

At Teignmouth Community School we study the OCR A Specification qualification. A Level Chemistry involves the investigation of substances, what they are made of, how they interact and what role they play in modern industrial society. In this course, you will explore all aspects of chemistry: physical, inorganic and organic as well as modern analytical techniques. You will develop a logical approach to problem solving as well as your ability to understand abstract principles. The course encourages imaginative and critical thinking and develops your skills in laboratory procedures. You will be expected to take a greater responsibility for your own learning - organising your workload and time management to ensure that you meet all deadlines.

### Independent Learning Task:

Over the summer holidays you are expected to complete the following independent learning task to show that you are committed to taking the course. This work will help your teachers to assess your prior knowledge and will help you and your teacher set personal targets for improvement. It will also contribute to the judgment made during the probation period in September.

**The work should be completed by your first Chemistry lesson in September.**

You will need to study the book provided and ensure that you complete all the practice questions.

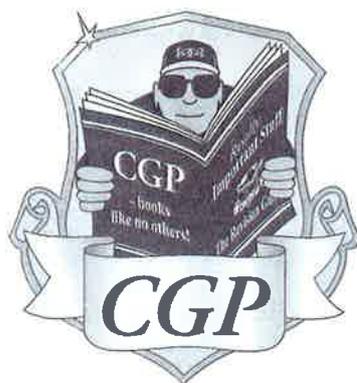
Make sure you revise the contents of the book thoroughly in preparation for a test during your first lesson in September.



## CHEMISTRY A: ASSESSMENT OVERVIEW

AS Chemistry A (H032) - First Exam June 2016					
ASSESSMENT OVERVIEW					
Paper		Marks	Duration	Weighting	
Paper 1	<b>Breadth in chemistry</b>		<b>70</b>	1 hr 30 mins	50%
	Section A	Multiple choice	20		
	Section B	Structured questions and extended response questions covering theory and practical skills	50		
Paper 2	<b>Depth in chemistry</b>		<b>70</b>	1 hr 30 mins	50%
	Structured questions and extended response questions, covering theory and practical skills		70		

A Level Chemistry A (H432) - First Exam June 2017					
ASSESSMENT OVERVIEW					
Paper		Marks	Duration	Weighting	
Paper 1	<b>Periodic table, elements and physical chemistry</b>		<b>100</b>	2 hr 15 mins	37%
	Section A	Multiple choice	15		
	Section B	Structured questions and extended response questions covering theory and practical skills	85		
Paper 2	<b>Synthesis and analytical techniques</b>		<b>100</b>	2 hr 15 mins	37%
	Section A	Multiple choice	15		
	Section B	Structured questions and extended response questions covering theory and practical skills	85		
Paper 3	<b>Unified chemistry</b>		<b>70</b>	1 hr 30 mins	26%
	Structured questions and extended response questions covering theory and practical skills		70		
Non-exam assessment	<b>Practical Endorsement for chemistry</b>		<b>Pass/Fail</b>	Non-exam assessment	Reported separately
	See pages 27 and 28. Teacher-assessed component common to Chemistry A and Chemistry B (Salters). Candidates complete a minimum of 12 practical activities to demonstrate practical competence. Performance reported separately to the A Level grade. Moderation details still to be confirmed by Ofqual at the time of going to press		0		



# Head Start to A-Level Chemistry

A-Level Chemistry is a **big step up** from GCSE... no doubt about that. But don't worry — this CGP book has been lovingly made to help you **hit the ground running** at the start of your A-Level (or AS-Level) course.

It recaps everything you'll need to remember from GCSE, and introduces some of the crucial concepts you'll meet at A-Level. For every topic, there are **crystal-clear study notes** and plenty of **practice questions** to test your skills.

What's more, it's fully up-to-date for the **new A-Level courses** starting in 2015.

## *What CGP is all about*

Our sole aim here at CGP is to produce the highest quality books — carefully written, immaculately presented and dangerously close to being funny.

Then we work our socks off to get them out to you — at the cheapest possible prices.

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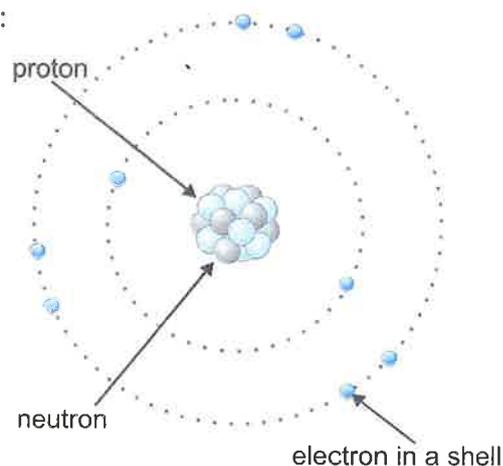
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# Atomic Structure

## 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  $\text{Al}^{3+}$  ion lost or gained?

The  $\text{Al}^{3+}$  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  $\text{Al}^{3+}$  must have **lost 3 electrons**.

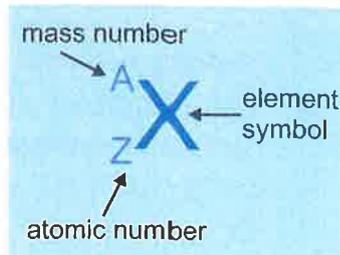
## Neutrons are the perfect criminals — they never get charged...

- 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  $^{63}\text{Cu}$  isotope has  $63 - 29 = 34$  neutrons.

The  $^{65}\text{Cu}$  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?

$$\begin{aligned}
 \text{Average mass} &= (\text{abundance of } ^{35}\text{Cl} \times 35 + \text{abundance of } ^{37}\text{Cl} \times 37) \div 100 \\
 &= [(75 \times 35) + (25 \times 37)] \div 100 \\
 &= (2625 + 925) \div 100 \\
 &= 3550 \div 100 \\
 &= \mathbf{35.5} \quad (\text{You can check your answer against a periodic table to see if it's right.})
 \end{aligned}$$

## 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  $\text{CaCl}_2$ .

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

$$\begin{aligned}
 M_r &= (1 \times 40.1) + (2 \times 35.5) \\
 &= \mathbf{111}
 \end{aligned}$$



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

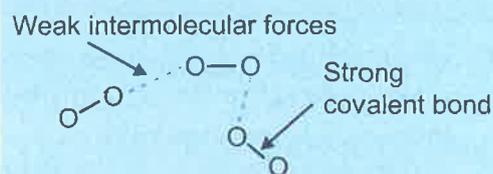
- 1) Find the relative atomic mass of lithium if its composition is 8%  $^6\text{Li}$  and 92%  $^7\text{Li}$ .
- 2) Find the relative atomic mass of carbon if its composition is 99%  $^{12}\text{C}$  and 1%  $^{13}\text{C}$ .
- 3) Find the relative atomic mass of silver if its composition is 52%  $^{107}\text{Ag}$  and 48%  $^{109}\text{Ag}$ .
- 4) Find the relative formula mass of sodium fluoride,  $\text{NaF}$ .
- 5) Find the relative formula mass of chloromethane,  $\text{CH}_3\text{Cl}$ .



# Intermolecular Bonding

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

- 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 ( $\text{H}_2\text{O}$ ) or oxygen ( $\text{O}_2$ ).
- 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).



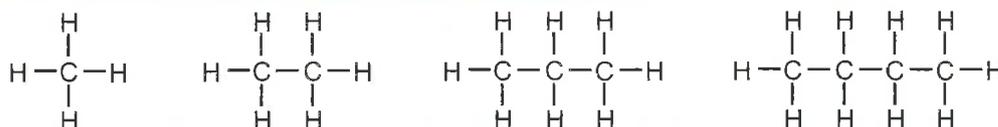
## The Strength of Intermolecular Bonds Affects Boiling Points

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:

- The number of **electrons** in a molecule: the **more** electrons there are, the **stronger** the intermolecular bonds between molecules.
- 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.

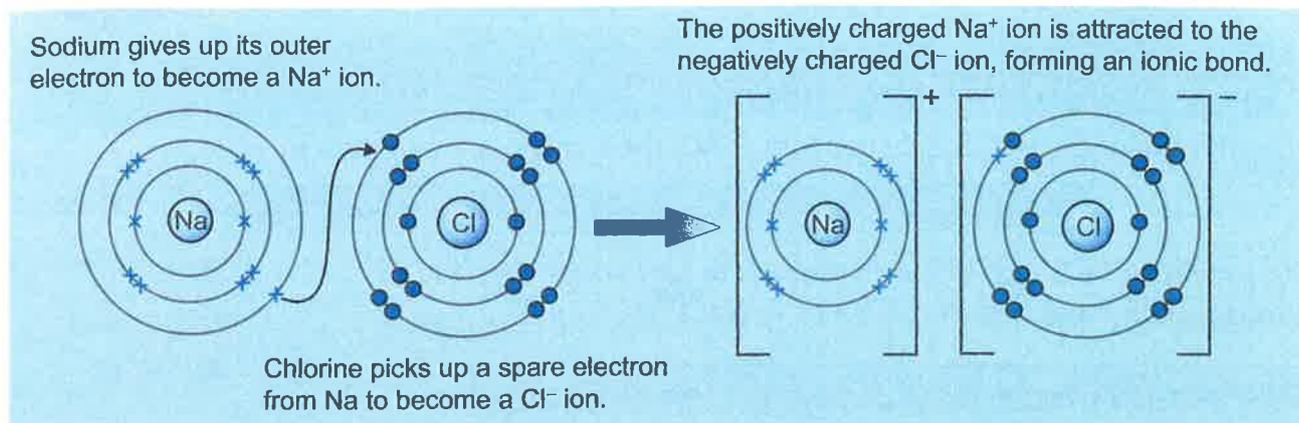
## This page was alright — we formed a sort of bond...

- Draw a diagram to show the different types of bonding in a sample of gaseous chlorine molecules ( $\text{Cl}_{2(\text{g})}$ ). What type of bond is the strongest?
- 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.)

# Ionic Bonding

## Ionic Bonds Involve the Transfer of Electrons

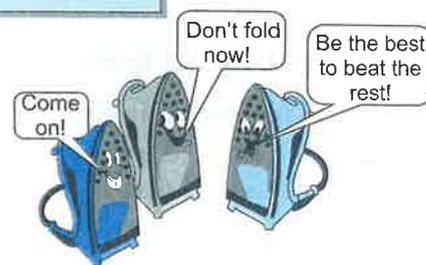
- 1) Ions form when **electrons** are transferred from **one atom** to **another**. Atoms that **lose electrons** form **positive ions** and atoms that **gain electrons** become **negative ions**.
- 2) These oppositely charged ions are **attracted** to each other by **electrostatic attraction**. When 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  $\text{Ca}^{2+}$  to  $\text{Cl}^-$  ions?

For the compound to be neutral it must contain **two  $\text{Cl}^-$  ions** ( $2 \times -1$ ) to **balance** the charge of **each  $\text{Ca}^{2+}$  ion** ( $1 \times +2$ ).  
So the ratio of  $\text{Ca}^{2+}$  ions to  $\text{Cl}^-$  ions in the compound must be **1:2**.  
The ionic formula will be  **$\text{CaCl}_2$** .

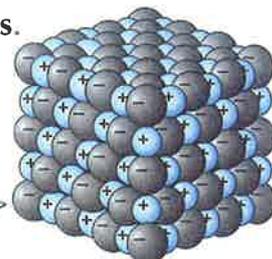
*I can't afford  $\text{Mg}^{2+}$  — the charge is just too high...*

- 1) Draw a diagram showing how a magnesium atom reacts with an oxygen atom to form magnesium oxide,  $\text{MgO}$ . Your diagram should show the electron transfer process.
- 2) In potassium oxide, what is the ratio of  $\text{K}^+$  ions to  $\text{O}^{2-}$  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** ( $\text{Mg}^{2+}\text{O}^{2-}$ ) will be **stronger** than those between the ions in **NaCl** ( $\text{Na}^+\text{Cl}^-$ ).

## 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.

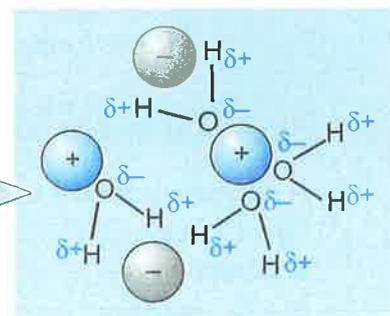
### Electrical conductivity

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 lattices...

- 1) Put these ionic compounds in order of melting point, highest to lowest: Lithium oxide ( $\text{Li}_2\text{O}$ ), Beryllium oxide ( $\text{BeO}$ ), Lithium fluoride ( $\text{LiF}$ ). Explain why you have put them in that order.
- 2) Explain why the ionic compound, potassium chloride ( $\text{KCl}$ ), can conduct electricity when molten or dissolved, but not when it is solid.

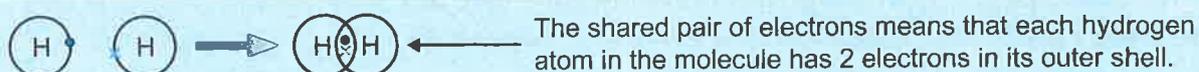
# Covalent Bonding

## Covalent Bonding Involves Shared Pairs of Electrons

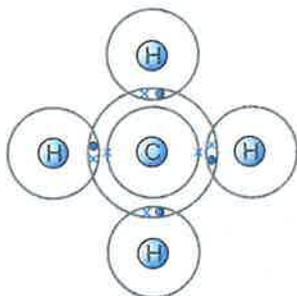
Ionic bonding only really works between elements that have to gain or lose one or two electrons to get a full outer shell. Elements with **half-full** shells have to do something different. These elements **share** their electrons with another atom so they've both got a full outer shell. Both positive 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.

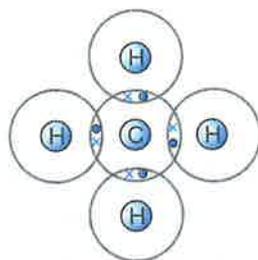
For example, two hydrogen atoms share a pair of electrons to form a covalent bond:



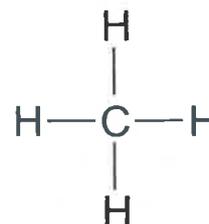
When a small number of atoms share electrons in this way, a small covalent molecule forms. Such molecules can be represented in several different ways:



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

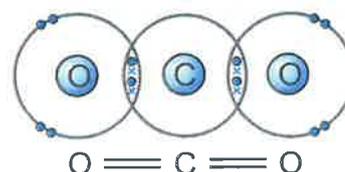


A more simple dot-and-cross diagram, showing only the outer shells of electrons.



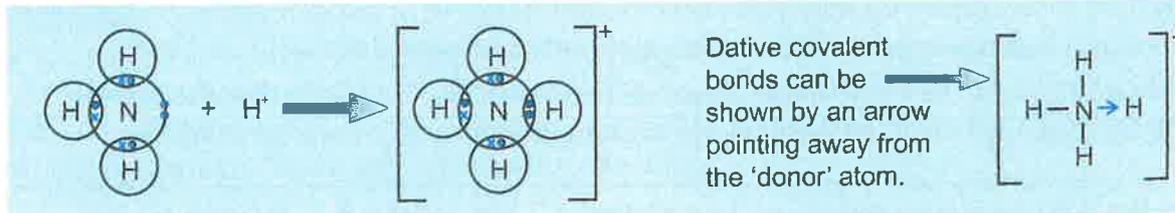
Each dash represents a single covalent bond (this is the most common notation).

If two atoms share **more than one** pair of electrons between them, then a **multiple covalent bond** can form. For example, in **carbon dioxide** ( $\text{CO}_2$ ), there are two  $\text{C}=\text{O}$  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 ( $\text{NH}_4^+$ ) there is a dative covalent bond formed from the nitrogen to a hydrogen ion ( $\text{H}^+$ ):



*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 ( $\text{Cl}_2$ )   b) water ( $\text{H}_2\text{O}$ )   c) ethane ( $\text{C}_2\text{H}_6$ )   d) oxygen ( $\text{O}_2$ )

# 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,  $\text{H}_2\text{O}$ ) or gases (e.g. oxygen,  $\text{O}_2$ ) at room temperature.

### Electrical Conductivity

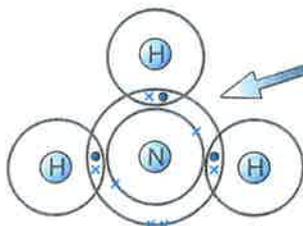
Small covalent molecules don't contain any of the **free charged** particles that are needed to carry an electric current. As a result they **cannot** conduct electricity — they're electrical **insulators**.

### 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 ( $\text{NH}_3$ ) 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**.

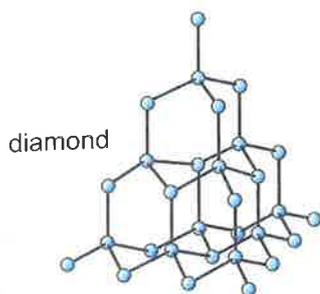
## Aisling had four satsumas at lunchtime. Harold had a lone pear...

- 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.

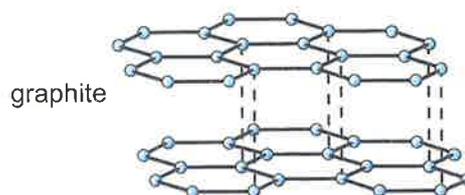
# Giant Covalent Structures

## 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?

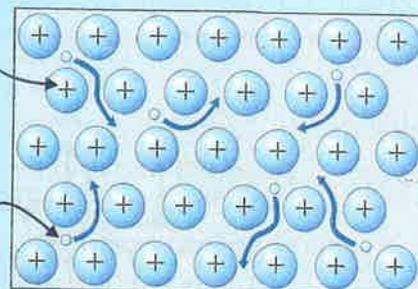
# Metallic Bonding

## 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**.

Metal atoms become positively charged when electrons are delocalised.

Free electrons move throughout the structure.



## 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  $\text{Mg}^{2+}$  ions with **two** delocalised electrons per atom. Sodium is made up of  $\text{Na}^+$  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 ( $\text{NaCl}$ ) is an ionic compound. Give one similarity and one difference 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.)

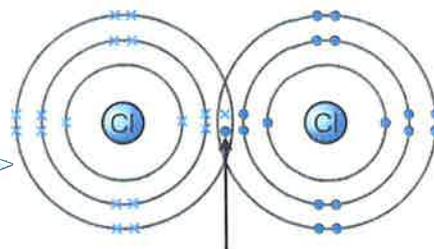
# Group 7

## 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. →



single covalent bond

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

- **Chlorine** (Cl<sub>2</sub>) is a **yellow/green gas** at room temperature.
- **Bromine** (Br<sub>2</sub>) is a **brown liquid** at room temperature.
- **Iodine** (I<sub>2</sub>) 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.



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

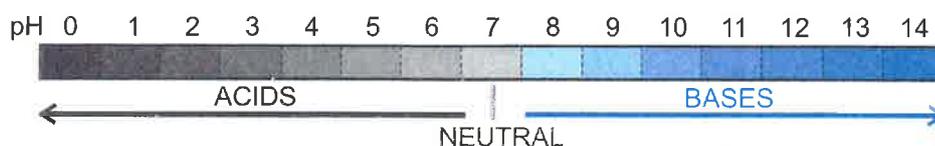
*I met a friend for coffee today — I said 'Hallo, Jen'...*

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

# Acids and Bases

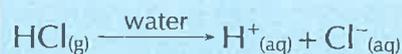
## The pH Scale

- 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**.
- Neutral** substances (such as water) have a pH of 7. They are neither acidic nor basic.

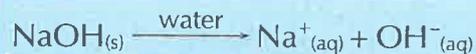


## Acids are Proton Donors and Bases are Proton Acceptors

- Acids are **proton donors**. They **release hydrogen ions** ( $\text{H}^+$ ) when mixed with water.



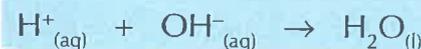
- The reverse happens with **bases** — they're proton acceptors so they **take  $\text{H}^+$  ions**. **Alkalis** are bases that are **soluble** in water. They release  **$\text{OH}^-$  ions** in solution.



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

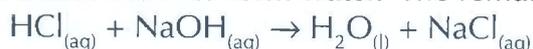


You can show neutralisation just in terms of  $\text{H}^+$  and  $\text{OH}^-$  ions. The hydrogen ions ( $\text{H}^+$ ) from the acid will react with hydroxide ions ( $\text{OH}^-$ ) from the base to produce water.



**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:



## Some Common Acids and Bases

Acid	Formula
Hydrochloric acid	HCl
Sulfuric acid	$\text{H}_2\text{SO}_4$
Nitric acid	$\text{HNO}_3$
Ethanoic acid	$\text{CH}_3\text{COOH}$

Base	Formula
Sodium hydroxide	NaOH
Potassium hydroxide	KOH
Ammonia	$\text{NH}_3$

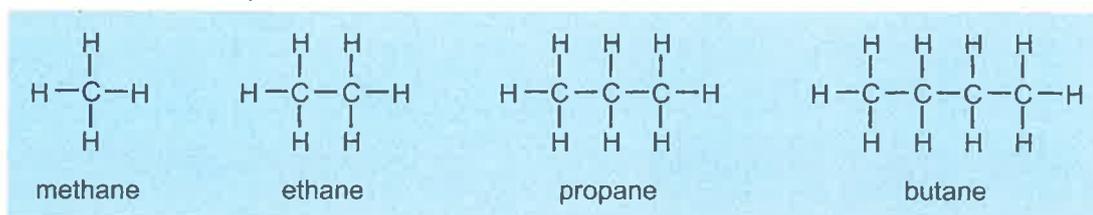
## *Siobhan always tells the truth, but Alka lies...*

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

# Alkanes

## Structure and Bonding in Alkanes

- 1) Alkanes are **hydrocarbons** — they **only** contain hydrogen and carbon atoms.
- 2) Alkanes contain **two types** of bond. All of the **carbon-carbon** bonds are **single covalent bonds**. All the other bonds are **carbon-hydrogen covalent bonds** (which are always single).
- 3) All of the available bonds have been formed, so we call alkanes **saturated** molecules.
- 4) **Carbon** always forms **four** covalent bonds and **hydrogen** makes **one** covalent bond.
- 5) The diagrams below show the structures of the first four **straight-chain alkanes**: methane, ethane, propane and butane.



It is important to realise that these structures are only **2D representations** of the **3D molecules**. The molecules are not rigid. There is **free rotation** around a carbon-carbon single bond. This means that the carbon chains are quite **flexible** and gives the molecules the ability to **change shape**, particularly as the chain length increases.

## Properties of Alkanes

The **bonds** in alkanes are very **strong** and it requires a **lot of energy** to break them. This can be used to explain some of their **properties**:

- 1) They are very **unreactive**.
- 2) They are **not** able to form **polymers**.
- 3) They **burn cleanly**, tending to undergo **complete combustion** to form **carbon dioxide** and **water** (see page 28). The flame is usually a faint blue colour.

For example, the combustion of ethane:



Also:

- 4) **Boiling point increases** as the **length** of the carbon chain increases.
- 5) **Viscosity** (resistance to flow) **increases** as chain length increases.
- 6) **Volatility** (ease of evaporation) **decreases** as chain length increases.

These last three properties are explained by the fact that the **attractive forces** between molecules get stronger as the chain length **increases** (page 9).



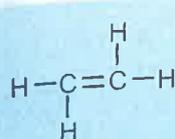
## Alkanes are like the weather in the UK — completely saturated...

- 1) Draw out the structures of the next two alkanes, pentane ( $\text{C}_5\text{H}_{12}$ ) and hexane ( $\text{C}_6\text{H}_{14}$ ).
- 2) a) Write out the molecular formulae of the first four alkanes.  
b) We can work out a general formula for the alkanes of the form  $\text{C}_n\text{H}_?$ , where  $n$  is the number of carbon atoms. Work out, in terms of  $n$ , what should be in place of the ?.
- 3) Write a balanced equation for the complete combustion of propane in oxygen.

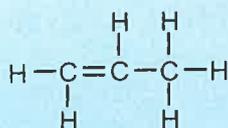
# Alkenes

## Structure and Bonding in Alkenes

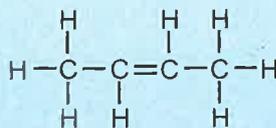
- 1) **Alkenes** are similar to alkanes in that they are also **hydrocarbons**. The difference is in the presence of a **carbon-carbon double covalent bond** (C=C) somewhere in the carbon chain.
- 2) This means **not** all possible single bonds have been made — these molecules are **unsaturated**.
- 3) As in all compounds the carbon atoms must have **four** bonds, and hydrogen only **one**.
- 4) The structures of the first three alkenes (ethene, propene and butene) are shown below:



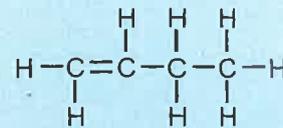
ethene



propene



butene



butene (again)

As you can see from butene, the presence of the C=C bond means that most alkenes have **more than one** possible structure. The C=C bond can be in various different **positions** along the chain.

Molecules with the same **molecular formula** but different **structures** are called **isomers**.

The C=C bond does not allow the same **free rotation** and flexibility around itself as a C–C bond. It is a **rigid** bond. But the rest of the carbon chain is the same as in an alkane molecule, so rotation is allowed around the **single** bonds.

## Properties of Alkenes

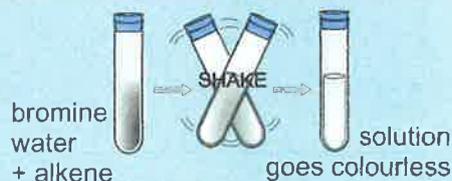
The presence of the C=C bond dictates the chemical properties of alkenes.

- 1) They are **reactive** compounds, undergoing many different types of chemical reaction.
- 2) They are used extensively to form **polymers**, e.g. poly(ethene) (see next page).
- 3) They **do not** burn **cleanly**, giving very **yellow** flames and lots of **soot**.

And as for alkanes, when you increase the **chain length** of an alkene:

- 4) The **boiling point** increases.
- 5) The **viscosity** increases.
- 6) The **volatility** decreases.

You can **test** for whether a compound is an alkene by adding it to **bromine water**. Alkenes **decolourise** bromine water, turning it from **orange** to **colourless**.



**'Sleeping Butene'** — coming soon to a cinema near you...

- 1) Draw out a structure for the next alkene: pentene (C<sub>5</sub>H<sub>10</sub>).
- 2) Draw out two alternative structures for hexene (C<sub>6</sub>H<sub>12</sub>).
- 3) Work out the general formula for the alkenes of the form C<sub>n</sub>H<sub>2n</sub>.

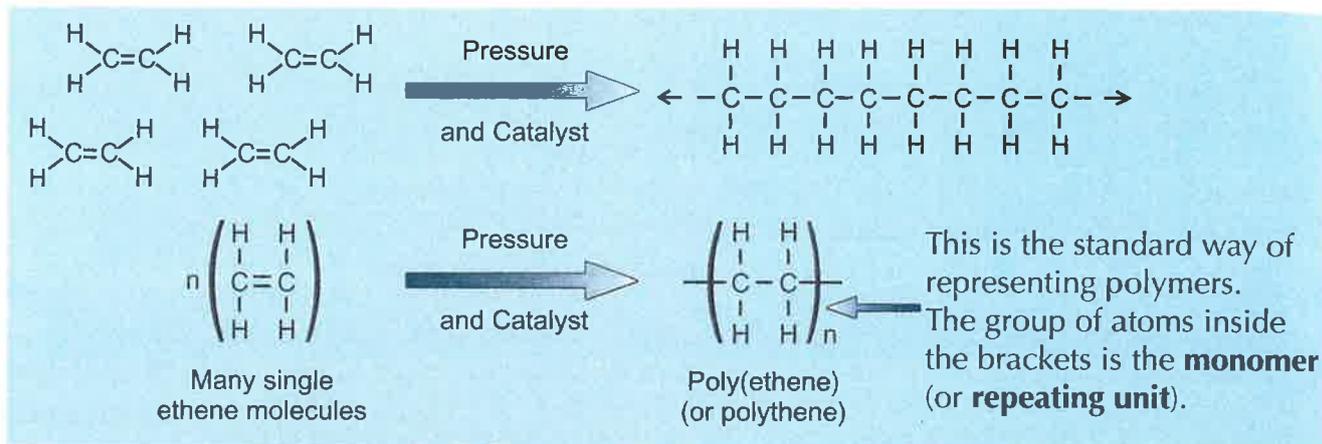
# Polymerisation

## Alkenes Can Form Polymers

The presence of the double bond in alkene molecules means that they are capable of forming **polymers**. A polymer is a long, chain-like molecule built up from lots of **repeating units**. In this case the repeating units, called **monomers**, are alkene molecules.

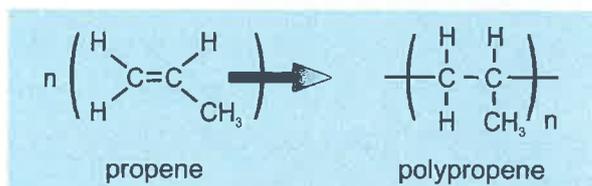
Under the right conditions (these depend on the alkene and the desired properties of the polymer), many small alkenes (like ethene and propene) will **open up** their double bonds and **link together** to form these long chain polymers.

The following example shows the formation of **poly(ethene)** (or **polythene** for short):



## Other Small Alkenes do a Similar Thing

1) **Propene** polymerises to form **polypropene**.

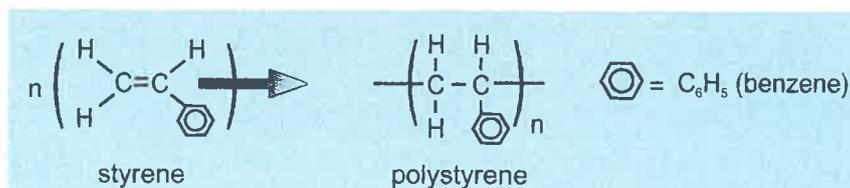


Pretty polymer.



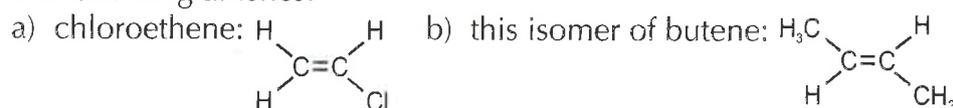
2) **Styrene**, which has a **benzene** ring in it, polymerises to form **polystyrene**.

(Benzene is just a ring of six carbon atoms in which the bonding electrons are shared between all six carbons.)



*I'd go on and on about how great polymers are, but it'd get repetitive...*

- 1) What property of alkenes allows them to form polymers?
- 2) Using the standard way of representing polymers, shown above, draw the polymers formed by the following alkenes:



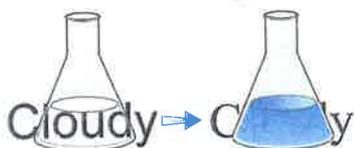
# Reaction Rates

## Measuring the Rate of a Reaction

The **rate** of reaction is just a measure of how **fast** a particular reaction is going. You need to know some of the ways that you can follow the rates of different reactions. They're all about measuring how fast the **reactants** are being **used up**, or measuring how fast the **products** of the reaction are **forming**.

There are lots of ways of measuring the rate of a reaction:

- 1) You can measure the **change in mass** that occurs during a reaction where gas is released as one of the products.



- 2) You can follow the **colour change** of a reaction. This includes precipitation reactions, where the solution turns cloudy as more of the product is made.

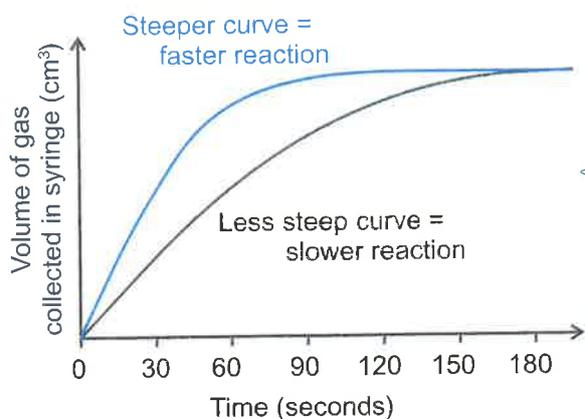
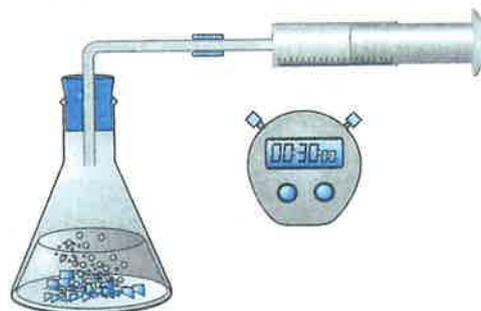
- 3) You can measure changes in **temperature** or **pH** that occur during the reaction.

- 4) You can measure the **volume of gas** produced during a reaction.

**EXAMPLE:** Measuring the rate of reaction between hydrochloric acid and magnesium metal.

magnesium + hydrochloric acid → magnesium chloride + hydrogen

- Use a **gas syringe** to collect the hydrogen gas that is given off during the reaction.
- Use a **stopwatch** to **time** the reaction.
- At **timed intervals**, say every 30 seconds, **record** how much hydrogen gas has been produced.



Plotting graphs lets you compare rates of reactions.

(Another way to measure the rate of this reaction would be to measure the decrease in **mass** as hydrogen gas is lost from the reaction container.)

## My rate of chocolate biscuit consumption is worryingly high...

- 1) Describe how you could measure the rates of the following reactions:
  - a) The endothermic reaction between citric acid and sodium bicarbonate to give carbon dioxide, water and a sodium salt.
  - b) The precipitation reaction between sodium thiosulfate and hydrochloric acid to form a sulfur precipitate, sulfur dioxide gas, sodium chloride and water.
  - c) The reaction between solid calcium carbonate and hydrochloric acid to produce calcium chloride and carbon dioxide gas.

# Collision Theory

## Particles Need to **Collide** in Order to React

Reaction rates are explained by **collision theory**. It's based on the idea that particles in liquids and gases are always **moving around** and **colliding** with each other.

Not every **collision** results in the particles **reacting**. The following **conditions** need to be right:

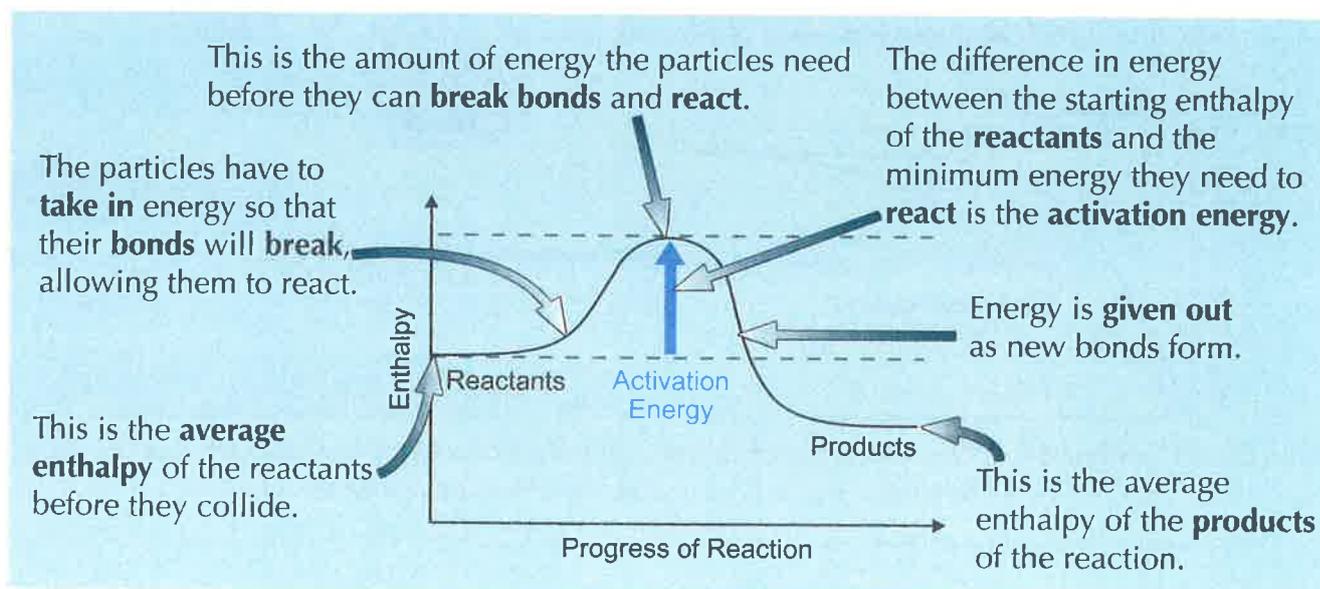
- The particles need to collide in the **right direction**. They need to be **facing** each other the right way.
- The particles need to collide with at least a certain **minimum** amount of **energy**.

Collision theory states that the **more collisions** there are, and the **more energy** these collisions have, the **more likely** particles are to react.

## Particles Need Enough **Energy** to React

- 1) The **minimum amount of kinetic** (movement) **energy** particles need to react is known as the **activation energy**. This energy is used to **break the bonds** to start the reaction.
- 2) Reactions with **low activation energies** often happen **pretty easily**, but reactions with **high activation energies** don't — you have to give the particles **extra energy** (e.g. by **heating** them).

To make this a bit clearer, here's an **enthalpy profile diagram** — enthalpy is just a fancy word for energy. These diagrams show how the **enthalpy** of the reacting particles changes over the course of a reaction (see page 41 for more on enthalpy changes).



## Toast and a large cup of tea — that's my morning activation energy...

- 1) Two particles in a reaction vessel collide but don't react. Give two reasons why the reaction may not have happened.
- 2) What is the activation energy of a reaction?
- 3) Draw an enthalpy profile diagram for a reaction. On your diagram, label the reactants, products and activation energy.

# Reaction Rates and Catalysts

## Changing the Rate of Reaction

The rate of reaction depends on **how often** particles **collide** (see page 32) and how **likely** the collisions are to be **successful**.

More frequent **successful** collisions mean a **faster** rate of reaction.

These factors all **increase** the rate of reaction:

- 1) **Increasing temperature** — the particles tend to have more kinetic energy. This means that they move around faster, and so are more likely to collide with each other **and** have enough energy to react.
- 2) **Increasing concentration (or pressure in gases)** — this means that the particles of reactant will be closer together, so they will be more likely to collide.
- 3) **Increasing the surface area of a solid reactant** — this increases the number of particles of the solid reactant able to come into contact with other reactants.

**EXAMPLE:** Predict whether magnesium dust or magnesium ribbon will react faster with hydrochloric acid.

Magnesium dust has a larger **surface area** than magnesium ribbon. Increasing the surface area of a solid reactant increases the rate of reaction, so **magnesium dust** will react faster than magnesium ribbon.

## Catalysts Speed Up Reactions

You met **activation energy** on the last page — it's just the **minimum** amount of energy needed for a reaction to happen.

A **catalyst increases** the **rate** of a reaction by **lowering** its activation energy.

A **catalyst** is any substance which changes the **rate** of a reaction, without being **changed** or **used up** itself.

Catalysts are also very **specific** — different reactions will only be sped up by **certain catalysts**.

There are loads of advantages to using catalysts:

- 1) Catalysts reduce the need for **high temperatures** and **pressures** in industrial reactions, like hydrocarbon cracking (see page 28) and ethanol production (see page 36). This makes these processes **cheaper** to run.
- 2) Using lower temperatures also means less **energy demand**, and so lower CO<sub>2</sub> emissions.

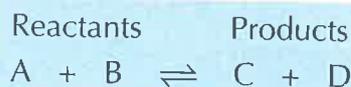
## Tabbys are number one on my cat list...

- 1) Describe two things you could do to increase the rate of a reaction between aqueous species.
- 2) Why does increasing the pressure increase the rate of a reaction between gases?
- 3) What's a catalyst?
- 4) Give two advantages of using a catalyst in industrial reactions.

# Reversible Reactions

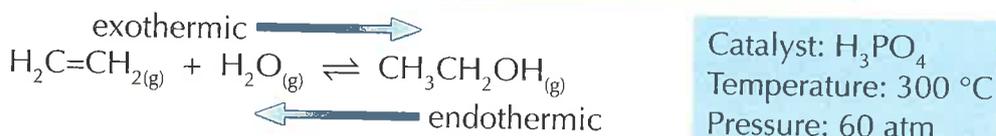
## Reversible Reactions Go Both Ways

In a reversible reaction, the **products** can react with each other and **change back** into the reactants.



So there are actually two reactions happening at once:  $A + B \rightarrow C + D$  and  $C + D \rightarrow A + B$ . This can affect the **yield** of a reaction, as some of the products will be converted **back** into reactants.

**EXAMPLE:** The industrial production of ethanol from ethene.



Because the reaction is reversible you **don't** get a **high yield** — some of the ethanol **converts back** to ethene and water. But you can keep **removing** and **recycling** any ethene that you have left, so you can still end up with a good overall yield.

## Reversible Reactions Reach an Equilibrium

If a reversible reaction is taking place in a **closed system** it will eventually reach a state of **equilibrium**.

$\longleftarrow$  A **closed system** is one where nothing can **get in** or **out**.

- 1) When a reaction **begins** there will be a **high concentration** of **reactants**, and a **low concentration** of **products** in the system. So the **forward** reaction will be **fast**, and the **reverse** reaction quite **slow**.
- 2) The concentration of **reactants** will gradually **decrease**, while the products build up. So the **forward** reaction will start to **slow down** while the **reverse** reaction **speeds up**.
- 3) After a while the forward reaction and the reverse reaction end up going at the **same rate**. From this point on the **concentration** of the **reactants** and **products won't change**.
- 4) This is called **dynamic equilibrium**. The forward and reverse reactions are **both still happening** — some reactant is being made into product, and some product is being made into reactant.
- 5) But since these processes are going at **exactly the same rate**, it seems as if nothing's happening.

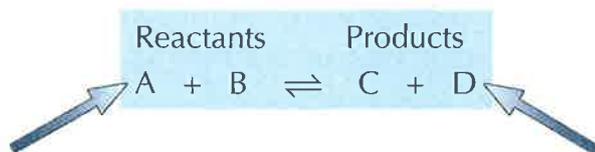
### Dynamic equilibrium — like walking up a down escalator...

- 1) Compare the rates of the forward and backward reactions of a reversible reaction at the following points:
  - a) At the start of the reaction.
  - b) At equilibrium.
- 2) What is dynamic equilibrium?

# Le Chatelier's Principle

## Position of Equilibrium

The **position** of equilibrium tells you the amount of **reactants compared** to the amount of **products** that are present when the reaction reaches an **equilibrium**.



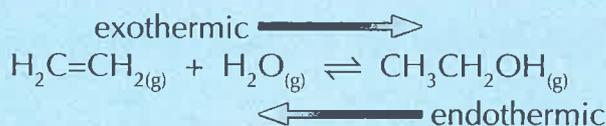
If the position of equilibrium lies on the **left-hand side**, there are **more reactants** than products in the reaction mixture.

If the position of equilibrium lies on the **right-hand side**, there are more **products** than reactants in the reaction mixture.

## Changing Conditions Changes the Equilibrium Position

**Altering the conditions** of a reversible reaction can **move** the position of equilibrium in one direction or the other. Careful control of the conditions can result in a higher yield (more of the products).

Look at the production of ethanol from ethene again as an example:



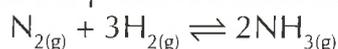
- If you increase the **pressure**, conditions will favour the forward reaction and **more ethanol** ( $\text{CH}_3\text{CH}_2\text{OH}$ ) will be formed. This is because there are **more molecules** of gas on the **left-hand side** than on the right-hand side — two molecules of  $\text{H}_2\text{C}=\text{CH}_2/\text{H}_2\text{O}$  react to form **only one** molecule of  $\text{CH}_3\text{CH}_2\text{OH}$ . This **reduces** the pressure.
- Raising the **temperature** favours the **reverse** reaction. This is because it's **endothermic** (see page 41) and **absorbs** the extra heat energy, **lowering** the temperature.
- Removing ethanol** from the container as it forms will push the equilibrium to the **right** to try and make up for the change in concentration between the reactants and products.

These observations can be summarised by an important rule known as **Le Chatelier's Principle**:

**A reversible reaction will move its equilibrium position to resist any change in the conditions.**

## Equilibrium reactions are so stubborn — always resisting change...

- You are making ethanol from ethene and steam using the reaction shown above. What will happen to the yield of ethanol if you increase the amount of steam in the reaction mixture?
- Ammonia is produced industrially using the following reversible reaction:



The forward reaction is exothermic and the backwards reaction is endothermic.

How will the position of the equilibrium change if you:

- Increase the temperature of the reaction?
- Remove some ammonia from the reaction?

# Equilibrium and Yield

## Deciding on the Best Conditions to Use

Thanks to Le Chatelier's principle (see page 35) you might think it would be **easy** to work out the **optimum conditions** for a reversible reaction. But in real life it's not quite that simple.

For most reversible reactions that are used on an industrial scale there are other factors, such as **cost** and **time**, that need to be taken into account.

Have a look at the conditions used for the production of ethanol from ethene again:



### Temperature:

- 1) **Lowering the temperature** would favour the forward reaction and so it should increase the **yield** of ethanol.
- 2) But lowering the temperature also means that fewer of the particles in the reaction mixture will have **enough energy** to react (see page 32). The particles will also be moving **more slowly**, so there will be **fewer collisions**. So lowering the temperature will **slow down** the **rate** of both the forward and reverse reactions.
- 3) A low temperature would make the forward reaction **too slow** to be useful. So a compromise temperature of **300 °C** is used.

### Pressure:

- 1) **Increasing the pressure** would favour the forward reaction and increase the **rate** of reaction (as the particles will be **closer together** so will collide **more frequently**). This would increase the **yield** of ethanol.
- 2) But producing high pressures uses a lot of **energy** and **costs** a lot of money. You'd need some pretty strong equipment to stand up to the high pressures too — and that would be expensive to buy and maintain.
- 3) To make the reaction economic, a moderately high pressure of **60 atm** is used.

### Concentration:

- 1) Ethanol is **removed** from the reaction vessel as it is produced.
- 2) This reduces the concentration of products so the equilibrium shifts to favour the **forwards reaction**. This **improves** the **yield** of ethanol.

### Catalyst:

- 1) Using a solid **phosphoric acid(V)** catalyst **increases** the rate of **both** the forward and the backward reactions.
- 2) The catalyst has **no effect** on the **position** of the equilibrium — it just means the equilibrium is reached **faster** and the **temperature** and **pressure** at which the reaction can happen, at a reasonable rate, are **reduced**.

*I should put a dodgy pun here, but I won't yield to the pressure...*

- 1) Explain why the reaction above is not run industrially at a temperature of 40 °C.
- 2) Explain why the reaction above is not run industrially at a pressure of 500 atm.

# 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  $6.02 \times 10^{23}$  particles.  
 $6.02 \times 10^{23} \text{ mol}^{-1}$  is known as **Avogadro's constant**.

The particles can be **anything** — e.g. atoms or molecules (or even giraffes).  
 So  $6.02 \times 10^{23}$  atoms of **carbon** is 1 mole of carbon,  
 and  $6.02 \times 10^{23}$  molecules of  $\text{CO}_2$  is 1 mole of  $\text{CO}_2$ .

No, I'm not getting on there.  
 That joke's far too obvious...



## 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  $12 \text{ g mol}^{-1}$ .  
 For  $\text{CO}_2$ ,  $M_r = 44.0$  so 1 mole of  $\text{CO}_2$  weighs **44 g** and the **molar mass** of  $\text{CO}_2$  is  $44 \text{ g mol}^{-1}$ .  
 So, **12.0 g** of **carbon** and **44.0 g** of  $\text{CO}_2$  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:

$$\text{Number of moles} = \frac{\text{Mass of substance (g)}}{\text{Molar mass (g mol}^{-1}\text{)}} \quad \leftarrow \text{g mol}^{-1} \text{ is the same as g/mol.}$$

**EXAMPLE:** How many moles of sodium oxide are present in 24.8 g of  $\text{Na}_2\text{O}$ ?

$$\text{Molar mass of Na}_2\text{O} = (2 \times 23.0) + (1 \times 16.0) = 62.0 \text{ g mol}^{-1}$$

$$\text{Number of moles of Na}_2\text{O} = 24.8 \text{ g} \div 62.0 \text{ g mol}^{-1} = \mathbf{0.400 \text{ moles}}$$

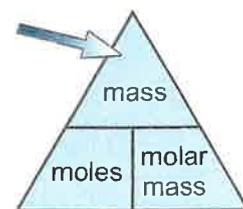
You can **rearrange** the formula above and use it to work out the mass of a substance or its relative formula mass (see page 3). It can help to remember this triangle:

**EXAMPLE:** What is the mass of 1.30 moles of magnesium oxide ( $\text{MgO}$ )?

$$\text{Molar mass of MgO} = (1 \times 24.3) + (1 \times 16.0) = 40.3 \text{ g mol}^{-1}$$

Rearranging the formula,  $\text{mass} = \text{moles} \times \text{molar mass}$

$$\text{So mass of MgO} = 1.30 \times 40.3 = \mathbf{52.4 \text{ g (3 s.f.)}}$$



*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  $\text{NaCl}$ ?
- 3) I have 54.0 g of water ( $\text{H}_2\text{O}$ ) and 84.0 g of iron ( $\text{Fe}$ ). 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 its 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

Number of moles of each compound:

C:  $85.7 \div 12.0 = 7.14$       H:  $14.3 \div 1.0 = 14.3$

Divide by the smallest number (7.14):

C:  $7.14 \div 7.14 = 1$       H:  $14.3 \div 7.14 = 2$

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

The empirical formula is **CH<sub>2</sub>**.

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

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

The empirical formula (CH<sub>2</sub>) has a relative mass of  $12.0 + 1.0 + 1.0 = 14.0$ .

The molecular formula has a relative mass of 42.0.

$42.0 \div 14.0 = 3$

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

C:  $1 \times 3 = 3$       H:  $2 \times 3 = 6$

The molecular formula is **C<sub>3</sub>H<sub>6</sub>**.

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...

- 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)
- 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)
- Find the percentage composition of oxygen in each of the following compounds:
  - Ethanol (C<sub>2</sub>H<sub>5</sub>OH).
  - Nitric acid (HNO<sub>3</sub>).
  - Propanone (C<sub>3</sub>H<sub>6</sub>O).

# Atom Economy

## A Higher Atom Economy Means Less Waste

- 1) Lots of reactions make **more than one product**. Some of them will be **useful**, but others will just be **waste**.
- 2) The **atom economy** of a reaction tells you how much of the **mass** of the reactants is converted into **useful products**, and how much is wasted during a reaction.

$$\text{atom economy} = \frac{\text{total } M_r \text{ of desired products}}{\text{total } M_r \text{ of all products}} \times 100$$

- 3) If a reaction has **100% atom economy** then **all** the atoms in the reactants have been turned into **useful** (desired) **products**. The higher the atom economy, the 'greener' the process.

**EXAMPLE:** Calculate the atom economy of the reaction to make hydrogen gas from methane and steam:  $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$ .

First identify the useful product, which in this reaction is hydrogen gas.

$$\text{atom economy} = \frac{\text{total } M_r \text{ of desired products}}{\text{total } M_r \text{ of all products}} \times 100$$

$$= \frac{M_r \text{ of H}_2}{(M_r \text{ of H}_2) + (M_r \text{ of CO})} \times 100$$

$$= \frac{3 \times (2 \times 1.0)}{3 \times (2 \times 1.0) + 1 \times (12.0 + 16.0)} \times 100 = \frac{6.0}{6.0 + 28.0} \times 100 = 17.6 \%$$

## High Atom Economy is Better in Industry

- 1) **Industrial reactions** are designed to be as **cheap** and **green** as possible. Generally, reactions with high atom economies are the **most efficient** processes as there is **minimal waste**.
- 2) The reactions with the **highest** atom economy are the ones that only have **one product**. These reactions have an atom economy of **100%**.
- 3) Reactions with low atom economies **use up resources** very quickly. They also make lots of **waste** materials that have to be **disposed** of somehow. That tends to make these reactions **unsustainable** — the raw materials run out and the waste has to go somewhere.
- 4) For the same reasons, low atom economy reactions aren't usually **profitable**. Raw materials are **expensive to buy**, and waste products can be expensive to **dispose of**.
- 5) The best way around the problem is to find a **use** for the waste products or to find a reaction with a **better** atom economy to make the same product.

## Atom (Economy) — upgrade to Superior for only £16.99...

- 1) a) Ethanol can be made from bromoethane in the following reaction:  
 $\text{CH}_3\text{CH}_2\text{Br} + \text{NaOH} \rightarrow \text{CH}_3\text{CH}_2\text{OH} + \text{NaBr}$   
 What is the atom economy of this reaction?  
 b) In industry, ethanol is made from ethene and steam using the following reaction:  
 $\text{CH}_2\text{CH}_2 + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{CH}_2\text{OH}$   
 Suggest why this reaction is used, rather than the reaction in part a).

# Endothermic and Exothermic Reactions

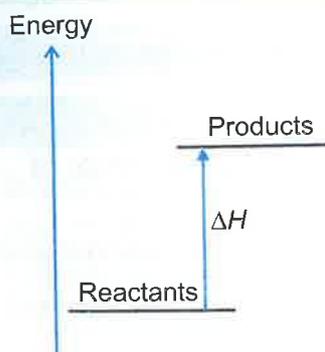
In an **exothermic** reaction, **heat** energy is **given out** (the room temperature rises).

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

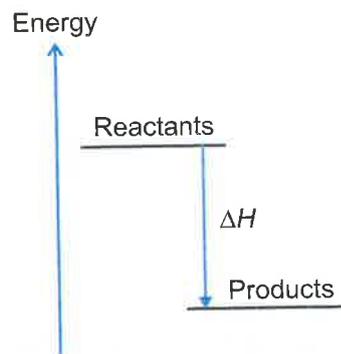
## Making and Breaking Bonds

- 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 ( $\Delta H$ ) 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 ( $\Delta H$ ) 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  $\Delta H$  is **positive**.



In an **exothermic** reaction, the reactants **release** energy to the surroundings. The products therefore have **less energy** than the reactants and  $\Delta H$  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:  

$$\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} \quad \Delta H = -2809 \text{ kJ mol}^{-1}$$
 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	All these values are in $\text{kJ mol}^{-1}$ .
O=O 498	H-H 436	C=O 743	
C-C 348	O-H 463		

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  $\text{CO}_2$  (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:  $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

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:

$$4 \text{ C-H bonds} = 4 \times 413 = 1652 \text{ kJ}$$

$$2 \text{ O=O bonds} = 2 \times 498 = 996 \text{ kJ}$$

$$\text{Total} = 2648 \text{ kJ}$$

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

$$2 \text{ C=O bonds} = 2 \times 743 = 1486 \text{ kJ}$$

$$4 \text{ O-H bonds} = 4 \times 463 = 1852 \text{ kJ}$$

$$\text{Total} = 3338 \text{ kJ}$$

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

$$\text{The overall energy change is: } 2648 - 3338 = \mathbf{-690 \text{ kJ mol}^{-1}}$$

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).



## Planning Experiments

### Make Sure You Plan Your Experiment Carefully

To get accurate and precise results from your experiments, you first need to plan them carefully...



- 1) Work out the **aim** of the experiment.
- 2) Identify the **variables** (see below).
- 3) Decide what **data** to collect.
- 4) Decide the right **equipment** to use.
- 5) Plan how to reduce any **risks** in your experiment.
- 6) Write out a **detailed method**.
- 7) Carry out **tests** to address the aim of your experiment.

### You Need to Control All the Variables

A **variable** is a quantity that might **change** during an experiment, for example temperature. There are two types of variables to know about when carrying out an experiment:

- The **independent variable** is the quantity that you **change**.
- The **dependent variable** is the thing that you measure.

When you plan an experiment you need to work out how you will **control** the variables so that the only one that changes is the one you're investigating — all the others are kept **constant**.

**EXAMPLE:** Measuring the effect of surface area on reaction rate.

In this experiment, the **independent variable** is the **surface area**, and the **dependent variable** is the **rate** of reaction.

Everything else, such as temperature and concentration, has to stay exactly the same between different experiments. Surface area is the only variable that you change.

### Choose the Right Equipment

You need to think carefully about selecting the right **equipment** for your experiment...

- 1) The equipment has to be **appropriate** for the specific experiment — for example, in an experiment where you're collecting a **gas** the equipment you use needs to be properly **sealed** so that the gas can't **escape**.
- 2) The equipment needs to be the right **size**.
- 3) The equipment needs to be the right level of **sensitivity** — for example, if you want to measure out 4.2 g of a compound, you'll need a balance that measures to at least the nearest 0.1 g, not the nearest gram.

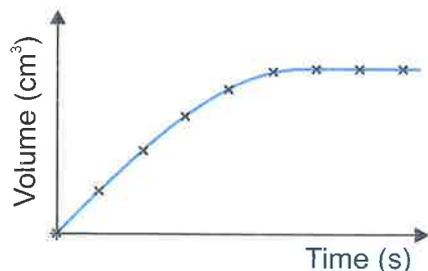
### Reduce Risk — and play poker instead...

- 1) A student is measuring the effect of temperature on the time taken for a lump of magnesium to react completely in a sample of concentrated hydrochloric acid.
  - a) What is the dependent variable in the student's experiment?
  - b) Name two variables that the student should control to make the experiment a fair test.

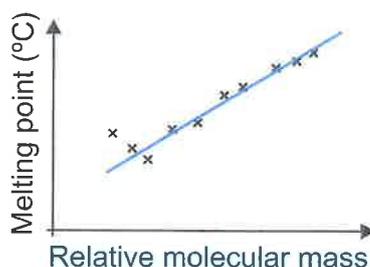
# Presenting and Interpreting Data

## You Can Represent Your Data in a Table or on a Graph

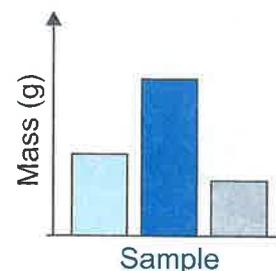
When you do an experiment, it's a good idea to set up a table to **record** your **results** in. Make sure you **include** enough **rows** and **columns** to **record all of the data** you need. Tables are good for **recording** data, but it can be easier to interpret your results if you **plot** them on a **graph**. Depending on the **type** of experiment, the **graph** you plot will vary:



**Line graphs** show how two sets of data are related.



**Scatter plots** show **trends** in data. Don't join all the points — just draw a **line of best fit**.



If one of your sets of data can be split into **groups**, draw a **bar graph**.

## Repeating an Experiment Makes Your Results More Reliable

- If you **repeat** an experiment, your results will usually **differ slightly** each time you do it. You can use the **mean** (or average) of the measurements to represent all these values. The more times you repeat the experiment the **more reliable** the average will be. To find the mean:

**Add together** all the data values then **divide** by the total number of values in the sample.

**EXAMPLE:** Calculate the mean result for the volume of hydrogen gas produced after 30 seconds in the reaction between hydrochloric acid and magnesium.

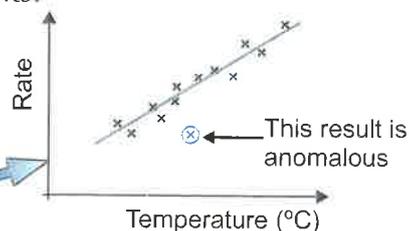
Run 1	Run 2	Run 3
23 cm <sup>3</sup>	22 cm <sup>3</sup>	25 cm <sup>3</sup>

There are **three** values in this sample, so to find the mean result, just add together the results and divide by three:

$$(23 + 22 + 25) \div 3 = \mathbf{23.3 \text{ cm}^3}$$

- Repeating experiments also lets you spot any **weird results** that stick out like a hedgehog in a tea cup. These are called **anomalous** results. For example — if one of the results above was only 5 cm<sup>3</sup>, then something probably went wrong. You should **ignore** the anomalous result when you calculate the mean.

- Anomalous results are really easy to spot on **scatter plots** and **line graphs** as they sit miles away from the line of best fit.



## *I was hoping for a nice result, but it ended up being mean...*

- Kay measured the volume of gas given off in a reaction. Her results were 22.0 cm<sup>3</sup>, 23.0 cm<sup>3</sup>, 22.0 cm<sup>3</sup>, 19.0 cm<sup>3</sup> and 24.0 cm<sup>3</sup>. Identify any anomalous results and calculate the mean.

# Conclusions and Error

## Measurements Always Have Uncertainty in Them

- 1) You saw on the last page that if you **repeat** an experiment then the results normally **vary** slightly, even if you do everything **exactly the same** each time. This is partly to do with the fact that there will always be a degree of **error** in any measurements you make. For example, if you use **measuring scales** that measure to the nearest **gram**, then whenever you weigh something, you might have **up to 0.5 g** more or less than the measured mass.
- 2) The **uncertainty** of a measurement is just the **maximum error** there could be — so in the example above it's **0.5 g**.
- 3) The **percentage error** of a measurement is just another way of showing the uncertainty. You calculate it using the following formula:

$$\text{percentage error} = \frac{\text{uncertainty}}{\text{measurement}} \times 100$$

**EXAMPLE:** Nikhil measures out 6.0 cm<sup>3</sup> of water in a measuring cylinder that has markings every 0.5 cm<sup>3</sup>. Calculate the percentage error of his measurement.

The measuring cylinder can be read to the nearest 0.5 cm<sup>3</sup>, so any measurement could be up to 0.25 cm<sup>3</sup> more than or less than the reading. So the uncertainty is 0.25 cm<sup>3</sup>.

$$\text{Percentage error} = \frac{\text{uncertainty}}{\text{measurement}} \times 100 = \frac{0.25 \text{ cm}^3}{6.0 \text{ cm}^3} \times 100 = 4.2\%$$

## Don't Jump to Conclusions

- 1) Collecting reliable data is important, but if the data doesn't answer the aim, it's not much use.
- 2) The **data** should **support** the conclusion. This may sound obvious, but it's easy to **jump** to conclusions. Conclusions should be **specific** — not make sweeping generalisations.
- 3) Conclusions often try to link changes in one variable with another. For your conclusion to be valid, you have to make sure all the other variables in the experiment were **controlled**.
- 4) And remember — **correlation doesn't** always mean **cause**.

**EXAMPLE:** Investigating whether chlorinated drinking water increases cancer risk.

Some studies claim that drinking chlorinated tap water increases the risk of some cancers. But it's hard to **control** all the **variables** between people who drink tap water and people who don't, so designing a fair test is very tricky.

But, even if some studies show a group of people who drink more chlorinated water are slightly more likely to get certain cancers, it doesn't mean that drinking chlorinated water **causes** cancer. There will be heaps of other differences between the groups of people. It could be due to any of them.

## You've reached the conclusion. Please travel safely and mind the gap...

- 1) Calculate the percentage error in each of the following measurements:
  - a) A mass of 1.4 g on a weighing scale that can measure to the nearest 0.1 g.
  - b) A time of 23 seconds on a clock that times to the nearest second.
  - c) A temperature of 10.6 °C on a thermometer that has markings every 0.2 °C.

# Answers

## Section 1 — The Structure of the Atom

### Page 1 — Atomic Structure

- 1 Protons and neutrons.
- 2 +2
- 3 -2

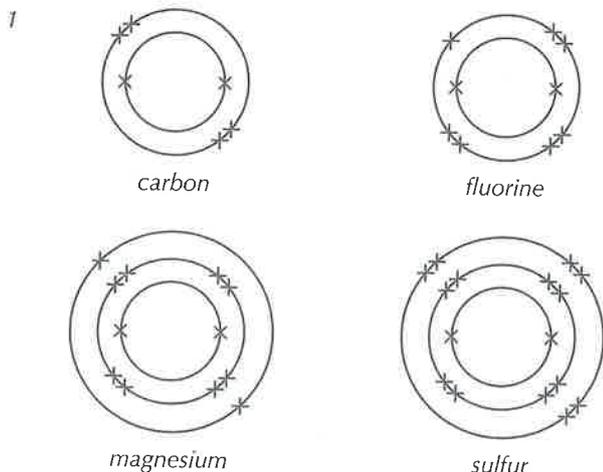
### Page 2 — Atomic Number, Mass Number and Isotopes

- 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



- 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

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  $\text{Na}_{(g)} \rightarrow \text{Na}^+_{(g)} + e^-$
- 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  $\text{SO}_3^{2-}$

### Page 8 — Oxidation Numbers

- 1 The oxidation number tells you how many electrons an atom has donated or accepted.
- 2  $\text{Al}^{3+}$ : +3     $\text{H}^+$ : +1    Ne: 0     $\text{O}^{2-}$ : -2
- 3 0

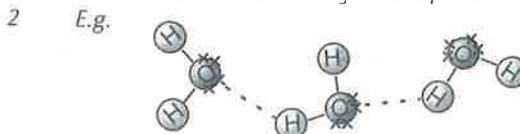
## Section 3 — Intermolecular Bonding

### Page 9 — Intermolecular Bonding

- 1 Weak intermolecular forces
- 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

- a) HF as it is polar, and  $\text{H}_2$  is non-polar.  
b)  $\text{H}_2\text{O}$  as it can form hydrogen bonds, and  $\text{H}_2\text{S}$  can't.  
c)  $\text{CH}_3\text{F}$  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  $\text{CH}_3\text{I}$  is non-polar.



## Section 4 — Bonding and Properties

### Page 11 — Ionic Bonding

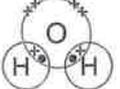
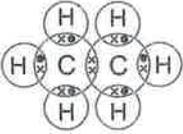
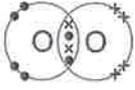
- 1
- 2 You need two  $\text{K}^+$  ions ( $2 \times +1$ ) to balance out each  $\text{O}^{2-}$  ion ( $1 \times -2$ ), so the ratio is 2:1. The ionic formula is  $\text{K}_2\text{O}$ .

# Answers

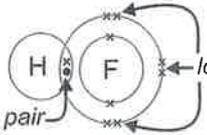
## Page 12 – Ionic Compounds

- BeO, Li<sub>2</sub>O, 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.
- When potassium chloride is solid, the K<sup>+</sup> and Cl<sup>-</sup> 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

- Chlorine: 
  - Water: 
  - Ethane: 
  - Oxygen: 

## Page 14 – Small Covalent Molecules

- 

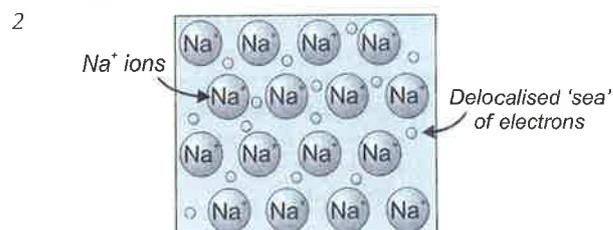
- For a small covalent compound to boil, you have to break the weak intermolecular bonds between the molecules rather than the strong covalent bonds between the atoms in a molecule. These intermolecular bonds don't need much energy to break, so nitrogen has a low boiling point and is a gas at room temperature.

## Page 15 – Giant Covalent Structures

- 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.
- 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

- Calcium is likely to have a higher melting point. This is because calcium is made up of a lattice of Ca<sup>2+</sup> ions, with two delocalised electrons per ion. Potassium however is made up of K<sup>+</sup> 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.



- 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

- 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.
- 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.
- Sodium chloride has ionic bonds.
  - The chloride of phosphorus has covalent bonds and is a small covalent molecules.

## Section 5 – Chemical Equations

### Page 19 – Writing and Balancing Equations

- Step 2: CH<sub>4</sub> + O<sub>2</sub> → CO<sub>2</sub> + H<sub>2</sub>O  
Step 3: CH<sub>4</sub> + 2O<sub>2</sub> → CO<sub>2</sub> + 2H<sub>2</sub>O  
(The Cs already balance, so put a 2 in front of H<sub>2</sub>O to balance the Hs. Now put a 2 in front of O<sub>2</sub> to balance the Os. Check that all still balances.)
- C<sub>2</sub>H<sub>5</sub>OH + 3O<sub>2</sub> → 2CO<sub>2</sub> + 3H<sub>2</sub>O
  - Ca(OH)<sub>2</sub> + 2HCl → CaCl<sub>2</sub> + 2H<sub>2</sub>O
- First balance the atoms:  
Cl<sub>2(g)</sub> + Fe<sup>2+</sup><sub>(aq)</sub> → 2Cl<sup>-</sup><sub>(aq)</sub> + Fe<sup>3+</sup><sub>(aq)</sub>  
Then balance the charges:  
Cl<sub>2(g)</sub> + 2Fe<sup>2+</sup><sub>(aq)</sub> → 2Cl<sup>-</sup><sub>(aq)</sub> + 2Fe<sup>3+</sup><sub>(aq)</sub>

# Answers

## Section 6 – Inorganic Chemistry

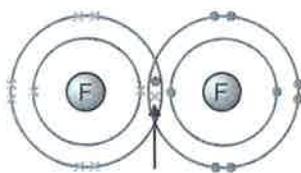
### Page 20 – Group 2

- 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



single covalent bond

### Page 22 – Acids and Bases

- 1  $\text{HNO}_3 + \text{KOH} \rightarrow \text{KNO}_3 + \text{H}_2\text{O}$
- 2 a)  $\text{H}_2\text{SO}_4 \xrightarrow{\text{water}} 2\text{H}^+ + \text{SO}_4^{2-}$   
b)  $\text{KOH} \xrightarrow{\text{water}} \text{K}^+ + \text{OH}^-$   
c)  $\text{HNO}_3 \xrightarrow{\text{water}} \text{H}^+ + \text{NO}_3^-$

## Section 7 – Organic Chemistry

### Page 23 – Organic Molecules

- 1 skeletal: displayed:
- 2  $\text{C}_3\text{H}_6\text{O}_2$

### Page 24 – Alkanes

- 1 pentane:   
hexane:

- 2 a)  $\text{CH}_4, \text{C}_2\text{H}_6, \text{C}_3\text{H}_8, \text{C}_4\text{H}_{10}$   
b) General formula:  $\text{C}_n\text{H}_{2n+2}$
- 3  $\text{CH}_3\text{CH}_2\text{CH}_3(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$

### Page 25 – Alkenes

- 1 e.g.
- 2 Two from:
- 3 General formula:  $\text{C}_n\text{H}_{2n}$

### 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:  $\text{C}_n\text{H}_{(2n+1)}\text{OH}$
- 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

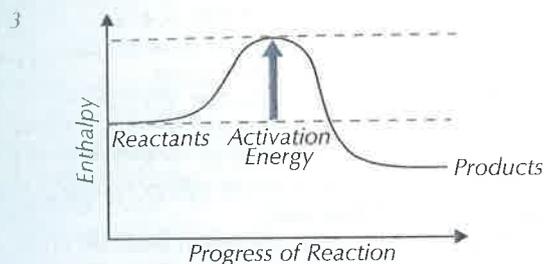
## Section 9 — Rates of Reaction

### Page 31 — Reaction Rates

- a) e.g. Measure the change in temperature / Measure the change in pH / Measure the loss of mass as  $\text{CO}_2$  is evolved / Measure the volume of  $\text{CO}_2$  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  $\text{SO}_2$  is evolved / Measure the volume of  $\text{SO}_2$  produced using a gas syringe.
- c) e.g. Measure the change in pH / Measure the loss of mass as  $\text{CO}_2$  is evolved / Measure the volume of  $\text{CO}_2$  produced using a gas syringe.

### Page 32 — Collision Theory

- The particles may have collided facing in the wrong direction or without enough energy to react.
- The activation energy is the minimum amount of kinetic energy particles need to react.



### Page 33 — Reaction Rates and Catalysts

- e.g. Increase the temperature / Increase the concentration of the reactants / Add a catalyst.
- The particles in the gas will be closer together so they're more likely to collide.
- A catalyst is a substance that speeds up the rate of reaction without being changed or used up itself.
- E.g. Catalysts make reactions cheaper to run and reduce their  $\text{CO}_2$  emissions.

## Section 10 — Equilibria

### Page 34 — Reversible Reactions

- 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.
- 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

- 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.
- 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

- 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.
- 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

- molar mass = mass  $\div$  moles  
 $= 68.6 \div 0.700 = 98.0 \text{ g mol}^{-1}$
- Molar mass  $\text{NaCl} = 23.0 + 35.5 = 58.5 \text{ g mol}^{-1}$   
moles = mass  $\div$  molar mass =  $117 \div 58.5 = 2.00 \text{ moles}$
- Molar mass water =  $16.0 + (2 \times 1.0) = 18.0 \text{ g mol}^{-1}$   
moles of water =  $54.0 \div 18.0 = 3.00 \text{ moles}$ .  
Molar mass of iron =  $55.8 \text{ g mol}^{-1}$   
moles of iron =  $84.0 \div 55.8 = 1.51 \text{ moles}$ .  
There are more moles of water.

### Page 38 — Determination of Formulae from Experiments

- a) Mass of each substance: Pb: 12.9 g    O: 1.00 g  
Number of moles of each substance:  
Pb:  $12.9 \div 207.2 = 0.0623 \text{ moles}$   
O:  $1.00 \div 16.0 = 0.0625 \text{ moles}$   
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:  $2.33 \div 55.8 = 0.0418 \text{ moles}$   
O:  $1.00 \div 16 = 0.0625 \text{ moles}$   
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 **Fe<sub>2</sub>O<sub>3</sub>**.
- Percentage composition of each substance:  
H: 4.30%    C: 26.1%    O: 69.6%  
Number of moles of each substance:  
H:  $4.3 \div 1.0 = 4.30 \text{ moles}$   
C:  $26.1 \div 12.0 = 2.18 \text{ moles}$   
O:  $69.6 \div 16.0 = 4.35 \text{ moles}$   
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 **CH<sub>2</sub>O<sub>2</sub>**.

# 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:  $52.2 \div 12.0 = 4.35$  moles  
 H:  $13.0 \div 1.0 = 13.0$  moles  
 O:  $34.8 \div 16.0 = 2.18$  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  
 $(2 \times 12.0) + (6 \times 1.0) + (1 \times 16.0) = 46.0$ . 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:  $92.3 \div 12.0 = 7.69$  moles  
 H:  $7.70 \div 1.0 = 7.70$  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  
 $12.0 + 1.0 = 13.0$ .  
 $78.0 \div 13.0 = 6.00$ , so there are six lots of the  
 empirical formula in the compound.  
 The molecular formula is  $C_6H_6$ .
- 3 a)  $\frac{1 \times 16.0}{(2 \times 12.0) + (6 \times 1.0) + (1 \times 16.0)} \times 100 = 34.8\%$   
 b)  $\frac{3 \times 16.0}{(1 \times 1.0) + (1 \times 14.0) + (3 \times 16.0)} \times 100 = 76.2\%$   
 c)  $\frac{1 \times 16.0}{(3 \times 12.0) + (6 \times 1.0) + (1 \times 16.0)} \times 100 = 27.5\%$

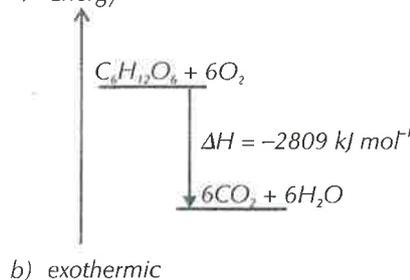
## Page 40 — Atom Economy

- 1 a)  $M_r C_2H_5OH = (2 \times 12.0) + (6 \times 1.0) + 16.0 = 46.0$   
 $M_r NaBr = 23.0 + 79.9 = 102.9$   
 $\frac{46.0}{46.0 + 102.9} \times 100 = 30.9\%$   
 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) Energy



## 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 =  $8 \times 413 = 3304$   
 2 C — C bonds =  $2 \times 348 = 696$   
 5 O = O bonds =  $5 \times 498 = 2490$   
 TOTAL = 6490  
 Step 2: Calculate the energy released by all the new  
 bonds formed between the product atoms:  
 6 C = O bonds =  $6 \times 743 = 4458$   
 8 O — H bonds =  $8 \times 463 = 3704$   
 TOTAL = 8162  
 Step 3: Find the overall value for the energy change:  
 $+6490 - 8162 = -1672 \text{ kJ mol}^{-1}$
- b) Step 1: Calculate the energy required to break all of the  
 bonds between the reactant atoms:  
 1 C — C bond =  $1 \times 348 = 348$   
 1 C — O bond =  $1 \times 360 = 360$   
 5 C — H bonds =  $5 \times 413 = 2065$   
 1 O — H bonds =  $1 \times 463 = 463$   
 3 O = O bond =  $3 \times 498 = 1494$   
 TOTAL = 4730  
 Step 2: Calculate the energy released by all the new  
 bonds formed between the product atoms  
 4 C = O bonds =  $4 \times 743 = 2972$   
 6 O — H bonds =  $6 \times 463 = 2778$   
 TOTAL = 5750  
 Step 3: Find the overall value for the energy change:  
 $+4730 - 5750 = -1020 \text{ kJ mol}^{-1}$
- c) Step 1: Calculate the energy required to break the H—H  
 and C=C bonds:  
 1 H — H bond =  $1 \times 436 = 436$   
 1 C = C bond =  $1 \times 612 = 612$   
 TOTAL = 1048  
 Step 2: Calculate the energy released by all the new  
 bonds formed between product atoms  
 2 C — H bonds =  $2 \times 413 = 826$   
 1 C — C bond =  $1 \times 348 = 348$   
 TOTAL = 1174  
 Step 3: Find the overall value for the energy change:  
 $+1048 - 1174 = -126 \text{ kJ mol}^{-1}$

## 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  $19.0 \text{ cm}^3$ .  
 The mean is:  $(22.0 + 23.0 + 22.0 + 24.0) \div 4 = 22.8 \text{ cm}^3$

### Page 45 — Conclusions and Error

- 1 a) The uncertainty of the weighing scales is 0.05 g, so the  
 percentage error =  $\frac{0.05}{1.4} \times 100 = 3.6\%$   
 b) The uncertainty of the clock is 0.5 s, so the  
 percentage error =  $\frac{0.5}{23} \times 100 = 2.2\%$   
 c) The uncertainty of the thermometer is  $0.1 \text{ }^\circ\text{C}$ , so the  
 percentage error =  $\frac{0.1}{10.6} \times 100 = 0.9\%$

