Formation of Ions

Ions crop up all over the place in chemistry. You're gonna have to be able to explain how they form and predict the charges of simple ions formed by elements in Groups 1, 2, 6 and 7. You'd better get on...

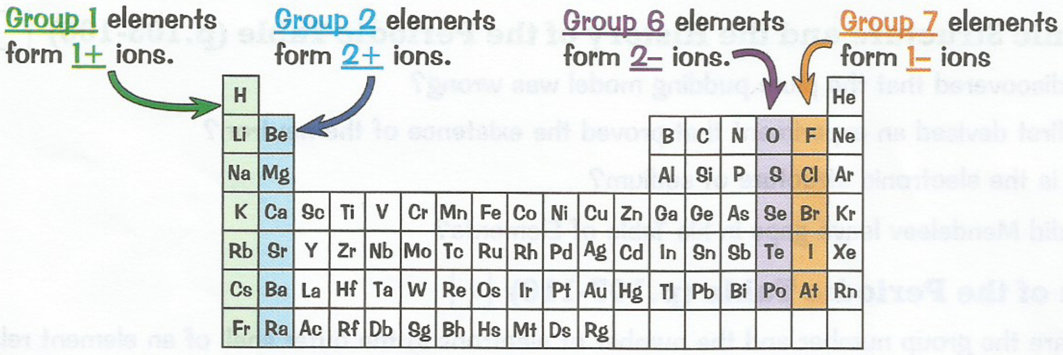
Ions are Made When Electrons are Transferred

1) Ions are charged particles — they can be single atoms (e.g. Cl^-) or groups of atoms (e.g. NO_3^-).
2) When atoms lose or gain electrons to form ions, all they're trying to do is get a full outer shell like a noble gas (also called a "stable electronic structure"). Atoms with full outer shells are very stable.
3) When metals form ions, they lose electrons from their outer shell to form positive ions.
4) When non-metals form ions, they gain electrons into their outer shell to form negative ions.
5) The number of electrons lost or gained is the same as the charge on the ion. E.g. If 2 electrons are lost the charge is 2+. If 3 electrons are gained the charge is 3–.

Remember that the noble gases are in Group 0 of the periodic table.

Groups 1 & 2 and 6 & 7 are the Most Likely to Form Ions

1) The elements that most readily form ions are those in Groups 1, 2, 6 and 7.
2) Group 1 and 2 elements are metals and they lose electrons to form positive ions (cations).
3) Group 6 and 7 elements are non-metals. They gain electrons to form negative ions (anions).
4) You don't have to remember what ions most elements form — nope, you just look at the periodic table.
5) Elements in the same group all have the same number of outer electrons. So they have to lose or gain the same number to get a full outer shell. And this means that they form ions with the same charges.



- A sodium atom (Na) is in Group 1 so it loses 1 electron to form a sodium ion (Na^+) with the same electronic structure as neon: $Na \rightarrow Na^+ + e^-$.
- A magnesium atom (Mg) is in Group 2 so it loses 2 electrons to form a magnesium ion (Mg^{2+}) with the same electronic structure as neon: $Mg \rightarrow Mg^{2+} + 2e^-$.
- A chlorine atom (Cl) is in Group 7 so it gains 1 electron to form a chloride ion (Cl^-) with the same electronic structure as argon: $Cl + e^- \rightarrow Cl^-$.
- An oxygen atom (O) is in Group 6 so it gains 2 electrons to form an oxide ion (O^{2-}) with the same electronic structure as neon: $O + 2e^- \rightarrow O^{2-}$.

Have a look back at page 104 for how to work out electronic structures.

I've got my ion you...

Some elements like to gain electrons, some elements like to lose electrons, but they all want to have a full outer shell. Poor little electron shells, all they want in life is to be full...

Q1 Explain why simple ions often have noble gas electronic structures. [2 marks]

Q1 Video Solution

Q2 Predict the charges of the ions formed by the following elements:
a) Bromine (Br) b) Calcium (Ca) c) Potassium (K) [3 marks]

Ionic Bonding

Time to find out how particles bond together to form compounds (bet you can't wait). There are three types of bonding you need to know about — ionic, covalent and metallic. First up, it's ionic bonds.

Ionic Bonding — Transfer of Electrons

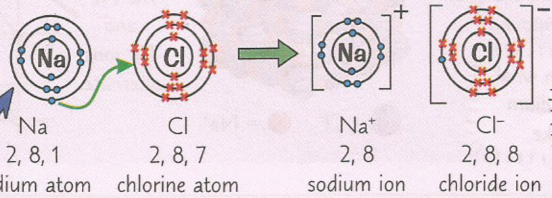
When a metal and a non-metal react together, the metal atom loses electrons to form a positively charged ion and the non-metal gains these electrons to form a negatively charged ion. These oppositely charged ions are strongly attracted to one another by electrostatic forces. This attraction is called an ionic bond.

Dot and Cross Diagrams Show How Ionic Compounds are Formed

Dot and cross diagrams show the arrangement of electrons in an atom or ion. Each electron is represented by a dot or a cross. So these diagrams can show which atom the electrons in an ion originally came from.

Sodium Chloride (NaCl)

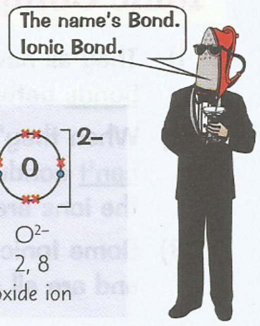
The sodium atom gives up its outer electron, becoming an Na^+ ion. The chlorine atom picks up the electron, becoming a Cl^- (chloride) ion.



Remember, you can work out how many electrons an atom will gain or lose from its group number.

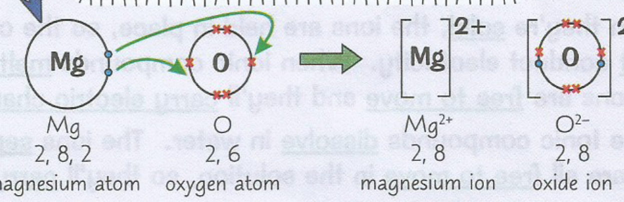
Here, the dots represent the Na electrons and the crosses represent the Cl electrons (all electrons are really identical, but this is a good way of following their movement).

Here we've only shown the outer shells of electrons on the dot and cross diagram — it makes it much simpler to see what's going on.



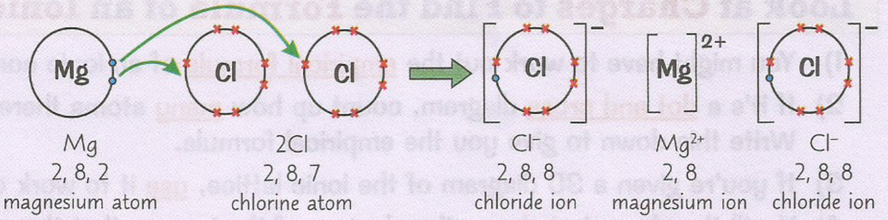
Magnesium Oxide (MgO)

The magnesium atom gives up its two outer electrons, becoming an Mg^{2+} ion. The oxygen atom picks up the electrons, becoming an O^{2-} (oxide) ion.



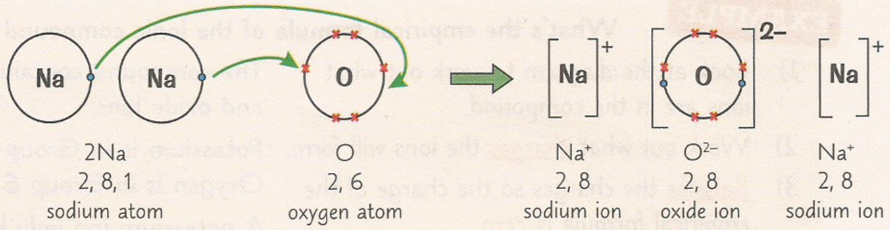
Magnesium Chloride ($MgCl_2$)

The magnesium atom gives up its two outer electrons, becoming an Mg^{2+} ion. The two chlorine atoms pick up one electron each, becoming two Cl^- (chloride) ions.



Sodium Oxide (Na_2O)

Two sodium atoms each give up their single outer electron, becoming two Na^+ ions. The oxygen atom picks up the two electrons, becoming an O^{2-} ion.



Dot and cross diagrams are useful for showing how ionic compounds are formed, but they don't show the structure of the compound, the size of the ions or how they're arranged. But hey-ho — nothing's perfect.

Any old ion, any old ion — any, any, any old ion...

You need to be able to describe how ionic compounds are formed using both words and dot and cross diagrams. It gets easier with practice, so here are some questions to get you started.

Q2 Video Solution

Q1 Describe, in terms of electron transfer, how sodium (Na) and chlorine (Cl) react to form sodium chloride (NaCl). [3 marks]

Q2 Draw a dot and cross diagram to show how potassium (a Group 1 metal) and bromine (a Group 7 non-metal) form potassium bromide (KBr). [3 marks]

Ionic Compounds

I'd take everything on this page with a pinch of salt if I were you... Ho ho ho — I jest, it's important really.

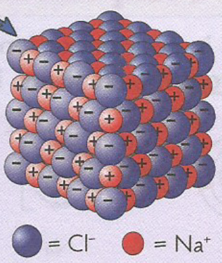
Ionic Compounds Have A Regular Lattice Structure

1) Ionic compounds have a structure called a giant ionic lattice.
2) The ions form a closely packed regular lattice arrangement and there are very strong electrostatic forces of attraction between oppositely charged ions, in all directions in the lattice.

The electrostatic attraction between the oppositely charged ions is ionic bonding.

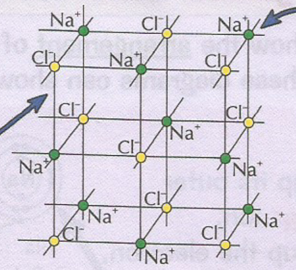
A single crystal of sodium chloride (table salt) is one giant ionic lattice. The Na^+ and Cl^- ions are held together in a regular lattice. The lattice can be represented in different ways...

This model shows the relative sizes of the ions, as well as the regular pattern of an ionic crystal, but it only lets you see the outer layer of the compound.



Make sure you learn what the structure of sodium chloride looks like.

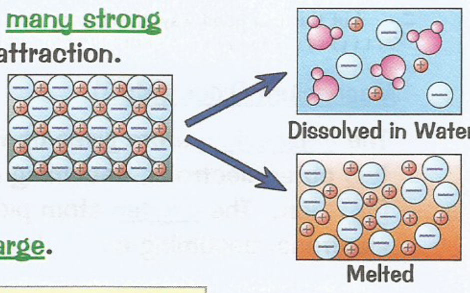
The Na^+ and Cl^- ions alternate.



This is a ball and stick model. It shows the regular pattern of an ionic crystal and shows how all the ions are arranged. It also suggests that the crystal extends beyond what's shown in the diagram. The model isn't to scale though, so the relative sizes of the ions may not be shown. Also, in reality, there aren't gaps between the ions.

Ionic Compounds All Have Similar Properties

1) They all have high melting points and high boiling points due to the many strong bonds between the ions. It takes lots of energy to overcome this attraction.
2) When they're solid, the ions are held in place, so the compounds can't conduct electricity. When ionic compounds melt, the ions are free to move and they'll carry electric charge.
3) Some Ionic compounds dissolve in water. The ions separate and are all free to move in the solution, so they'll carry electric charge.



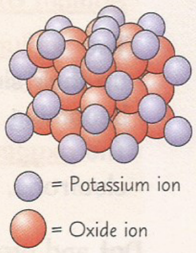
Look at Charges to Find the Formula of an Ionic Compound

1) You might have to work out the empirical formula of an ionic compound from a diagram of the compound.
2) If it's a dot and cross diagram, count up how many atoms there are of each element. Write this down to give you the empirical formula.
3) If you're given a 3D diagram of the ionic lattice, use it to work out what ions are in the ionic compound.
4) You'll then have to balance the charges of the ions so that the overall charge on the compound is zero.

EXAMPLE

What's the empirical formula of the ionic compound shown on the right?

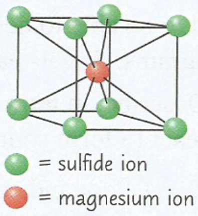
1) Look at the diagram to work out what ions are in the compound.
   The compound contains potassium and oxide ions.
2) Work out what charges the ions will form.
   Potassium is in Group 1 so forms 1+ ions.
3) Balance the charges so the charge of the empirical formula is zero.
   Oxygen is in Group 6 so forms 2– ions.
   A potassium ion only has a 1+ charge, so you'll need two of them to balance out the 2– charge of an oxide ion. The empirical formula is K_2O.



Giant ionic lattices — all over your chips...

Here's where you can get a little practice working out formulas for ionic compounds.

Q1 The structure of an ionic compound is shown on the right.
   a) Predict, with reasoning, whether the compound has a high or a low melting point. [2 marks]
   b) Explain why the compound can conduct electricity when molten. [1 mark]
   c) Use the diagram to find the empirical formula of the compound. [3 marks]



Q1 Video Solution

Covalent Bonding

Some elements bond ionically (see page 113) but others form strong covalent bonds. This is where atoms share electrons with each other so that they've got full outer shells.

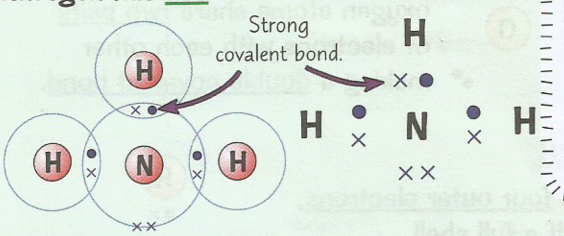
Covalent Bonds — Sharing Electrons

1) When non-metal atoms bond together, they share pairs of electrons to make covalent bonds.
2) The positively charged nuclei of the bonded atoms are attracted to the shared pair of electrons by electrostatic forces, making covalent bonds very strong.
3) Atoms only share electrons in their outer shells (highest energy levels).
4) Each single covalent bond provides one extra shared electron for each atom.
5) Each atom involved generally makes enough covalent bonds to fill up its outer shell. Having a full outer shell gives them the electronic structure of a noble gas, which is very stable.
6) Covalent bonding happens in compounds of non-metals (e.g. H_2O) and in non-metal elements (e.g. Cl_2).

There are Different Ways of Drawing Covalent Bonds

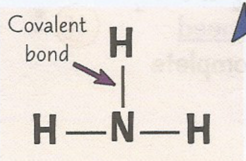
1) You can use dot and cross diagrams to show the bonding in covalent compounds.
2) Electrons drawn in the overlap between the outer orbitals of two atoms are shared between those atoms.
3) Dot and cross diagrams are useful for showing which atoms the electrons in a covalent bond come from, but they don't show the relative sizes of the atoms, or how the atoms are arranged in space.

Nitrogen has five outer electrons...

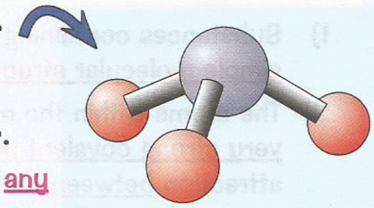


...so it needs to form three covalent bonds to make up the extra three electrons needed.

You don't have to draw the orbitals in these diagrams. The important thing is that you get all the dots and crosses in the right places.



4) The displayed formula of ammonia (NH_3) shows the covalent bonds as single lines between atoms.
5) This is a great way of showing how atoms are connected in large molecules. However, they don't show the 3D structure of the molecule, or which atoms the electrons in the covalent bond have come from.
6) The 3D model of ammonia shows the atoms, the covalent bonds and their arrangement in space next to each other. But 3D models can quickly get confusing for large molecules where there are lots of atoms to include. They don't show where the electrons in the bonds have come from, either.

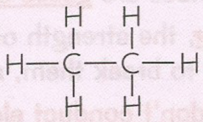


7) You can find the molecular formula of a simple molecular compound from any of these diagram by counting up how many atoms of each element there are.

A molecular formula shows you how many atoms of each element are in a molecule.

EXAMPLE

A diagram of the molecule ethane is shown on the right. Use the diagram to find the molecular formula of ethane.



In the diagram, there are two carbon atoms and six hydrogen atoms. So the molecular formula is C_2H_6.

Sharing is caring...

There's a whole page of dot and cross diagrams for other covalent molecules yet to come, but make sure you can draw the different diagrams that can be used to show the bonding in ammonia on this page first.

Q1 Draw a dot and cross diagram to show the bonding in a molecule of ammonia (NH_3). [2 marks]

Q1 Video Solution

Simple Molecular Substances

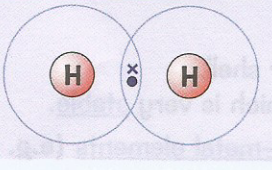
These molecules might be simple, but you've still gotta know about them. I know, the world is a cruel place.

Learn These Examples of Simple Molecular Substances

Simple molecular substances are made up of molecules containing a few atoms joined together by covalent bonds. Here are some common examples that you should know...

Hydrogen, H_2

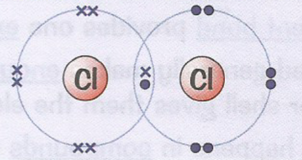
Hydrogen atoms have just one electron. They only need one more to complete the first shell...



...so they often form single covalent bonds, either with other hydrogen atoms or with other elements, to achieve this.

Chlorine, Cl_2

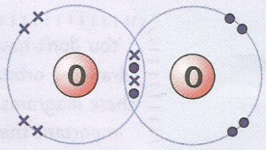
Each chlorine atom needs just one more electron to complete the outer shell...



...so two chlorine atoms can share one pair of electrons and form a single covalent bond.

Oxygen, O_2

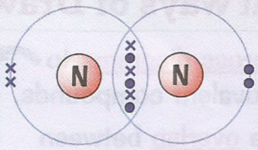
Each oxygen atom needs two more electrons to complete its outer shell...



...so in oxygen gas two oxygen atoms share two pairs of electrons with each other making a double covalent bond.

Nitrogen, N_2

Nitrogen atoms need three more electrons...

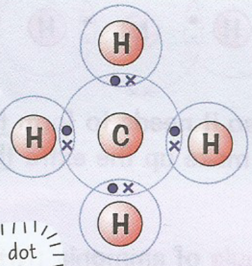


...so two nitrogen atoms share three pairs of electrons to fill their outer shells. This creates a triple bond.

Methane, CH_4

Carbon has four outer electrons, which is half a full shell.

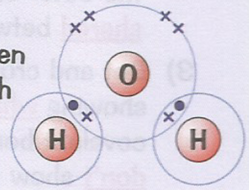
It can form four covalent bonds with hydrogen atoms to fill up its outer shell.



Make sure you can also draw the dot and cross diagram of ammonia, NH_3, which is on the previous page.

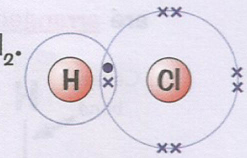
Water, H_2O

In water molecules, the oxygen shares a pair of electrons with two H atoms to form two single covalent bonds.

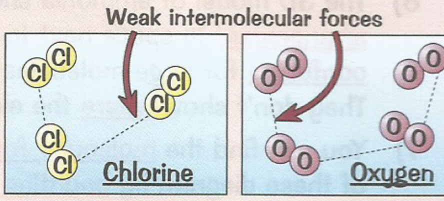


Hydrogen Chloride, HCl

This is very similar to H_2 and Cl_2. Again, both atoms only need one more electron to complete their outer shells.



Properties of Simple Molecular Substances



1) Substances containing covalent bonds usually have simple molecular structures, like the examples above.
2) The atoms within the molecules are held together by very strong covalent bonds. By contrast, the forces of attraction between these molecules are very weak.
3) To melt or boil a simple molecular compound, you only need to break these feeble intermolecular forces and not the covalent bonds. So the melting and boiling points are very low, because the molecules are easily parted from each other.
4) Most molecular substances are gases or liquids at room temperature.
5) As molecules get bigger, the strength of the intermolecular forces increases, so more energy is needed to break them, and the melting and boiling points increase.
6) Molecular compounds don't conduct electricity, simply because they aren't charged, so there are no free electrons or ions.

May the intermolecular force be with you...

Never forget that it's the weak forces between molecules that are broken when a simple molecular substance melts.

Q1 Explain why oxygen, O_2, is a gas at room temperature. [1 mark]

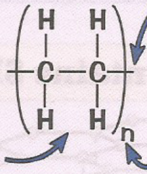
Q2 Explain why nitrogen, N_2, doesn't conduct electricity. [1 mark]

Polymers and Giant Covalent Structures

Wouldn't it be simply marvellous if only simple molecular substances had covalent bonds, and it was now time to put your feet up? Well it's not like that. Polymers and giant covalent substances also have covalent bonds.

Polymers Are Long Chains of Repeating Units

1) In a polymer, lots of small units are linked together to form a long molecule that has repeating sections.
2) All the atoms in a polymer are joined by strong covalent bonds.
3) Instead of drawing out a whole long polymer molecule (which can contain thousands or even millions of atoms), you can draw the shortest repeating section, called the repeating unit, like this:

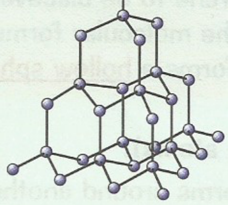


4) To find the molecular formula of a polymer, write down the molecular formula of the repeating unit in brackets, and put an 'n' outside.
5) So for poly(ethene), the molecular formula of the polymer is $(C_2H_4)_n$.
6) The intermolecular forces between polymer molecules are larger than between simple covalent molecules, so more energy is needed to break them. This means most polymers are solid at room temperature.
7) The intermolecular forces are still weaker than ionic or covalent bonds, so they generally have lower boiling points than ionic or giant molecular compounds.

Giant Covalent Structures Are Macromolecules

1) In giant covalent structures, all the atoms are bonded to each other by strong covalent bonds.
2) They have very high melting and boiling points as lots of energy is needed to break the covalent bonds between the atoms.
3) They don't contain charged particles, so they don't conduct electricity — not even when molten (except for a few weird exceptions such as graphite, see next page).
4) The main examples that you need to know about are diamond and graphite, which are both made from carbon atoms only, and silicon dioxide (silica).

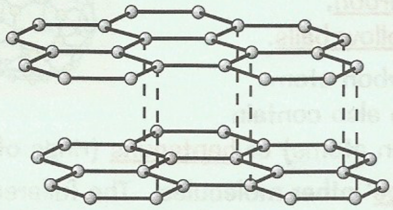
Diamond:



Each carbon atom forms four covalent bonds in a very rigid giant covalent structure.

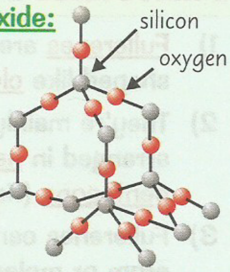
There's more about diamond and graphite, as well as other types of carbon structure, on the next page.

Graphite:



Each carbon atom forms three covalent bonds to create layers of hexagons. Each carbon atom also has one delocalised (free) electron.

Silicon dioxide:



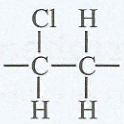
Sometimes called silica, this is what sand is made of. Each grain of sand is one giant structure of silicon and oxygen.

What do you call a vehicle made of sand? Sili-car...

To melt or boil a simple molecular substance or a polymer, only the weakish intermolecular forces need to be broken. To melt or boil a giant covalent substance, you have to break very strong covalent bonds.

Q1 Video Solution

Q1 The repeating unit of poly(chloroethene) is shown on the right. What's the molecular formula of poly(chloroethene)? [1 mark]

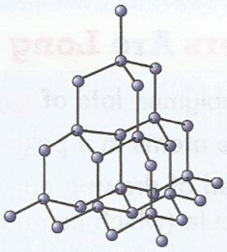


Q2 Predict, with reasoning, whether diamond or poly(ethene) has a higher melting point. [3 marks]

Allotropes of Carbon

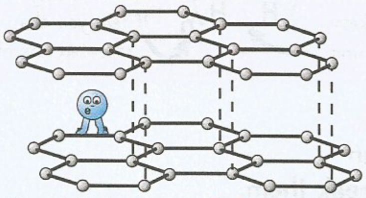
Allotropes are different structural forms of the same element in the same physical state. Carbon's got lots...

Diamond is Very Hard



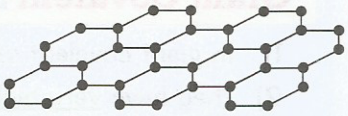
1) Diamond has a giant covalent structure, made up of carbon atoms that each form four covalent bonds. This makes diamond really hard.
2) Those strong covalent bonds take a lot of energy to break and give diamond a very high melting point.
3) It doesn't conduct electricity because it has no free electrons or ions.

Graphite Contains Sheets of Hexagons



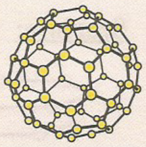
1) In graphite, each carbon atom only forms three covalent bonds, creating sheets of carbon atoms arranged in hexagons.
2) There aren't any covalent bonds between the layers — they're only held together weakly, so they're free to move over each other. This makes graphite soft and slippery, so it's ideal as a lubricating material.
3) Graphite's got a high melting point — the covalent bonds in the layers need loads of energy to break.
4) Only three out of each carbon's four outer electrons are used in bonds, so each carbon atom has one electron that's delocalised (free) and can move. So graphite conducts electricity and thermal energy.

Graphene is One Layer of Graphite



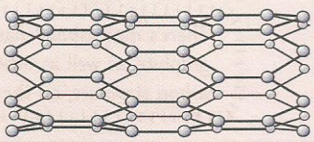
Graphene is a sheet of carbon atoms joined together in hexagons.
The sheet is just one atom thick, making it a two-dimensional substance.
The network of covalent bonds makes it very strong. It's also incredibly light, so can be added to composite materials to improve their strength without adding much weight.
Like graphite, it contains delocalised electrons so can conduct electricity through the whole structure. This means it has the potential to be used in electronics.

Fullerenes Form Spheres and Tubes



Buckminsterfullerene was the first fullerene to be discovered. It's got the molecular formula C_{60} and forms a hollow sphere.

1) Fullerenes are molecules of carbon, shaped like closed tubes or hollow balls.
2) They're mainly made up of carbon atoms arranged in hexagons, but can also contain pentagons (rings of five carbon atoms) or heptagons (rings of seven carbon atoms).
3) Fullerenes can be used to 'cage' other molecules. The fullerene structure forms around another atom or molecule, which is then trapped inside. This could be used to deliver a drug into the body.
4) Fullerenes have a huge surface area, so they could help make great industrial catalysts — individual catalyst molecules could be attached to the fullerenes. Fullerenes also make great lubricants.



Fullerenes can form nanotubes — tiny carbon cylinders.
The ratio between the length and the diameter of nanotubes is very high.
Nanotubes can conduct both electricity and thermal energy (heat).
They also have a high tensile strength (they don't break when they're stretched).
Technology that uses very small particles such as nanotubes is called nanotechnology. Nanotubes can be used in electronics or to strengthen materials without adding much weight, such as in tennis racket frames.

Greetings in the Caribbean — they're 'allo-tropical...

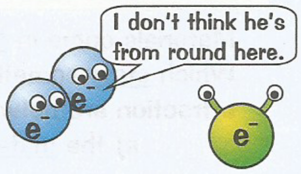
Before you go on, make sure you can explain the properties of all these allotropes of carbon.

Q1 Give three uses of fullerenes. [3 marks]

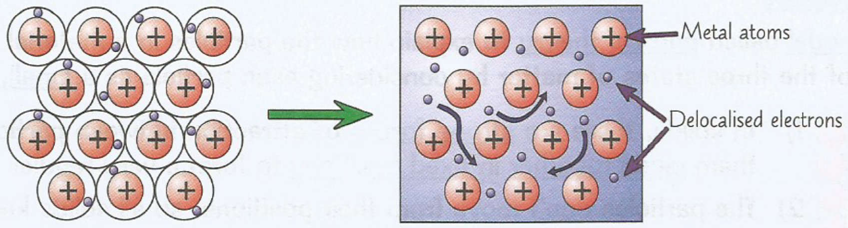
Metallic Bonding

Ever wondered what makes metals tick? Well, either way, this is the page for you.

Metallic Bonding Involves Delocalised Electrons



1) Metals also consist of a giant structure.
2) The electrons in the outer shell of the metal atoms are delocalised (free to move around). There are strong forces of electrostatic attraction between the positive metal ions and the shared negative electrons.
3) These forces of attraction hold the atoms together in a regular structure and are known as metallic bonding. Metallic bonding is very strong.



4) Substances that are held together by metallic bonding include metallic elements and alloys (see below).
5) It's the delocalised electrons in the metallic bonds which produce all the properties of metals.

Most Metals are Solid at Room Temperature

The electrostatic forces between the metal atoms and the delocalised sea of electrons are very strong, so need lots of energy to be broken.

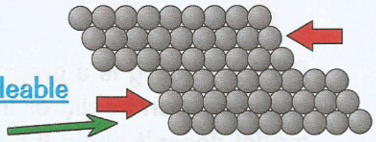
This means that most compounds with metallic bonds have very high melting and boiling points, so they're generally solid at room temperature.

Metals are Good Conductors of Electricity and Heat

The delocalised electrons carry electrical charge and thermal (heat) energy through the whole structure, so metals are good conductors of electricity and heat.

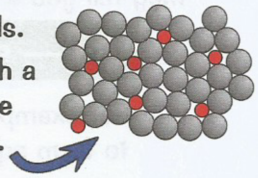
Most Metals are Malleable

The layers of atoms in a metal can slide over each other, making metals malleable — this means that they can be bent or hammered or rolled into flat sheets.



Alloys are Harder Than Pure Metals

1) Pure metals often aren't quite right for certain jobs — they're often too soft when they're pure so are mixed with other metals to make them harder. Most of the metals we use everyday are alloys — a mixture of two or more metals or a metal and another element. Alloys are harder and so more useful than pure metals.
2) Different elements have different sized atoms. So when another element is mixed with a pure metal, the new metal atoms will distort the layers of metal atoms, making it more difficult for them to slide over each other. This makes alloys harder than pure metals.



I saw a metal on the bus once — he was the conductor...

If your knowledge of metals is still feeling a bit delocalised, the questions below will help...

Q1 Copper is a metallic element. Describe and explain what property of copper makes it suitable for using in electrical circuits. [2 marks]

Q2 Suggest why an alloy of copper, rather than pure copper, is used to make hinges for doors. [1 mark]

States of Matter

Better get your thinking hat on, as states of matter really... err.. matter. You'll need to imagine the particles in a substance as little snooker balls. Sounds strange, but it's useful for explaining lots of stuff in chemistry.

The Three States of Matter — Solid, Liquid and Gas

Materials come in three different forms — solid, liquid and gas. These are the three states of matter. Which state something is at a certain temperature (solid, liquid or gas) depends on how strong the forces of attraction are between the particles of the material. How strong the forces are depends on THREE THINGS:

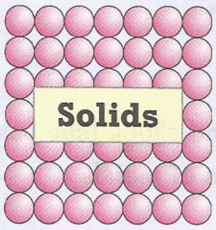
a) the material (the structure of the substance and the type of bonds holding the particles together),

b) the temperature,

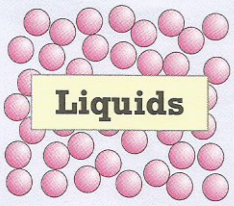
c) the pressure.

The particles could be atoms, ions or molecules.

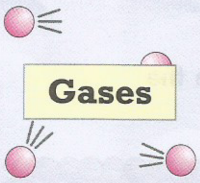
You can use a model called particle theory to explain how the particles in a material behave in each of the three states of matter by considering each particle as a small, solid, inelastic sphere.



1) In solids, there are strong forces of attraction between particles, which holds them close together in fixed positions to form a very regular lattice arrangement.
2) The particles don't move from their positions, so all solids keep a definite shape and volume, and don't flow like liquids.
3) The particles vibrate about their positions — the hotter the solid becomes, the more they vibrate (causing solids to expand slightly when heated).



1) In liquids, there's a weak force of attraction between the particles. They're randomly arranged and free to move past each other, but they tend to stick closely together.
2) Liquids have a definite volume but don't keep a definite shape, and will flow to fill the bottom of a container.
3) The particles are constantly moving with random motion. The hotter the liquid gets, the faster they move. This causes liquids to expand slightly when heated.



1) In gases, the force of attraction between the particles is very weak — they're free to move and are far apart. The particles in gases travel in straight lines.
2) Gases don't keep a definite shape or volume and will always fill any container.
3) The particles move constantly with random motion. The hotter the gas gets, the faster they move. Gases either expand when heated, or their pressure increases.

Particle theory is a great model for explaining the three states of matter, but it isn't perfect. In reality, the particles aren't solid or inelastic and they aren't spheres — they're atoms, ions or molecules. Also, the model doesn't show the forces between the particles, so there's no way of knowing how strong they are.

State Symbols Tell You the State of a Substance in an Equation

You saw on page 99 how a chemical reaction can be shown using a word equation or symbol equation. Symbol equations can also include state symbols next to each substance — they tell you what physical state the reactants and products are in:

(s) — solid (l) — liquid (g) — gas (aq) — aqueous

'Aqueous' means 'dissolved in water'.

For example, aqueous hydrochloric acid reacts with solid calcium carbonate to form aqueous calcium chloride, liquid water and carbon dioxide gas:

$$2HCl_{(aq)} + CaCO_{3(s)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Phew, what a page — particle-ularly gripping stuff...

I think it's pretty clever the way you can explain all the differences between solids, liquids and gases with just a page full of pink snooker balls. Anyway, that's the easy bit. The not-so-easy bit is learning it all.

Q1 Substance A does not have a definite shape or volume. What state is it in? [1 mark]

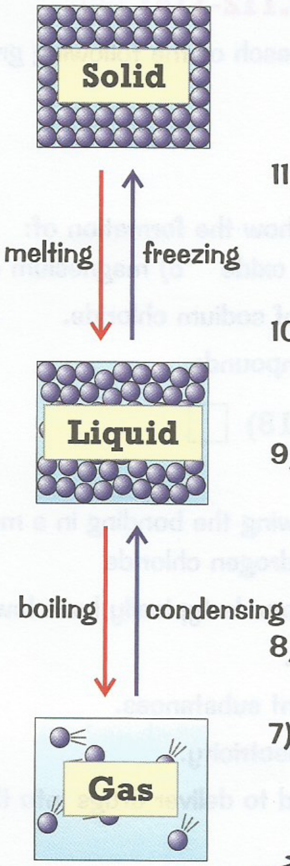
Changing State

This page is like a game show. To start, everyone seems nice and solid, but turn up the heat and it all changes.

Substances Can Change from One State to Another

Physical changes don't change the particles — just their arrangement or their energy.

1) When a solid is heated, its particles gain more energy.
2) This makes the particles vibrate more, which weakens the forces that hold the solid together.
3) At a certain temperature, called the melting point the particles have enough energy to break free from their positions. This is called MELTING and the solid turns into a liquid.
4) When a liquid is heated, again the particles get even more energy.
5) This energy makes the particles move faster, which weakens and breaks the bonds holding the liquid together.
6) At a certain temperature, called the boiling point, the particles have enough energy to break their bonds. This is BOILING (or evaporating). The liquid becomes a gas.



7) As a gas cools, the particles no longer have enough energy to overcome the forces of attraction between them.
8) Bonds form between the particles.
9) At the boiling point, so many bonds have formed between the gas particles that the gas becomes a liquid. This is called CONDENSING.
10) When a liquid cools, the particles have less energy, so move around less.
11) There's not enough energy to overcome the attraction between the particles, so more bonds form between them.
12) At the melting point, so many bonds have formed between the particles that they're held in place. The liquid becomes a solid. This is FREEZING.

So, the amount of energy needed for a substance to change state depends on how strong the forces between particles are. The stronger the forces, the more energy is needed to break them, and so the higher the melting and boiling points of the substance.

You Have to be Able to Predict the State of a Substance

You might be asked to predict what state a substance is in at a certain temperature. If the temperature's below the melting point of substance, it'll be a solid. If it's above the boiling point, it'll be a gas. If it's in between the two points, then it's a liquid.

The bulk properties such as the melting point of a material depend on how lots of atoms interact together. An atom on its own doesn't have these properties.

EXAMPLE

Which of the molecular substances in the table is a liquid at room temperature (25 °C)?

	melting point	boiling point
oxygen	–219 °C	–183 °C
nitrogen	–210 °C	–196 °C
bromine	–7 °C	59 °C

Oxygen and nitrogen have boiling points below 25 °C, so will both be gases at room temperature.

So the answer's bromine. It melts at –7 °C and boils at 59 °C. So, it'll be a liquid at room temperature.

Some people are worth melting for...

Make sure you can describe what happens to particles, and the forces between them, as a substance is heated and cooled. Then learn all the technical terms, and you'll sound like a states of matter pro.

Q1 Ethanol melts at –114 °C and boils at 78 °C. Predict the state that ethanol is in at:
a) –150 °C b) 0 °C c) 25 °C d) 100 °C [4 marks]

Q1 Video Solution

Revision Questions for Topic C2

Now you've finished Topic C2, I bet I can guess what you're after next. Some questions to test how much of this topic you can remember...

- Try these questions and tick off each one when you get it right.
- When you're completely happy with a sub-topic, tick it off.

For even more practice, try the Retrieval Quiz for Topic C2 — just scan this QR code!

Topic C2 Quiz

Ions and Ionic Compounds (p.112-114)

1) What type of ion do elements from each of the following groups form?
   a) Group 1
   b) Group 7
2) Describe how an ionic bond forms.
3) Sketch dot and cross diagrams to show the formation of:
   a) sodium chloride b) magnesium oxide c) magnesium chloride d) sodium oxide
4) Describe the structure of a crystal of sodium chloride.
5) List the main properties of ionic compounds.

Covalent Substances (p.115-118)

6) Describe how covalent bonds form.
7) Sketch dot and cross diagrams showing the bonding in a molecule of:
   a) hydrogen b) water c) hydrogen chloride
8) Explain why simple molecular compounds typically have low melting and boiling points.
9) Describe the structure of a polymer.
10) Give three examples of giant covalent substances.
11) Explain why graphite can conduct electricity.
12) Explain how fullerenes could be used to deliver drugs into the body.

Metallic Bonding (p.119)

13) What is metallic bonding?
14) List three properties of metals and explain how metallic structure causes each property.
15) Explain why alloys are harder than pure metals.

States of Matter (p.120-121)

16) Name the three states of matter.
17) What is the state symbol of an aqueous substance?
18) What is the name of the temperature at which a liquid becomes a gas?
19) How does the strength of the forces between particles influence the temperature at which a substance changes state?