CGP

Head Start to A-Level Chemistry

46

Bridging the gap between GCSE and A-Level

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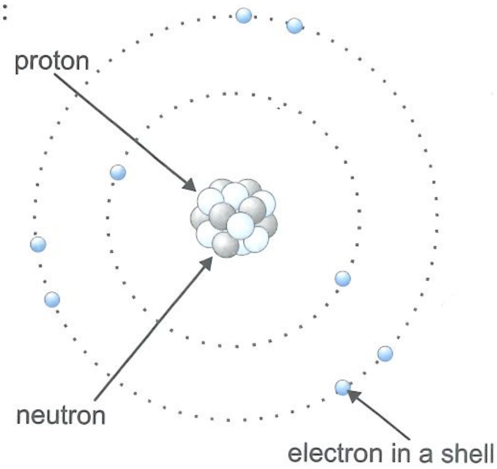
Atomic Structure

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?

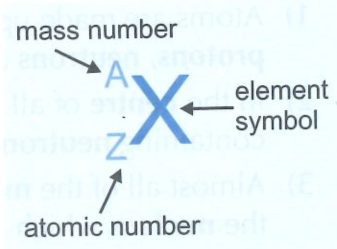
The Al^{3+} ion has a charge of **+3**, so there must be **3 more protons** than **electrons**.
Ions are formed when **electrons** are lost or gained, so Al^{3+} must have **lost 3 electrons**.

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 $^{37}Cl \times 37$) ÷ 100

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

$= (2625 + 925) \div 100$

$= 3550 \div 100$

$=$ **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.

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

$=$ **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.

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+}

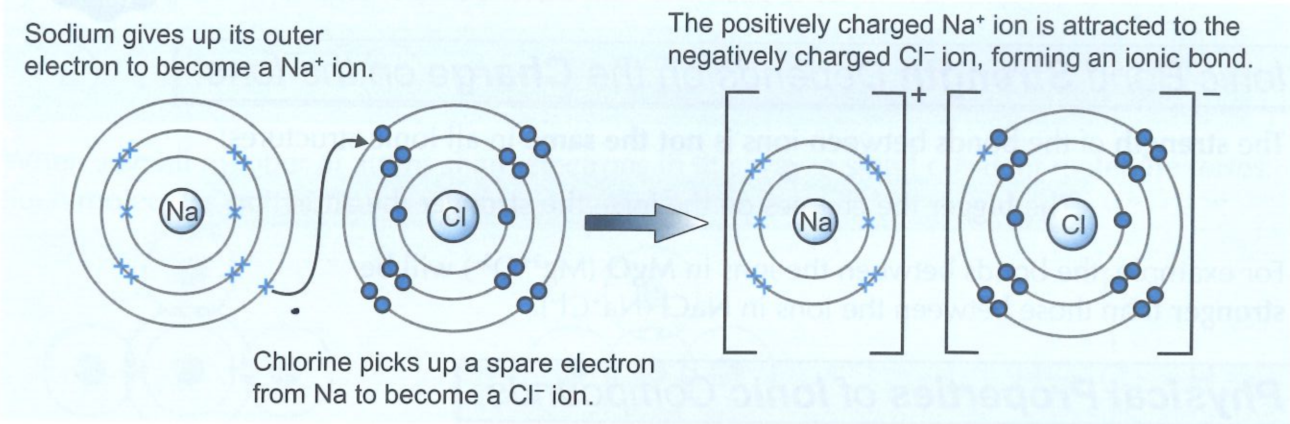
I never ask for an anion's opinion — they're always so negative...

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.

Ionic Bonding

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

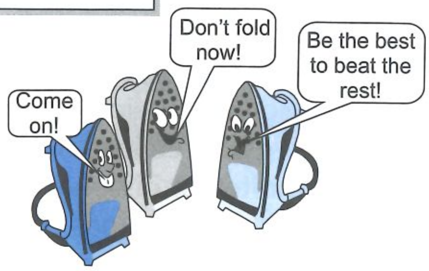
1) Ions form when **electrons** are transferred from **one atom** to **another**. Atoms that **lose electrons** form **positive ions** and atoms that **gain electrons** become **negative ions**.
2) These oppositely charged ions are **attracted** to each other by **electrostatic attraction**. When this 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 **$CaCl_2$**.

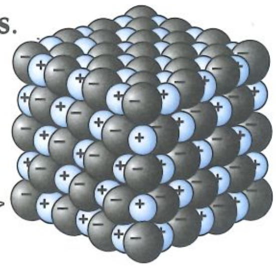
I can't afford Mg^{2+} — the charge is just too high...

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

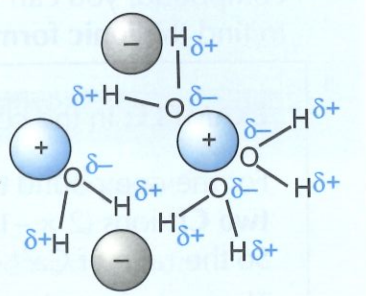
In order to **melt** a solid, the forces holding the particles together have to be **overcome**.
In an ionic solid, these bonds are very **strong**, so a **large** amount of energy is required to break them. So, ionic compounds have very **high** melting points.

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*

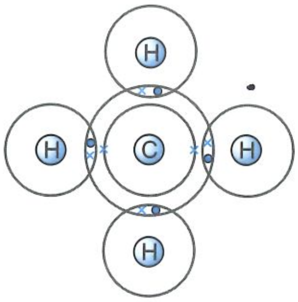
Ionic bonding only really works between elements that have to gain or lose one or two electrons to get a full outer shell. Elements with **half-full** shells have to do something different. These elements **share** their electrons with another atom so they've both got a full outer shell. Both positive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

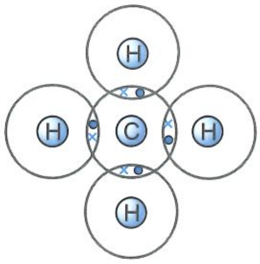
A covalent bond is a **shared pair** of electrons.

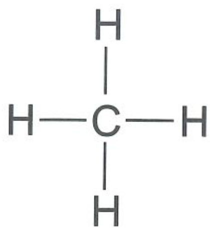
For example, two hydrogen atoms share a pair of electrons to form a covalent bond:

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:





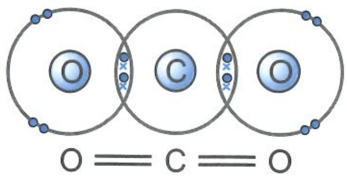


Dots represent electrons from the Hs and crosses represent electrons from C.

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

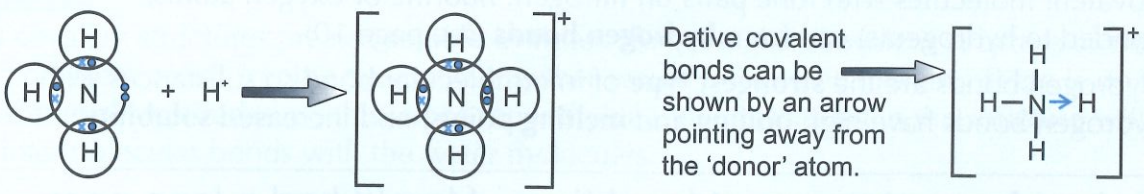
Each dash represents a single covalent bond (this is the most common notation).

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^+):



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

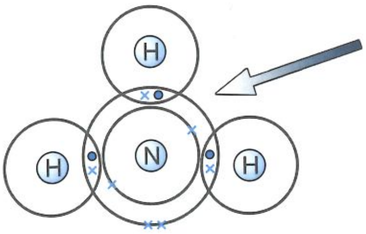
Small covalent molecules don't contain any of the **free charged** particles that are needed to carry an electric current. As a result they **cannot** conduct electricity — they're electrical **insulators**.

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...

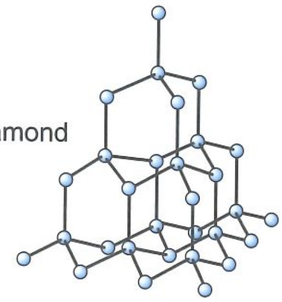
1) Draw a dot-and-cross diagram to show the bonding in hydrogen fluoride (HF). Label the bonding electrons and lone pairs of electrons.
2) Explain why nitrogen is a gas at room temperature, despite the nitrogen atoms in each molecule being strongly bonded to each other.

Giant Covalent Structures

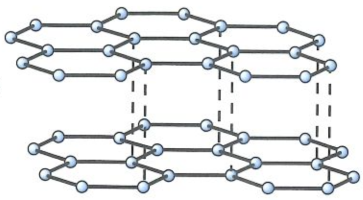
Giant Covalent Structures

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**.



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



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

Properties of Giant Covalent Structures

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

Melting points

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

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

Electrical conductivity

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

Solubility

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

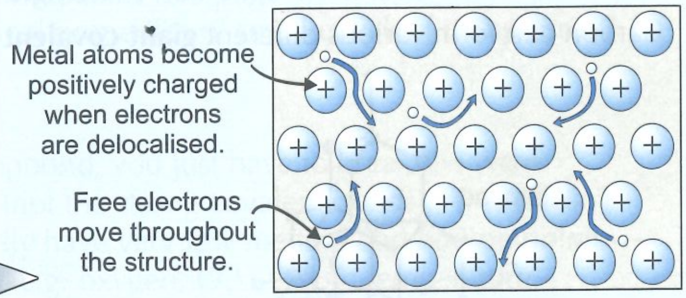
Diamonds — don't mess with 'em — they're well 'ard...

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

Metallic Bonding

Metals have Giant Structures Too

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



Properties of Metals

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

Melting points

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

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

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

Electrical conductivity

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

Solubility

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

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

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

Trends in Properties Across the Periodic Table

Structure and Bonding Change Across the Periodic Table

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

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

Trends Across Period Three

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

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

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

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

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

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

Trending now — #Arewedoneyet? #Dontworryitstheendofthesection...

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

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:

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

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 $2 \times +1 =$ **+2**.
- On the right hand side, the sodium ion contributes +1, so the charge is $1 \times +1 =$ **+1**.

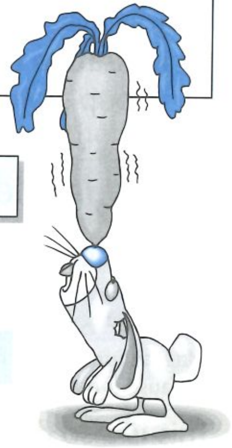
To get the charges to balance, you need another positive charge on the right-hand side. One way of doing this is by adding another sodium ion to the products:

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

Now check that the number of atoms still balances:
The Hs balance, but there are 2Nas 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: $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$.

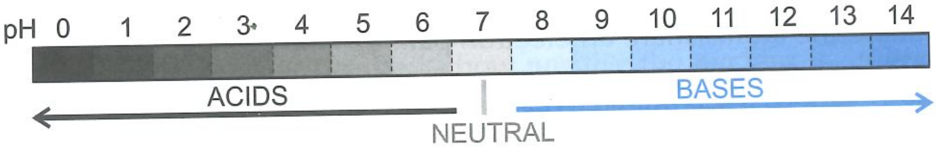
Hold one ear and stare at something still — it'll help you balance...

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.

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.

The Mole

A Mole is a Number of Particles

If you had a sample of a substance, and you wanted to **count** the number of atoms that were in it, you'd have to use some very **big numbers**, and spend a very long time counting. So you need a **unit** to describe the **amount** of a substance that you have — that unit is the **mole**.

One mole of a substance contains **6.02×10^{23}** particles.
6.02×10^{23} mol^{-1} is known as **Avogadro's constant**.

The particles can be **anything** — e.g. atoms or molecules (or even giraffes).
So 6.02×10^{23} atoms of **carbon** is 1 mole of carbon,
and 6.02×10^{23} molecules of CO_2 is 1 mole of CO_2.

Molar Mass is the Mass of One Mole

One mole of atoms or molecules has a **mass in grams** equal to the **relative formula mass** (A_r or M_r) of that substance.

For **carbon**, A_r = **12.0** so 1 mole of carbon weighs **12 g** and the **molar mass** is **12 g mol^{-1}**.
For **CO_2**, M_r = **44.0** so 1 mole of CO_2 weighs **44 g** and the **molar mass** of CO_2 is **44 g mol^{-1}**.
So, **12.0 g** of **carbon** and **44.0 g** of **CO_2** must contain the **same number of particles**.

You can use molar mass in calculations to work out how many moles of a substance you have.
Just use this formula:

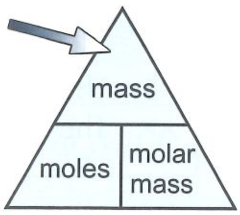
$$\text{Number of moles} = \frac{\text{Mass of substance (g)}}{\text{Molar mass (g mol}^{-1})}$$

g mol^{-1} is the same as g/mol.

EXAMPLE: How many moles of sodium oxide are present in 24.8 g of Na_2O?

Molar mass of Na_2O = $(2 \times 23.0) + (1 \times 16.0) = 62.0$ g mol^{-1}
Number of moles of Na_2O = 24.8 g ÷ 62.0 g mol^{-1} = **0.400 moles**

You can **rearrange** the formula above and use it to work out the mass of a substance or its relative formula mass (see page 3). It can help to remember this triangle:



EXAMPLE: What is the mass of 1.30 moles of magnesium oxide (MgO)?

Molar mass of MgO = $(1 \times 24.3) + (1 \times 16.0) = 40.3$ g mol^{-1}
Rearranging the formula, mass = moles × molar mass
So mass of MgO = 1.30×40.3 = **52.4 g** (3 s.f.)

Avocado's constant: how much I need to satisfy my guacamole craving...

1) Find the molar mass of sulfuric acid, given that 0.700 moles weighs 68.6 g.
2) How many moles of sodium chloride are present in 117 g of NaCl?
3) I have 54.0 g of water (H_2O) and 84.0 g of iron (Fe). Do I have more moles of water or of iron?

Determination of Formulae from Experiments

Empirical and *Molecular* Formulae

The **empirical formula** of a compound is the **simplest ratio** of the atoms of each element in the compound.

The **molecular formula** of a compound gives the **actual number** of atoms of each element in the compound.

For example, a compound with the molecular formula C_2H_6 has the empirical formula CH_3. The **ratio** of the atoms is one C to every three Hs.

Calculating *Empirical* Formulae

Often, the only way to find out the formula of a compound is through **experimentation** and **calculation**. You can calculate the formula of a compound from the **masses** of the **reactants**. Here is a simple set of rules to follow when calculating a formula:

1) Write the **mass** or **percentage mass** of each element.
2) Find the number of **moles** of each substance by dividing by the atomic or molecular mass.
3) Divide all answers by the **smallest** answer.
4) If required: multiply to make up to **whole numbers**.
5) Use the **ratio** of atoms to write the formula (this gives the empirical formula).

EXAMPLE: Find the formula of an oxide of aluminium formed from 9.00 g aluminium and 8.00 g oxygen.

1) First write down the mass of each substance:
   Al: 9.00 g   O: 8.00 g
2) Divide the mass by the atomic masses to find the number of moles of each substance:
   Al: $9.00 \div 27.0 = 0.333$ moles   O: $8.00 \div 16.0 = 0.500$ moles
3) Divide by the smallest number, which is 0.333:
   Al: $0.333 \div 0.333 = 1.00$ O: $0.5 \div 0.333 = 1.50$
4) Multiply by 2 to give whole numbers:
   Al: $1.00 \times 2 = 2$ O: $1.50 \times 2 = 3$
5) The ratio of Al : O is **2 : 3**.
   The empirical formula is Al_2O_3.

Roman empirical formula — 1 Caesar, 3 gladiators & 8 straight roads...

1) Find the empirical formulae of the following oxides:
   a) An oxide containing 12.9 g of lead to every 1.00 g of oxygen.
   b) An oxide containing 2.33 g of iron to every 1.00 g of oxygen.
   (Relative atomic mass values: Pb = 207.2, O = 16.0, Fe = 55.8)
2) Calculate the empirical formula of the carboxylic acid that is comprised of 4.30% hydrogen, 26.1% carbon and 69.6% oxygen.
   (Relative atomic mass values: H = 1.0, C = 12.0, O = 16.0)

Calculation of Molecular Formulae

Use the *Relative Formula Mass* to Work Out the Molecular Formula

To find the **molecular formula** from the **empirical formula**, you need to know the **relative formula mass** (see page 3) of the compound. This will usually be given to you in the question. Read through the example below and then try the questions.

EXAMPLE: Calculate the molecular formula of a hydrocarbon molecule if the compound contains 85.7% carbon and it's relative formula mass is 42.0.

First calculate the empirical formula:

In 100 g of the compound, there will be:

C: 85.7 g H: (100 g – 85.7 g) = 14.3 g

Hydrocarbons only contain carbon and hydrogen, so any mass that isn't carbon will be hydrogen.

Number of moles of each compound:

C: 85.7 ÷ 12.0 = 7.14 H: 14.3 ÷ 1.0 = 14.3

Divide by the smallest number (7.14):

C: 7.14 ÷ 7.14 = 1 H: 14.3 ÷ 7.13 = 2

So the ratio of C:H is **1:2**.

The empirical formula is $\mathbf{CH_2}$.

Calculate how many multiples of the empirical formula the molecular formula contains:

The empirical formula (CH_2) has a relative mass of 12.0 + 1.0 + 1.0 = 14.0.

The molecular formula has a relative mass of 42.0.

42.0 ÷ 14.0 = 3

To find the molecular formula, multiply each of the values in the empirical formula by 3:

C: 1 × 3 = 3 H: 2 × 3 = 6

The molecular formula is $\mathbf{C_3H_6}$.

The example above uses **percentage compositions** rather than the **mass** of each element in the compound. You can calculate the **percentage composition** yourself using the formula:

$$\text{percentage composition of element X} = \frac{\text{total mass of element X in compound}}{\text{total mass of compound}} \times 100\%$$

The percentage composition of my fridge is 80% cheese & 20% juice...

1) Calculate the molecular formula of a compound containing 52.2% carbon, 13.0% hydrogen and 34.8% oxygen if the relative formula mass of the compound is 46.0.
(Relative atomic mass values: C = 12.0, H = 1.0, O = 16.0)
2) Calculate the molecular formula of a hydrocarbon with a relative formula mass of 78.0 that contains 92.3% carbon.
(Relative atomic mass values: C = 12.0, H = 1.0)
3) Find the percentage composition of oxygen in each of the following compounds:
   a) Ethanol (C_2H_5OH).
   b) Nitric acid (HNO_3).
   c) Propanone (C_3H_6O).

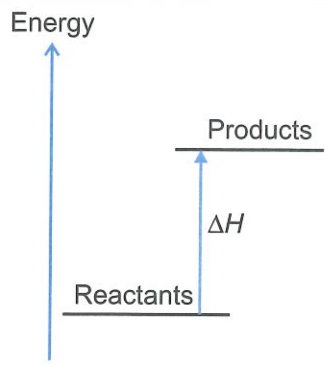
Endothermic and Exothermic Reactions

In an **exothermic** reaction, **heat** energy is **given out** (the room temperature rises).
In an **endothermic** reaction, **heat** energy is **taken** from the surroundings (the room temperature drops).

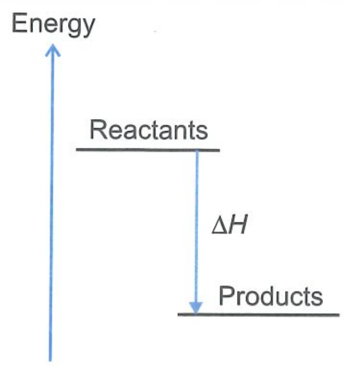
Making and Breaking Bonds

1) It takes energy to **break bonds**. When two atoms joined by a bond are **separated**, the energy required to do this must be provided from the surroundings.
2) However, energy is **released** when bonds are made. When two atoms become **joined together** by forming a bond, energy is **released** to the surroundings.
3) In a reaction, if more energy is taken in to break bonds than is given out when bonds are made, the process is **endothermic** — it will take in heat energy. The overall **enthalpy change** of the reaction (ΔH) is **positive**.
4) But, if more energy is given out when bonds are made than is taken in when bonds are broken, the process is **exothermic** — it will give out heat energy. The overall **enthalpy change** of an exothermic reaction (ΔH) is **negative**.

Reactions can be Represented by Energy Level Diagrams



In an **endothermic** reaction, the reactants **take in** energy from the surroundings. The products therefore have **more energy** than the reactants, and ΔH is **positive**.



In an **exothermic** reaction, the reactants **release** energy to the surroundings. The products therefore have **less energy** than the reactants and ΔH is **negative**.

After that I think I need a cup of tea. It'll help improve my energy level...

1) Are the following reactions exothermic or endothermic?
   a) burning coal
   b) sodium hydrogencarbonate + hydrochloric acid (temperature drops)
   c) acid + hydroxide (gets hotter)
   d) methane + steam (cools as they react)
2) a) Draw an energy level diagram for the following reaction:
   $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O \qquad \Delta H = -2809\ \text{kJ mol}^{-1}$
   You should label the products, reactants and enthalpy change on your diagram.
   b) Is the reaction in part a) endothermic or exothermic?

Bond Energy

Average Bond Energy

Bonds between **different atoms** require different amounts of **energy** to break them. When the **same two atoms** bond in the same way, the amount of energy needed is always about the same. The average bond energy values for some common bonds are given below:

C–H 413	C–O 360	C=C 612
O=O 498	H–H 436	C=O 743
C–C 348	O–H 463	

All these values are in kJ mol^{-1}.

The values tell you that:

e.g. It takes 413 kJ of energy to break 1 mole of C–H bonds.

It takes $463 \times 2 = 926$ kJ to break 1 mole of water (which has 2 O–H bonds per molecule) into oxygen and hydrogen atoms.

$743 \times 2 = 1486$ kJ are released when 1 mole of CO_2 (which has 2 C=O bonds) forms.

Calculating the **Change** in Energy

When a reaction takes place, the change in energy is simply:

sum of energy required to break old bonds – sum of energy released by new bonds formed

EXAMPLE: Calculate the energy change involved when 1 mole of methane burns in oxygen.

The equation for the reaction is: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

This tells you that 1 mole of methane reacts with 2 moles of oxygen to form 1 mole of carbon dioxide and 2 moles of water.

Step 1: Calculate the energy required to break all of the bonds between the reactant atoms:

4 C–H bonds $= 4 \times 413 = 1652$ kJ

2 O=O bonds $= 2 \times 498 = 996$ kJ

Total = 2648 kJ

Step 2: Calculate the energy released by all the new bonds formed in the products:

2 C=O bonds $= 2 \times 743 = 1486$ kJ

4 O–H bonds $= 4 \times 463 = 1852$ kJ

Total = 3338 kJ

Step 3: Combine the two values to give the overall value for the energy change:

The overall energy change is: 2648 – 3338 = **–690 kJ mol^{-1}**.

The negative sign shows that energy is being released to the surroundings, indicating that this is an **exothermic** reaction. This is expected, since this is a combustion reaction.

Ian Fleming was like an exothermic reaction — he made lots of Bonds...

1) Calculate the energy change of the following reactions:
(Use the values for the average bond energies given at the top of the page).
   a) burning 1 mole of propane $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
   b) burning 1 mole of ethanol $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
   c) hydrogenation of 1 mole of ethene $C_2H_4 + H_2 \rightarrow C_2H_6$

Answers

Section 1 — The Structure of the Atom

Page 1 — Atomic Structure

1 Protons and neutrons.

2 +2

3 −2

Page 2 — Atomic Number, Mass Number and Isotopes

1 $A - Z = 31 - 15 =$ **16**

2 Two isotopes of the same element have the same number of protons and electrons but different numbers of neutrons.

3 All three isotopes have 6 protons and 6 electrons. Carbon-12 has (12 − 6) = 6 neutrons, carbon-13 has (13 − 6) = 7 neutrons and carbon-14 has (14 − 6) = 8 neutrons.

Page 3 — Relative Atomic Mass

1 $[(8 \times 6) + (92 \times 7)] \div 100 =$ **6.92**

2 $[(99 \times 12) + (1 \times 13)] \div 100 =$ **12.01**

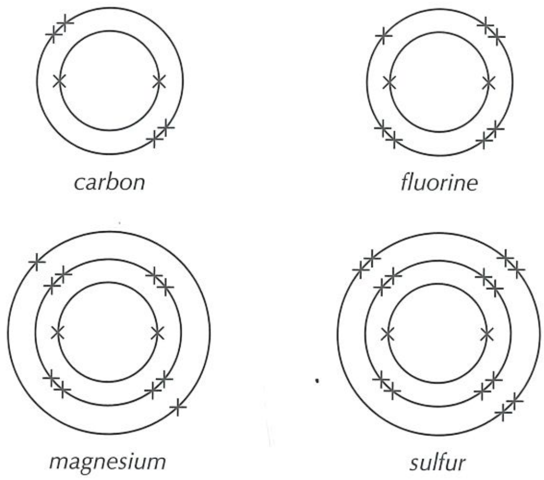
3 $[(52 \times 107) + (48 \times 109)] \div 100 =$ **107.96**

4 $23.0 + 19.0 =$ **42.0**

5 $12.0 + (3 \times 1.0) + 35.5 =$ **50.5**

Page 4 — Electronic Structure

1



2 Lithium: 2,1 Sodium: 2,8,1
Potassium: 2,8,8,1 Beryllium: 2,2
Magnesium: 2,8,2 Calcium: 2,8,8,2

3 Oxygen: $1s^2\ 2s^2\ 2p^4$
Chlorine: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Page 5 — The Periodic Table

1

s-block elements	p-block elements
caesium	phosphorus
potassium	aluminium
calcium	sulfur
barium	

2 E.g. They have the same number of electrons in their outer shell. / They react in similar ways.

Section 2 — Formation of Ions

Page 6 — Ionisation Energy

1 $Na_{(g)} \rightarrow Na^+_{(g)} + e^-$

2 Nuclear charge, the distance of the electron from the nucleus and shielding by inner electrons.

3 Magnesium
Fluorine
Oxygen

Page 7 — Formation of Ions

1 +1

2 Group 7

3 SO_3^{2-}

Page 8 — Oxidation Numbers

1 The oxidation number tells you how many electrons an atom has donated or accepted.

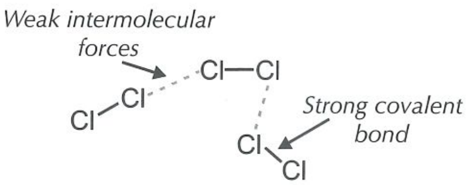
2 Al^{3+}: +3 H^+: +1 Ne: 0 O^{2-}: −2

3 0

Section 3 — Intermolecular Bonding

Page 9 — Intermolecular Bonding

1

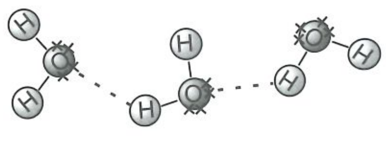


2 The trend should show an increase in boiling points as size of the alkane increases.
E.g. pentane: 36 °C, hexane: 69 °C, heptane: 98 °C, octane: 126 °C.

Page 10 — Polarity

1 a) HF as it is polar, and H_2 is non-polar.
b) H_2O as it can form hydrogen bonds, and H_2S can't.
c) CH_3F as fluorine is much more electronegative than carbon so it will be a polar molecule. Iodine is less electronegative / iodine and carbon have similar electronegativities, so CH_3I is non-polar.

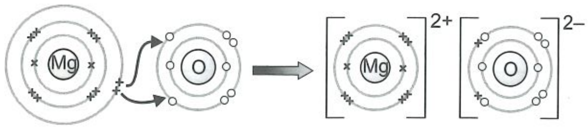
2 E.g.



Section 4 — Bonding and Properties

Page 11 — Ionic Bonding

1



2 You need two K^+ ions (2 × +1) to balance out each O^{2-} ion (1 × −2), so the ratio is 2:1. The ionic formula is K_2O.

Answers

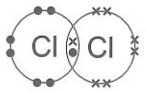
Page 12 — Ionic Compounds

1 BeO, Li_2O, LiF
The higher the charges on the ions, the stronger the bonds between them. The stronger the bonds, the higher the melting point. Beryllium oxide is formed from ions which both have charges of magnitude 2. Lithium oxide is formed from oxide ions with a −2 charge and lithium ions with only a +1 charge. The ions in lithium fluoride both have charges of magnitude one. Therefore the strongest bonds will be in beryllium oxide, followed by lithium oxide, then lithium fluoride.

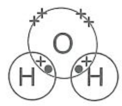
2 When potassium chloride is solid, the K^+ and Cl^- ions are held together in an ionic lattice, so they're not free to move and conduct electricity. When molten or dissolved, the ions separate, so they're free to move and able to carry a current.

Page 13 — Covalent Bonding

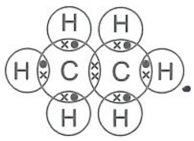
1 a) Chlorine:



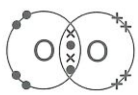
b) Water:



c) Ethane:

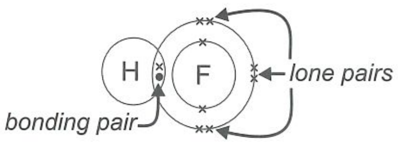


d) Oxygen:



Page 14 — Small Covalent Molecules

1



2 For a small covalent compound to boil, you have to break the weak intermolecular bonds between the molecules rather than the strong covalent bonds between the atoms in a molecule. These intermolecular bonds don't need much energy to break, so nitrogen has a low boiling point and is a gas at room temperature.

Page 15 — Giant Covalent Structures

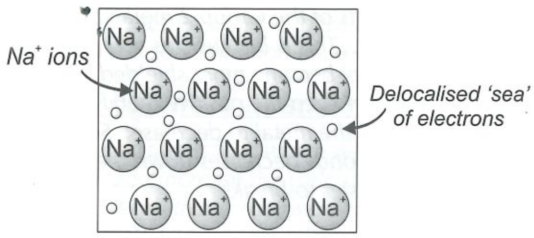
1 You could use solubility, though not all ionic compounds are soluble in water so this test may prove inconclusive. Melt and check to see if the molten substance conducts electricity— if it does it's probably ionic. You need to be careful of graphite though, which is a giant covalent molecule that is able to conduct electricity. To get around this, you could test the conductivity of the crystal in its solid form as well. While solid graphite will conduct, ionic salts only conduct electricity when molten or in solution.

2 In sodium chloride, the energy required to break the strong ionic bonds is provided when the ions become surrounded by water molecules. Fewer strong bonds are replaced by many more weaker bonds. In the case of a giant covalent molecule there is no way to get the energy required to break the many strong covalent bonds between atoms, so diamond doesn't dissolve.

Page 16 — Metallic Bonding

1 Calcium is likely to have a higher melting point. This is because calcium is made up of a lattice of Ca^{2+} ions, with two delocalised electrons per ion. Potassium however is made up of K^+ ions and only one delocalised electron per ion. So the bonding in calcium is stronger, and will require more energy to be broken leading to a higher melting point.

2



3 Similarities, e.g. high melting and boiling points / both conduct electricity when molten.
Differences, e.g. solid sodium chloride is soluble in water, sodium is not / solid sodium chloride is an electrical insulator but solid sodium conducts electricity.

Page 17 — Trends in Properties Across the Periodic Table

1 The melting points of the oxides of sodium, magnesium and aluminium are all high. Also, all three are conductors of electricity when molten. These properties clearly point towards the oxides being ionic.
Silicon dioxide also has a high melting point but is a non-conductor when molten so it has a giant covalent structure.
Oxides of phosphorus and sulfur have low melting points and are non-conductors. These are therefore likely to be small covalent molecules.

2 The melting points will be high up to silicon (which forms strong covalent bonds) then drop down to phosphorus and sulfur which are small covalent molecules.

3 a) Sodium chloride has ionic bonds.
b) The chloride of phosphorus has covalent bonds and is a small covalent molecules.

Section 5 — Chemical Equations

Page 19 — Writing and Balancing Equations

1 Step 2: $CH_4 + O_2 \rightarrow CO_2 + H_2O$
Step 3: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
(The Cs already balance, so put a 2 in front of H_2O to balance the Hs. Now put a 2 in front of O_2 to balance the Os. Check that all still balances.)

2 a) $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
b) $Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

3 First balance the atoms:
$Cl_{2\,(g)} + Fe^{2+}_{(aq)} \rightarrow 2Cl^-_{(aq)} + Fe^{3+}_{(aq)}$
Then balance the charges:
$Cl_{2\,(g)} + 2Fe^{2+}_{(aq)} \rightarrow 2Cl^-_{(aq)} + 2Fe^{3+}_{(aq)}$

Answers

Section 6 — Inorganic Chemistry

Page 20 — Group 2

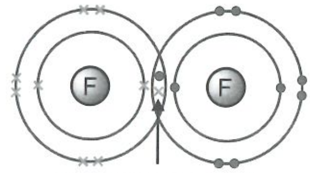
1 Reactivity increases down Group 2, so you'd expect magnesium to react very slowly with cold water to produce hydrogen gas and magnesium hydroxide. You'd expect strontium to react vigorously with cold water to produce hydrogen gas and strontium hydroxide.

2 The boiling points of the Group 2 metals will decrease down the group. This is because, as you go down the Group, the nuclei become more shielded. This causes the attraction between the positive metal ions and the free electrons in the metal to decrease. So the strength of the metallic bonds decreases down the group, making them easier to break.

Page 21 — Group 7

1 a) Chlorine is more reactive than bromine, so it displaces the bromide in solution producing chloride and bromine.
b) Iodine is less reactive than chlorine, so no reaction takes place.
c) Iodine is less reactive than bromine, so no reaction takes place.
d) Chlorine is more reactive than iodine, so it displaces the iodide in solution producing chloride and iodine.

2



Page 22 — Acids and Bases

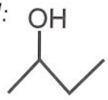
1 $HNO_3 + KOH \rightarrow KNO_3 + H_2O$

2 a) $H_2SO_4 \xrightarrow{\text{water}} 2H^+ + SO_4^{2-}$
b) $KOH \xrightarrow{\text{water}} K^+ + OH^-$
c) $HNO_3 \xrightarrow{\text{water}} H^+ + NO_3^-$

Section 7 — Organic Chemistry

Page 23 — Organic Molecules

1 skeletal:



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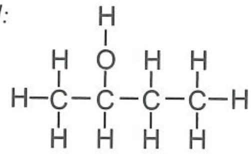
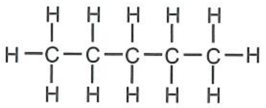


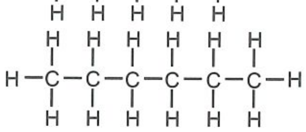
2 $C_3H_6O_2$

Page 24 — Alkanes

1 pentane:



hexane:

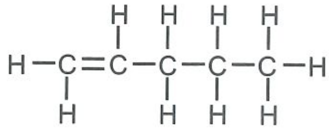


2 a) CH_4, C_2H_6, C_3H_8, C_4H_{10}.
b) General formula: C_nH_{2n+2}

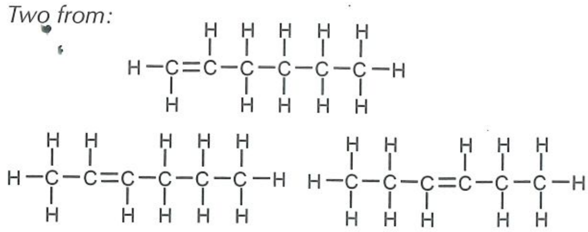
3 $CH_3CH_2CH_{3\,(g)} + 5O_{2\,(g)} \rightarrow 3CO_{2\,(g)} + 4H_2O_{(g)}$

Page 25 — Alkenes

1 e.g.



2 Two from:

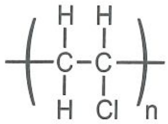


3 General formula: C_nH_{2n}

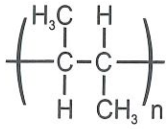
Page 26 — Polymerisation

1 Alkenes have a double bond that can open and link to other monomers.

2 a)

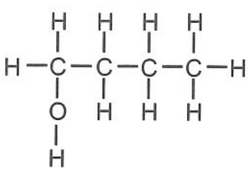


b)

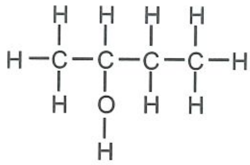


Page 27 — Alcohols

1



and



2 General formula: $C_nH_{(2n+1)}OH$

3 Ethanol will have a higher melting point than ethane because it is able to form hydrogen bonds. These require more energy to break than the intermolecular bonds that form between nonpolar molecules such as ethane.

Section 8 — Chemical Reactions

Page 30 — Reaction Types

1 a) combustion, exothermic, oxidation, redox, (reduction)
b) displacement, oxidation, precipitation, redox, reduction, substitution
c) exothermic, neutralisation, oxidation, substitution, redox, reduction
d) addition, hydrogenation, (oxidation, redox, reduction)

Answers

Section 9 — Rates of Reaction

Page 31 — Reaction Rates

1 a) e.g. Measure the change in temperature / Measure the change in pH / Measure the loss of mass as CO_2 is evolved / Measure the volume of CO_2 produced using a gas syringe.

b) e.g. Measure how long it takes for the solution to go cloudy / Measure the loss of mass as SO_2 is evolved / Measure the volume of SO_2 produced using a gas syringe.

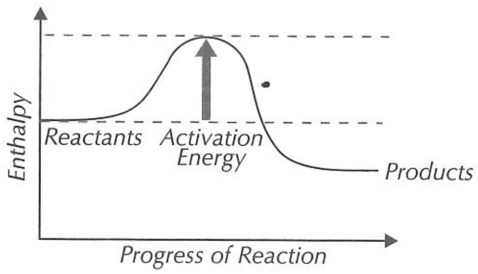
c) e.g. Measure the change in pH / Measure the loss of mass as CO_2 is evolved / Measure the volume of CO_2 produced using a gas syringe.

Page 32 — Collision Theory

1 The particles may have collided facing in the wrong direction or without enough energy to react.

2 The activation energy is the minimum amount of kinetic energy particles need to react.

3



Page 33 — Reaction Rates and Catalysts

1 e.g. Increase the temperature / Increase the concentration of the reactants / Add a catalyst.

2 The particles in the gas will be closer together so they're more likely to collide.

3 A catalyst is a substance that speeds up the rate of reaction without being changed or used up itself.

4 E.g. Catalysts make reactions cheaper to run and reduce their CO_2 emissions.

Section 10 — Equilibria

Page 34 — Reversible Reactions

1 a) At the start of the reaction, the rate of the forward reaction is faster than the rate of the backwards reaction.

b) At equilibrium, the rates of the forward and backward reactions are the same.

2 Dynamic equilibrium is where the forwards and backwards reactions of a reversible reaction are going at the same rate. This means that although both reactions are still happening, the concentrations of the products and reactants don't change.

Page 35 — Le Chatelier's Principle

1 Increasing the amount of steam will increase the concentration of particles on the left of the equation (which will also increase the pressure on the left hand side), and move the position of equilibrium to the right, increasing the yield of ethanol.

2 a) The equilibrium will move to the left to favour the endothermic reaction.

b) The equilibrium will move to the right to try and increase the concentration of ammonia.

Page 36 — Equilibrium and Yield

1 The temperature is low, which would favour the forward reaction, and increase the yield of ethanol.
But it is so low that the forward reaction rate will be much too slow to be economic.

2 The pressure is high, which would favour the forward reaction, and increase the yield of ethanol.
But such a high pressure would be very expensive to maintain, making the reaction uneconomic.

Section 11 — Calculations

Page 37 — The Mole

1 molar mass = mass ÷ moles
= 68.6 ÷ 0.700 = **98.0 g mol^{-1}**

2 Molar mass NaCl = 23.0 + 35.5 = 58.5 g mol^{-1}
moles = mass ÷ molar mass = 117 ÷ 58.5 = **2.00 moles**

3 Molar mass water = 16.0 + (2 × 1.0) = 18.0 g mol^{-1}
moles of water = 54.0 ÷ 18.0 = 3.00 moles.
Molar mass of iron = 55.8 g mol^{-1}
moles of iron = 84.0 ÷ 55.8 = 1.51 moles.
There are more moles of water.

Page 38 — Determination of Formulae from Experiments

1 a) Mass of each substance: Pb: 12.9 g O: 1.00 g
Number of moles of each substance:
Pb: 12.9 ÷ 207.2 = 0.0623 moles
O: 1.00 ÷ 16.0 = 0.0625 moles
Divide by the smallest number (0.0623):
Pb: 1.00 O: 1.00
The ratio of Pb : O is 1 : 1. The empirical formula is **PbO**.

b) Mass of each substance: Fe: 2.33 g O: 1.00 g
Number of moles of each substance:
Fe: 2.33 ÷ 55.8 = 0.0418 moles
O: 1.00 ÷ 16 = 0.0625 moles
Divide by the smallest number (0.0418):
Fe: 1.00 O: 1.50
Multiply by two to give whole numbers: Fe: 2 O: 3
The ratio of Fe : O is 2 : 3. The empirical formula is $\mathbf{Fe_2O_3}$.

2 Percentage composition of each substance:
H: 4.30% C: 26.1% O: 69.6%
Number of moles of each substance:
H: 4.3 ÷ 1.0 = 4.30 moles
C: 26.1 ÷ 12.0 = 2.18 moles
O: 69.6 ÷ 16.0 = 4.35 moles
Divide by the smallest number (2.18):
H: 1.97 C: 1.00 O: 2.00
The ratio of H : C : O is 2 : 1 : 2.
The empirical formula is $\mathbf{CH_2O_2}$.

Answers

Page 39 — Calculation of Molecular Formulae

1 Mass of each substance in 100 g:
C: 52.2 g H: 13.0 g O: 34.8 g
Number of moles of each substance:
C: 52.2 ÷ 12.0 = 4.35 moles
H: 13.0 ÷ 1.0 = 13.0 moles
O: 34.8 ÷ 16.0 = 2.18 moles
Divide by the smallest number (2.18):
C: 2.00 H: 5.96 O: 1.00
Ratio of C:H:O is 2:6:1.
The empirical formula is C_2H_6O.
The empirical formula has a relative mass of
(2 × 12.0) + (6 × 1.0) + (1 × 16.0) = 46.0. This is the same as the relative formula mass of the compound, so the molecular formula is also **C_2H_6O**.

2 Mass of each substance in 100 g:
C: 92.3 g H: 7.70 g
Number of moles of each substance:
C: 92.3 ÷ 12.0 = 7.69 moles
H: 7.70 ÷ 1.0 = 7.70 moles
Divide by the smallest number (7.69):
C: 1.00 H: 1.00
So the ratio of C:H is 1:1. The empirical formula is CH.
The empirical formula has a relative formula mass of
12.0 + 1.0 = 13.0.
78.0 ÷ 13.0 = 6.00, so there are six lots of the empirical formula in the compound.
The molecular formula is **C_6H_6**.

3 a) $\frac{1\times16.0}{(2\times12.0)+(6\times1.0)+(1\times16.0)}\times100 = \mathbf{34.8\%}$

b) $\frac{3\times16.0}{(1\times1.0)+(1\times14.0)+(3\times16.0)}\times100 = \mathbf{76.2\%}$

c) $\frac{1\times16.0}{(3\times12.0)+(6\times1.0)+(1\times16.0)}\times100 = \mathbf{27.5\%}$

Page 40 — Atom Economy

1 a) M_r C_2H_5OH = (2 × 12.0) + (6 × 1.0) + 16.0 = 46.0
M_r NaBr = 23.0 + 79.9 = 102.9
$\frac{46.0}{46.0+102.9}\times100 = \mathbf{30.9\%}$

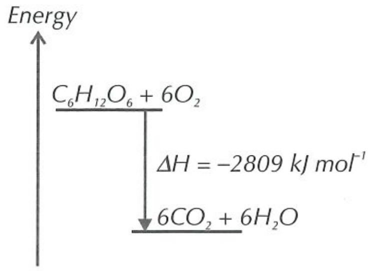
b) This reaction has an atom economy of 100%, so will be cheaper and greener to run.

Section 12 — Enthalpy

Page 41 — Endothermic and Exothermic Reactions

1 a) exothermic
b) endothermic
c) exothermic
d) endothermic

2 a)



b) exothermic

Page 42 — Bond Energy

1 a) Step 1: Calculate the energy required to break all of the bonds between the reactant atoms:

8 C — H bonds	=	8 × 413 = 3304
2 C — C bonds	=	2 × 348 = 696
5 O = O bonds	=	5 × 498 = 2490
		TOTAL = 6490

Step 2: Calculate the energy released by all the new bonds formed between the product atoms:

6 C = O bonds	=	6 × 743 = 4458
8 O — H bonds	=	8 × 463 = 3704
		TOTAL = 8162

Step 3: Find the overall value for the energy change:
+6490 − 8162 = **−1672 kJ mol^{-1}**

b) Step 1: Calculate the energy required to break all of the bonds between the reactant atoms:

1 C — C bond	=	1 × 348 = 348
1 C — O bond	=	1 × 360 = 360
5 C — H bonds	=	5 × 413 = 2065
1 O — H bonds	=	1 × 463 = 463
3 O = O bond	=	3 × 498 = 1494
		TOTAL = 4730

Step 2: Calculate the energy released by all the new bonds formed between the product atoms

4 C = O bonds	=	4 × 743 = 2972
6 O — H bonds	=	6 × 463 = 2778
		TOTAL = 5750

Step 3: Find the overall value for the energy change:
+4730 − 5750 = **−1020 kJ mol^{-1}**

c) Step 1: Calculate the energy required to break the H–H and C=C bonds:

1 H — H bond	=	1 × 436 = 436
1 C = C bond	=	1 × 612 = 612
		TOTAL = 1048

Step 2: Calculate the energy released by all the new bonds formed between product atoms

2 C — H bonds	=	2 × 413 = 826
1 C — C bond	=	1 × 348 = 348
		TOTAL = 1174

Step 3: Find the overall value for the energy change:
+1048 − 1174 = **−126 kJ mol^{-1}**

Section 13 — Investigating & Interpreting

Page 43 — Planning Experiments

1 a) The time taken for the lump of magnesium to react.
b) E.g. The mass of magnesium used / The concentration of hydrochloric acid / The volume of hydrochloric acid / The surface area of the magnesium.

Page 44 — Presenting and Interpreting Data

1 The anomalous result is 19.0 cm^3.
The mean is: (22.0 + 23.0 + 22.0 + 24.0) ÷ 4 = **22.8 cm^3**

Page 45 — Conclusions and Error

1 a) The uncertainty of the weighing scales is 0.05 g, so the percentage error = $\frac{0.05}{1.4}\times100$ = **3.6%**

b) The uncertainty of the clock is 0.5 s, so the percentage error = $\frac{0.5}{23}\times100$ = **2.2%**

c) The uncertainty of the thermometer is 0.1 °C, so the percentage error = $\frac{0.1}{10.6}\times100$ = **0.9%**