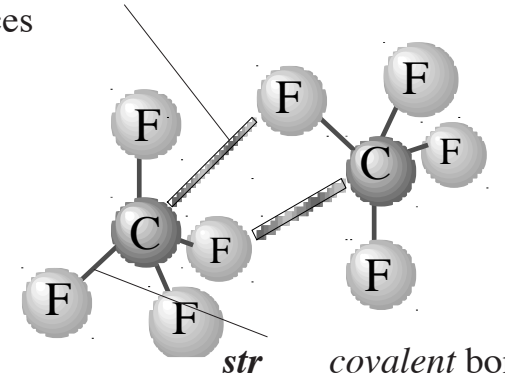
Carbon tetrafluoride is a ***cov mol comp*** .

By ***sha electrons***, both the ***car*** and the ***fluo*** atoms achieve the ***extra stab*** of a ***full out shell***.

The ***cov bonds***, however, go no further than the 5 atoms making up each ***dis molecule***.

"Dot-and-cross" diagram	*Structural formula*	*Molecular formula and name*
		*carbon tetrafluoride*
		
		*phosphorus trichloride*
		
		

we *London Dispersion* forces



str *covalent* bonds

The *intramolecular* forces (intra- = within), that join the ***fluo*** atoms to the ***car*** atom, are ***strong cov*** ***bonds***.

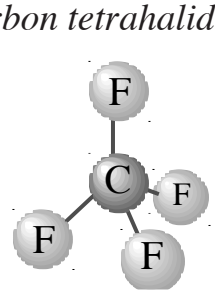
Once ***coo*** enough to *solidify*, ***weak London Dispersion*** forces will exist ***bet*** molecules (*intermolecular* forces).

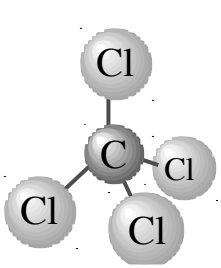
It is only these ***we*** forces that are broken when ***mol*** ***compounds*** like CF_4 melt, hence their ***very l*** ***mel*** ***points***.

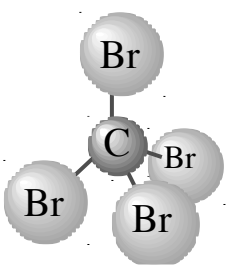
Even when solid, substances like CF_4 are ***not str*** - again, this is because only the ***weak London Dispersion*** forces need to be broken.

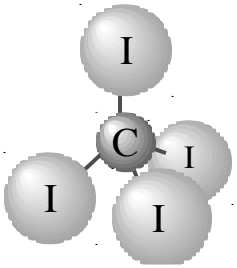
Effect of Increasing Atomic Size

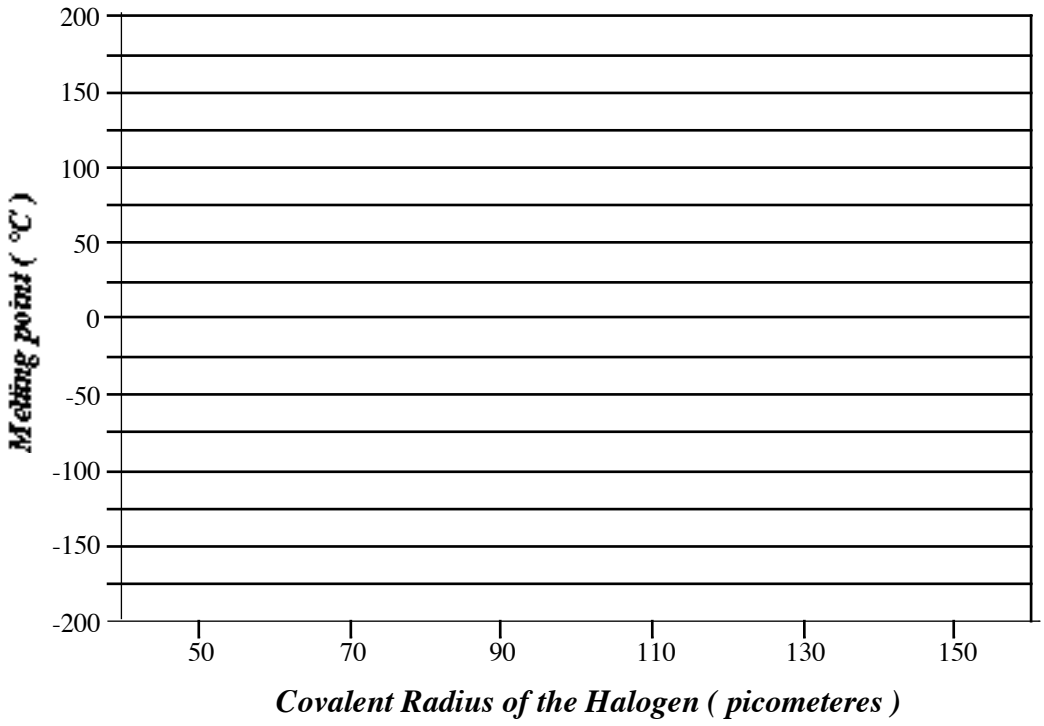
This activity looks at the effect that increasing atomic size has on the melting points of the carbon tetrahalides.







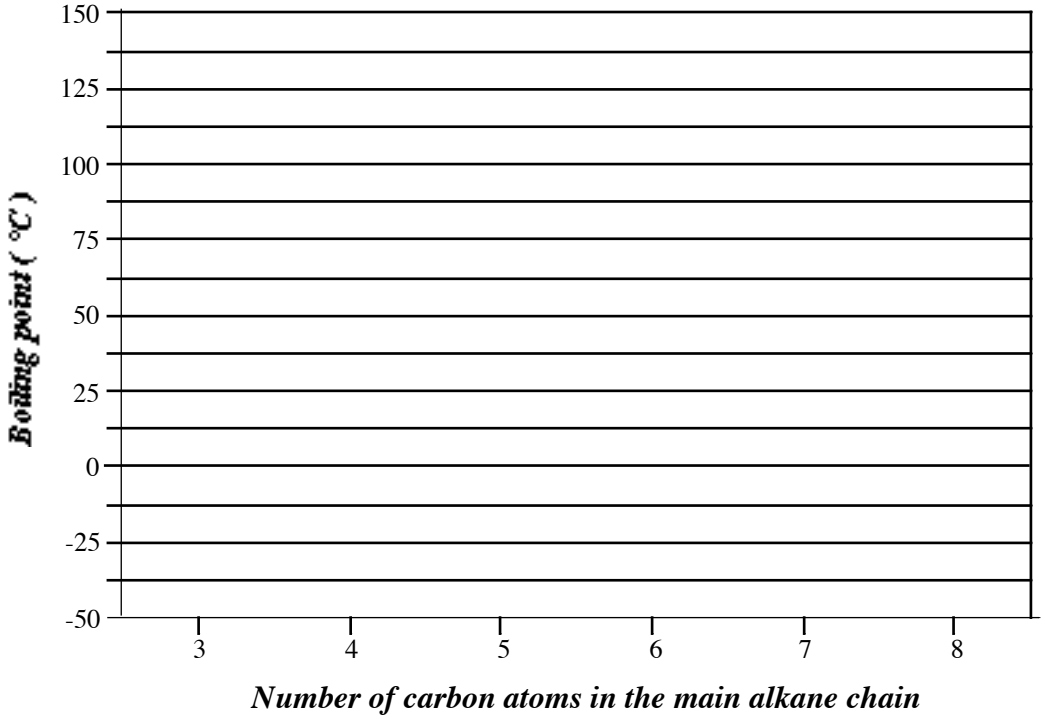




As the ***si*** of the ***hal*** ***atom*** in these compounds , the ***melt*** ***point*** .
This is because the ***elec*** in ***lar*** ***atoms*** are ***more lik*** to 'wobble' leading to ***stro*** *van der Waal's* forces ***bet*** molecules.

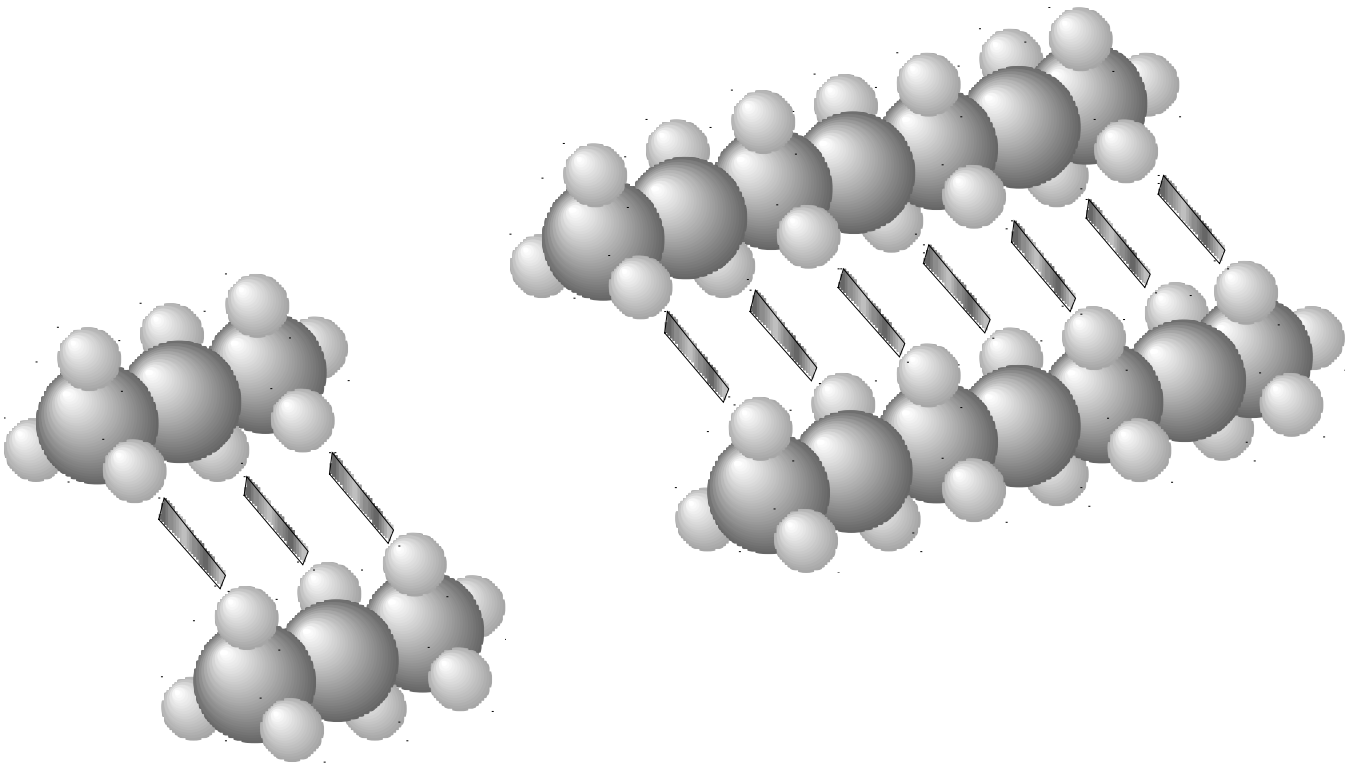
Effect of Increasing the Number of Atoms

This activity looks at the effect that increasing the number of atoms has on the boiling points of the alkanes.



As the ***num*** of ***car*** atoms in the molecules , the ***melt*** ***point*** .

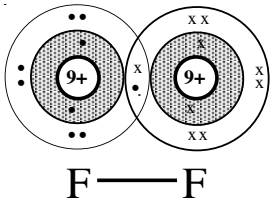
This is because ***lar*** ***num*** of atoms produce ***more temp*** ***dip*** and, therefore, ***mo*** *van der Waal's forces* ***bet*** molecules.

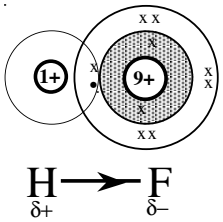


1.9 Polar Covalent Bonds

This lesson is about ***the unequal sharing of electrons in covalent bonds***

Unequal Sharing





This activity explores the effect that unequal sharing of electrons has on the atoms with a covalent bond.

In a ***mol*** like F_2, both atoms are ***exactly the sa*** . They have ***equ attraction*** for the *bonding pair* of ***elec*** .

The ***elec*** are ***equ sha*** . This is a ***pure cov*** bond.

A ***fluo*** atom has a ***stron attraction*** for ***elec*** than a ***hydr*** atom, The *bonding pair* is ***pulled clo*** to the ***fluo*** .

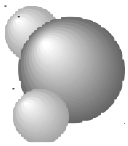
The *fluorine* becomes ***sligh neg*** (δ–), while the *hydrogen* becomes ***sligh pos*** (δ+). A ***perm dipole*** is set up. This is a ***polar cov*** bond.

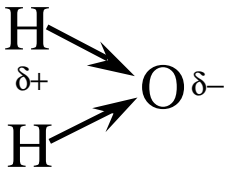
1.10 Polar-Polar Attractions

This lessonlooks in more detail at how the shape of a molecule can affect its overall polarity, and how this, in turn, can affect the forces acting between molecules.

Shape & Polarity

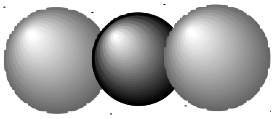
This activity considers how the shape of a molecule can affect its overall polarity.

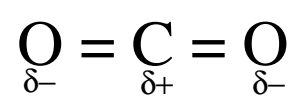




The O—H bond is a ***po*** ***cov*** ***bond***, with the ***slightly neg*** and the ***slightly pos*** .

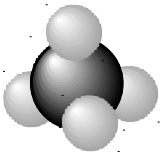
Its ***shape*** results in a ***neg*** ***end*** and a ***pos*** ***end*** so it is a ***po*** ***molecule***.

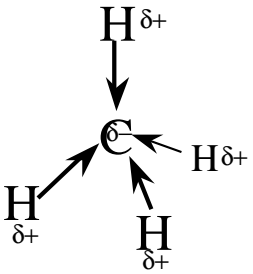




The C = O bond is a ***po*** ***cov*** ***bond***, with the ***slightly neg*** and the ***slightly pos*** .

Its ***shape*** <u>*doesn't*</u> produce a ***neg*** ***end*** and a ***pos*** ***end***, it is <u>*not*</u> a ***po*** ***molecule***.

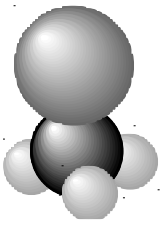


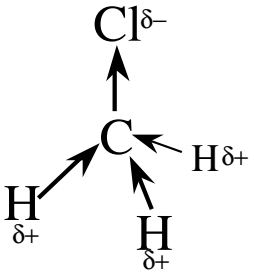


The C — H bond is a ***po*** ***cov*** ***bond***, with the ***slightly neg*** and the ***slightly pos*** .

Its ***shape*** <u>*doesn't*</u> produce a ***neg*** ***end*** and a ***pos*** ***end***, it is <u>*not*</u> a ***po*** ***molecule***.

In ***symmetrical*** molecules, like this, the ***dip*** *cancel* each other out





The C — H bond is a ***po*** ***cov*** ***bond***, with the ***slightly neg*** and the ***slightly pos*** .

The C—Cl bond is even ***more po*** with the ***slightly neg*** and the ***slightly pos*** .

Overall this <u>*does*</u> produce a ***neg*** ***end*** and a ***pos*** ***end*** so it is a ***po*** ***molecule***.

Solubilities

This activity compares the intermolecular forces between polar and non-polar molecular compounds in terms of their ability to dissolve in different solvents

Polar→polar attractions*are intermolecular forces of attractions that exist between molecules which have a* ***permanent*** *overall dipole.*

Polar→polar attractions are ***perm*** and, therefore, ***stro*** than the ***temp*** ***dipoles*** set up in non-polar molecules, *London Dispersion forces*.

Molecular Compound	***Polarity***	***Solubility in heptane*** *(non-polar)*	***Solubility in propanone*** *(polar)*
hexane C_6H_{14}	non-polar		
toluene $CH_3C_6H_5$	nearly non-polar		
trichloroethane CH_3CCl_3	polar		
ethanal CH_3CHO	polar		

When the ***inter attr*** are fairly ***sim***, then ***mol*** can ***mix*** easily be ***miscible*** - '*like dissolves like*'. If the molecules of one substance have much ***stro attr*** then they will tend to form ***sep layers***, be ***immiscible***.

Comparing Boiling Points

This activity compares the boiling points of polar and non-polar molecular compounds

It is only fair to compare molecules of

- ***similar size***
- ***similar shape***
- ***similar numbers of electrons***

iodine chloride
b.p. °C

bromine
b.p. °C

Both molecules are of ***sim si***, ***iden sha*** and have ***sim num of elec*** ( ). Both molecules have ***sim London Dispersion forces***.

Iodine choride has a ***hi boiling point*** because of the ***ex perm po attractions*** that exist.

Both molecules are of ***sim si*** ,***sim sha*** and have ***sim num of elec*** ( ). Both molecules have ***sim London Dispersion forces***.

Methanal has a ***hi boiling point*** because of the ***ex perm po attractions*** that exist.

methanal
b.p. °C

ethane
b.p. °C

ether
b.p. °C

propane
b.p. °C

Both molecules are of ***sim si***, ***sim sha*** and have ***sim num of elec*** ( ). Both molecules have ***sim London Dispersion forces***.

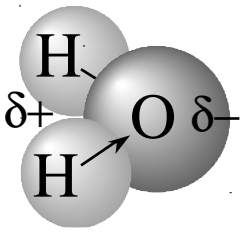
Ether has a ***hi boiling point*** because of the ***ex perm po attractions*** that exist.

The ***stro*** the ***forces of attr*** that exist between molecules, the ***hi*** the boiling point. The ***attr*** that exist between moleules with ***perm*** ***dipoles*** are ***stro*** than London Dispersion forces (***temp*** ***→ind*** ***dipoles***).

Polar Molecules

This activity is about polar molecules and the property which can be used to detect when they are present in a liquid.

*A **polar molecule** is a molecule which has a **permanent** dipole (a permanent **slight positive** charge on one side and a permanent **slight negative** charge on the other)*

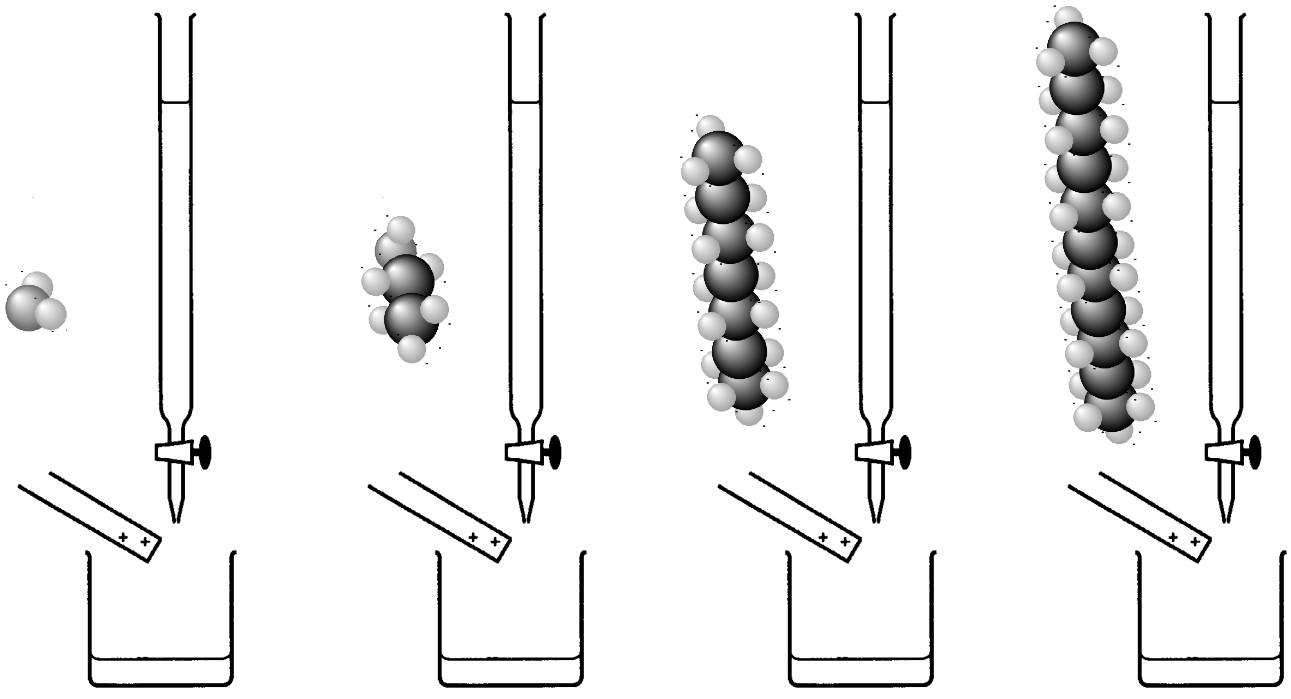


A ***wa*** molecule is a good example of a ***pol*** ***molecule***.

It is ***pol*** , first of all, because the ***gre*** ***electronegativity*** of the ***ox*** atom ( ) compared to the ***hy*** atom ( ), makes the O—H bonds ***pol*** .

Secondly, and just as important; the ***shape*** of the water molecule means that *one side* has a ***sli*** ***neg*** ***charge*** while the *other side* is ***sli*** ***pos*** . This makes water a ***polar molecule***.

Liquids which contain polar molecules can be detected using a charged rod. A fine jet of liquid from a burette will be attracted to the rod if the liquid contains polar molecules.



water ***alcohol*** *(ethanol)* ***white spirit*** *(alkane)* ***paraffin*** *(alkane)*

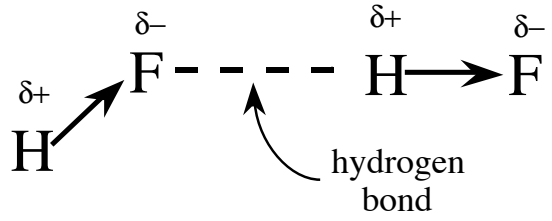
The O—H bond is ***po*** , and the ***shapes*** of ***wa*** and ***eth*** molecules are suitable, so these are ***po*** ***mol*** and were ***attr*** towards the ***char*** rod.

In ***alk*** , the *bonds* and the *molecule shapes* are ***not*** able to produce ***po*** ***mol*** , so they are ***not attr*** towards the ***char*** rod.

1.11 Hydrogen Bonding

This lesson looks at hydrogen bonding, how it arises, and its effect on properties such as melting point, boiling point, and viscosity.

Hydrogen Bonds

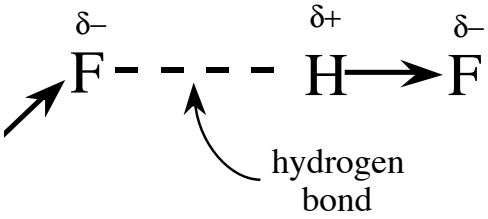


This activity is about the extra bonding strength of the hydrogen bond and the situations in which it can arise.

Molecules such as ammonia (NH_3), water (H_2O) and hydrogen fluoride (HF) have stronger attractions than London Dispersion forces or polar→polar attractions. These stronger attractions are called hydrogen bonds.

Hydrogen bonds are only set up where a ***strongly po cov bond*** already exists between a ***hydr atom*** and a ***sm***, strongly ***electron attr*** atom. There are only 3 really suitable bonds:

$N^{\delta-}—H^{\delta+}$ $O^{\delta-}—H^{\delta+}$ $F^{\delta-}—H^{\delta+}$

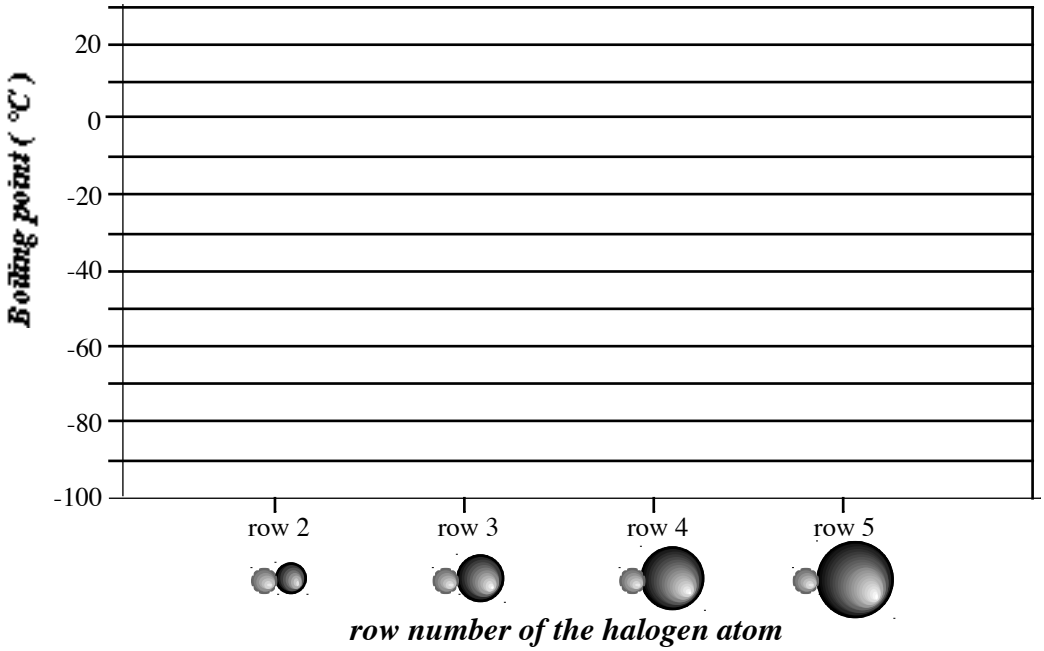


These bond must also ***line up*** so that the ***hydr*** is in a ***str line*** between ***two sli neg atoms***: this leads to much ***stro*** than normal attractions called ***hydrogen bonds***.

Weak Bonding Forces	*Intermediate strength*	*Very Strong Forces*
London Dispersion forces *polar→polar attractions*	*Hydrogen bonds*	*Covalent bonds* *Ionic bonds* *Metallic bonds*

Boiling Points

This activity examines the boiling points of the hydrides of group 7 elements in relation to the intermolecular forces present.

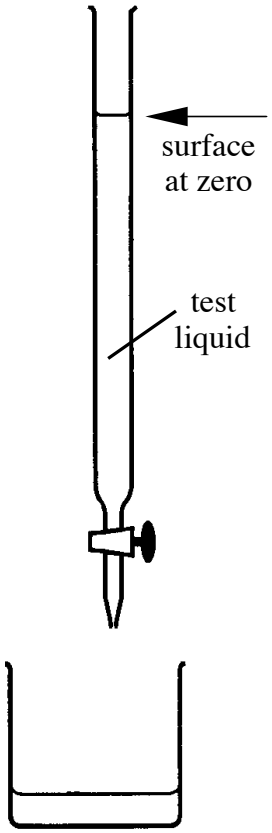


As ***molecular size dec*** we expect the ***boiling point to*** as ***attr*** , especially *London Dispersion forces*, ***dec*** .

We would ***expect***, therefore that the boiling point for *hydrogen fluoride* would be about - °C. It is in fact, 19°C, about ° ***hi*** ***than expected***.

The ***extra en*** is needed to overcome the ***much str*** ***hydr*** ***bonds*** that exist ***bet*** ***molecules*** of *hydrogen fluoride*.

Viscosity & Bonding



This activity examines the effect of hydrogen bonding on the viscosity of a liquid by timing how long two liquids take to drain from a burette.

Two molecules of ***similar si*** , ***similar sh*** and ***num*** ***of elec*** are used to ensure that *London Dispersion forces* would be ***very sim*** .

Molecule X *(propanone)*	***Molecule Y*** *(propan-2-ol)*
polar *hydrogen bonding* ***not*** *present* *seconds to drain from 0 to 50*	*polar* *hydrogen bonding* ***is*** *present* *seconds to drain from 0 to 50*

Propan-2-ol is ***mo*** ***viscous*** than *propanone* because the H—O bond is ***very po*** and the ***slightly pos*** *hydrogen* atom ($H^{\delta+}$) can end up in a ***str*** ***line*** between two ***slightly neg*** *oxygen* atoms ($O^{\delta-}$).

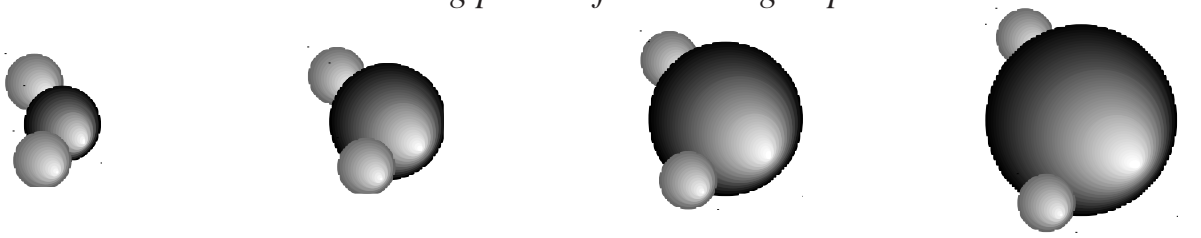
The ***extra attr*** (*hydrogen bonds*) set up ***bet*** the molecules of *propan-2-ol* make the molecules tend to ***sti*** ***together*** and make it a ***mo*** ***viscous*** liquid

1.12 Water

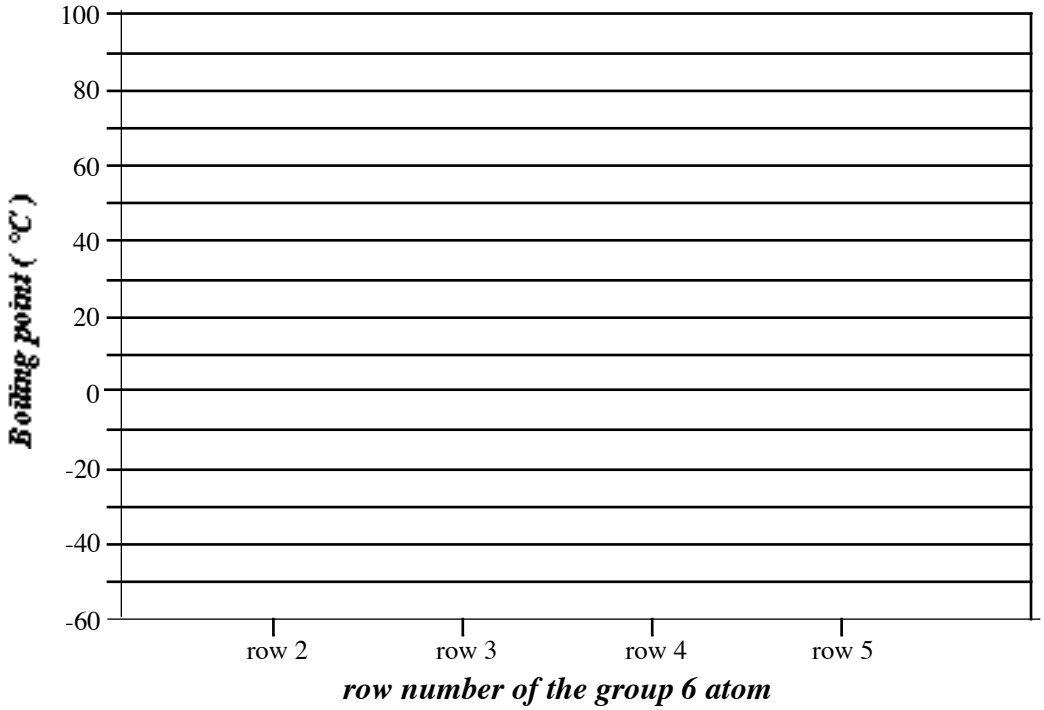
This lesson is about water molecules and the intermolecular forces that act between them.

Boiling Point

This activity considers the boiling point of water in relation to the boiling points of the other group 6 elements.



H_2O H_2S H_2Se H_2Te



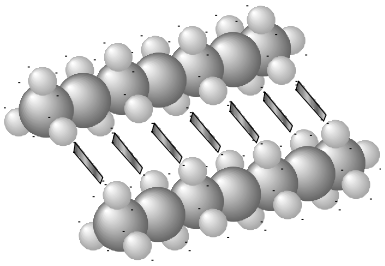
As ***molecular size dec*** we expect the ***boiling point to*** as ***attr*** , especially *Van der Waal's forces*, ***dec*** .

We would ***expect***, therefore that the boiling point for *water* would be about - °C. It is in fact, 100°C, about ° ***hi*** ***than expected***.

The ***extra en*** is needed to overcome the ***much str*** ***hydr*** ***bonds*** that exist ***bet*** ***molecules*** of *water*.

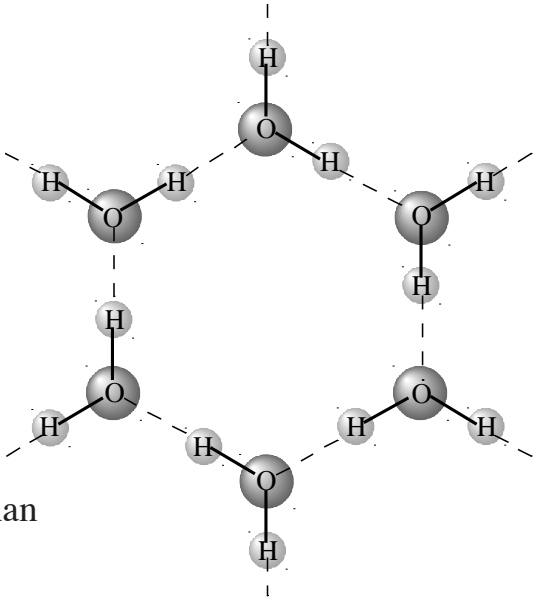
Structure of Ice

This activity looks at the structure and bonding in ice and how this explains it unusual density.



Normally ***mol*** ***solids*** are ***den*** than their own ***liquids***. This is because the ***incr*** ***attr*** (van der Waal's) pull the molecules ***clo*** ***and clo*** together as the solid forms.

For example, a piece of ***can*** ***wax*** will sink when dropped into liquid wax.



Ice ***flo*** on water because it is ***le*** ***dense***. The strongest ***attr*** ***bet*** the water molecules are the ***hydr*** ***bonds***.

Hydrogen bonds are ***str*** , making ice a *strong structure* (think of the Titanic !).

Hydrogen bonding also makes the water molecules line up so that the ***hydr*** atoms ($H^{\delta+}$) are in a ***stra*** ***line*** between two ***ox*** atoms ($O^{\delta-}$).

The water molecules end up ***fur*** ***apart*** in *solid ice* than in *liquid water* - so ice is ***less de*** than water.

1.13 Ionic Compounds

This topic looks more closely at ions and the properties of ionic compounds

Atomic Ions

This activity is about the formation of atomic ions from neutral atoms.

Li^{+} Ca^{2+} Fe^{3+} Sn^{4+} P^{3-} S^{2-} F^{-}

Element	***Atom electron arrangement***	***Ion electron arrangement***	***Ion symbol***
magnesium	2 , 8 , 2	2 , 8	Mg^{2+}
sodium			
chlorine			
bromine			
oxygen			
aluminium			
nitrogen			

Me ***atoms*** tend to form ***pos*** ***ions*** because they have , , or 3 electrons in their ***out shell***. They will try to ***lo*** these electrons to reach the ***nea*** ***noble gas*** electron arrangement. Their ***pos*** ***charges*** (in the ***nuc***) outnumbers their ***elec*** so they become, overall, ***pos*** .

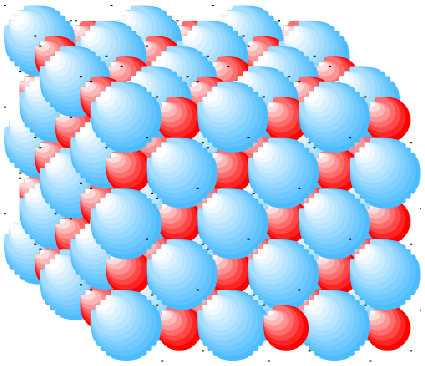
Non-me atoms tend to form ***neg ions*** because they have , , or 7 electrons in their ***out shell***. They will try to ***ga*** extra electrons to reach the ***nea noble gas*** electron arrangement. Their ***pos charges*** (in the ***nuc***) are outnumbered by their ***elec*** so they become, overall, ***neg*** .

For an atom to ***lo electrons***, another atom must be available to ***ga the electrons***. A ***me atom*** cannot ***lo electrons*** if there is not a ***n -metal*** atom available. Simple, two element, ***io compounds*** must always contain a ***met*** and a ***n metallic*** element.

Simple Ionic Structures

This activity looks more closely at the structure and bonding in a simple ionic compound.

The ***ionic bond*** is the ***electrostatic attr*** between a ***pos*** and a ***neg*** ion.

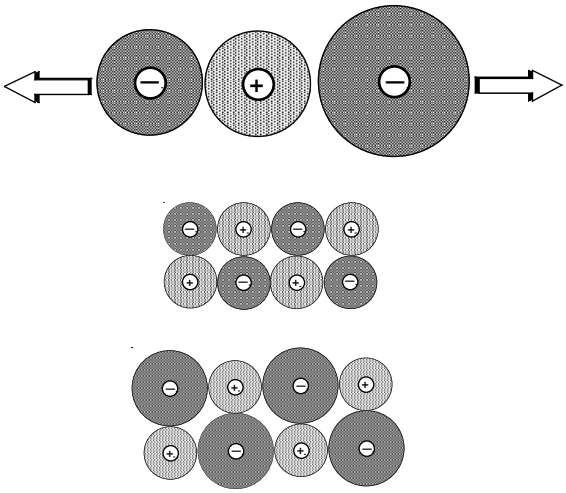


An ***ion*** is able to ***attr*** , and be attracted by, ***sev ions*** in ***all dir*** .

In *sodium chloride*, each Na^+ ***ion*** attracts Cl^- ***ions***.

Each Cl^- is also able to attract Na^+. Overall, there are exactly ***equ numbers*** of the two ions, hence the formula ***NaCl***.

The ionic bonding *extends throughout the structure* (***net***) and a *regular repeating pattern* also exists (***latt***).



The ***stre*** of an ***io bond*** is affected by two factors:-

i) the ***amount*** of ***pos /neg*** charge
ii) the ***size*** of the ions

A ***bi ion*** has a ***wea attr*** than a ***sma*** ion with the <u>*same*</u> *charge*:- the charge is spread over a ***lar vol*** , weakening its effect.

This is why, for example, *sodium fluoride* (°C) has a higher MPt than *sodium iodide* (°C):- the ionic bond in NaF is stronger than in NaI.

1.14 Dissolving

Dissolving in water

This activity looks at the type of substances which dissolve readily in water and explains it in terms of intermolecular forces

Water is made up of ***small po molecules*** with ***hyd bond***s the main ***in molecular forces***, even in the liquid state.

For another substance to ***diss*** in water it has to able to produce ***in molecular forces*** of ***sim stre*** .

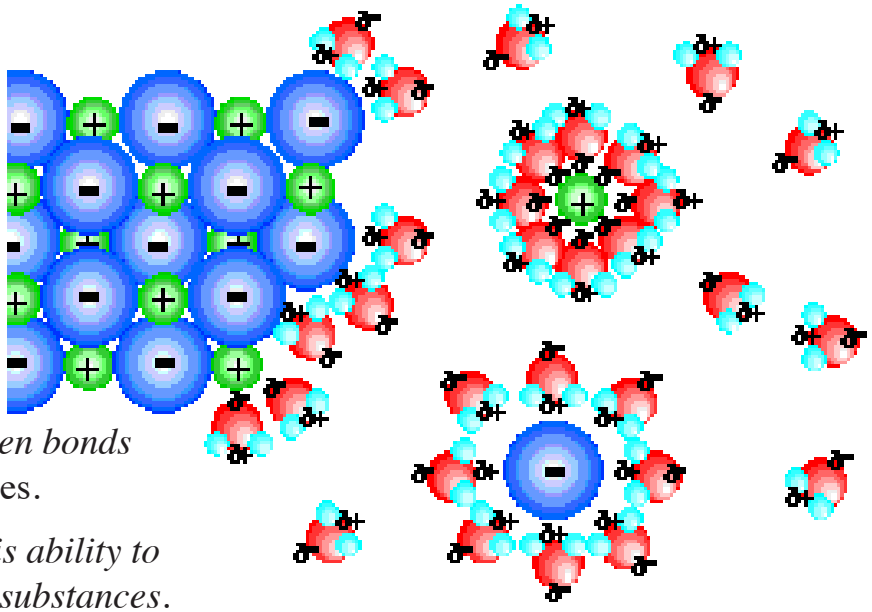
Molecular Compound	*Polarity*	*Hydrogen bonding present*	*Ability to form hydrogen bonds with water*	*Solubility in water*
hexane C_6H_{14}	non-polar	✗	✗	
toluene $CH_3C_6H_5$	nearly non-polar	✗	✗	
trichloroethane CH_3CCl_3	polar	✗	✗	
ethanal CH_3CHO	polar	✗	✓	
propan-2-ol $CH_3CHOHCH_3$	polar	✓	✓	

Miscible liquids, those that mix and ***diss*** in water, must have ***hydr bonds*** already or the *ability to form* ***hydr bonds*** with water molecules.

Some ***io substances*** can also ***diss*** in water.

The ***po*** nature of the water ***mol*** allows it to ***attr*** both ***pos*** and ***neg*** ions.

The ***attr*** between water ***mol*** and these ***io*** are ***sim*** in ***stre*** to the *hydrogen bonds* that exist ***bet*** water molecules.



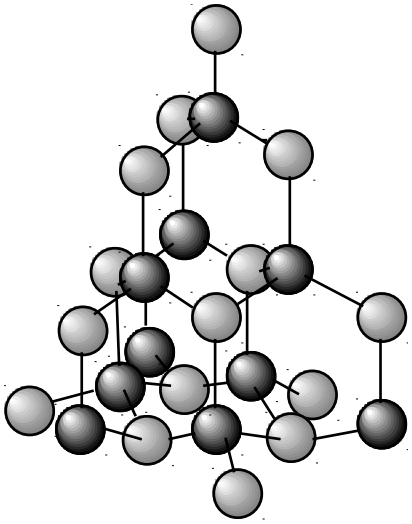
Water is reasonably unique in this ability to dissolve both ionic and covalent substances.

1.15 Covalent Network Compounds

This topic looks at compounds which have a covalent network structure

Silicon Carbide

This activity looks at the bonding and structure of the compound ***silicon carbide*** *(also known as* ***carborundum****).*



Silicon carbide is a ***cov net comp*** :-

comp - two ***diff*** atoms are ***bon*** together

net - the covalent bonding *extends throughout structure*

covalent - both are ***n -metals*** and will ***overlap she*** and ***sh electrons***

The ***sil*** and ***car*** atoms are ***bon*** together by the forces of ***attr*** that exist between the ***sha pairs*** of ***ele*** and the ***nuclei*** of the sharing atoms.

By ***sha electrons***, both the ***silicon*** and the ***carbon*** atoms achieve the ***extra stab*** of a ***full out shell***.

Like all ***net*** , the structure of ***silicon carbide*** goes on 'forever'. However, there is a simple ***rep pattern*** Si C Si C Si C etc, so it can also be described as a ***covalent latt*** and its ***form*** is simply ***SiC***.

Silicon Dioxide

This activity looks at the structure and bonding in ***silicon dioxide*** *(also known as* ***silica****)*

Silicon oxide is a ***cov net comp*** :-

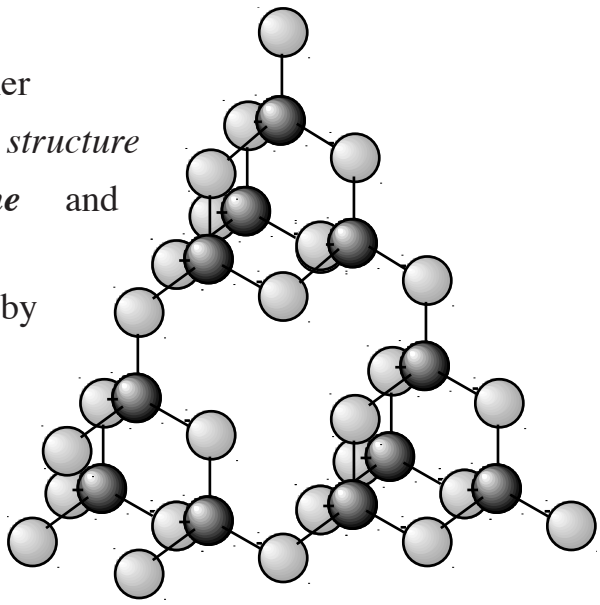
comp - two ***diff*** atoms are ***bon*** together

net - the covalent bonding *extends throughout structure*

covalent - both are ***n -metals*** and will ***overlap she*** and ***sh electrons***

The ***sil*** and ***ox*** atoms are ***bon*** together by the forces of ***attr*** that exist between the ***sha pairs*** of ***ele*** and the ***nuclei*** of the sharing atoms.

By sharing electrons, both the silicon and the oxygen atoms achieve the extra stability of a full outer shell.

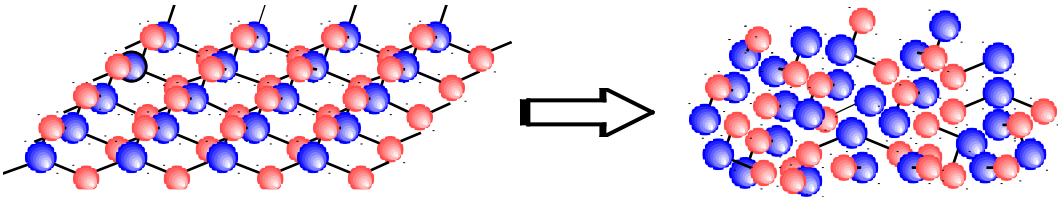


Like all ***net*** , the structure of ***silicon dioxide*** goes on 'forever'. However, there is a simple ***rep pattern*** so it can also be described as a ***covalent latt*** . Each ***sili*** atom is bonded to ***4 ox*** while each ***ox*** is bonded to only ***2 sili*** atoms. Overall, there are ***tw*** as many *oxygen* atoms as *silicon*,and its ***form*** is simplified to ***SiO_2***.

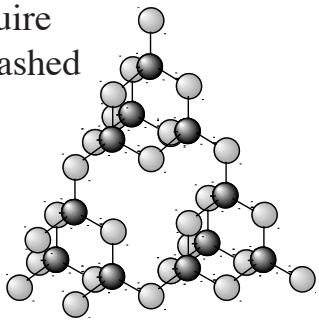
Properties & Uses

Silicon carbide (carborundum) and silicon dioxide (silica) have the typical properties of covalent network substances.

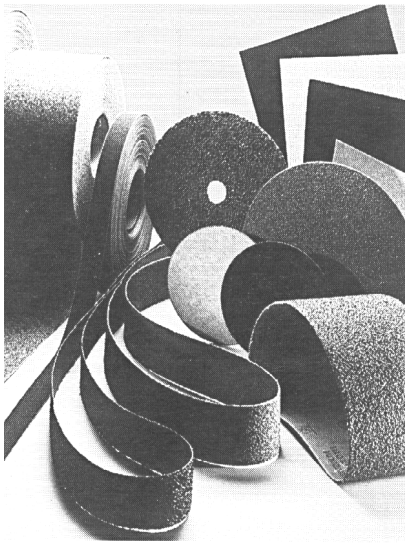
Very High Melting points - a ***lar num*** of ***very str cov bonds*** require to be ***bro*** before ***covalent net*** can melt.



Very Hard - a ***lar num*** of ***very str cov bonds*** require to be ***bro*** before ***covalent net*** can be smashed apart .



Very Unreactive - a ***lar num*** of ***very str cov bonds*** require to be ***bro*** before ***covalent net*** can react with other chemicals.



Uses of silicon carbide	***Related Property***
furnace linings, crucibles	
abrasive papers, cutting discs	
Uses of silicon dioxide	***Related Property***
building material	
abrasive papers, cutting discs	
making glass	

1.17 Redox Agents

This lesson is about reactions involving reduction and oxidation.

Definitions

Oxi ***Is Lo*** of electrons by a reactant.

Gaining ***oxy*** / losing ***hydr*** (*increasing* O:H ***rat***)
can also be described as ***Oxi*** :-

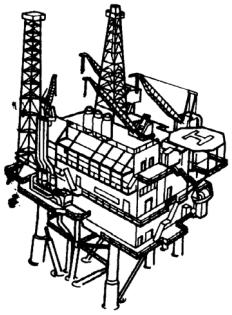
e.g. *aldehyde* (—CHO) → *acid* (—COOH)

e.g *alcohol* (—CH_2OH) → *aldehyde* (—CHO)

Red ***Is Ga*** of electrons by a reactant.

Losing ***oxy*** /gaining ***hydr*** (*decreasing* O:H ***rat***)
can also be described as ***Red*** :-

e.g. *ketone* (—CO—) → *alcohol* (—CHOH—)



OIL RIG

Oxi and ***Red*** always go together so we call these reactions ***REDOX***.

Oxidising & Reducing Agents

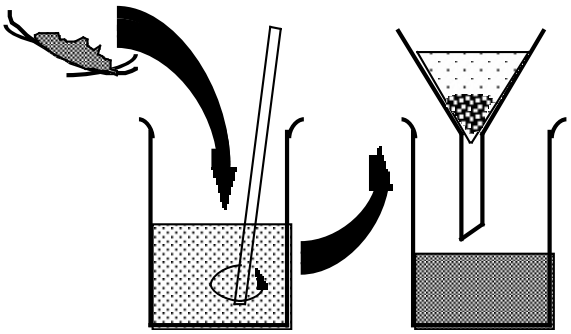
This activity considers how the term agent can be applied to oxidation & reduction reactions

*An **oxidising agent** is a chemical which acts on another substance to oxidise it*

*A **reducing agent** is a chemical which acts on another substance to reduce it*

For example, during the ***displ*** of *copper* by *zinc*, the following reaction occurs.

$$Zn_{(s)} + Cu^{2+}(NO_3^-)_{2\,(aq)} \rightarrow Cu_{(s)} + Zn^{2+}(NO_3^-)_{2\,(aq)}$$



Oxi : $Zn_{(s)} \rightarrow$ $_{(aq)}$ + 2e

By losing electrons ***zinc at*** are *causing the reduction* of Cu^{2+} ions so the ***zinc at*** are the ***red*** ***ag***

Red : $_{(aq)}$ + 2e → $Cu_{(s)}$

By gaining electrons ***copper io*** are *causing the oxidation* of Zn atoms so the ***copper io*** are the ***ox*** ***ag***

During the Organic Unit you will be asked to learn the following ***oxi*** reactions:

pri alcohol → ***ald*** → ***ac*** and ***sec*** alcohol → ***ket***

A number of ***oxi*** ***agents*** can be used; each time the ***ag*** will itself be ***red*** as they ***accept electrons***

e.g.	***acidified dichr***	$Cr_2O_7^{2-}$	+	$14H^+$	+	6e	→	$2Cr^{3+}$ + $7H_2O$
	feh /ben	Cu^{2+}			+	e	→	Cu^+
	toll reagent	Ag^+			+	e	→	Ag

*An **oxidising agent** will be **reduced** when it reacts*

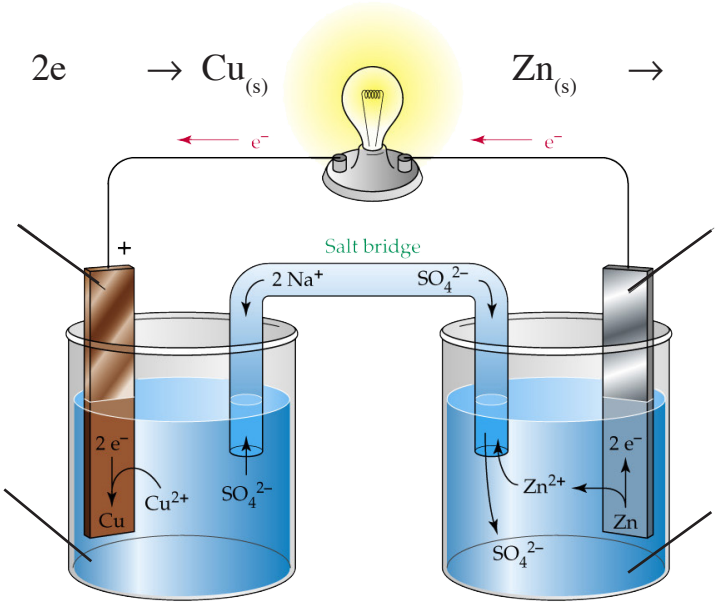
*A **reducing agent** will be **oxidised** when it reacts*

Donating & Accepting Electrons

This activity demonstrates electron transfer in a redox reaction in an electrochemical cell

Red ***Ox***

$_{(aq)}$ + 2e → $Cu_{(s)}$ $Zn_{(s)}$ → $_{(aq)}$ + 2e



By gaining electrons ***copper io*** are *causing the oxidation* of Zn atoms so the ***copper io*** are the ***ox*** ***ag***

By losing electrons ***zinc at*** are *causing the reduction* of Cu^{2+} ions so the ***zinc at*** are the ***red*** ***ag***

By *gaining* electrons ***copper io*** are *accepting electrons* from the Zn atoms.

By *losing electrons* ***zinc at*** are *donating electrons* to the Cu^{2+} ions.

In an electrochemical cell electrons flow from the reducing agent to the oxidising agent.

*An **oxidising agent** will **accept electrons** when it reacts*

*A **reducing agent** will **donate electrons** when it reacts*

Electrochemical Series

This activity is about how the electrochemical series can be used to write both reduction and oxidation equations

Help is available in the Data Booklet in the form of the ***Electro Series***.

Equations are always written as ***Red*** reactions, but can be ***rev*** for ***Oxi*** reactions. Generally speaking, each equation represents the following change:

Oxi agent + ***e*** → ***Red agent***

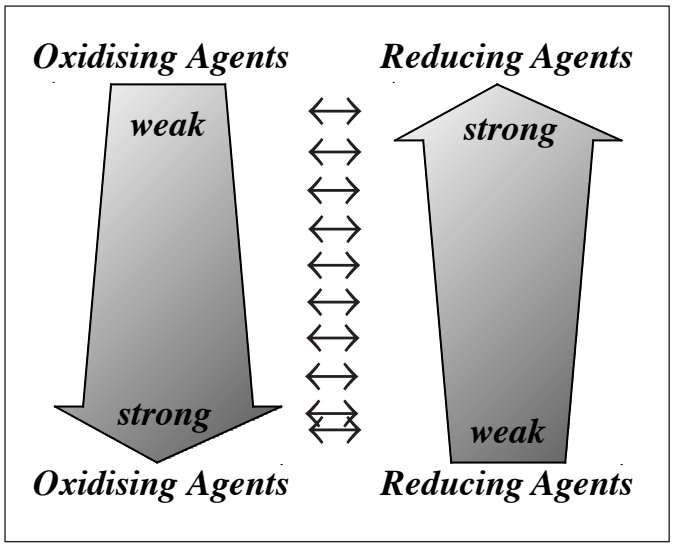
e.g. K^+ + e → K

This reaction, from near the ***top*** of the ***electro series***, is *very unlikley* to go ***as wri*** so we can describe ***pota ions*** as ***very we oxidising agents*** while ***potassium at*** would make ***very str reducing agents***.

Oxidising agent + ***e*** → ***Reducing agent***

e.g. $Cr_2O_7^{2-}$ + $14H^+$ + 6e → $2Cr^{3+}$ + $7H_2O$

This reaction, from near the ***bottom*** of the ***electro series***, is *very likely* to go ***as wri*** so we can describe ***dichr ions*** as ***very str oxidising agents*** while ***chro (III) ions*** would make ***very we reducing agents***.



A useful *'rule-of-thumb'* that is worth remembering is that the

reaction lower in the series goes as written while the

reaction further up will be reversed.

e.g a simple ***displ*** reaction, eg Zn/Cu

+ 2e → ***as wri***

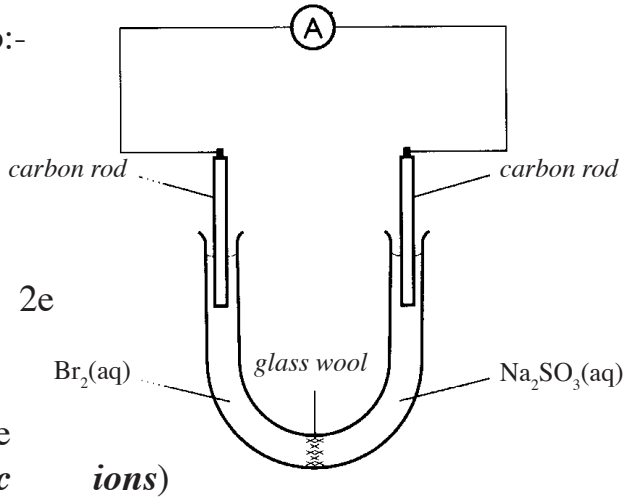
→ + 2e ***rev***

In this example, ***bro*** is ***lo*** in the series so:-

+ 2e →

While ***sulph ions*** are found ***hig*** in the series so:-

+ H_2O → + $2H^+$ + 2e



(Sodium ions are so far up the series that it is safe to assume that they will never react; they are ***spec ions***)

1.18 Redox Equations

This lesson looks at how the overall equation for a redox reaction can be worked out from the ion-electron equations for the oxidation and reduction changes involved.

Combining Simple Ion-Electron Equations

This activity looks at the conditions for combining ion-electron equations and how this can lead to an overall balanced equation for a redox reaction.

Before ion-electron equations can be combined, the following two conditions must be satisfied :-

One equation must be a ***red*** (*as written* in *electrochemical series*) ,
One equation must be an ***ox*** (*reverse* of equation in series)

and The ***num*** ***of elec*** in each equation must be the ***sa*** .

So

1. *Identify and copy the two ion-electron equations from the Electrochemical series.*

   *(**lower** in series goes **as written, higher** in series is **reversed**)*

2. *Multiply each ion-electron equation as necessary to ensure*

   number of electrons lost = number of electrons gained

3. *Combine by writing everything on the left of both ion-electron equations on the left of RedOx equation, similarly for the right.*

   *(**electrons should be equal and cancel each other out**)*

Example 1: *Aluminium metal and silver ions (displacement).*

→ (***oxidation***)

→ (***reduction***)

→ (***combined RedOx***)

Example 2: *oxidation of iron(II) ions by chlorine (a strong oxidising agent).*

→ (***oxidation***)

→ (***reduction***)

→ (***combined RedOx***)

More Complex Examples

This activity is about combining ion-electron equations which have ions containing oxygen

These examples appear more complicated because they also involve ***H^+ ions*** and ***H_2O molecules*** - so they are usually ***only able to react when in solution*** .

(Remember that, from your work on acids, water always contains a small proportion of H^+ ions, though we often add some acid to help these reactions go faster).

Many of these ***ion-electron equations*** are available in the ***Electrochemical Series*** in your ***Data Booklet*** and can be dealt with in exactly the same way.

Example 3: *oxidation of sulphite ions, SO_3^{2-}, by acidifed permanganate solution, MnO_4^-, (an extremely strong oxidising agent).*

SO_3^{2-} + → + + (***oxidation***)

MnO_4^- + + → + (***reduction***)

+ + → + + (***RedOx***)

Oxidising Agent: MnO_4^- is a powerful *oxidising agent* and can oxidise your skin, turning it brown so care must be taken when using permanganate solutions.

Reducing Agent: SO_3^{2-} is a useful *reducing agent* and can be seen on the labels of many foodstuffs - it slows down the oxidation of food and is labelles as a '***preservative***'.

Ion-Electron Equations From Scratch

This activity is about how ion-electron equations can be written without using a data booklet

As you have seen, many ***oxi*** /***red*** reactions need more than just ***elect*** to keep them ***bal*** . We must also introduce and balance, as necessary, ***H^+ ions*** and ***H_2O molecules***.

However, every balanced equation (even an ion-electron equation) must still have the *same number of each type of atom* on both sides and this must be done ***before*** adding H^+ *ions* and H_2O *molecules* and ***before*** we *add in electrons* to ensure that the *overall charge* on both sides is *the same*.

e.g ***dichr*** ***ions*** *(orange)* are strong ***oxi*** ***agents*** because they are ***easily red*** to form ***chro*** ***(III) ions*** *(green)*.

Step 1 $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ *(unbalanced)*

Step 2 there must be an *equal number of chromiums* on both sides so,

$Cr_2O_7^{2-} \rightarrow \mathbf{2}\ Cr^{3+}$

Step 3 there must be an *equal number of oxygens* on both sides so ***wat*** ***is added***,

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + \mathbf{7\ H_2O}$

Step 4 there must be an *equal number of hydrogens* on both sides so ***ions are added***,

$Cr_2O_7^{2-} + \mathbf{14\ H^+} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Step 5 there must be an *equal amount of charge* on both sides so ***elec*** ***are added***,

before (2— and 14 + = **12 +**) (2 times 3 + = **6+**)

$Cr_2O_7^{2-} + 14\ H^+ + \mathbf{6e} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

after (2— and 14 + and 6 — = **6 +**) (2 times 3 + = **6+**)

(Most ***redox*** reactions will only take place in ***sol*** , so there will be plenty of H_2O and H^+ available when needed.The pH of a solution often ***cha*** during a Redox reaction. If a large number of H^+ ions are needed then ***ac*** may be added e.g. *'acidified dichromate'*.)

Example 4: *Thiosulphate ion*, $\mathbf{S_2O_3^{2-}}$, being *oxidised* to the *sulphite ion*, $\mathbf{SO_3^{2-}}$

$S_2O_3^{2-} \rightarrow SO_3^{2-}$

1.19 Redox Titrations

This lesson introduces how titration can be used to determine the concentration of one of the reactants in a redox reaction.

The Redox Reaction

This activity considers the overall equation for the oxidation of iron (II) by permanganate ions.

original $Fe^{2+} \rightarrow$ (***oxidation***)

balanced $Fe^{2+} \rightarrow$ (***oxidation***)

MnO_4^- + + $\rightarrow$ + (***reduction***)

MnO_4^- + Fe^{2+} + $\rightarrow$ + + (***RedOx***)

As shown by the overall equation, a ***large num*** of ions are used up during the ***RedOx*** reaction. Though there are always some ions present in water, the ***conc*** is ***very low*** and the ***ra of the reaction*** would be ***very sl*** - too slow for the ***rapid col change*** needed during a ***titr***. Therefore, ***ac*** is always added to ensure a ***rapid reaction*** - the ***solution is acidified***.

The Redox Reaction

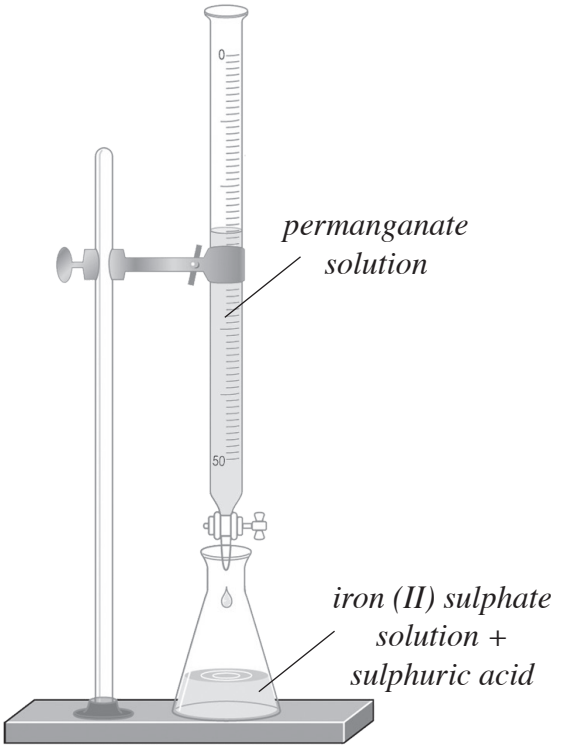
In this activity, an iron (II) sulphate solution of unknown concentration is titrated with potassium permanganate solution of known concentration.

Exactly 20 cm^3 of *iron (II) sulphate* solution was ***pip*** into the flask, and ***aci*** using *sulphuric acid*.

Permanganate solution (***pur***), of ***known conc***, was added from the ***bur***. As the MnO_4^- ions reacted with the Fe^{2+} ions, the ***pur*** colour disappears.

More *permanganate* solution was added, eventually drop by drop, until a ***faint pink/purple colour persisted***.

The strength of the MnO_4^- ***col*** means that even a single *excess* drop produces a visible colour so no need for a separate ***ind*** to detect end-point.



The ***tit*** was repeated until ***con*** ***results*** were obtained - at least two results within 0.2 cm^3 of each other.

Non-con results are discarded and the ***con*** results are ***av*** .

volume of 0.02 M $KMnO_4$ required to exactly oxidise 20 cm^3 of $FeSO_4$

=

Attempt number	*Starting volume (ml)*	*Final volume (ml)*	*Volume added(ml)*
1			
2			
3			
4			

The Calculation

This activity is about calculating the amount of the iron (II) ions in the solution of iron (II) sulphate.

Chemistry calculations can be done in a variety of ways - the 'best' method is often a matter of personal choice. However, the ***nature of the information provided*** *and the* ***format of the final answer required*** *can tip the balance.*

Method 1 ***provided:*** overall RedOx equation
required: concentration of iron(II) solution

n_{ox} = number of moles of oxidising agent

$$\frac{V_{ox}\ C_{ox}}{n_{ox}} = \frac{V_{red}\ C_{red}}{n_{red}}$$

n_{red} = number of moles of reducing agent

C_{ox} = known, V_{ox} = known
C_{ox} = 0.02 M, V_{ox} = average titre

V_{red} = known, C_{red} = calculate
V_{red} = 20 cm^3, C_{red} = calculate

$n_{ox} = 1$ $n_{red} = 5$

$MnO_4^- + 5\ Fe^{2+} + 8\ H^+ \rightarrow Mn^{2+} + 5\ Fe^{3+} + 4\ H_2O$ (***RedOx***)

Method 2 ***provided:*** individual ion-electron equations (data book)
required: concentration of iron(II) solution

p_{ox} = 'power' of the oxidising agent

$$V_{ox}\ C_{ox}\ p_{ox} = V_{red}\ C_{red}\ p_{red}$$

p_{red} = 'power' of the reducing agent

C_{ox} = known, V_{ox} = known
C_{ox} = 0.02 M, V_{ox} = average titre

V_{red} = known, C_{red} = calculate
V_{red} = 20 cm^3, C_{red} = calculate

$p_{red} = 1$
$p_{ox} = 5$

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

$MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

Method 3 ***provided:*** overall RedOx equation
required: mass of iron present

C_{ox} = known, V_{ox} = known
C_{ox} = 0.02 M, V_{ox} = average titre (*must be converted to litres*)

Step 1: *Start with 'known' chemical and calculate number of moles added from burette.*

$$n_{ox} = C_{ox} \times V_{ox}$$

Step 2: *Use balanced equation to convert from n_{ox} to n_{red}*

MnO_4^- + 5 Fe^{2+} + 8 H^+ → Mn^{2+} + 5 Fe^{3+} + 4 H_2O (***RedOx***)
1 mole 5 moles

$$n_{red} = n_{ox} \times 5$$

Step 3: *Use gram formula mass of iron to calculate mass of iron present.*

$$mass = n_{red} \times 55.8\ g$$