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:

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

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.

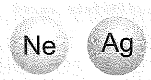
Roman Numerals Tell you the *Oxidation Number*

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 = $FeCl_3$

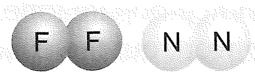
There are Some *Rules* About *Oxidation Numbers*

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



Uncombined elements.
Oxidation number = 0

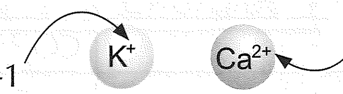
2) Elements that are bonded to identical atoms also have an oxidation number of **0**.



Elements bonded to identical elements.
Oxidation number = 0

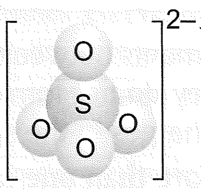
3) The oxidation number of an ion made up of just one atom is the same as its **charge** (the little number to the **right** of the symbol).

Oxidation number = +1



Oxidation number = +2

4) For **molecular ions** (ions that are made up of more than one atom) the **oxidation number** of the whole ion is the same as the **overall charge** of the ion.



Overall oxidation number is −2.

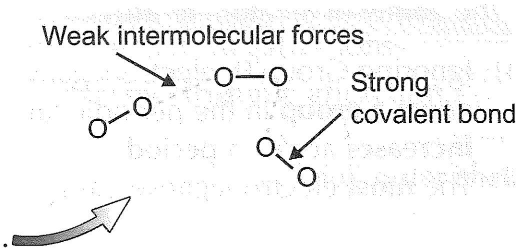
Caesar the day — and get to know your Roman numerals...

1) What does the oxidation number tell you about an atom?
2) Give the overall oxidation numbers for the following species:
Al^{3+}, NH_4^+, 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).



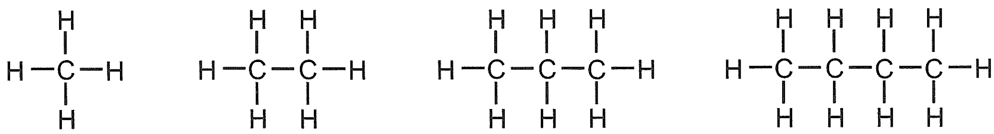
The ***Strength*** *of Intermolecular Bonds Affects* ***Boiling Points***

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...

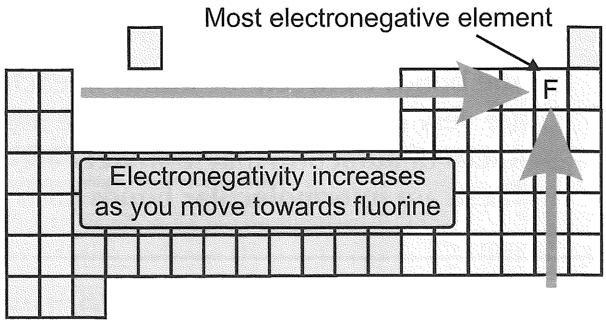
1) Draw a diagram to show the different types of bonding in a sample of gaseous chlorine molecules ($Cl_{2(g)}$). What type of bond is the strongest?
2) Use the data in the example above to predict the boiling points of the next four members of the alkane series. They are called pentane, hexane, heptane 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**.

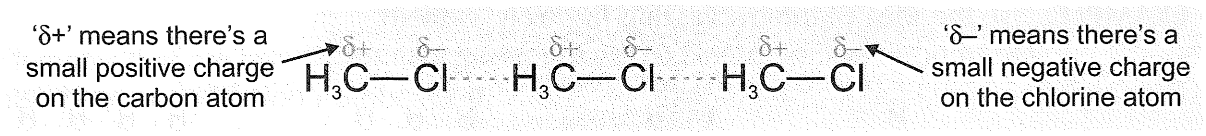


Polar Bonds** Can Affect the **Strength** of **Intermolecular Forces

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

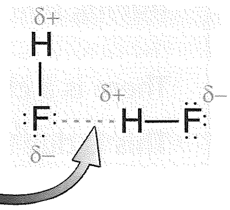
The reason for the dramatic increase in boiling point is that the chlorine atom **polarises** the molecule, making one end **slightly positive** and the other **slightly negative**. The **oppositely charged** ends of **different** molecules **attract** each other, so more energy is required to separate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

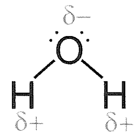
This is because the hydrogen atoms are strongly **polarised** by the very electronegative fluorine, oxygen or nitrogen atoms. These slightly positive hydrogen atoms are attracted to the lone pair of electrons on a F, O or N atom in a **nearby molecule** to form an attraction known as a **hydrogen bond**.



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.



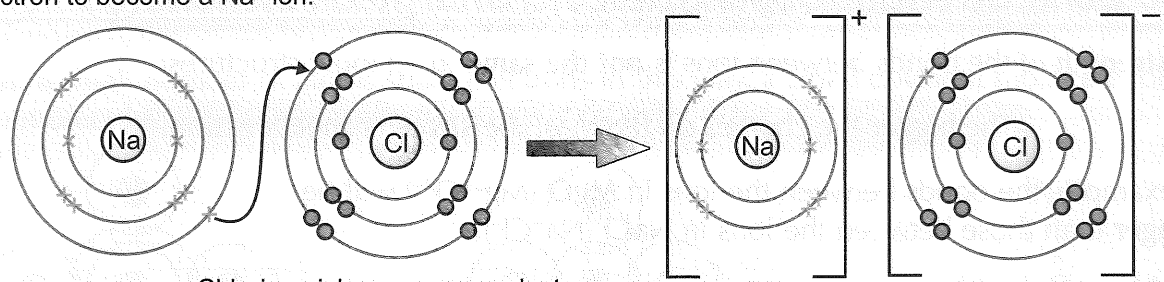
Ionic Bonding

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):

Sodium gives up its outer electron to become a Na^+ ion.

The positively charged Na^+ ion is attracted to the negatively charged Cl^- ion, forming an ionic bond.

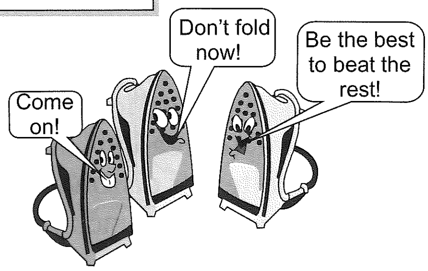


Chlorine picks up a spare electron from Na to become a Cl^- ion.

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?