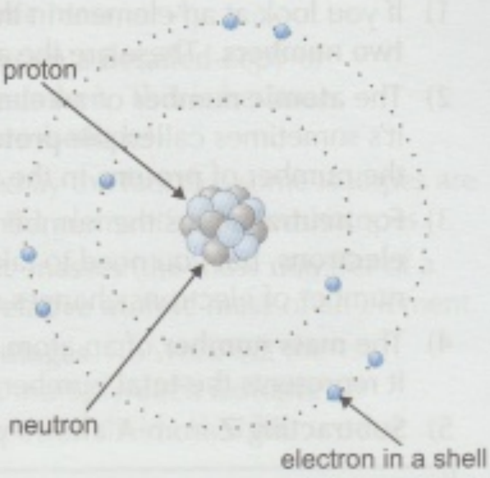
Atomic Structure

What Are Atoms Like?



1) Atoms are made up of **three** types of **subatomic particle**: **protons**, **neutrons** and **electrons**.
2) In the **centre** of all atoms is a **nucleus** containing **neutrons** and **protons**.
3) Almost all of the **mass** of the atom is contained in the **nucleus** which has an overall **positive** charge. The positive charge arises because each of the **protons** in the nucleus have a **+1** charge.
4) The **neutrons** in the nucleus have a very similar **mass** to the protons but they are **uncharged**.
5) **Electrons** are much **smaller** and **lighter** than either the neutrons or protons. They have a **negative charge** (–1) and **orbit** the nucleus in **shells** (or energy levels).
6) There's an **attraction** between the **protons** in the nucleus and the **electrons** in the shells.
7) The nucleus is **tiny** compared with the total volume occupied by the whole atom.
8) The **volume** occupied by the **shells** of the electrons determines the **size** of the atom.

Here's a round up of the **properties** of the subatomic particles:

Particle	Relative Mass	Relative Charge
Proton	1	+1
Neutron	1	0
Electron	$\frac{1}{2000}$	–1

What is the Charge on an Atom?

The overall charge on an atom is **zero**.

This is because each **+1** charge from a **proton** in the nucleus is **cancelled out** by a **–1** charge from an **electron**.

If an atom **loses** or **gains** electrons it becomes **charged**. These charged particles are called **ions**.

EXAMPLE: How many electrons has an Al^{3+} ion lost or gained?

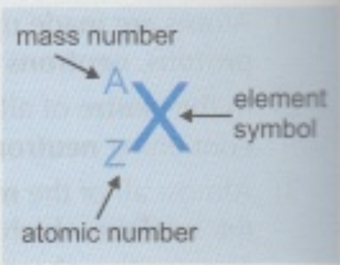
The Al^{3+} ion has a charge of **+3**, so there must be **3 more protons** than **electrons**.
Ions are formed when **electrons** are lost or gained, so Al^{3+} must have **lost 3 electrons**.

Neutrons are the perfect criminals — they never get charged...

1) Which subatomic particles are found in the nucleus?
2) What is the charge on an ion formed when an atom loses two electrons?
3) What is the charge on an ion formed when an atom gains two electrons?

Atomic Number, Mass Number and Isotopes

Atomic and *Mass* Numbers



1) If you look at an element in the periodic table, you'll see it's given **two numbers**. These are the **atomic number** and the **mass number**.
2) The **atomic number** of an element is given the symbol **Z.** It's sometimes called the **proton number** as it represents the number of **protons** in the nucleus of the element.
3) For **neutral** atoms the number of **protons equals** the number of **electrons**, but you need to take care when considering ions as the number of electrons changes when an ion forms from an atom.
4) The **mass number** of an atom is given the symbol **A.** It represents the **total** number of **neutrons** and **protons** in the nucleus.
5) **Subtracting Z** from **A** allows you to calculate the number of **neutrons** in the nucleus.

EXAMPLE: Use the periodic table to complete the following information about sodium.

Element	Symbol	Z	A	No. Protons	No. Neutrons	No. Electrons
Sodium			23			

The periodic table tells you that the **symbol** for sodium is **Na** and **Z** is **11**.

The number of **protons** in sodium is the same as the **atomic number**, which is **11**.

You work out the number of **neutrons** by **subtracting** Z from A: 23 – 11 = **12**.

The number of **electrons** is the **same** as the number of protons, which is **11**.

Isotopes

1) Atoms of the same **element** always have the same number of **protons**, so they'll always have the same **atomic number**, but their **mass numbers** can **vary** slightly.
2) Atoms of the same **element** with different **mass numbers** are called **isotopes**.
3) Isotopes have the same number of **protons** but different numbers of **neutrons** in their nuclei.

EXAMPLE: Copper has an atomic number of 29. Its two main isotopes have mass numbers of 63 and 65. How many neutrons does each of the isotopes have?

The ^{63}Cu isotope has 63 – 29 = **34 neutrons**.

The ^{65}Cu isotope has 65 – 29 = **36 neutrons**.

Finding the number of neutrons — it's as easy as knowing your A – Z...

1) Use the periodic table to work out how many neutrons are in a neutral phosphorus atom.
2) In terms of the numbers of subatomic particles, state two similarities and one difference between two isotopes of the same element.
3) Three neutral isotopes of carbon have mass numbers 12, 13 and 14. State the numbers of protons, neutrons and electrons in each.

Relative Atomic Mass

Calculating the *Relative Atomic Mass*

1) The average mass of an element is called its **relative atomic mass**, or A_r.
2) When you look up the **relative atomic mass** of an element on a **detailed** copy of the periodic table, you'll see that it isn't always a **whole number**. This is because the value given is the **average** mass number of two or more **isotopes**.
3) The **value** of the relative atomic mass is further complicated by the fact that some isotopes are **more abundant** than others. It's a **weighted average** of all the element's different isotopes.
4) You can use the **relative abundances** and **relative isotopic masses** (the mass number of a single, specific isotope) of each isotope to work out the **relative atomic mass** of an element.
5) Relative abundances of isotopes are often given as **percentages**. To work out the **relative atomic mass** of an element, all you need to do is multiply **each isotopic mass** by its **relative abundance**, add all the values together and divide by **100**.

EXAMPLE: What is the relative atomic mass of chlorine given that 75% of atoms have an atomic mass of 35 and 25% of atoms have an atomic mass of 37?

Average mass = (abundance of $^{35}Cl \times 35$ + abundance of $^{37}Cl \times 37) \div 100$

$= [(75 \times 35) + (25 \times 37)] \div 100$

$= (2625 + 925) \div 100$

$= 3550 \div 100$

$= \mathbf{35.5}$ (You can check your answer against a periodic table to see if it's right.)

Calculating the *Relative Formula Mass*

If you **add up** the relative atomic masses of all the atoms in a chemical formula, you get the **relative formula mass**, or M_r, of that compound.

(If the compound is molecular, you might hear the term relative molecular mass used instead, but it means pretty much the same.)

EXAMPLE: Calculate the relative formula mass of $CaCl_2$.

Ca has an atomic mass of 40.1 and Cl has an atomic mass of 35.5.

$M_r = (1 \times 40.1) + (2 \times 35.5)$

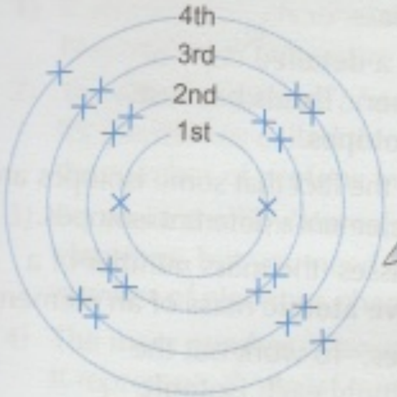
$= \mathbf{111}$

Together, my brother and I weigh 143 kg — it's our relative mass...

1) Find the relative atomic mass of lithium if its composition is 8% ^{6}Li and 92% ^{7}Li.
2) Find the relative atomic mass of carbon if its composition is 99% ^{12}C and 1% ^{13}C.
3) Find the relative atomic mass of silver if its composition is 52% ^{107}Ag and 48% ^{109}Ag.
4) Find the relative formula mass of sodium fluoride, NaF.
5) Find the relative formula mass of chloromethane, CH_3Cl.

Electronic Structure

Electrons are Arranged in *Energy Shells*



1) Electrons orbit the nucleus in **shells** (also called **energy levels**).
2) You can draw concentric **circles** to represent the different **shells**. Then add **crosses** to represent the **electrons** at each level.
3) For example, this diagram shows the energy levels, for an atom with 20 electrons, filling up with electrons. It has two electrons in the first shell, eight in the second shell, eight in the third shell and two in the fourth shell. (Remember you should always **start** filling the **innermost levels** first.)

Here's another way to show electron arrangements using simple notation:

An atom with 6 electrons:	2, 4
An atom with 11 electrons:	2, 8, 1
An atom with 20 electrons:	2, 8, 8, 2

The first number tells you how many electrons are in the first shell, the second number tells you how many electrons are in the second shell, and so on.

Energy Levels are *Split* into *Subshells*

1) Energy levels can be **split** into **sub-levels** called **subshells**. The first three subshells are called '**s**', '**p**' and '**d**'. They can each hold a different **number** of electrons.
2) The first energy level has **one subshell** — an 's' level. So the first energy level can contain up to **2 electrons**.

Subshell	Maximum electrons
s	2
p	6
d	10

3) At GCSE you learnt that the second energy level can contain up to **8 electrons**. It's actually split into **2 sub-levels**. **Two** of the electrons are in an '**s**' level and the remaining **six** are in a '**p**' level. If you combine the 2 's' electrons with the 6 'p' electrons you get a total of 8.
4) Electrons generally start by filling the **energy level** with the **lowest energy**. So the **first** energy level will be completely **filled** before any electrons go into the **second** energy level. Within an energy level, electrons will fill the **subshells** in the order **s**, then **p**, then **d**.
5) As well as telling you how many electrons are in each **shell**, the **electron configuration** of an atom also tells you what **subshells** the electrons are in. For an atom with 10 electrons:

The big number tells you the energy level. $1s^2\ 2s^2\ 2p^6$ The little number tells you how many electrons are in that subshell.

The letter tells you the subshell.

I'm trying to be calm, but my energy level is too high...

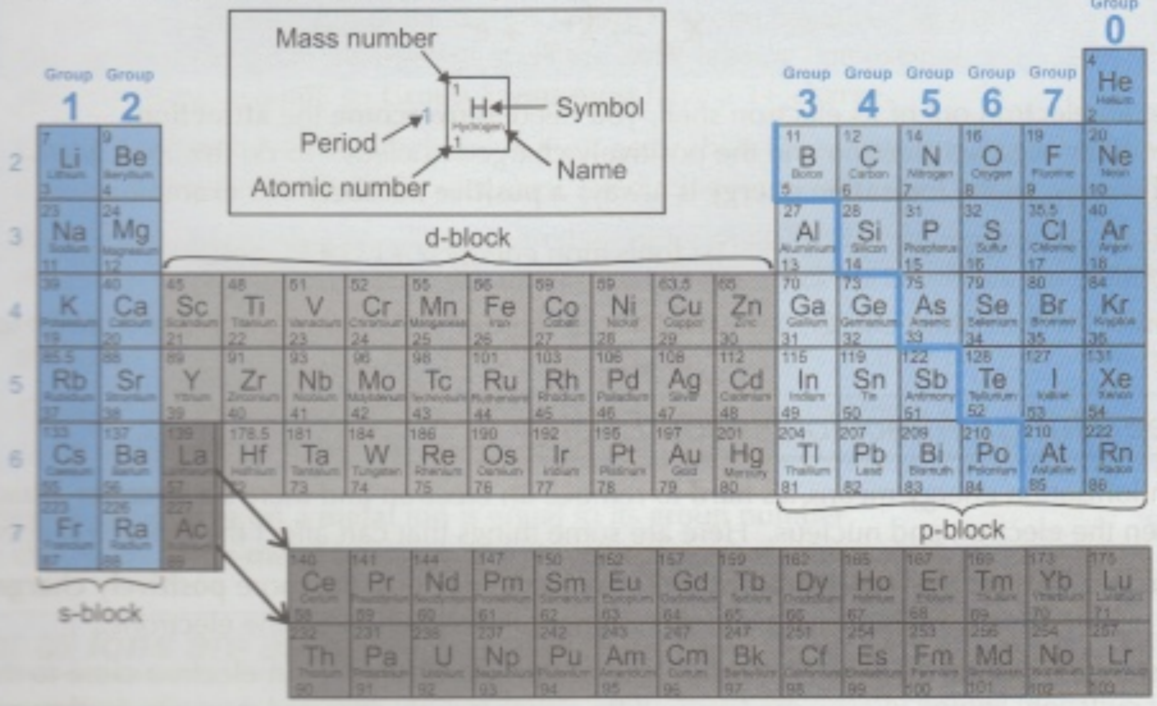
1) Draw diagrams to show the electron arrangements of the following elements: carbon, fluorine, magnesium, sulfur.
2) Use the simple notation shown above to write the electron arrangements of these elements: lithium, sodium, potassium, beryllium, magnesium, calcium.
3) Give the electron configurations of oxygen and chlorine.

The Periodic Table

*The **Periodic Table***

The periodic table contains:

- All of the elements in order of atomic number.
- Vertical groups of elements which have similar properties.
- Horizontal rows of elements called periods.



*What Are **Groups** and **Periods**?*

Chemical reactions involve atoms **reacting** to gain a **full outer shell** of electrons.
All of the elements in a group have the same **number** of **electrons** in their **outer shell**.
As a result, the elements in a group **react** in a **similar** way.

The properties of elements in the same **period change gradually** as you move from one side of the periodic table to the other.

*The **Periodic Table** is **Split** into **Blocks***

1) As well as being split into groups and periods, the periodic table has four **blocks**. You only need to worry about two of them at the moment though — the '**s**' block and the '**p**' block.
2) Groups 1 and 2 are called the **s-block** elements. Their outer electrons are in energy levels called **s subshells**. S subshells can accommodate up to 2 electrons (see page 4).
3) Groups 3 to 0 are called the **p-block** elements. Their outer electrons are in energy levels called **p subshells**. P subshells can accommodate up to 6 electrons (see page 4).
4) There's always one exception. **Helium** (He) is an **s-block** element, even though it's in **Group 0**. Its electron configuration is $1s^2$.

The mystery of the periodic table? It's elementary, my dear Watson...

1) Sort the following elements into a table to show which ones are from the s-block, and which are from the p-block: caesium, potassium, phosphorus, calcium, aluminium, barium and sulfur.
2) Give one similarity between elements that are in the same group.

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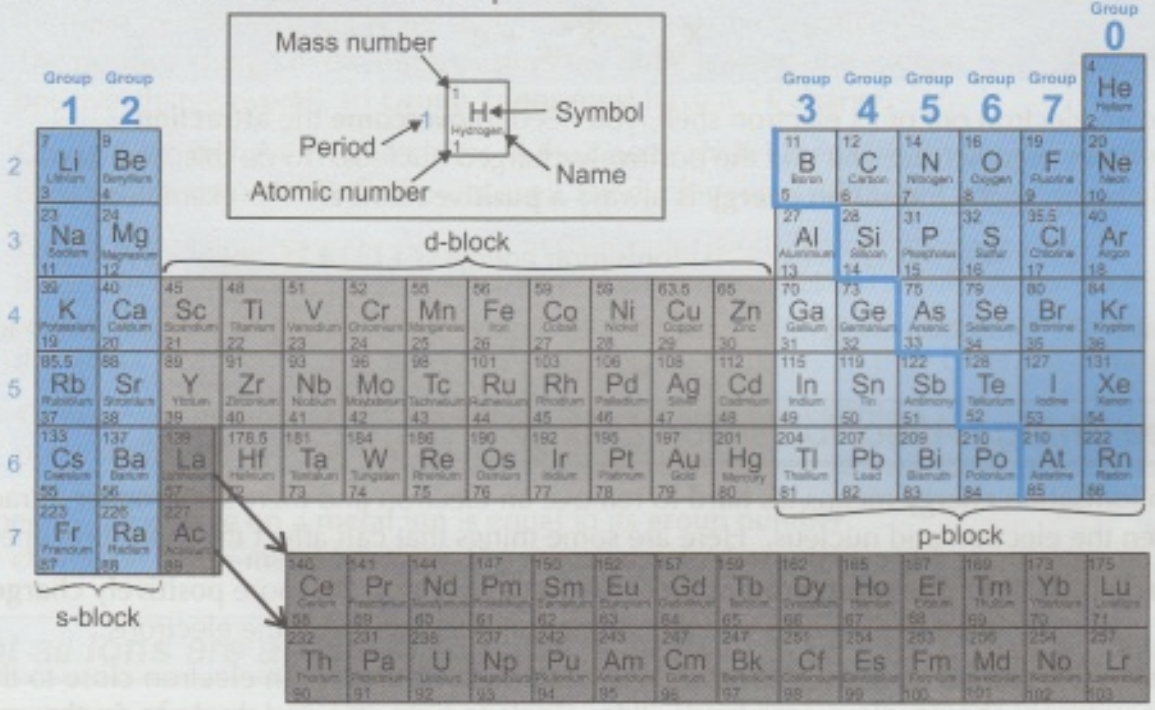

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Ionisation Energy

When Atoms **Lose Electrons** *they are* **Ionised**

When electrons have been removed from an atom or molecule, it's been **ionised**.
The energy you need to remove the **first outer electron** is called the **first ionisation energy**.

The first ionisation energy is the energy needed to **remove 1 electron** from **each atom** in **1 mole** of **gaseous** atoms to form 1 mole of gaseous 1+ ions (see page 37 for more on moles).

$$X_{(g)} \rightarrow X^{+}_{(g)} + e^{-}$$

To take an electron out of its electron shell, you need to **overcome** the **attraction** between the negative electron and the positively charged nucleus. To do this, you have to **add** energy, so the **ionisation energy** is always a **positive number**. For example:

$O_{(g)} \rightarrow O^{+}_{(g)} + e^{-}$ 1st ionisation energy = +1314 kJ mol^{-1}

The **lower** the ionisation energy, the **easier** it is to remove the outer electron and form an ion.

Three Things *Affect Ionisation Energy*

A high ionisation energy means it's **hard** to remove an electron and there's a **stronger** attraction between the electron and nucleus. Here are some things that can affect the ionisation energy:

1) **Nuclear charge**: The **more protons** there are in the nucleus, the more **positively charged** the nucleus is and the **stronger** the attraction for the electrons.
2) **Distance from the nucleus**: Attraction decreases with **distance**. An electron **close** to the nucleus will be more strongly attracted than one **further away**.
3) **Shielding**: Electrons in shells **closer** to the nucleus can **stop** the outer electrons from feeling the **full force** of the nuclear charge. The inner electrons are said to **shield** the outer electrons from the nucleus. More inner electrons mean more shielding, so a **weaker attraction** for the **outer electrons** and a **lower ionisation energy**.

The Periodic Table Shows **Trends** *in Ionisation Energies*

1) Ionisation energy **decreases** down a **group**. This is because as you go down a group, each element has **one more** electron shell than the one above — so the distance between the **nucleus** and the **outer shell increases**. There will also be more **shielding** from the larger number of **inner electrons**. So overall, going down a group the **attraction** between the nucleus and the outer electrons **decreases**.
2) Ionisation energy generally **increases** across a **period**. There are **more protons** in the nucleus so there's a **higher nuclear charge**. Electrons are also going into the **same shell**, so the **distance** from the nucleus and the amount of **shielding** by inner electrons doesn't change much. So overall, the attraction between the nucleus and the electrons **increases**.

Let it go, let it go, lose electrons from my outer shell...

1) Write an equation to show the first ionisation of sodium.
2) What three things can affect ionisation energy?
3) For the following pairs of elements, decide which will have the higher first ionisation energy: Magnesium and Calcium, Lithium and Fluorine, Oxygen and Sulfur.

Formation of Ions

Elements in the s-block and the p-block form Simple Ions

Most elements in the **s-block** and the **p-block** form ions with **full outer electron shells.** This means you can **predict** what ion an element will form by looking at the **periodic table** — just follow through the reasoning below:

- **Group 1** atoms have **one electron** in their outer shell. The **easiest way** for them to achieve a full outer shell is to **lose** that one negative electron. The positive charge in the nucleus stays the same leaving one excess positive charge overall, so **Group 1 ions** must have a **1+ charge**.
- **Group 2** atoms have **two electrons** in their outer shell. They **lose** these two negative electrons to get a **stable** (full) outer shell, producing ions with a **2+** charge.
- **Group 6** elements have six electrons in their outer shell. Rather than releasing all six of these electrons (which would take a lot of energy) they **pick up** two electrons from their surroundings to complete their outer shell. The positive charge in the nucleus stays the same, so Group 6 ions have **two extra** negative charges — they carry a **2– charge.**
- **Group 7** atoms need to pick up **one** extra electron to get a stable outer shell, so they form ions with a charge of **1–**.

Generally the charge on a **metal ion** is equal to its **group number**.
The charge on a **non-metal ion** is equal to its **group number minus eight**.

Not all Ions are as Simple

Some **groups of atoms** can also exist as stable ions. These are usually **anions** (negative ions) like sulfate and carbonate (one of the few exceptions being ammonium with a 1+ charge). It is harder to work out the charges on these than in the case of the simple ions above.

It is useful to **learn** the charges on the most common of these **molecular ions**:

+1	–2	–1
NH_4^+ (ammonium)	SO_4^{2-} (sulfate)	OH^- (hydroxide)
	CO_3^{2-} (carbonate)	NO_3^- (nitrate)
	SO_3^{2-} (sulfite)	HCO_3^- (hydrogencarbonate)
		CN^- (cyanide)

Transition metals (the block of elements between Groups 2 and 3) also form ions. They are **positive** (like all metal ions) but they **do not** form ions with a full outer shell of electrons. This means you can't predict the charges in the same way as you can with the s-block metals. Most transition metals can form **more than one** ion. The different charges are called '**oxidation numbers**' of the element (see page 8). The common ones that you should be aware of are:

Fe^{2+}, Fe^{3+}, Cu^{2+}, Co^{2+}, Ni^{2+}, Zn^{2+} and Cr^{3+}

I never ask for an anion's opinion — they're always so negative...

1) What is the charge on a sodium ion?
2) Which Group typically forms 1– ions?
3) What is the formula of a sulfite ion? Remember to include the overall charge on the ion.

Oxidation Numbers

Oxidation Numbers Tell you the Charge on an Atom

When atoms **react** or **bond** to other atoms, they can **lose** or **gain** electrons. The **oxidation number** tells you how many electrons an atom has donated or accepted when it's reacted. You may also see **oxidation numbers** called **oxidation states**, but they're the same thing.

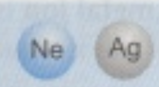
Roman Numerals Tell you the Oxidation Number

Roman numerals can be used to show what oxidation number a certain element has. You'll probably remember your Roman numerals from Maths, where (I) = +1, (II) = +2, (III) = +3 and so on. The Roman numerals are written **after** the name of the **element** they correspond to.

In iron(II) chloride, iron has an oxidation number of +2. Formula = $FeCl_2$
In iron(III) chloride, iron has an oxidation number of +3. Formula = $FeCl_3$

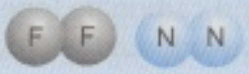
There are Some Rules About Oxidation Numbers

1) Elements that aren't bonded to anything else all have an oxidation number of **0**.



Uncombined elements.
Oxidation number = 0

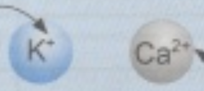
2) Elements that are bonded to identical atoms also have an oxidation number of **0**.



Elements bonded to identical elements.
Oxidation number = 0

3) The oxidation number of an ion made up of just one atom is the same as its **charge** (the little number to the **right** of the symbol).

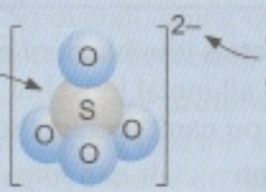
Oxidation number = +1



Oxidation number = +2

4) For **molecular ions** (ions that are made up of more than one atom) the **overall charge** of the whole ion is equal to the **sum** of the **oxidation numbers** of the individual atoms or ions.

Sulfur has an oxidation number of +6, and each oxygen has an oxidation number of –2.



Overall charge is –2.

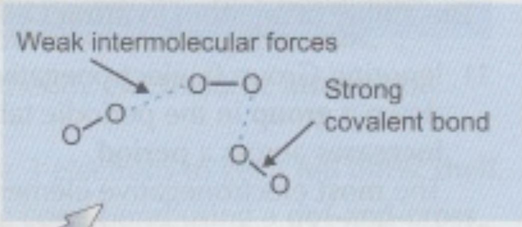
Caesar the day — and get to know your Roman numerals...

1) What does the oxidation number tell you about an atom?
2) Give the oxidation number of each of the following atoms/ions:
Al^{3+}, H^+, Ne, O^{2-}.
3) What is the oxidation number of an atom of chlorine in Cl_2?

Intermolecular Bonding

***Intermolecular Bonds** Form Between **All** Molecules*

1) Some compounds are made up of **simple molecules** — these are just groups of a **few atoms** joined together by **covalent bonds** (see page 13). For example, water (H_2O) or oxygen (O_2).
2) The bonds **between** the **atoms** in each molecule are very **strong**. By contrast, there are very **weak** forces of attraction that form **between** the **molecules**. These are **intermolecular bonds** (also called intermolecular forces).



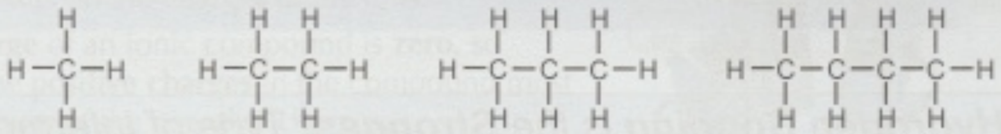
*The **Strength** of Intermolecular Bonds Affects **Boiling Points***

When simple molecular substances **melt** or **boil**, it's the **intermolecular bonds** that get broken — not the much stronger covalent bonds. The **stronger** the intermolecular bonds, the more **energy** is needed to break them, so the **higher** the boiling or melting point will be.

Two things that can affect the strength of intermolecular forces are:

1) The number of **electrons** in a molecule: the **more** electrons there are, the **stronger** the intermolecular bonds between molecules.
2) The **surface area** of the molecule: the **larger** the surface area over which intermolecular bonds can act, the **stronger** the intermolecular bonds between molecules.

EXAMPLE: Use the idea of intermolecular bonds to explain the trend in boiling points of the following alkanes.



Alkane:	Methane	Ethane	Propane	Butane
Boiling point:	–161 °C	–89 °C	–42 °C	0 °C

There is a clear trend showing that as the molecules get **larger** their boiling point **increases**.

This is due to the fact that the larger molecules have a greater **surface area**, so there is stronger intermolecular bonding. The larger molecules also have more **electrons** — this further increases the strength of the intermolecular bonds that form between molecules.

This page was alright — we formed a sort of bond...

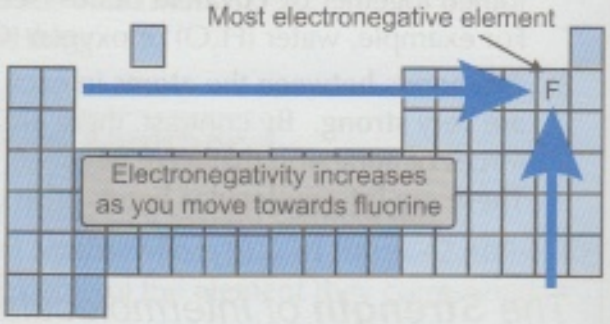
1) Draw a diagram to show the different types of bonding in a sample of gaseous chlorine molecules ($Cl_{2(g)}$). What type of bond is the strongest?
2) Use the data in the example above to predict the boiling points of the next four members of the alkane series. They are called pentane, hexane, heptane and octane. (Bear in mind that, in Chemistry, the first member of a series does not always provide an ideal example.)

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**.

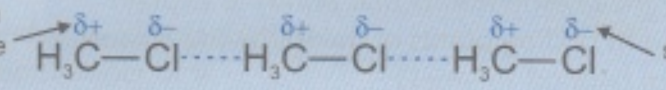


Polar Bonds** Can Affect the **Strength** of **Intermolecular Forces

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.

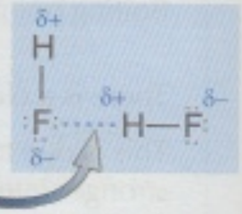
Boiling point of methane (CH_4)	= −161 °C
Boiling point of chloromethane (CH_3Cl)	= −24 °C

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.



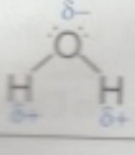
***Hydrogen Bonding** is the **Strongest** Type of Intermolecular Force*

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...

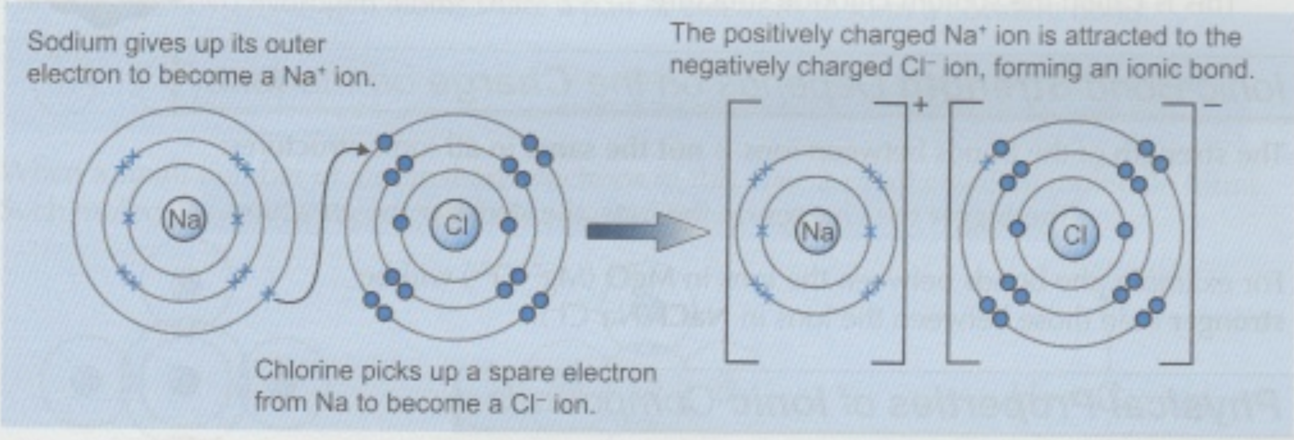
1) For the following pairs of molecules, predict with reasoning which has the higher boiling point:
   a) H_2 and HF, b) H_2O and H_2S, c) CH_3F and CH_3I.
2) Water is a polar molecule. Draw a diagram showing three water molecules attracted together. You should use dotted lines to indicate forces between 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 **$CaCl_2$**.

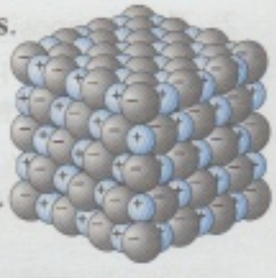
I can't afford Mg^{2+} — the charge is just too high...

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

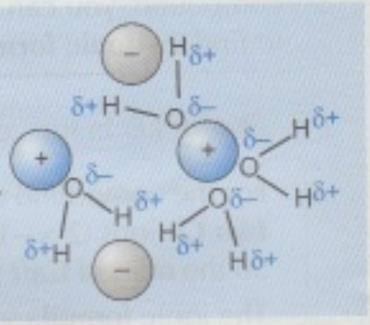
In order to **melt** a solid, the forces holding the particles together have to be **overcome**. In an ionic solid, these bonds are very **strong**, so a **large** amount of energy is required to break them. So, ionic compounds have very **high** melting points.

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.

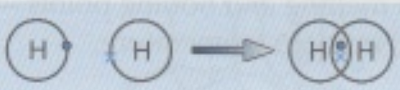
Covalent Bonding

***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**.

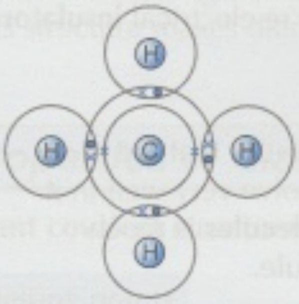
A covalent bond is a **shared pair** of electrons.

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

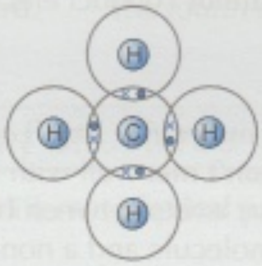


The shared pair of electrons means that each hydrogen atom in the molecule has 2 electrons in its outer shell.

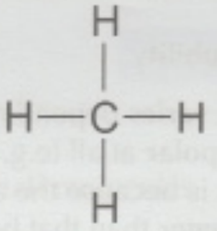
When a small number of atoms share electrons in this way, a small covalent molecule forms. Such molecules can be represented in several different ways:



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

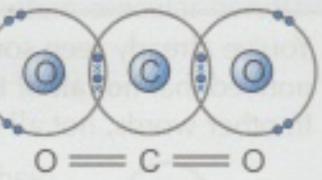


A more simple dot-and-cross diagram, showing only the outer shells of electrons.



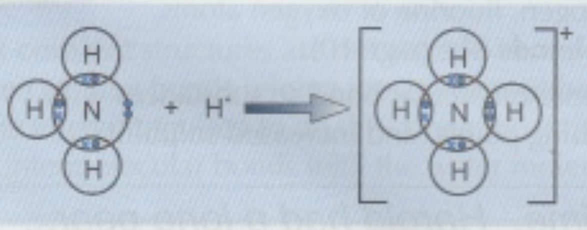
Each dash represents a single covalent bond (this is the most common notation).

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:

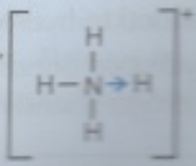


***Dative** Covalent Bonding*

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

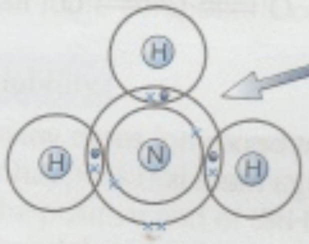
Small covalent molecules don't contain any of the **free charged** particles that are needed to carry an electric current. As a result they **cannot** conduct electricity — they're electrical **insulators**.

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.

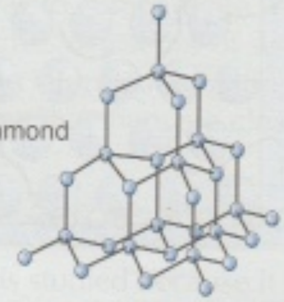
Giant Covalent Structures

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**.

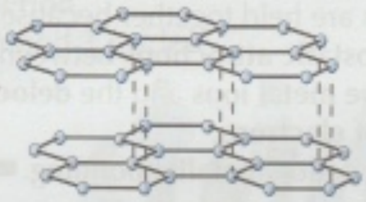
Carbon can form various different **giant covalent structures** such as **diamond** and **graphite**.

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.

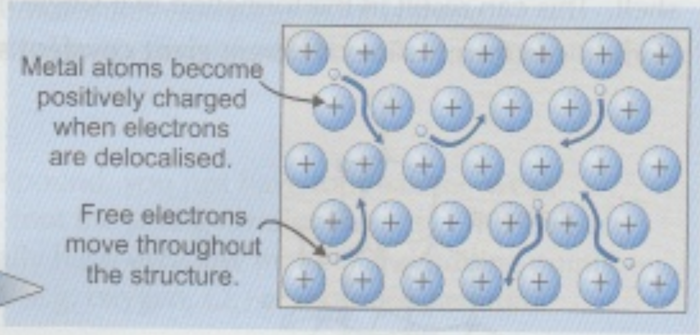
Diamonds — don't mess with 'em — they're well 'ard...

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.**

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

...ing whether potassium or calcium will have a higher melting point.

...the bonding in a sample of sodium.

...e, whilst sodium chloride (NaCl) is an ionic compound.

...nce between the physical properties of these substances.

Writing and Balancing Equations

Reaction Equations Show How Chemicals React Together

1) A reaction equation shows what happens during a chemical reaction. The **reactants** are shown on the **left hand side**, and the **products** on the **right hand side**.
2) **Word equations** just give the **names** of the components in the reaction. For example: propane + oxygen → carbon dioxide + water
3) **Symbol equations** give the chemical formulae of all the different components. They show all the **atoms** that take part in the reaction, and how they rearrange. For example: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
4) Symbol equations have to **balance** — there has to be the **same number** of each **type** of atom on each side of the equation. The big numbers in front of each substance tell you how much of that particular thing there has to be for all the atoms to balance.

Writing Balanced Equations

To write a balanced symbol equation for a reaction there are 4 simple steps:

1) Write out the **word equation** first.
2) Write the correct **formula** for each substance below its name.
3) Go through each element in turn, making sure the **number of atoms** on each **side** of the equation **balances**. If your equation isn't balanced, you can only add more atoms by adding **whole reactants** or **products**.
4) If you changed any numbers, do step 3 again, and repeat until **all** the elements **balance**.

Doing the third step:
If the atoms in the equation don't balance you **can't** change the **molecular formulae** — only the numbers in **front** of them.
For example: $CaO + HCl \rightarrow CaCl_2 + H_2O$
There are **two Cl** atoms on the **right-hand side** of the equation, so we need to have **two HCl** on the **left-hand** side: $CaO + 2HCl \rightarrow CaCl_2 + H_2O$
This also doubles the number of **hydrogen atoms** on the left-hand side, so that the hydrogens **balance** as well.

EXAMPLE: Write a balanced equation for the reaction of magnesium with hydrochloric acid.

Step 1 — Write the word equation:
magnesium + hydrochloric acid → magnesium chloride + hydrogen

Step 2 — Write the symbol equation: $Mg + HCl \rightarrow MgCl_2 + H_2$

Step 3 — Go through the equation and balance the elements one by one:
$Mg + 2HCl \rightarrow MgCl_2 + H_2$
(the Mgs balance, but there are different amounts of H and Cl on each side. Put a 2 in front of HCl to balance the Hs and Cls. Check everything still balances.)

Writing and Balancing Equations

*In **Ionic Equations** Make Sure the **Charges** Balance*

1) In some reactions, particularly those in solution, not all the particles take part in the reaction.
2) **Ionic equations** are chemical equations that just show the **reacting particles**.
3) As well as having the same number of **atoms** of each element on each side of the equation, in ionic equations you need to make sure the **charge** is the same on both sides.

EXAMPLE: Balance the following ionic equation: $Na + H^+ \rightarrow Na^+ + H_2$

First, balance the **number of atoms** of each element using the method on the last page:

$$Na + 2H^+ \rightarrow Na^+ + H_2$$

Then check the **charge** is the same on both sides of the equation:

- On the left hand side, each H^+ ion contributes +1, so the charge is $2 \times +1 = \mathbf{+2}$.
- On the right hand side, the sodium ion contributes +1, so the charge is $1 \times +1 = \mathbf{+1}$.

To get the charges to balance, you need another positive charge on the right-hand side.
One way of doing this is by adding another sodium ion to the products:

$$Na + 2H^+ \rightarrow \mathbf{2}Na^+ + H_2$$

Now check that the number of atoms still balances:
The Hs balance, but there are 2 Nas on the right-hand side, and only one on the left.
So put a 2 in front of the left-hand side Na:

$$\mathbf{2}Na + 2H^+ \rightarrow 2Na^+ + H_2$$

The atoms **and** charges on each side balance, so that's your final answer.

*Chemical Equations Sometimes Include **State Symbols***

State symbols show the **physical state** that a substance is in.
The state symbols you need to know about are in the box below:

(l) — liquid (g) — gas (s) — solid (aq) — aqueous (dissolved in water)

So the balanced equation for the reaction between hydrochloric acid and magnesium, including state symbols is: $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$.

Hold one ear and stare at something still — it'll help you balance...

1) Write a balanced symbol equation for the combustion of methane (CH_4) in oxygen.
   Step 1 has been done for you.
   Step 1: Methane + oxygen → carbon dioxide + water
2) Write balanced symbol equations for the following reactions.
   a) The complete combustion of ethanol (C_2H_5OH) in oxygen (O_2) to give carbon dioxide (CO_2) and water (H_2O).
   b) The reaction of calcium hydroxide ($Ca(OH)_2$) with hydrochloric acid (HCl) to give calcium chloride ($CaCl_2$) and water (H_2O).
3) Balance the following ionic equation: $Cl_2 + Fe^{2+} \rightarrow Cl^- + Fe^{3+}$.
   Include state symbols given that Cl_2 is a gas and everything else is aqueous.

Group 2

Trend in **Reactivity** Down the Group

During their reactions, Group 2 metals **donate** their **two outer electrons** to another atom. The reactivity of Group 2 metals depends on how **easily** the outer electrons can be donated. The **easier** the electrons can be donated, the **more reactive** the metal will be. You will find that:

Reactivity **increases** as you go **down** Group 2.

To see why, think about the factors that affect how strongly an electron is held by the nucleus:

1) The first is the **positive nuclear charge** — how **positive** the nucleus is. A **greater** nuclear charge provides a **stronger** force of attraction between the nucleus and electrons, and makes it more difficult for the atom to donate its outer electrons. As you go down the group, the nuclear charge **increases** as more **protons** are added to the nucleus, so if this was the **only** factor, reactivity would decrease down Group 2. But that **isn't** the case.
2) The second factor is that in **larger atoms**, the outer electrons are **further away** from the nucleus. The electrostatic attraction **decreases** in strength with **distance** from the source.
3) The third factor is **electron shielding**. As the atoms in Group 2 get **larger**, the number of **full electron shells** round the nucleus **increases**. These negative charges **shield** the two outer electrons from the attraction of the positive nucleus.

The increase in the **distance** between the outer electrons and the nucleus, and the increased **shielding** as you go down the Group, far **outweigh** the increase in nuclear charge.

4) You may have noticed that these are the **same factors** that affect the **ionisation enthalpy** (page 6). This is because both the reactivity of Group 2 and ionisation are to do with **removing electrons**.

Trend in the **Melting Points** of Group 2 Metals

You can see from the table that:

As you go **down** Group 2, **melting point decreases.**

Magnesium doesn't fit in with the general trend. It behaves a bit oddly because it has a slightly different structure to the other Group 2 metals.

	Melting Point (°C)
Beryllium (Be)	1278
Magnesium (Mg)	651
Calcium (Ca)	839
Strontium (Sr)	769
Barium (Ba)	727

This is also due to the increase in **electron shielding** as you go down the group.

Group 2 metals, like all other metals, are held together in a lattice structure by **metallic bonds** (page 16).

The **strength** of the metallic bonds depend on how strong the **attraction** is between the positive ions and the free electrons. The **more shielded** the positive nuclei are, the **weaker** the attraction will be, and so the **less energy** will be required to break the bond and melt the metal.

I love the Group 2 Metals — they're really trendy...

1) The following are descriptions of the reactions of Be and Ca with cold water. Use them to predict the reactions of Mg and Sr.
   - Beryllium will not react with cold water at all.
   - Calcium reacts steadily with cold water to produce hydrogen gas and calcium hydroxide.
2) Predict, with reasoning, the trend in boiling points of the Group 2 metals.

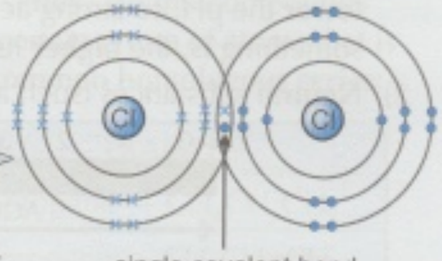
Group 7

Some **General Properties** of Group 7 Elements

Group 7 elements all have **7 electrons** in their outer shell. As a result these elements either:

1) Form **ionic compounds** by gaining an **extra electron** or,
2) **share** a pair of electrons and form a **covalent bond**.

In their elemental state, the halogens bond **covalently**, forming diatomic molecules (two atoms joined with a single covalent bond). In each case the atoms share an electron pair.



The halogen elements all have **coloured vapours**:

- **Chlorine** (Cl_2) is a **yellow/green gas** at room temperature.
- **Bromine** (Br_2) is a **brown liquid** at room temperature.
- **Iodine** (I_2) is a **grey solid** at room temperature (and sublimes to produce a **purple** vapour).

As you go down the group, the **melting points** and the **boiling points** of the elements **increase**. This is because the **strength** of the weak intermolecular bonds **between** molecules **increases** as the number of **electrons** in the molecules increases (see page 9).

Trend in **Reactivity** Down Group 7

During their reactions, Group 7 elements accept an **extra electron** from another atom. The reactivity of Group 7 elements depends on how **strongly** the nucleus can attract electrons. The **stronger** the attraction, the **more reactive** the element will be.

Reactivity **decreases** as you go **down** Group 7.

1) As with the Group 2 elements, **nuclear charge increases** as you go down the group. A greater nuclear charge will attract the extra electron required to fill the outer shell more strongly. This works to increase the reactivity of the elements as you go down the group.
2) However, as the atoms get bigger, the **extra shells** of electrons **shield** the nuclear charge more effectively. So the nucleus is **less able** to attract the extra electron the atom wants.

In Group 7 this **shielding** outweighs the effect of increasing nuclear charge. The elements at the **top** of the group are best able to attract an extra electron, and are more **reactive**.

Group 7 Reactivity and **Displacement Reactions**

You can show the relative reactivity of the Group 7 elements using **displacement reactions**. If you mix a **halogen** with a solution containing **halide ions**, a **more reactive** halogen will **displace** a **less reactive** halide ion (one below it in the group) from solution. For example:

Fluorine is more reactive than chlorine.

$$F_{2\,(aq)} + 2Cl^-_{(aq)} \rightarrow Cl_{2\,(aq)} + 2F^-_{(aq)}$$

The chloride ions have been **displaced** from the solution.

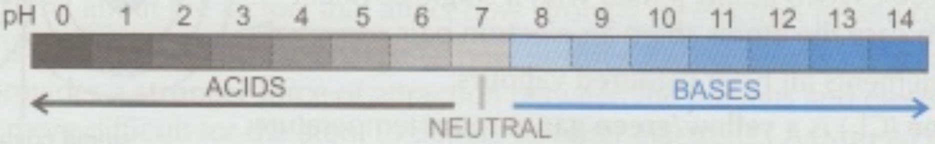
I met a friend for coffee today — I said 'Hallo, Jen'...

1) Predict, with reasoning, what would happen if you mixed the following halogens and halide solutions.
   a) Cl_2 and Br^-, b) I_2 and Cl^-, c) I_2 and Br^-, d) Cl_2 and I^-.
2) Draw a diagram to show the bonding between atoms in a fluorine molecule.

Acids and Bases

The pH Scale

1) The **pH** scale goes from **0** to **14** and measures how **acidic** or **basic** something is. **Acids** have a pH **less** than 7, while **bases** have a pH **greater** than 7. The **more acidic** something is, the **lower** the pH, so strong acids have a pH of between **0** and **1**. By contrast, the more **basic** something is, the **higher** its pH will be. **Strong bases** have a pH of **14**.
2) **Neutral** substances (such as water) have a pH of **7**. They are neither acidic nor basic.



Acids are **Proton Donors** and Bases are **Proton Acceptors**

1) Acids are **proton donors**. They **release hydrogen ions** (H^+) when mixed with water.

$$HCl_{(g)} \xrightarrow{\text{water}} H^+_{(aq)} + Cl^-_{(aq)}$$

2) The reverse happens with **bases** — they're proton acceptors so they **take H^+ ions**. **Alkalis** are bases that are **soluble** in water. They release **OH^-** ions in solution.

$$NaOH_{(s)} \xrightarrow{\text{water}} Na^+_{(aq)} + OH^-_{(aq)}$$

3) When an acid reacts with an alkali, a **salt** and **water** are formed — this is called a **neutralisation** reaction.

$$\text{acid} + \text{alkali} \rightarrow \text{salt} + \text{water}$$

You can show neutralisation just in terms of H^+ and OH^- ions. The hydrogen ions (H^+) from the acid will react with hydroxide ions (OH^-) from the base to produce water.

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

EXAMPLE: Write a balanced equation for the reaction between hydrochloric acid (HCl) and sodium hydroxide (NaOH).

This reaction is a **neutralisation reaction** — a hydrogen ion from HCl combines with a hydroxide ion from NaOH to form water. The remaining ions combine to form the salt:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$$

Some **Common Acids** and **Bases**

Acid	Formula
Hydrochloric acid	HCl
Sulfuric acid	H_2SO_4
Nitric acid	HNO_3
Ethanoic acid	CH_3COOH

Base	Formula
Sodium hydroxide	$NaOH$
Potassium hydroxide	KOH
Ammonia	NH_3

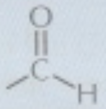
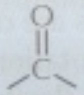
Siobhan always tells the truth, but Alka lies...

1) Write a balanced equation for the reaction between nitric acid and potassium hydroxide.
2) Write equations to show what happens when the following substances are mixed with water:
a) sulfuric acid, b) potassium hydroxide, c) nitric acid.

Organic Molecules

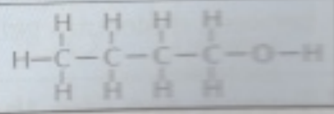
There are Lots of **Families** of Compounds in Organic Chemistry

Organic Chemistry is the study of organic compounds — these are just substances that contain **carbon**. Carbon compounds can be split up into different **groups** which have similar **properties** and **react** in similar ways. These groups are called **homologous series**. All the compounds in a homologous series contain the same **functional group** — a certain group of atoms that is responsible for the **properties** of the molecule. Here are some **common homologous series**:

HOMOLOGOUS SERIES	FUNCTIONAL GROUP	EXAMPLE
alkanes	-C-C-	propane — $CH_3CH_2CH_3$
alkenes	-C=C-	propene — $CH_3CH{=}CH_2$
alcohols	-OH	ethanol — CH_3CH_2OH
aldehydes		ethanal — CH_3CHO
ketones		propanone — CH_3COCH_3
carboxylic acids	-COOH	ethanoic acid — CH_3COOH

There are Different Ways of **Representing** a Molecule's **Structure**

Chemists have a few different ways of representing an organic molecule's **formula**. Here are a few ways that you'll need to be able to interpret:

FORMULA	WHAT IT SHOWS YOU	FORMULA FOR BUTANOL (an alcohol)
General formula	This describes **any member** in a homologous series. The number of carbons is represented by 'n' and the number of hydrogens in terms of 'n'.	$C_nH_{2n+1}OH$ (this is true for all alcohols.)
Molecular formula	This shows the number of **atoms** of each **element** in a molecule.	$C_4H_{10}O$
Structural formula	This shows the molecule **carbon by carbon**, with all attached hydrogens and functional groups.	$CH_3CH_2CH_2CH_2OH$
Skeletal formula	The **bonds** of the carbon skeleton are drawn, with any **functional groups**. The carbon atoms and attached hydrogens aren't shown.	
Displayed formula	All the **atoms** and **bonds** are drawn to show how the molecule is arranged.	

Organic Chemistry — no pesticides, no added sugars, no flavourings...

1) Draw the skeletal and displayed formulae for the molecule with the structural formula $CH_3CHOHCH_2CH_3$.
2) What is the molecular formula of the compound with the structural formula CH_3CH_2COOH?