4.1 – Ionic Bonding

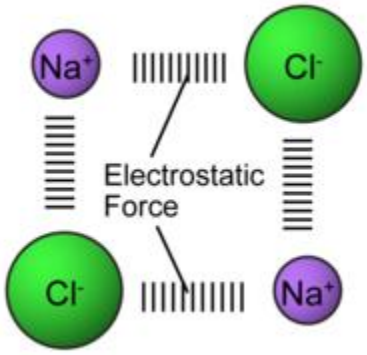
4.1.1 - Describe the ionic bond as the electrostatic attraction between oppositely charged ions

Ions are formed when electrons are transferred from a metal atom to a non-metal atom in order to produce a full outer shell for both ions. The metal will have a **positive charge**, whilst the non-metal will have a **negative charge**.

These opposite charges create and electrostatic attraction between the ions, causing them to form a neutral lattice. The charges of the ions in the lattice will cancel each other out. Ionic bonds are very strong, so ionic compounds have high melting points.

Ions have different charges, depending on how many electrons they lost or gained to form their stable configuration. Ions with a positive charge are called **cations**, and ions with a negative charge are called **anions**. The transition metals are able to form ions of more than one charge.

Ionic compounds are often called **salts**.

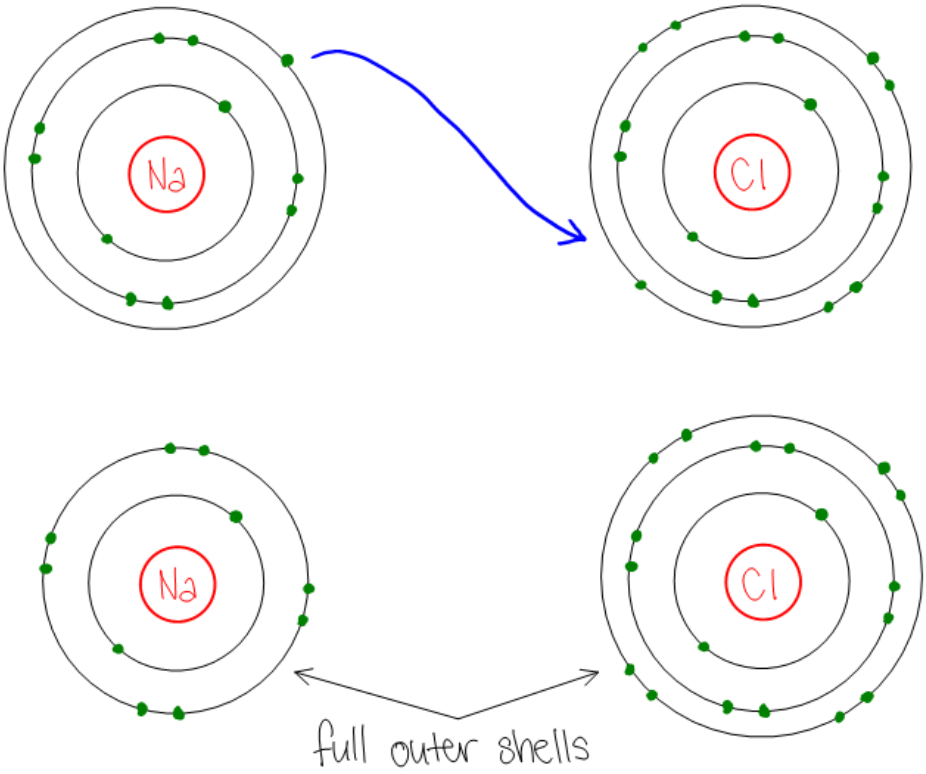


4.1.2 - Describe how ions can be formed as a result of electron transfer

The electrons in the outer shell of an atom are called the **valence electrons**. During electron transfer, the valence electrons of the metal atom are donated to the non-metal atom in order to fill its outer shell.

The result of this is that the metal becomes a **positively charged cation**, and the non-metal becomes a **negatively charged anion**.

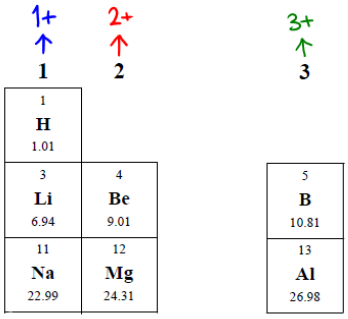
The cation loses its entire valence shell, so the ion will be smaller than the atom. The anion increases in size because of the repulsion of the additional electrons.



The configuration of the ions in the lattice will depend on the size of the ions. The ratio of ions will depend on the charges of the ions. However, it will always work out that the ions neutralise each other.

4.1.3 - Deduce which ions will be formed when element in groups 1, 2 and 3 lose electrons

Looking at the periodic table, there are trends in the ions that form in each group.



In groups 1, 2 and 3, the atoms form positive ions by losing their electrons. Their charge is the same as their group number.

Note that Boron **does not** form an ion, as it does not form an octet by doing so.

4.1.4 - Deduce which ions will be formed when elements from groups 5, 6 and 7 gain electrons

Likewise, the atoms in groups 5, 6 and 7 form ions based on the group they are in.

7 **N** 14.01	8 **O** 16.00	9 **F** 19.00
15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45

These trends occur because atoms in the same group have the same number of valence electrons in their outer shell. Therefore, when they lose or gain electrons, the atoms in the same group will lose or gain the same number of them.

4.1.5 - State that transition elements can form more than one ion

Transition metals have a complex electron arrangement, which means that they form more than one type of ion. They still form cations. For example:

$$Fe \rightarrow Fe^{2+} \text{ or } Fe^{3+}$$

4.1.6 - Predict whether a compound of two elements would be ionic from the position of the elements in the periodic table or from their electronegativity values

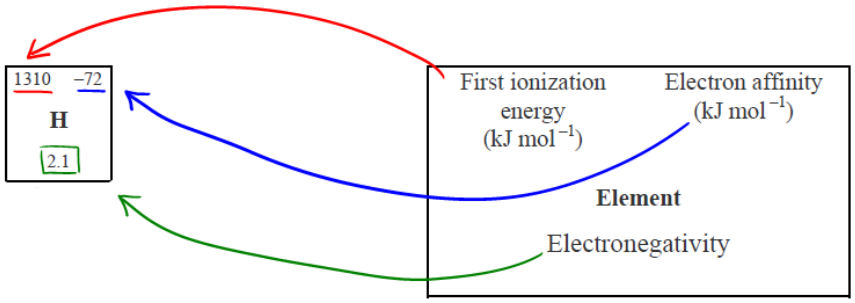
The non-metals occur on the right side of the periodic table [the white ones below]. The metals are on the left [highlighted in yellow]. An ionic compound can only form between a metal and a non-metal.

1	2											3	4	5	6	7	0
1 H 1.01				Atomic number Element Atomic mass													2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.71	29 Cu 63.55	30 Zn 65.37	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.30
55 Cs 132.91	56 Ba 137.34	57 † La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.21	77 Ir 192.22	78 Pt 195.09	79 Au 196.97	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.98	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 ‡ Ac (227)															

†	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 146.92	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
‡	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (254)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Each atom has an electronegativity value assigned to it. If the two atoms forming the compound have a difference in their electronegativities of over 1.8, then their bonding would be ionic.

For example, the electronegativity of Lithium is 1.0, whilst the electronegativity of Chlorine is 3.0. The difference is 2.0, which is greater than 1.8. Therefore, they will form an ionic bond.



4.1.7 - State the formula of the common polyatomic ions formed by non-metals in periods 2 and 3

A polyatomic ion is when more than one element forms a single ion. The formula of the ionic compound is still found the same way, by neutralising the charge on each ion. Polyatomic ions are simply treated as a single ion.

Examples include:

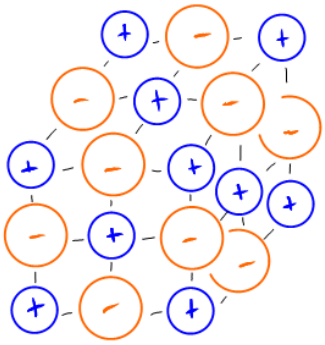
SO_4^{2-} CO_3^{2-} PO_4^{3-} OH^- NO_3^- NH_4^+ H_3O^+ HSO_4^- HCO_3^- CN^- CH_3COO^- MnO_4^-

If more than one polyatomic ion occurs in a compound, then brackets are placed around it to indicate it is a separate entity:

$Al(OH)_3$

4.1.8 - Describe the lattice structure of ionic compounds

Ionic compounds exist in a regular pattern called a lattice structure, or **ionic lattice**. This can contain millions of ions that extend in all three dimensions. There is no fixed number of ions that can be involved, however the ratio of positive to negative ions must be the same as in the **empirical formula** to ensure that all the charges on the ions are neutralised.



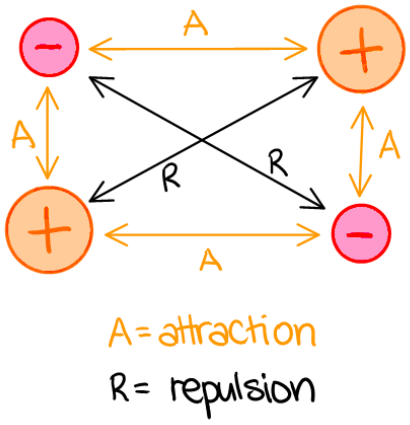
For the most stable arrangement, positively charged ions are packed as closely as possibly to the negatively charged ions, whilst ions of the same charge are as far apart as possible. This **maximises electrostatic attraction between the ions**, while minimising repulsion. Many

different arrangements can be generated to do this, which depends on the size of the ions and their ratio. It will result in the lattice structure for that compound.

4.2 - Covalent Bonding

4.2.1 - Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively-charged nuclei

Outer shell electrons interact and rearrange themselves into a **more stable arrangement** that has lower chemical energy.

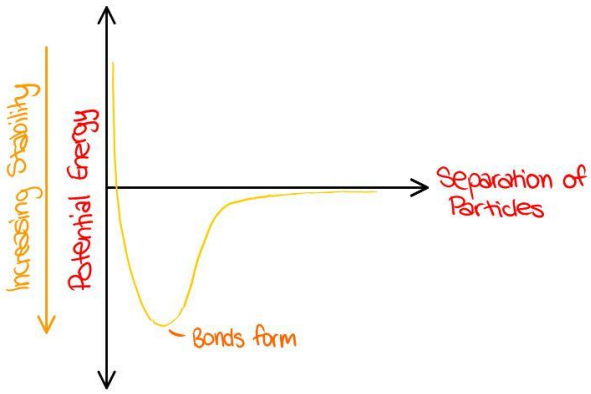


The **positively-charged nucleus** of an atom is attracted to the **negatively charged electrons**. When two atoms come together to form a covalent bond, the nuclei will be attracted to the electron pairs of the other atom.

However, there is **repulsion** between all the electrons, as they have the same charge. The same is true for the nuclei, which also repel each other.

To maintain the covalent bond, a balance must be reached between attraction and repulsion.

A **molecule** can be defined as a discrete group of non-metal atoms covalently bonded to one another. Molecules contain atoms in a set ratio.



4.2.2 - Describe how the covalent bond is formed as a result of electron sharing

There is significant overlap in the radii of the atoms when a molecule is formed.

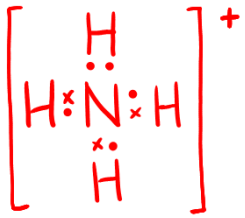
As the two atoms approach each other, electrostatic attractions and repulsions occur between the nuclei and electrons.

Covalent bonds involve the sharing of the electrons. The electrons form **bonding pairs**. When only one pair occupies the space between the electrons, this is a single covalent bond. Any other pairs of valence electrons are called **non-bonding pairs**, or lone pairs. These will help to determine the shape of the molecule, which in turn affects its properties.

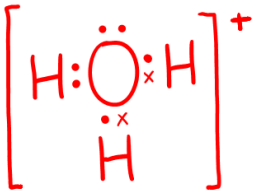
When both the electrons in a bonding pair come from the same atom, they form a **dative covalent bond**, such as in CO, NH_4^+ and H_3O^+

CO

NH_4^+



H_3O^+



This sharing of electrons allows each of the atoms to fill their outer shell.

4.2.3 - Deduce the Lewis (electron dot) structures of molecules and ions for up to four electron pairs on each atom

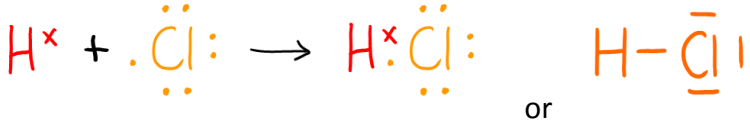
This can be done using dots, crosses or lines

Electron shell diagrams, also called **Lewis or electron dot structures**, can be constructed for covalently bonded molecules. In these, all the valence electrons are drawn, as they form part of the bonding, including the non-bonding electrons. They are used to show how a full outer shell is obtained.

In Lewis structures, the valence electrons are represented by **dots or crosses**. Pairs of dots or crosses represent non-bonding pairs, while a dot and a cross represent a bond.

For diatomic molecules, there are only two atoms bonded to fill the outer shell. In chlorine, which has seven valence electrons, has a single bond

In HCl, only a single bond is required. In both cases, the other six electrons do not take part, as they are non-bonding valence electrons.



Group six elements have six electrons so they must be sharing another two electrons to have their full outer shell. For the diatomic molecule 0_2, a **double bond** is formed to fill its shell.

Elements in group five need a triple covalent bond to get the eight electrons. This leaves only a single pair of non- bonding electrons.

Other examples include:

CO_2 **HCN** **C_2H_4** **C_2H_2**

Lewis structure can be drawn for any molecules, which become useful in determining their shape, according to this procedure:

1. Determine how many valence electrons there are in each atom in the molecule
2. Find how many electrons are required to fill the valence shell - the number of bonds the atom will form
3. Draw electron dot diagrams, pairing up all the electrons except the number that will be used to form bonds.
    - The atom with the most bonds will be the central atom
4. Arrange the outer atoms around the central atom so that their single dots are near the central atom
    - Pair up single electrons between the central atom and outer atoms to form covalent bonds
5. Each atom in the structure should now have a total of 8 valence electrons, except hydrogen, which has 2.

4.2.4 - State and explain the relationship between the number of bonds, bond length and bond strength

Number of Bonds	Bond Length	Bond Strength
Single bond	Long	Strong
Double bond	Shorter	Stronger
Triple bond	Shortest	Strongest

Bond length decreases as there are more electron pairs involved, causing greater attractive force between the two nuclei.

Bond strength increases because more energy is required to break them.

For example, looking at the bonds between carbon atoms:

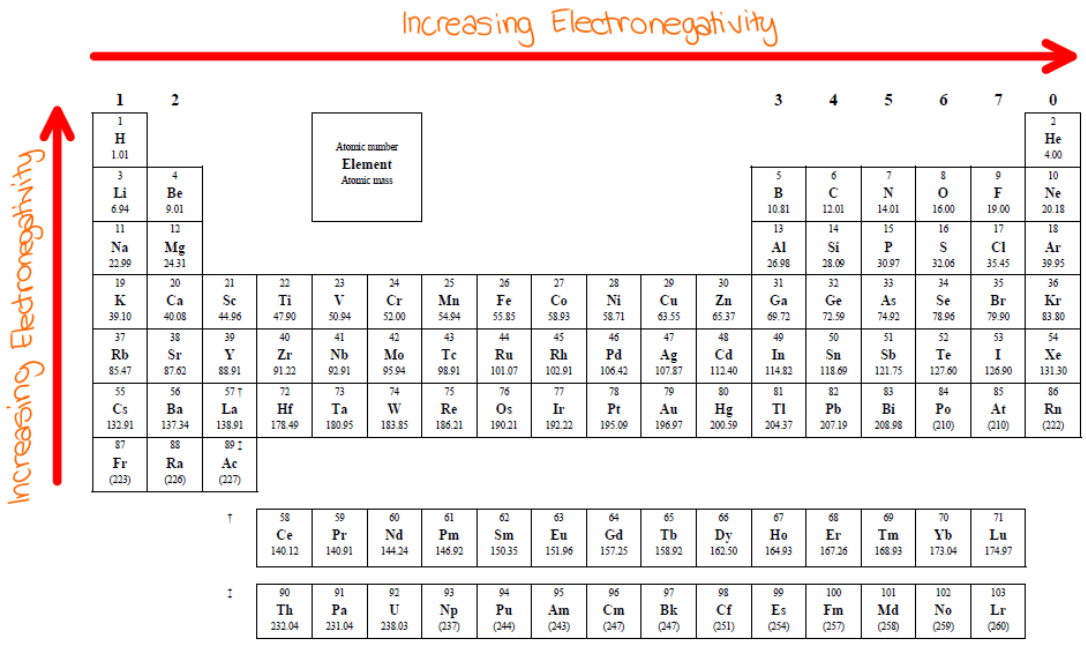
Number of Bonds	**Bond Length (pm)**	**Bond Strength** Dissociation Enthalpy (kJ mol^{-1})
Single bond	154	364
Double bond	134	602
Triple bond	120	835

In carboxylic acid:

Bond	**Bond Length (pm)**	**Bond Strength** Dissociation Enthalpy (kJ mol^{-1})
C=O	120	799
C-O	143	358

4.2.5 - Predict whether a compound of two elements would be covalent from the position of the elements on the periodic table or from their electronegativity values

When two or more different elements are bonded, the sharing of the electrons is not exactly equal. This is because their electronegativity values are different. Electronegativity is a measure of the ability of an atom to attract electron in a bond.

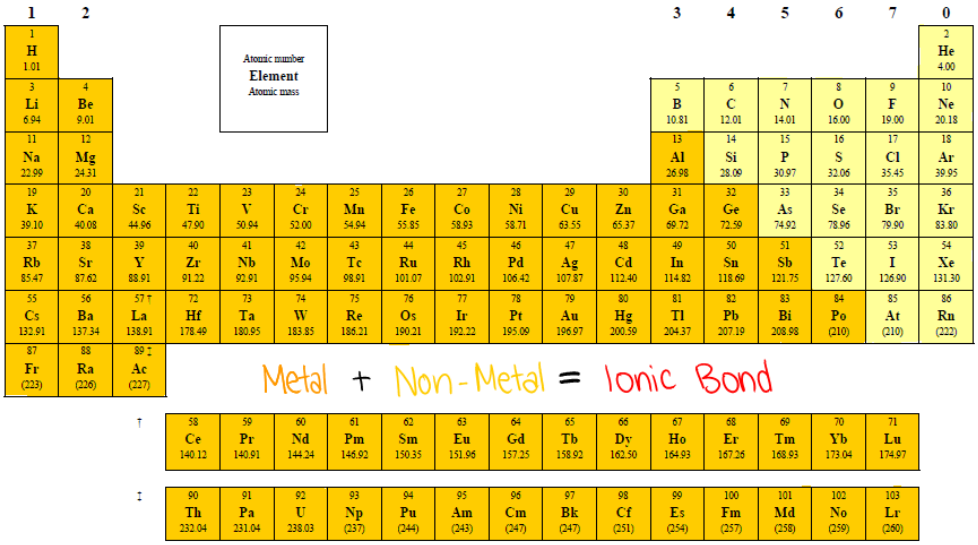


Electronegativity is the highest at the top of the period table, as these have their valence electrons closer to the nuclei. As we move down, the other hand, the valence shell becomes further away. Moving across, the nuclear charge increases, and so does the attraction between the nucleus and valence electrons.

So, the electronegativity **increases as we move up and across** the periodic table. However, the noble gases have an undefined electronegativity, as they already have a full shell.

- Non-metals have higher electronegativity than metals
    - This leads to ionic bonding.

If two non-metals bond, it will be covalent because the electronegativity values are closer together.

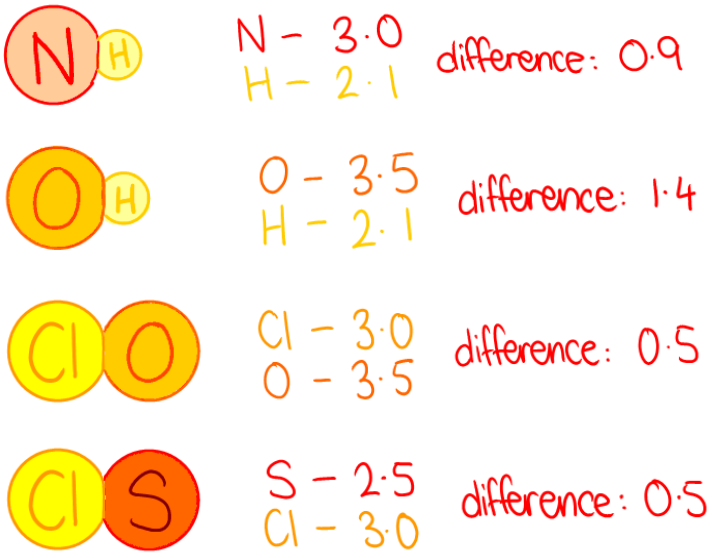


4.2.6 - Predict the relative polarity of bonds from electronegativity values

Every element has a different electronegativity. The polarity of bonds is determined by the difference in electronegativity of the constituent atoms.

Polar molecules have a slight charge on each end.

Difference in Electronegativity	Bond Type
0.0 - 0.4	Non-Polar Covalent
0.5 - 2.0	Polar Covalent
> 2.0	Ionic



4.2.7 - Predict the shape and bond angles for species with four, three and two negative charge centres on the central atom using the valence shell electron pair repulsion theory (VSEPR)

The **structural formula** is the most useful representation of molecules. The actual shape of the molecule is shown, with the electron pair drawn as simple lines, though the non-bonding pairs can also be shown as two dots. The shape has an important role in the chemical and physical properties of the molecules.

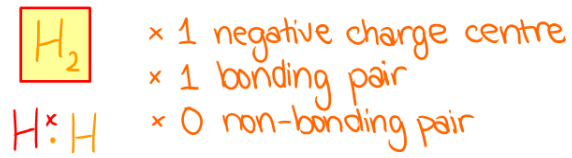
Valence Shell Electron Pair Repulsion theory (VSEPR) is based on the fact that each pair of electrons will be repelled from the others, causing them to move as far away from them as possible in the three-dimensional space. The electrostatic repulsion of pairs determines the geometry of the atoms in the molecule. The space between the electron pairs usually goes in this order

non-bonding/non-bonding > non-bonding/bonding > bonding/bonding

The shape is determined by the number of bonding and non-bonding pairs of electrons on the central atom. Although non-bonding pairs affect the shape, they are not considered when naming the shape of the molecule.

A **negative charge centre** or region refers to pairs of electron on the central atom, including both bonding and non-bonding pairs. Double or triple bonds are still counted as one charge centre.

Diatomic (linear)

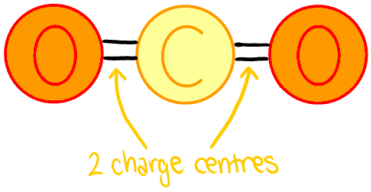


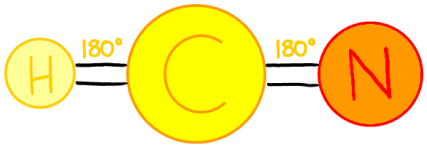
Other examples:

HCl **O_2** **N_2**

Linear

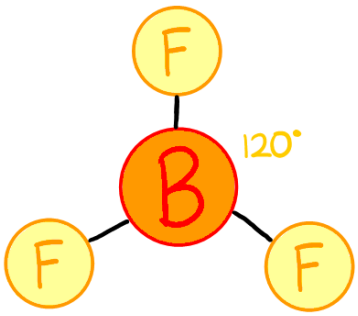
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
2	4	0	180^o

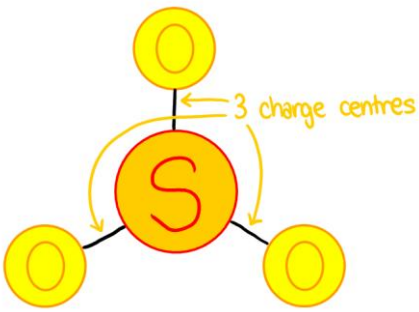




Trigonal Planar

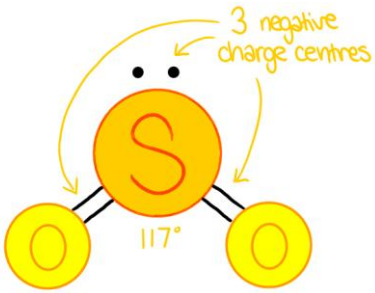
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
3	3	0	120^o





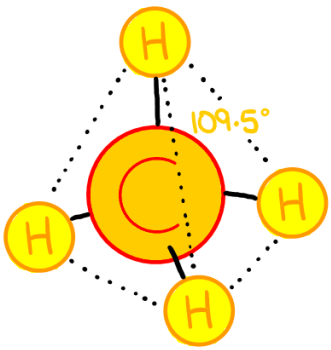
V-Shaped

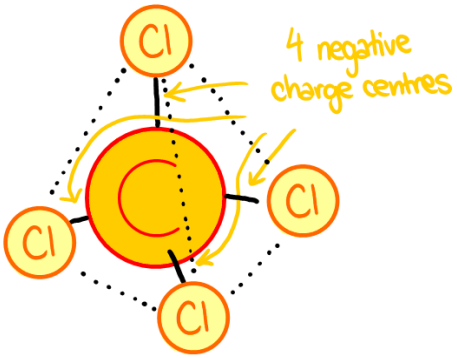
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
3	4	1	117^{o}



Tetrahedral

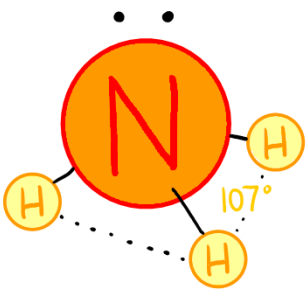
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
4	4	0	109.5^{o}

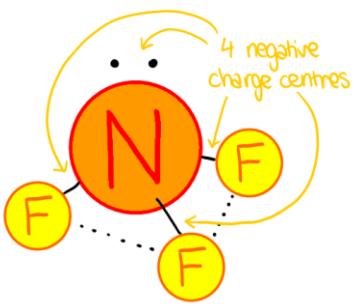




Trigonal Pyramidal

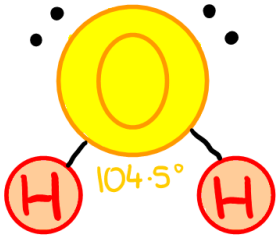
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
4	3	1	107^{o}

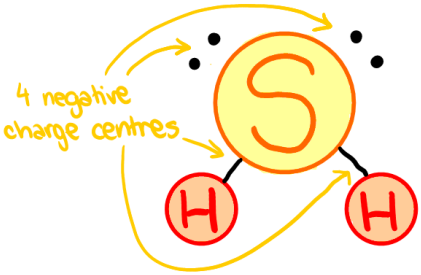




Bent Linear

Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
4	2	2	104.5^{o}





4.2.8 - Predict whether or not a molecule is polar from its molecular shape and bond polarities

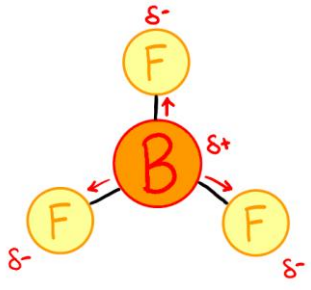
For a molecule to be non-polar, the bonded atoms must be within 0.4 of each other in electronegativity. The electrons would be evenly shared between them. Molecules with **permanent dipoles** that cancel each other out are also non- polar.

If the electrons are shared unevenly, the bond is polar, or a permanent dipole. The symbol delta δ is used to indicate the slight charge that result, especially **δ-** and **δ+**

The electrons are more likely to be found around the atom with the negative dipole.

In **diatomic molecules**, one part will be more negative, as one atom attracts the electrons towards it more.

When there is more than one polar covalent bond, the shape must be considered, as dipoles can cancel each other out. Any molecule that is perfectly symmetrical with be non-polar overall.



4.2.9 - Describe and compare the structure and bonding in the three allotropes of carbon (diamond, graphite and C_{60} fullerene)

The different forms of carbon are known as **allotropes**, and they have vastly different appearance and physical properties. The way they are covalently bonded makes them so different.

As well as the three below, carbon may bond to form

- Carbon nanotubes
- Lonsdaleites
- Chaoite (existence disputed)
- Fullerite

And many others, whose existence has not yet been confirmed

Diamond

In diamond, **all the valence electrons of each carbon atoms are bonded** in a giant covalent network lattice structure. As a result, it cannot conduct electricity. The bonding forces are all very strong, making it very hard. It also has a high sublimation point of 3550^{o}C, at which point it turns straight into a gas.

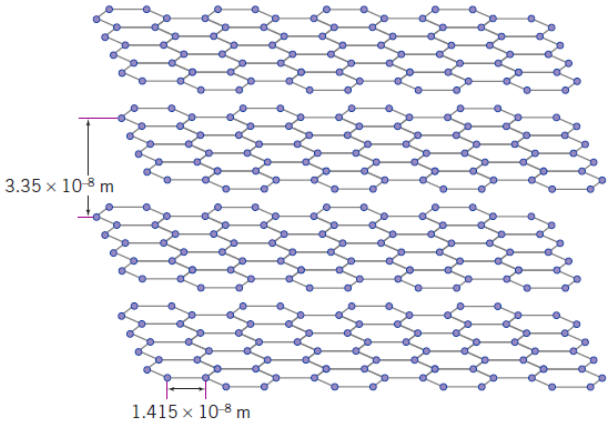
The lattice is very regular, so diamond can be cut in very specific direction known as **cleavage panes**, the lines along which the atoms in the lattice align perfectly. Diamond cutters must identify these panes to optimise the beauty of the cut diamond.

Diamonds are the hardest known naturally occurring substance. Small or flawed stones, known as industrial diamonds, are used on the tips of cutting equipment like drills.

The uses of diamonds include:

- Jewellery
- Cutting tools
- Drills

Graphite



In the structure of graphite, each carbon is **bonded to three others**, forming layers of hexagons. The free electrons explain why it is such a good conductor of electricity. The layers have weak forces between them, making it easy for them to slide past each other. Graphite has a high melting point of 3730°C, so it can still **lubricate** at the high temperatures inside engines and other moving mechanical structures. Its uses include:

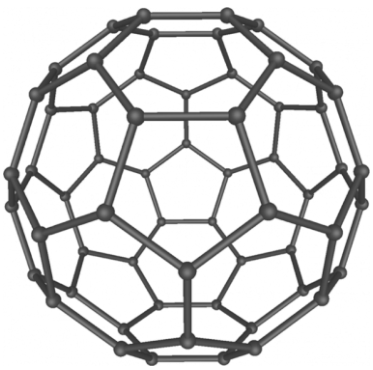
- Good conductor of electricity
- Lubricant
- Additive to rubber and plastic to make them more flexible
- Moderator in nuclear reactors
- High-strength composite material that is light and flexible
    - Tennis racquets
    - Fishing rods
- Pencils
- Electrodes

Fullerenes (Buckyballs)

Buckyballs have a roughly spherical group of covalently bonded carbons arranged in hexagons and pentagons, similar to a soccer ball.

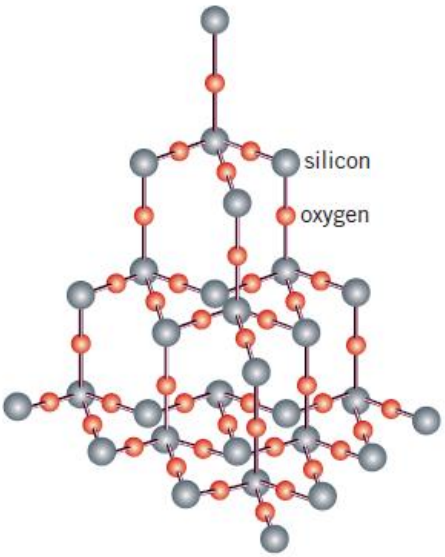
Each carbon atom is bonded to three others, leaving **one free electron per carbon atom**. This enables fullerenes to conduct electricity. They have many uses including

- Superconductivity
- Broad-spectrum lasers
- Catalyst
- Medical uses
- Electromagnetic devices



4.2.10 - Describe the structure of and bonding in silicon and silicon dioxide

Silicon is a member of **group 4** on the periodic table, enabling it to make **four covalent bonds** with other silicon atoms. As silicon is larger than carbon, its bond length is longer. Less energy is required to break a Si-Si bond, so it is more reactive than diamond.



Silicon dioxide (or silica) is a major constituent of sand and is **used to make glass**. It also has a network lattice structure that is made up of alternating silicon and oxygen atoms. It will occur in sandstone, silica sand or quartzite. Its crystal form is quartz. Silica is one of the most abundant oxide materials in the Earth's crust.

Each oxygen atom is covalently bonded to two silicon atoms, whilst each silicon atom is bonded to four oxygen atoms.

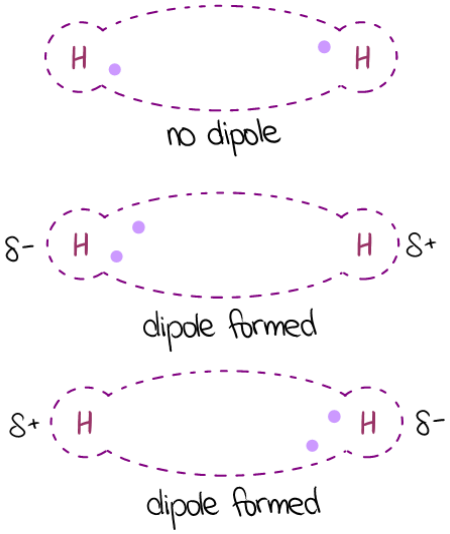
4.3 - Intermolecular Forces

4.3.1 - Describe the types of intermolecular forces (attractions between molecules that have temporary dipoles or hydrogen bonding) and explain how they arise from the structural features of molecules

Intermolecular forces are the bonds that form between molecules, its strength determined by the electrostatic attraction between them. This is dependent on the **size of the molecules** and their **polarity**. Intermolecular forces determine the melting and boiling points of the substance.

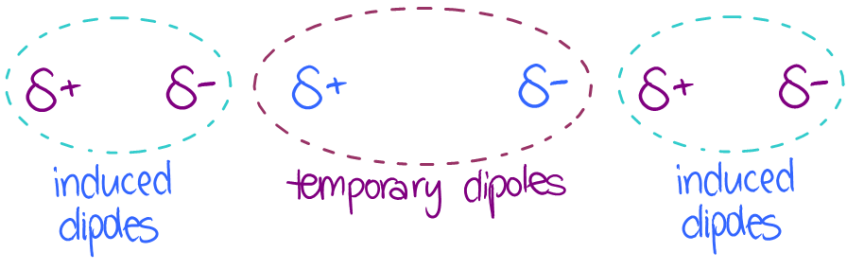
Van der Waals' Forces

These are weak forces between **all molecules** when temporary dipoles form.



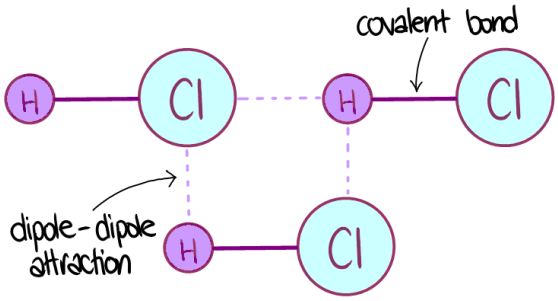
Electrons are not always symmetrically arranged, creating the instantaneous dipoles which in turn can induce a charge in surrounding molecules. This causes a weak overall attractive force. All molecules have Van der Waals' forces, regardless of the other intermolecular forces between them.

Lager molecules have more pronounced forces because there are more electrons. The magnitude of the dipoles increases, causing higher boiling and melting points.

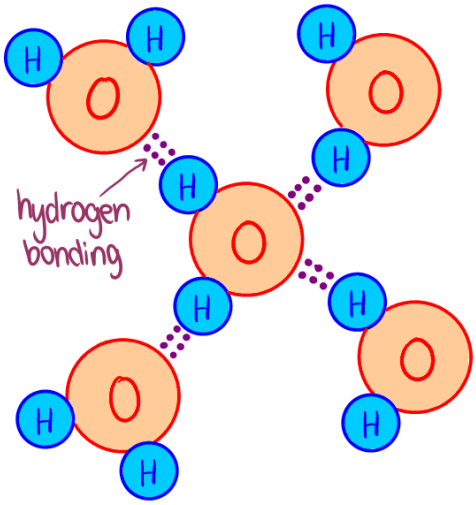


Permanent Dipoles

All **polar molecules** have permanent dipoles; with same ends have either a positive or negative charge. This slight charge creates <u>electrostatic attraction</u> between the molecules which is stronger than temporary dipoles, as the **dipoles are larger**.



Therefore, polar molecules will have <u>higher melting and boiling points</u> than non-polar molecules because their intermolecular forces are stronger.



Hydrogen Bonding

Hydrogen bonding is a **stronger dipole-dipole attraction**. It is the result of the exposure of the hydrogen nucleus, since its electrons are pulled towards the other atom. This results in a strong, permanent <u>positive dipole</u> on the hydrogen end of the molecule.

Molecules with hydrogen bonding are highly polar and have strong dipoles. This only occurs between molecules with **H-F**, **H-O** and **H-N** bonds.

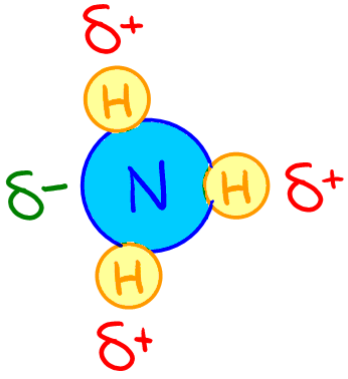
The strong positive end is attracted to the negative end of another molecule.

Hydrogen bonding in **water** explains its unusual properties:

- High melting point and boiling point
- High surface tension

Ammonia also has a high melting point

- Weaker bond than in water due to smaller electronegativity of nitrogen



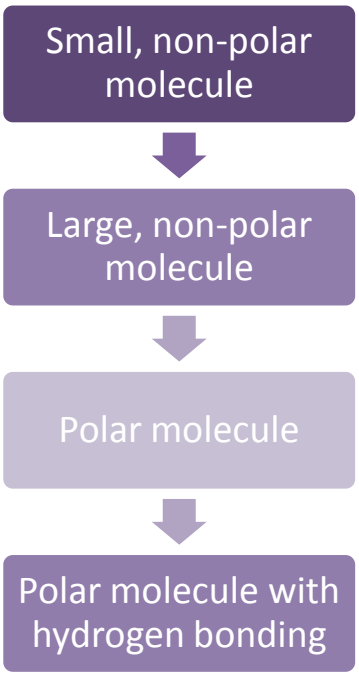
4.3.2 - Describe and explain how intermolecular forces affect the boiling points of substances

Hydrogen bonds are very strong, meaning that more energy is required to break them. This raises the melting and boiling point.

Van der Waals forces (temporary dipoles) are present in all molecules, but larger molecule with more electrons will have larger dipoles. The increased attraction means that larger molecules have a higher boiling point. In addition, more linear molecules have larger temporary dipoles, giving them a higher boiling point.

Permanent dipoles in polar molecules are stronger, requiring more energy to overcome and giving them a higher boiling point.

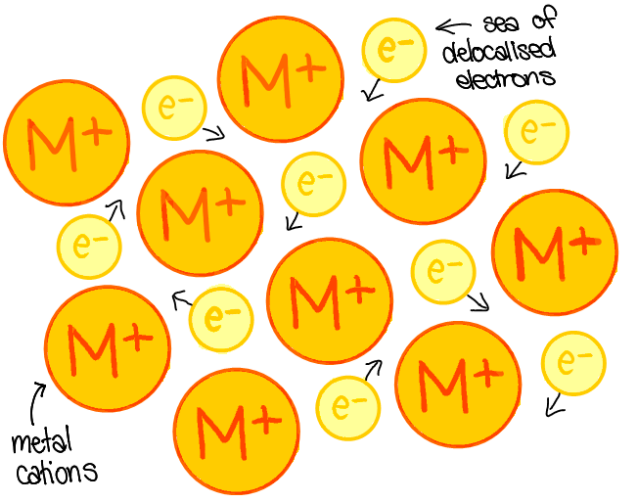
Boiling point essentially increases by:



4.4 - Metallic Bonding

4.4.1 - Describe the metallic bonds as the electrostatic attraction between a lattice of positive ions and delocalised electrons

Metals form a **3D lattice of cations** surrounded by a 'sea' of **delocalised electrons**. Only the valence electrons become delocalised. The electrons in this arrangement reduce repulsion between the ions.



Malleability - the ability to bend without breaking

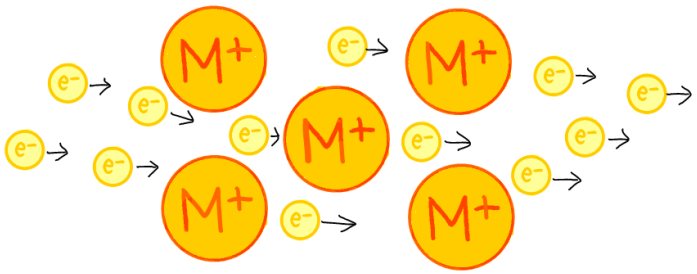
Ductility - the ability to be drawn into a wire

By releasing the valence electrons, the metal ions have noble gas configuration. The electrons move about freely between the cations, maintaining a neutral charge all the time.

4.4.2 - Explain the electrical conductivity and malleability of metals

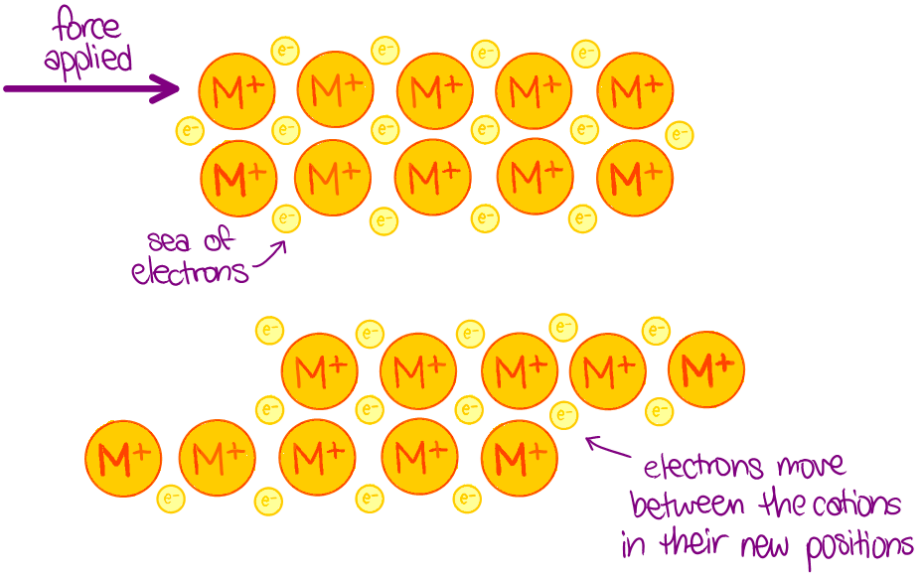
Electrical conductivity

This is a result of the **delocalised electrons**, move freely and respond to the application of a potential difference. The cations will vibrate and cause a barrier to the smooth flow of electrons. Some energy is lost, causing the metal to heat when electricity passes though.



Malleability

Since the cations are surrounded by a sea of delocalised electrons, the electrons are able to move about when the metal is bent, and prevent it from breaking.



4.5 – Physical Properties

4.5.1 - Compare and explain the properties of substances resulting from different types of bonding

Type of bonding	Ionic	Metallic	Covalent molecular	Covalent network lattice
Melting and boiling point	Very high	Medium to high	Low	Very high
Why	Packed closely in the lattice. Bond is strong and require a large amount of energy to break the bond	Depends on the size and charge of the ions. Metallic bonding is stronger than intermolecular forces	Weak bonds between molecules	Covalent bonds between all the atoms require large amounts of energy to break
Volatility	Low	Usually low (exception Mercury)	High	Low
Why	Solid at room temperature and have a high boiling point. Difficult to vaporise	Solid at room temperature, high boiling point. Mercury is a liquid.	Easy to break the weak intermolecular forces between molecules.	Usually solids at room temperature and high boiling points.
Electrical conductivity	Conduct when dissolved in water or in the liquid state	All conduct in the solid state	Polar molecules conduct when they are dissolved. Non-polar ones do not conduct at all	Graphite is the only one that will conduct electricity
Why	Only free to move in these conditions	Sea of delocalised electrons allow electrical current to be conducted	Polar molecules form ions when they dissolve	Delocalised electrons in graphite. Other do not have free particles or electrons
Solubility	Some will dissolve in a polar solvent, while others will not dissolve at all	Some metals react with water to make alkaline solutions, others will not dissolve at all	Polar molecules dissolve in polar substances, non-polar molecules dissolve in non-polar substances	Do not dissolve in any solvents
Why	The attraction to the solvent must be greater than the attraction to the other ions	Ones that react are excellent reducing agent, and undergo a redox reaction with water. The rest have bonds that are too strong for the water to break	Polar molecules are ionised as they dissolve. Non-polar molecules bond using Van der Waals' forces	The covalent bonding is not influenced by the solvents