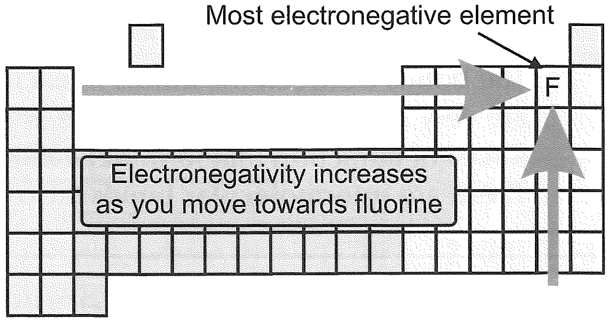
Polarity

Some Atoms Attract **Bonding Electrons** More **Strongly**

The ability of an atom to **attract** electrons in a **covalent bond** is called its **electronegativity**.

1) Ignoring Group 0, electronegativity **decreases** down a **group** in the periodic table, and **increases** across a **period**.
   The **most** electronegative element is **fluorine**.
2) In a bond between two **different** elements with different **electronegativities**, the **bonding electrons** will be attracted **more strongly** towards the atom with the **higher** electronegativity. This makes the bond **polar**.



Polar Bonds Can Affect the **Strength** of **Intermolecular Forces**

If you substitute a chlorine atom for one of the hydrogen atoms in a methane molecule, it has a marked effect on the boiling point.

Boiling point of methane (CH_4)	= –161 °C
Boiling point of chloromethane (CH_3Cl)	= –24 °C

The reason for the dramatic increase in boiling point is that the chlorine atom **polarises** the molecule, making one end **slightly positive** and the other **slightly negative**. The **oppositely charged** ends of **different** molecules **attract** each other, so more energy is required to separate them. This results in an increase in boiling point.

'δ+' means there's a small positive charge on the carbon atom

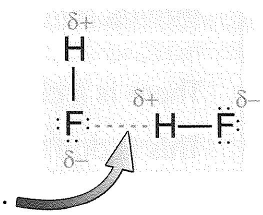
$\overset{\delta+}{H_3C}—\overset{\delta-}{Cl}\cdots\overset{\delta+}{H_3C}—\overset{\delta-}{Cl}\cdots\overset{\delta+}{H_3C}—\overset{\delta-}{Cl}$

'δ–' means there's a small negative charge on the chlorine atom

Hydrogen Bonding is the **Strongest** Type of Intermolecular Force

Molecules that contain a **fluorine**, **oxygen** or **nitrogen** atom **bonded** to a **hydrogen** atom can form strong intermolecular bonds.

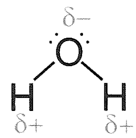
This is because the hydrogen atoms are strongly **polarised** by the very electronegative fluorine, oxygen or nitrogen atoms. These slightly positive hydrogen atoms are attracted to the lone pair of electrons on a F, O or N atom in a **nearby molecule** to form an attraction known as a **hydrogen bond**.



Hydrogen bonds are the strongest type of intermolecular attraction, though they are not as strong as either an ionic or a covalent bond (see next section).

Polar Bond — the Arctic's answer to 007...

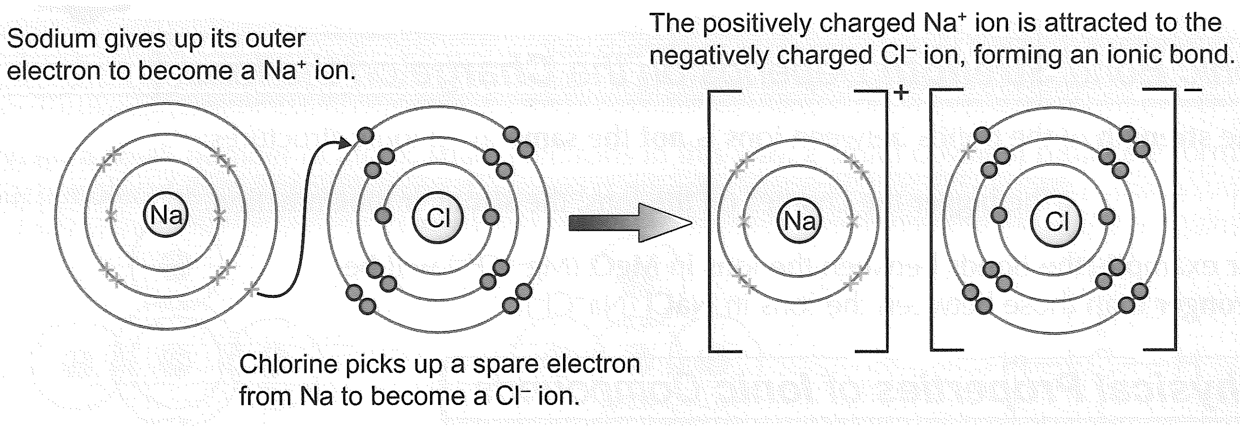
1) For the following pairs of molecules, predict with reasoning which has the higher boiling point:
   a) H_2 and HF, b) H_2O and H_2S, c) CH_3F and CH_3I.
2) Water is a polar molecule. Draw a diagram showing three water molecules attracted together. You should use dotted lines to indicate forces between atoms in different molecules. The shape of a water molecule is shown on the right.



Ionic Bonding

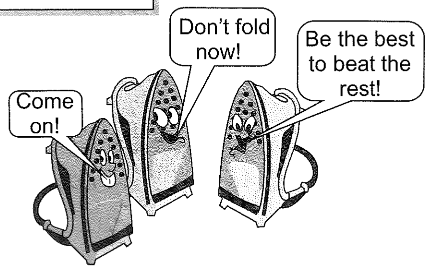
Ionic Bonds *Involve the* ***Transfer*** *of Electrons*

1) Ions form when **electrons** are transferred from **one atom** to **another**. Atoms that **lose electrons** form **positive ions** and atoms that **gain electrons** become **negative ions**.
2) These oppositely charged ions are **attracted** to each other by **electrostatic attraction**. When this happens, an **ionic bond** is formed.
3) The simplest ions form when atoms lose or gain 1, 2 or 3 electrons to get a **full outer shell**.
4) You can show the transfer of electrons to form an ionic compound using a **dot-and-cross** diagram. For example, sodium and chlorine will react to form sodium chloride (NaCl):



5) In the example above, the **dots** represent the electrons that come from the chlorine atom, and the **crosses** represent the electrons that come from the sodium atom.

You Can Find The ***Ratio*** *of* ***Positive*** *to* ***Negative*** *Ions*



1) The **ratio** of positive ions to negative ions in an ionic compound depends on the **charges** of the ions.
2) The **overall charge** of an ionic compound is **zero**, so the **sum** of all the **positive charges** in the compound must be **equal** to the **sum** of the **negative charges**.
3) If you know the **individual charges** of each of the ions in a compound, you can work out their **ratio**. You can use this to find the **ionic formula** of the compound.

> **EXAMPLE:** In the compound calcium chloride, what is the ratio of Ca^{2+} to Cl^- ions?
>
> For the compound to be neutral it must contain **two Cl^-** ions (2×-1) to **balance** the charge of **each Ca^{2+}** ion ($1 \times +2$).
> So the ratio of Ca^{2+} ions to Cl^- ions in the compound must be **1 : 2**.
> The ionic formula will be **$CaCl_2$**.

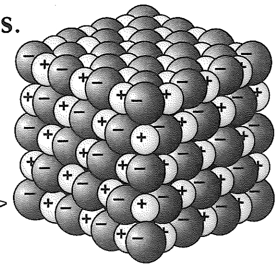
I can't afford Mg^{2+} — the charge is just too high...

1) Draw a diagram showing how a magnesium atom reacts with an oxygen atom to form magnesium oxide, MgO. Your diagram should show the electron transfer process.
2) In potassium oxide, what is the ratio of K^+ ions to O^{2-} ions? What is the ionic formula?

Ionic Compounds

Ionic Bonds Produce Giant Ionic Structures

1) Ionic bonds do not work in any particular direction.
The electrostatic attraction is just as strong in **all directions** around the ion.
2) This means that when ionic compounds form, they produce **giant lattices**.
3) The lattice is a closely packed **regular** array of ions, with each negative ion **surrounded** by positive ions and vice versa.
The **forces** between the **oppositely charged** ions are very **strong**.
4) **Sodium chloride** forms a lattice like this one.
This is called the sodium chloride structure.



Ionic Bond Strength Depends on the Charge on the Ions

The **strength** of the bonds between ions is **not the same** in all ionic structures:

The **bigger** the charges on the ions, the **stronger** the attraction.

For example, the bonds between the ions in **MgO** ($Mg^{2+}O^{2-}$) will be **stronger** than those between the ions in **NaCl** ($Na^{+}Cl^{-}$).

Physical Properties of Ionic Compounds

Melting points

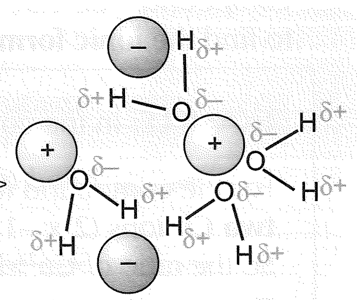
In order to **melt** a solid, the forces holding the particles together have to be **overcome**. In an ionic solid, these bonds are very **strong**, so a **large** amount of energy is required to break them. So, ionic compounds have very **high** melting points.

Electrical conductivity

In their solid form, ionic compounds are electrical **insulators** (they don't conduct electricity).
They have **no free ions** or electrons to carry electric current.
When **molten** or **dissolved**, the ions **separate** and are **free** to move and conduct electricity.
So **all** ionic compounds **conduct** electricity when **molten** or **dissolved**.

Solubility

In many cases ionic compounds are **soluble** in water.
This happens because water is a **polar** molecule (see page 10) — the positive end of the molecule points towards the negative ions and the negative end towards the positive ions.
Although **lots of energy** is required to break the strong bonds within the lattice, it is provided by the formation of **many weak bonds** between the water molecules and the ions in solution.



Rabbits love studying ionic compounds — all those giant lettuces...

1) Put these ionic compounds in order of melting point, highest to lowest: Lithium oxide (Li_2O), Beryllium oxide (BeO), Lithium fluoride (LiF). Explain why you have put them in that order.
2) Explain why the ionic compound, potassium chloride (KCl), can conduct electricity when molten or dissolved, but not when it is solid.

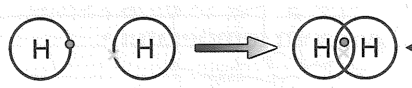
Covalent Bonding

Covalent Bonding *Involves* ***Shared Pairs*** *of Electrons*

Ionic bonding only really works between elements that have to gain or lose one or two electrons to get a full outer shell. Elements with **half-full** shells have to do something different.
These elements **share** their electrons with another atom so they've both got a full outer shell.
Both positive nuclei are **attracted** to the shared pair of electrons.
This results in the formation of **covalent bonds**.

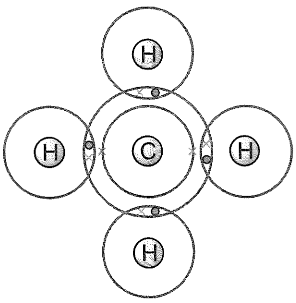
A covalent bond is a **shared pair** of electrons.

For example, two hydrogen atoms share a pair of electrons to form a covalent bond:

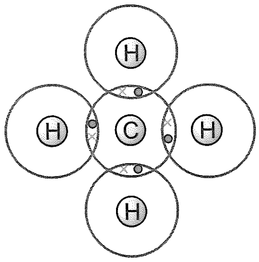


The shared pair of electrons means that each hydrogen atom in the molecule has 2 electrons in its outer shell.

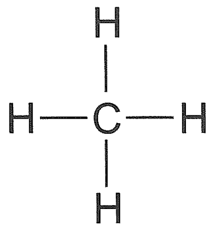
When a small number of atoms share electrons in this way, a small covalent molecule forms. Such molecules can be represented in several different ways:



Dots represent electrons from the Hs and crosses represent electrons from C.

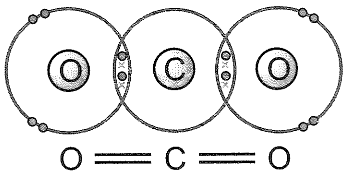


A more simple dot-and-cross diagram, showing only the outer shells of electrons.



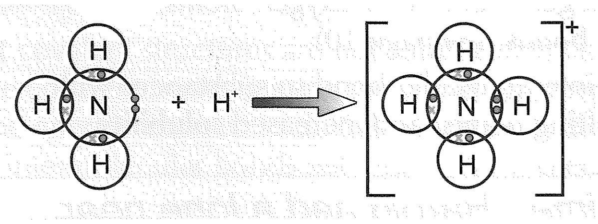
Each dash represents a single covalent bond (this is the most common notation).

If two atoms share **more than one** pair of electrons between them, then a **multiple covalent bond** can form. For example, in **carbon dioxide** (CO_2), there are two C=O double bonds:

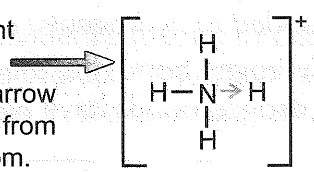


Dative *Covalent Bonding*

In **dative covalent bonds**, **both** of the shared electrons in the covalent bond come from the **same atom**. For example, in the ammonium ion (NH_4^+) there is a dative covalent bond formed from the nitrogen to a hydrogen ion (H^+):



Dative covalent bonds can be shown by an arrow pointing away from the 'donor' atom.



Friendly atom with GSOH WLTM special someone to share a bond with...

1) Draw simple 'dot-and-cross' diagrams to show the bonding in the following molecules:
a) chlorine (Cl_2) b) water (H_2O) c) ethane (C_2H_6) d) oxygen (O_2)

Small Covalent Molecules

***Properties** of Small Covalent Compounds*

Small covalent compounds are made up of **lots** of small covalent molecules.
There are **strong covalent bonds** between the **atoms** in each molecule, but very **weak**, **intermolecular bonds** between the individual molecules (see page 9). It is these intermolecular bonds that determine the physical properties of small covalent compounds.

Melting points

In order to **melt** (or boil) a small covalent compound, you just have to break the **weak** intermolecular bonds **between** the molecules (not the strong covalent bonds). This doesn't need much energy, so small molecules normally have very **low** melting and boiling points — they're often liquids (e.g. water, H_2O) or gases (e.g. oxygen, O_2) at room temperature.

Electrical Conductivity

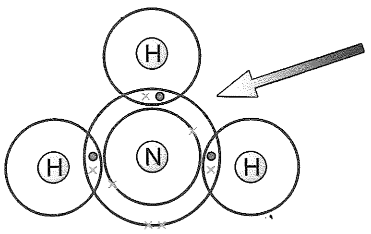
Small covalent molecules don't contain any of the **free charged** particles that are needed to carry an electric current. As a result they **cannot** conduct electricity — they're electrical **insulators**.

Solubility

This **varies** depending on the **type** of molecule. Small covalent molecules that are **not polar** at all (e.g. hydrocarbons) **don't mix** well with water, or dissolve very well in it. This is because the attractive force that exists between **two water molecules** is much **stronger** than that between a water molecule and a non-polar molecule.
Small covalent molecules that are **polar** or can form **hydrogen bonds** (see page 10) **can** dissolve in water.

Lone Pairs** Can Affect the **Physical Properties

1) You've already seen some examples of small covalent molecules on page 13. You may have noticed that **not all** of the electron pairs around the central atom are bonding electrons. In other words, not all of the electrons are **shared** between the atoms in the molecule.



2) In ammonia (NH_3) there are **4 electron pairs** around the central nitrogen atom. **Three** of these electron pairs are called **bonding pairs** as they are **shared** between the **nitrogen** and **hydrogen atoms**. The **fourth** electron pair is **not shared** between the atoms in the molecule. This is called a **lone pair**.
3) Covalent molecules with lone pairs on nitrogen, fluorine or oxygen atoms, bonded to hydrogen(s) can form **hydrogen bonds** (see page 10).
4) Hydrogen bonds are the **strongest type** of intermolecular bond so substances with hydrogen bonds have high **boiling** and **melting** points, and increased **solubility**.

Aisling had four satsumas at lunchtime. Harold had a lone pear...

1) Draw a dot-and-cross diagram to show the bonding in hydrogen fluoride (HF). Label the bonding electrons and lone pairs of electrons.
2) Explain why nitrogen is a gas at room temperature, despite the nitrogen atoms in each molecule being strongly bonded to each other.

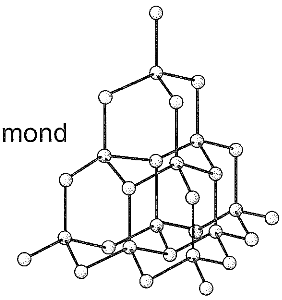
Giant Covalent Molecules

Giant Covalent Structures

Carbon is ideally placed to share electrons and form covalent bonds, because it has a **half-full** outer shell. Carbon atoms can share their electrons with four other carbons to gain a full outer shell. This can result in the formation of a single massive carbon molecule — a **giant structure**.

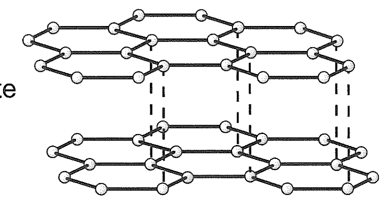
Carbon can form various different **giant covalent structures** such as **diamond** and **graphite**.

diamond



Each carbon atom forms **four** covalent bonds in a very **rigid** structure.
This structure makes diamond very **hard**.

graphite



Each carbon atom forms **three** covalent bonds in the same **plane**. This results in a series of **layers** which can **slide** over each other.
The fourth electron from each carbon atom is **free**.

Properties of Giant Covalent Structures

Giant covalent structures have some different **physical properties** from small molecules.

Melting points

Unlike small molecules, melting points are **extremely high**, as all of the atoms are held together by **strong covalent bonds**. These millions of covalent bonds need to be **broken** to allow the atoms within the structure to move freely, which requires a lot of energy.

This contrasts with small molecules where no covalent bonds (only intermolecular bonds) need to be broken in order for the substance to melt.

Electrical conductivity

Giant covalent structures are **electrical insulators**. This is because they don't contain **charged particles**, and the atoms aren't free to move.
Even a **molten** covalent compound will not conduct electricity.
Graphite is the only exception to this, as the loosely held **electrons** between the layers of atoms can move through the solid structure. Graphite conducts in both its solid and liquid forms.

Solubility

Giant covalent structures are **not soluble** in water. To get a giant covalent structure to dissolve, all the covalent bonds joining the atoms together would need to be **broken**. There is no way to get the energy required to do this, since the individual **neutral atoms** in the structure will **not** form intermolecular bonds with the water molecules.

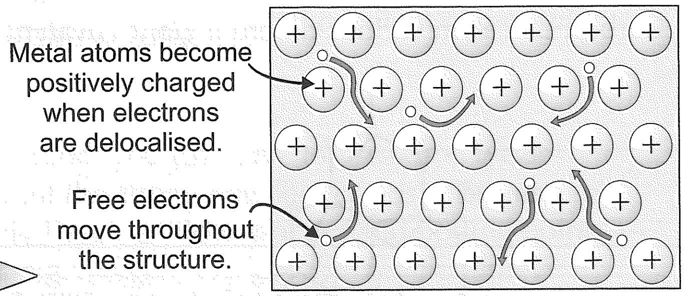
Diamonds — don't mess with 'em — they're well 'ard...

1) Devise a series of tests that would allow you to distinguish between two unknown crystalline solids, one of which is an ionic compound and the other a giant covalent structure.
2) Why won't diamond dissolve in water when sodium chloride will?

Metallic Bonding

Metals have Giant Structures Too

1) In a metal, the **outer electrons** from each atom are **delocalised** (they're not stuck on one atom) — this leaves **positive metal ions.**
2) The positive metal ions are arranged regularly in a **giant structure**, surrounded by a 'sea' of delocalised electrons.
3) Metals are held together because of the **electrostatic attractions** between the **positive metal ions** and the **delocalised 'sea'** of **electrons.**
   This is called **metallic bonding.**



Properties of *Metals*

Metallic bonding explains the **physical properties** of metals:

Melting points

Metals generally have **high** melting points. This is because a lot of energy is required to overcome the **strong metallic bonding** between the particles.

The **more** electrons that are **delocalised** from **each atom**, the **stronger** the bonding will be and the **higher** the melting point.

> **EXAMPLE:** Predict, with reasoning, whether magnesium or sodium will have a higher melting point.
>
> Magnesium is made up of **Mg^{2+}** ions with **two** delocalised electrons per atom.
> Sodium is made up of **Na^{+}** ions and only **one** delocalised electron per atom.
>
> So **magnesium** will have a **higher melting point** than sodium, because the metallic bonds will be **stronger** and require **more energy** to break.

Electrical conductivity

The **delocalised electrons** in metals are **free to move** around and can carry a **current**. This makes metals **good electrical conductors**.

Solubility

The **strong metallic bonds** mean that metals are generally **insoluble**.

Metallica bonds — friendships based on a love of '80s rock music...

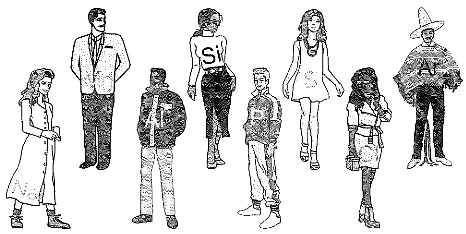
1) Predict, with reasoning whether potassium or calcium will have a higher melting point.
2) Draw a diagram to show the bonding in a sample of sodium.
3) Sodium has a metallic structure, whilst sodium chloride (NaCl) is an ionic compound. Give one similarity and one difference between the physical properties of these substances.

Trends in Properties Across the Periodic Table

***Structure** and **Bonding Change** Across the Periodic Table*

You should have seen from this section how much the **properties** of a compound depend on its **bonding** and **structure**. You also know that the type of bonding that occurs depends on the **number of electrons** in the outer shells of the elements making up the compound, and so their **positions** in the periodic table.

A good way to compare the way that different elements bond is by looking at the properties of a series of similar compounds across a **period**.
Look at the information in the table below about all the Period 3 oxides.
You can see that there are clear **patterns** in the data.



Trends** Across **Period Three

(Period 3 is studied because it is a simple case.
There are no d-block elements to confuse matters.)

The table below shows some of the physical properties of Period 3 oxides.
The final row has been deduced from these physical properties.

	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_2
State (at room temperature and standard pressure)	solid	solid	solid	solid	solid	gas
Melting point (°C) (at standard pressure)	1275	2800	2072	1650	570	–73
Electrical conductivity (when molten)	good	good	good	none	none	none
Bonding	ionic lattice	ionic lattice	ionic lattice	giant covalent structure	small covalent molecule	small covalent molecule

You can see from this data that there is a change in the properties of the Period 3 oxides as you move from left to right across the periodic table.

The trend is from **ionic** bonding to **small covalent** molecules via a **giant covalent structure**.

These trends across a period are **more subtle** than the trends going down a group that you saw at GCSE. However they are extremely useful as they allow you to make **predictions** about the reactions and properties of unknown compounds. There are of course **exceptions** to the rules/trends, but on the whole they allow links between physical properties and atomic structure to be made.

Trending now — #Arewedoneyet? #Dontworryitstheendofthesection...

1) Explain how the data in the first three rows of the table above supports the idea that the bonding type changes from ionic to covalent as you move across Period 3.
2) Use the information on Period 3 oxides to predict the trend in the melting points of the elements as you go across Period 3.
3) Predict the type of bonding you would expect in the chlorides of:
   a) sodium b) phosphorus