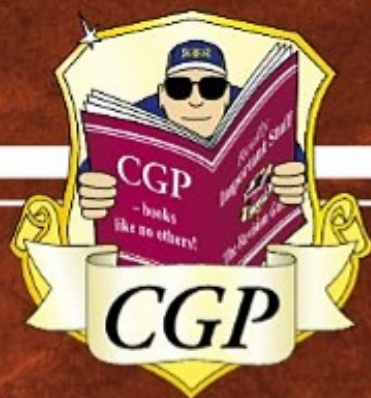


*CGP*



# **GCSE Chemistry**

For AQA (Grade 9-1)

## **The Revision Guide**

# What to Expect in the Exams

Before you get cracking with your revision, here's a handy guide to what you'll have to face in the exams — and the special features of this book that we've included especially to help you. You're welcome.

## 1 Topics are Covered in Different Papers

For GCSE Chemistry, you'll sit two exam papers at the end of your course.

You're expected to know the basic concepts of chemistry in both papers.

Paper	Time	No. of marks	Topics Assessed
1	1 hr 45 mins	100	1, 2, 3, 4 and 5
2	1 hr 45 mins	100	6, 7, 8, 9 and 10

## 2 There are Different Question Types

In each exam, you'll be expected to answer a mixture of multiple choice questions, structured questions, questions that have short, closed answers as well as open response questions.

For some open response questions, you'll be marked on the overall quality of your answer, not just its scientific content. So...

Always make sure:

- You answer the question fully.
- You include detailed, relevant information.
- Your answer is clear and has a logical structure.

## 3 You'll be Tested on your Maths...

At least 20% of the total marks for GCSE Chemistry will come from questions that test your maths skills.

For these questions, always remember to:

### EXAMPLE:

Look out for these worked examples in this book — they show you maths skills you'll need in the exam.

- Show your working — you could get marks for this, even if your final answer's wrong.
- Check that the units of your answer are the same as the ones they asked for in the question.
- Make sure your answer is given to an appropriate number of significant figures.

## 4 ...and on your Practical Skills

Whenever one of the required practical activities crops up in this book, it's marked up like this...

**PRACTICAL**

...and there's a whole section on Practical Skills on pages 107-110.

- GCSE Chemistry contains 8 required practical activities that you'll do during the course. You can be asked about these, and the practical skills involved in them, in the exams.
- At least 15% of the total marks will be for questions that test your understanding of the practical activities and practical skills.
- For example, you might be asked to comment on the design of an experiment (the apparatus and method), make predictions, analyse or interpret results... Pretty much anything to do with planning and carrying out the investigations.

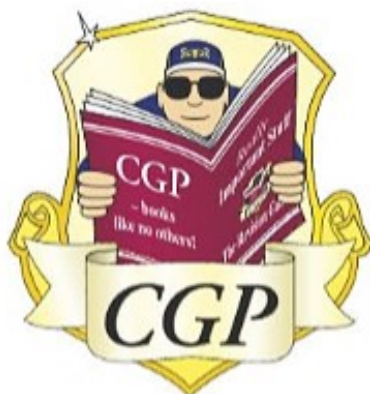
## 5 You'll need to know about Working Scientifically

Working Scientifically is all about how science is applied in the outside world by real scientists.

For example, you might be asked about ways that scientists communicate an idea to get their point across without being biased, or about the limitations of a scientific theory.

Working Scientifically is covered on pages 2-11.

You need to think about the situation that you've been given and use all your scientific savvy to answer the question. Always read the question and any data you've been given really carefully before you start writing your answer.



## **Chemistry revision? CGP has the solution...**

No doubt about it, GCSE Chemistry is a tough subject. Luckily, this CGP book is crammed full of info, from facts and theory to practical skills — and what's more, there are practice questions on each page to help you sharpen up your exam skills.

### **CGP — still the best! 😊**

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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# The Scientific Method

This section isn't about how to 'do' science — but it does show you the way most scientists work.

## Scientists Come Up With **Hypotheses** — Then **Test** Them

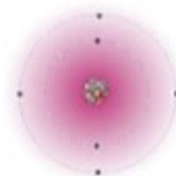
- 1) Scientists try to explain things. They start by observing something they don't understand.
- 2) They then come up with a hypothesis — a possible explanation for what they've observed.
- 3) The next step is to test whether the hypothesis might be right or not. This involves making a prediction based on the hypothesis and testing it by gathering evidence (i.e. data) from investigations. If evidence from experiments backs up a prediction, you're a step closer to figuring out if the hypothesis is true.



About 100 years ago, scientists hypothesised that atoms looked like this.

## Several Scientists Will **Test** a Hypothesis

- 1) Normally, scientists share their findings in peer-reviewed journals, or at conferences.
- 2) Peer-review is where other scientists check results and scientific explanations to make sure they're 'scientific' (e.g. that experiments have been done in a sensible way) before they're published. It helps to detect false claims, but it doesn't mean that findings are correct — just that they're not wrong in any obvious way.
- 3) Once other scientists have found out about a hypothesis, they'll start basing their own predictions on it and carry out their own experiments. They'll also try to reproduce the original experiments to check the results — and if all the experiments in the world back up the hypothesis, then scientists start to think the hypothesis is true.
- 4) However, if a scientist does an experiment that doesn't fit with the hypothesis (and other scientists can reproduce the results) then the hypothesis may need to be modified or scrapped altogether.



After more evidence was gathered, scientists changed their hypothesis to this.

## If **All** the **Evidence** Supports a Hypothesis, It's **Accepted** — For Now

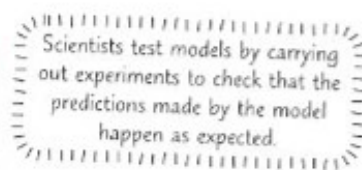
- 1) Accepted hypotheses are often referred to as theories. Our currently accepted theories are the ones that have survived this 'trial by evidence' — they've been tested many times over the years and survived.
- 2) However, theories never become totally indisputable fact. If new evidence comes along that can't be explained using the existing theory, then the hypothesising and testing is likely to start all over again.



Now we think it's more like this.

## Theories Can Involve **Different Types** of Models

- 1) A representational model is a simplified description or picture of what's going on in real life. Like all models, it can be used to explain observations and make predictions. E.g. the Bohr model of an atom is a simplified way of showing the arrangement of electrons in an atom (see p.19). It can be used to explain trends down groups in the periodic table.
- 2) Computational models use computers to make simulations of complex real-life processes, such as climate change. They're used when there are a lot of different variables (factors that change) to consider, and because you can easily change their design to take into account new data.
- 3) All models have limitations on what they can explain or predict. E.g. ball and stick models (a type of spatial model) can be used to show how ions are arranged in an ionic compound. One of their limitations is that they don't show the relative sizes of the ions (see p.30).



Scientists test models by carrying out experiments to check that the predictions made by the model happen as expected.

## I'm off to the zoo to test my hippo-thesis...

The scientific method has developed over time, and many people have helped to develop it. From Aristotle to modern day scientists, lots of people have contributed. And many more are likely to contribute in the future.

# Communication & Issues Created by Science

Scientific developments can be great, but they can sometimes raise more questions than they answer...

## It's Important to **Communicate Scientific Discoveries to the General Public**

Some scientific discoveries show that people should change their habits, or they might provide ideas that could be developed into new technology. So scientists need to tell the world about their discoveries.

Technologies are being developed that make use of fullerenes (see p.34). These include drug delivery systems for use in medicine. Information about these systems needs to be communicated to doctors so they can use them, and to patients, so they can make informed decisions about their treatment.

## Scientific Evidence can be Presented in a Biased Way

- 1) Reports about scientific discoveries in the media (e.g. newspapers or television) aren't peer-reviewed.
- 2) This means that, even though news stories are often based on data that has been peer-reviewed, the data might be presented in a way that is over-simplified or inaccurate, making it open to misinterpretation.
- 3) People who want to make a point can sometimes present data in a biased way. (Sometimes without knowing they're doing it.) For example, a scientist might overemphasise a relationship in the data, or a newspaper article might describe details of data supporting an idea without giving any evidence against it.

## Scientific Developments are Great, but they can Raise Issues

Scientific knowledge is increased by doing experiments. And this knowledge leads to scientific developments, e.g. new technologies or new advice. These developments can create issues though. For example:

**Economic issues:** Society can't always afford to do things scientists recommend (e.g. investing in alternative energy sources) without cutting back elsewhere.

**Personal issues:** Some decisions will affect individuals. For example, someone might support alternative energy, but object if a wind farm is built next to their house.

**Social issues:** Decisions based on scientific evidence affect people — e.g. should fossil fuels be taxed more highly? Would the effect on people's lifestyles be acceptable...

**Environmental issues:** Human activity often affects the natural environment. For example, building a dam to produce electricity will change the local habitat so some species might be displaced. But it will also reduce our need for fossil fuels, so will help to reduce climate change.

## Science Can't Answer Every Question — Especially Ethical Ones

- 1) We don't understand everything. We're always finding out more, but we'll never know all the answers.
- 2) In order to answer scientific questions, scientists need data to provide evidence for their hypotheses.
- 3) Some questions can't be answered yet because the data can't currently be collected, or because there's not enough data to support a theory.
- 4) Eventually, as we get more evidence, we'll answer some of the questions that currently can't be answered, e.g. what the impact of global warming on sea levels will be. But there will always be the "Should we be doing this at all?"-type questions that experiments can't help us to answer...

Think about new drugs which can be taken to boost your 'brain power'.

- Some people think they're good as they could improve concentration or memory. New drugs could let people think in ways beyond the powers of normal brains.
- Other people say they're bad — they could give you an unfair advantage in exams. And people might be pressured into taking them so that they could work more effectively, and for longer hours.



## Tea to milk or milk to tea? — Totally unanswerable by science...

Science can't tell you whether or not you should do something. That's for you and society to decide. But there are tons of questions science might be able to answer, like where life came from and where my superhero socks are.

# Risk

By reading this page you are agreeing to the **risk** of a paper cut or severe drowsiness...

## Nothing is Completely Risk-Free

- 1) A **hazard** is something that could **potentially cause harm**.
- 2) All hazards have a **risk** attached to them — this is the **chance** that the hazard will cause harm.
- 3) The risks of some things seem pretty **obvious**, or we've known about them for a while, like the risk of causing **acid rain** by polluting the atmosphere, or of having a **car accident** when you're travelling in a car.
- 4) **New technology** arising from **scientific advances** can bring **new risks**, e.g. scientists are unsure whether **nanoparticles** that are being used in cosmetics and sunscreen might be harming the cells in our bodies. These risks need to be considered **alongside** the **benefits** of the technology, e.g. improved sun protection.
- 5) You can estimate the **size** of a risk based on **how many times** something happens in a big sample (e.g. 100 000 people) over a given **period** (e.g. a year). For example, you could assess the risk of a driver crashing by recording how many people in a group of 100 000 drivers crashed their cars over a year.
- 6) To make **decisions** about activities that involve **hazards**, we need to take into account the **chance** of the hazard causing harm, and how **serious** the **consequences** would be if it did. If an activity involves a hazard that's **very likely** to cause harm, with **serious consequences** if it does, it's considered **high risk**.

## People Make Their Own Decisions About Risk

- 1) Not all risks have the same **consequences**, e.g. if you chop veg with a sharp knife you risk cutting your finger, but if you go scuba-diving you risk death. You're much **more likely** to cut your finger during half an hour of **chopping** than to die during half an hour of **scuba-diving**. But most people are happier to accept a higher **probability** of an accident if the **consequences** are **short-lived** and fairly **minor**.
- 2) People tend to be more willing to accept a risk if they **choose** to do something (e.g. go scuba diving), compared to having the risk **imposed** on them (e.g. having a nuclear power station built next door).
- 3) People's **perception** of risk (how risky they **think** something is) isn't always **accurate**. They tend to view **familiar** activities as **low-risk** and **unfamiliar** activities as **high-risk** — even if that's not the case. For example, cycling on roads is often **high-risk**, but many people are happy to do it because it's a **familiar** activity. Air travel is actually pretty **safe**, but a lot of people perceive it as **high-risk**.
- 4) People may **over-estimate** the risk of things with **long-term** or **invisible** effects, e.g. ionising radiation.

## Investigations Can be Hazardous

- 1) Hazards from science experiments might include:

- **Microorganisms**, e.g. some bacteria can make you ill.
- **Chemicals**, e.g. sulfuric acid can burn your skin and alcohols catch fire easily.
- **Fire**, e.g. an unattended Bunsen burner is a fire hazard.
- **Electricity**, e.g. faulty electrical equipment could give you a shock.



- 2) Part of planning an investigation is making sure that it's **safe**.
- 3) You should always make sure that you **identify** all the hazards that you might encounter. Then you should think of ways of **reducing the risks** from the hazards you've identified. For example:

- If you're working with **sulfuric acid**, always wear gloves and safety goggles. This will reduce the risk of the acid coming into contact with your skin and eyes.
- If you're using a **Bunsen burner**, stand it on a heat proof mat. This will reduce the risk of starting a fire.

You can find out about potential hazards by looking in textbooks, doing some Internet research, or asking your teacher.

## Not revising — an unacceptable exam hazard...

The world's a dangerous place, but if you can recognise hazards, decide how to reduce their risks, and be happy to accept some risks, you can still have fun. Just maybe don't go skydiving with a great white shark on Friday 13th.

# Designing Investigations

Dig out your lab coat and dust down your badly-scratched safety goggles... it's **investigation time**.

## Investigations Produce Evidence to Support or Disprove a Hypothesis

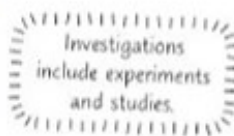
- 1) Scientists **observe** things and come up with **hypotheses** to explain them (see p.2). You need to be able to do the same. For example:

**Observation:** People have big feet and spots. **Hypothesis:** Having big feet causes spots.

- 2) To **determine** whether or not a hypothesis is **right**, you need to do an **investigation** to gather evidence. To do this, you need to use your hypothesis to make a **prediction** — something you think **will happen** that you can test. E.g. people who have bigger feet will have more spots.
- 3) Investigations are used to see if there are **patterns** or **relationships** between **two variables**, e.g. to see if there's a pattern or relationship between the variables 'number of spots' and 'size of feet'.

## Evidence Needs to be Repeatable, Reproducible and Valid

- 1) **Repeatable** means that if the **same person** does an experiment again using the **same methods** and equipment, they'll get **similar results**.
- 2) **Reproducible** means that if **someone else** does the experiment, or a **different** method or piece of equipment is used, the results will still be **similar**.
- 3) If data is **repeatable** and **reproducible**, it's **reliable** and scientists are more likely to **have confidence** in it.
- 4) **Valid results** are both repeatable and reproducible AND they **answer the original question**. They come from experiments that were designed to be a **FAIR TEST**...



## To Make an Investigation a Fair Test You Have to Control the Variables

- 1) In a lab experiment you usually **change one variable** and **measure** how it affects **another variable**.
- 2) To make it a fair test, **everything else** that could affect the results should **stay the same** — otherwise you can't tell if the thing you're changing is causing the results or not.
- 3) The variable you **CHANGE** is called the **INDEPENDENT** variable.
- 4) The variable you **MEASURE** when you change the independent variable is the **DEPENDENT** variable.
- 5) The variables that you **KEEP THE SAME** are called **CONTROL** variables.

You could find how **temperature** affects **reaction rate** by measuring the **volume of gas** formed over time. The **independent variable** is the **temperature**. The **dependent variable** is the **volume of gas** produced. **Control variables** include the **concentration** and **amounts** of reactants, the **time period** you measure, etc.

- 6) Because you can't always control all the variables, you often need to use a **control experiment**. This is an experiment that's kept under the **same conditions** as the rest of the investigation, but **doesn't** have anything **done** to it. This is so that you can see what happens when you don't change anything at all.

## The Bigger the Sample Size the Better

- 1) Data based on **small samples** isn't as good as data based on large samples. A sample should **represent** the **whole population** (i.e. it should share as many of the characteristics in the population as possible) — a small sample can't do that as well. It's also harder to spot **anomalies** if your sample size is too small.
- 2) The **bigger** the sample size the **better**, but scientists have to be **realistic** when choosing how big. For example, if you were studying the effects of a chemical used to sterilise water on the people drinking it, it'd be great to study **everyone** who was drinking the water (a huge sample), but it'd take ages and cost a bomb. It's more realistic to study a thousand people, with a mixture of ages, gender, and race.

## This is no high street survey — it's a designer investigation...

Not only do you need to be able to plan your own investigations, you should also be able to look at someone else's plan and decide whether or not it needs improving. Those examiners aren't half demanding.

# Collecting Data

You've designed the perfect investigation — now it's time to get your hands mucky and collect some data.

## Your Data Should be **Repeatable, Reproducible, Accurate and Precise**

- 1) To check repeatability you need to repeat the readings and check that the results are similar. You need to repeat each reading at least three times.
- 2) To make sure your results are reproducible you can cross check them by taking a second set of readings with another instrument (or a different observer).
- 3) Your data also needs to be **ACCURATE**. Really accurate results are those that are really close to the true answer. The accuracy of your results usually depends on your method — you need to make sure you're measuring the right thing and that you don't miss anything that should be included in the measurements. E.g. estimating the amount of gas released from a reaction by counting the bubbles isn't very accurate because you might miss some of the bubbles and they might have different volumes. It's more accurate to measure the volume of gas released using a gas syringe (see p.107).
- 4) Your data also needs to be **PRECISE**. Precise results are ones where the data is all really close to the mean (average) of your repeated results (i.e. not spread out).



Brian's result was a curate.

Repeat	Data set 1	Data set 2
1	12	11
2	14	17
3	13	14
Mean	13	14

Data set 1 is more precise than data set 2.

## Your **Equipment** has to be **Right for the Job**

- 1) The measuring equipment you use has to be sensitive enough to measure the changes you're looking for. For example, if you need to measure changes of 1 cm<sup>3</sup> you need to use a measuring cylinder or burette that can measure in 1 cm<sup>3</sup> steps — it'd be no good trying with one that only measures 10 cm<sup>3</sup> steps.
- 2) The smallest change a measuring instrument can detect is called its **RESOLUTION**. E.g. some mass balances have a resolution of 1 g, some have a resolution of 0.1 g, and some are even more sensitive.
- 3) Also, equipment needs to be calibrated by measuring a known value. If there's a difference between the measured and known value, you can use this to correct the inaccuracy of the equipment.

## You Need to Look out for **Errors and Anomalous Results**

- 1) The results of your experiment will always vary a bit because of **RANDOM ERRORS** — unpredictable differences caused by things like human errors in measuring. The errors when you make a reading from a burette are random. You have to estimate or round the level when it's between two marks — so sometimes your figure will be a bit above the real one, and sometimes it will be a bit below.
- 2) You can reduce the effect of random errors by taking repeat readings and finding the mean. This will make your results more precise.
- 3) If a measurement is wrong by the same amount every time, it's called a **SYSTEMATIC ERROR**. For example, if you measured from the very end of your ruler instead of from the 0 cm mark every time, all your measurements would be a bit small. Repeating the experiment in the exact same way and calculating a mean won't correct a systematic error.
- 4) Just to make things more complicated, if a systematic error is caused by using equipment that isn't zeroed properly, it's called a **ZERO ERROR**. For example, if a mass balance always reads 1 gram before you put anything on it, all your measurements will be 1 gram too heavy.
- 5) You can compensate for some systematic errors if you know about them though, e.g. if your mass balance always reads 1 gram before you put anything on it you can subtract 1 gram from all your results.
- 6) Sometimes you get a result that doesn't fit in with the rest at all. This is called an **ANOMALOUS RESULT**. You should investigate it and try to work out what happened. If you can work out what happened (e.g. you measured something totally wrong) you can ignore it when processing your results.

If there's no systematic error, then doing repeats and calculating a mean can make your results more accurate.

## Watch what you say to that mass balance — it's very sensitive...

Weirdly, data can be really precise but not very accurate. For example, a fancy piece of lab equipment might give results that are really precise, but if it's not been calibrated properly those results won't be accurate.

# Processing and Presenting Data

Processing your data means doing some **calculations** with it to make it **more useful**. Once you've done that, you can present your results in a nice **chart** or **graph** to help you **spot any patterns** in your data.

## Data Needs to be Organised

- 1) Tables are dead useful for **organising data**.
- 2) When you draw a table **use a ruler** and make sure **each column** has a **heading** (including the **units**).

## You Might Have to Process Your Data

- 1) When you've done repeats of an experiment you should always calculate the **mean** (average). To do this **add together** all the data values and **divide** by the total number of values in the sample.
- 2) You might also need to calculate the **range** (how spread out the data is). To do this find the **largest** number and **subtract** the **smallest** number from it.

Ignore anomalous results when calculating these.

### EXAMPLE:

The results of an experiment to find the mass of gas lost from two reactions are shown below. Calculate the mean and the range for the mass of gas lost in each reaction.

Test tube	Repeat 1 (g)	Repeat 2 (g)	Repeat 3 (g)	Mean (g)	Range (g)
A	28	37	32	$(28 + 37 + 32) \div 3 = 32$	$37 - 28 = 9$
B	47	51	60	$(47 + 51 + 60) \div 3 = 53$	$60 - 47 = 13$

## Round to the Lowest Number of Significant Figures

The **first significant figure** of a number is the first digit that's **not zero**. The second and third significant figures come **straight after** (even if they're zeros). You should be aware of significant figures in calculations.

- 1) In **any** calculation, you should round the answer to the **lowest number of significant figures** (s.f.) given.
- 2) Remember to write down **how many** significant figures you've rounded to after your answer.
- 3) If your calculation has multiple steps, **only** round the **final** answer, or it won't be as accurate.

### EXAMPLE:

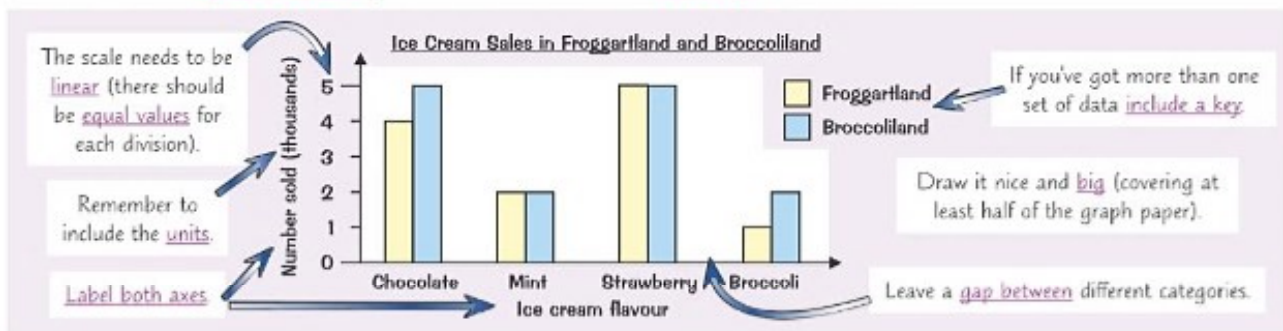
The volume of one mole of gas is  $24.0 \text{ dm}^3$  at room temperature and pressure. How many moles are there in  $4.6 \text{ dm}^3$  of gas under the same conditions?

$$\text{No. of moles of gas} = \underset{2 \text{ s.f.}}{4.6 \text{ dm}^3} \div \underset{3 \text{ s.f.}}{24.0 \text{ dm}^3} = 0.19166... = 0.19 \text{ mol (2 s.f.)}$$

Final answer should be rounded to 2 s.f.

## If Your Data Comes in Categories, Present It in a Bar Chart

- 1) If the independent variable is **categoric** (comes in distinct categories, e.g. alkane chain length, metals) you should use a **bar chart** to display the data.
- 2) You also use them if the independent variable is **discrete** (the data can be counted in chunks, where there's no in-between value, e.g. number of protons is discrete because you can't have half a proton).
- 3) There are some **golden rules** you need to follow for **drawing** bar charts:



## If Your Data is Continuous, Plot a Graph

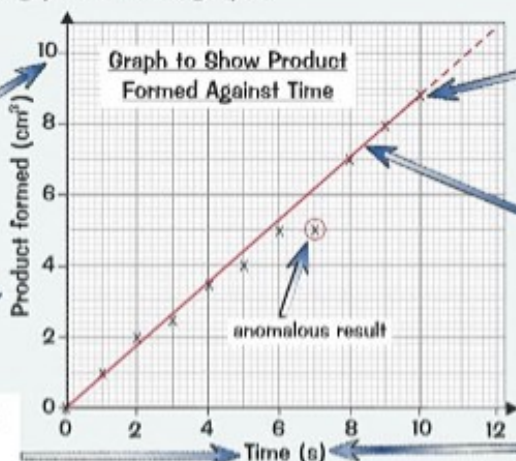
If both variables are **continuous** (numerical data that can have any value within a range, e.g. length, volume, temperature) you should use a **graph** to display the data.

Here are the rules for plotting points on a graph:

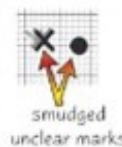
Use the biggest data values you've got to draw a **sensible scale** on your axes. Here, the highest amount of product formed is **8.8 cm<sup>3</sup>**, so it makes sense to label the y-axis up to **10 cm<sup>3</sup>**.

The **dependent** variable goes on the **y-axis** (the **vertical** one).

The **independent** variable goes on the **x-axis** (the **horizontal** one).



To plot points, use a sharp pencil and make **neat little crosses** (don't do blobs).



If you're asked to draw a **line** (or **curve**) of **best fit**, draw a line **through** or as **near** to as **many points as possible**, ignoring any **anomalous results**. **Don't** join the crosses up.

Draw it nice and **big** (covering at least half of the graph paper).

Remember to include the **units**.

## Graphs Can Give You a Lot of Information About Your Data

- The **gradient** (slope) of a graph tells you how quickly the **dependent variable** changes if you change the **independent variable**.

$$\text{gradient} = \frac{\text{change in } y}{\text{change in } x}$$

This **graph** shows the **volume of gas** produced in a reaction against **time**. The graph is **linear** (it's a straight line graph), so you can simply calculate the **gradient** of the line to find out the **rate of reaction**.

- To calculate the gradient, pick **two points** on the line that are easy to read and a **good distance** apart.
- Draw a line down** from one of the points and a **line across** from the other to make a **triangle**. The line drawn down the side of the triangle is the **change in y** and the line across the bottom is the **change in x**.

Change in  $y = 6.8 - 2.0 = 4.8 \text{ cm}^3$     Change in  $x = 5.2 - 1.6 = 3.6 \text{ s}$

$$\text{Rate} = \text{gradient} = \frac{\text{change in } y}{\text{change in } x} = \frac{4.8 \text{ cm}^3}{3.6 \text{ s}} = 1.3 \text{ cm}^3/\text{s}$$

The units of the gradient are (units of  $y$ )/(units of  $x$ ).  $\text{cm}^3/\text{s}$  can also be written as  $\text{cm}^3\text{s}^{-1}$ .

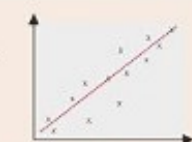


You can use this method to calculate other rates from a graph not just the rate of a reaction. Just remember that a rate is how much something changes over time, so  $x$  needs to be the time.

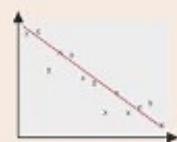
- To find the **gradient of a curve** at a **certain point**, draw a **tangent** to the curve at that point and then find the **gradient of the tangent**. See page 71 for details on how to do this.
- The **intercept** of a graph is where the line of best fit crosses one of the **axes**. The **x-intercept** is where the line of best fit crosses the **x-axis** and the **y-intercept** is where it crosses the **y-axis**.

## Graphs Show the Relationship Between Two Variables

- You can get **three** types of **correlation** (relationship) between variables:
- Just because there's correlation, it doesn't mean the change in one variable is **causing** the change in the other — there might be **other factors** involved (see page 10).



**POSITIVE correlation:**  
as one variable **increases**  
the other **increases**.



**INVERSE (negative) correlation:**  
as one variable **increases**  
the other **decreases**.



**NO correlation:**  
**no relationship** between  
the two variables.

## I love eating apples — I call it core elation...

Science is all about finding relationships between things. And I don't mean that chemists gather together in corners to discuss whether or not Devini and Sebastian might be a couple... though they probably do that too.

# Units and Equations

Graphs and maths skills are all very well, but the numbers don't mean much if you can't get the **units** right.

## S.I. Units Are Used All Round the World

- 1) It wouldn't be all that useful if I defined volume in terms of **bath tubs**, you defined it in terms of **egg-cups** and my pal Sarwat defined it in terms of **balloons** — we'd never be able to compare our data.
- 2) To stop this happening, scientists have come up with a set of **standard units**, called S.I. units, that all scientists use to measure their data. Here are some S.I. units you'll see in chemistry:

Quantity	S.I. Base Unit
mass	kilogram, kg
length	metre, m
time	second, s
amount of a substance	mole, mol

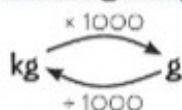
## Scaling Prefixes Can Be Used for Large and Small Quantities

- 1) Quantities come in a huge **range** of sizes. For example, the volume of a swimming pool might be around 2 000 000 000 cm<sup>3</sup>, while the volume of a cup is around 250 cm<sup>3</sup>.
- 2) To make the size of numbers more **manageable**, larger or smaller units are used. These are the **S.I. base unit** (e.g. metres) with a **prefix** in front:

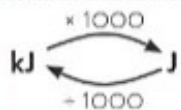
prefix	tera (T)	giga (G)	mega (M)	kilo (k)	deci (d)	centi (c)	milli (m)	micro (μ)	nano (n)
multiple of unit	10 <sup>12</sup>	10 <sup>9</sup>	1 000 000 (10 <sup>6</sup> )	1000	0.1	0.01	0.001	0.000001 (10 <sup>-6</sup> )	10 <sup>-9</sup>

- 3) These **prefixes** tell you **how much bigger** or **smaller** a unit is than the base unit. So one **kilometre** is **one thousand** metres.
- 4) To **swap** from one unit to another, all you need to know is what number you have to **divide** or **multiply** by to get from the original unit to the new unit — this is called the **conversion factor**.
  - To go from a **bigger unit** (like m) to a **smaller unit** (like cm), you **multiply** by the conversion factor.
  - To go from a **smaller unit** (like g) to a **bigger unit** (like kg), you **divide** by the conversion factor.
- 5) Here are some conversions that'll be useful for GCSE chemistry:

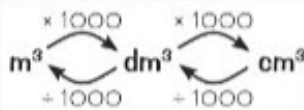
Mass can have units of kg and g.



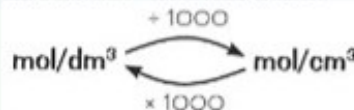
Energy can have units of J and kJ.



Volume can have units of m<sup>3</sup>, dm<sup>3</sup> and cm<sup>3</sup>.



Concentration can have units of mol/dm<sup>3</sup> and mol/cm<sup>3</sup>.



## Always Check The Values Used in Equations Have the Right Units

- 1) Formulas and equations show **relationships** between **variables**.
- 2) To **rearrange** an equation, make sure that whatever you do to **one side** of the equation you also do to the **other side**.
 

You can find the **number of moles** of something using the equation:  $\text{moles} = \text{mass} \div \text{molar mass}$ .

You can **rearrange** this equation to find the **mass** by **multiplying each side** by molar mass to give:  $\text{mass} = \text{moles} \times \text{molar mass}$ .
- 3) To use a formula, you need to know the values of **all but one** of the variables. **Substitute** the values you do know into the formula, and do the calculation to work out the final variable.
- 4) Always make sure the values you put into an equation or formula have the **right units**. For example, you might have done a titration experiment to work out the concentration of a solution. The volume of the solution will probably have been measured in cm<sup>3</sup>, but the equation to find concentration uses volume in dm<sup>3</sup>. So you'll have to **convert** your volume from cm<sup>3</sup> to dm<sup>3</sup> before you put it into the equation.
- 5) To make sure your units are **correct**, it can help to write down the **units** on each line of your **calculation**.

## I wasn't sure I liked units, but now I'm converted...

It's easy to get in a muddle when converting between units, but there's a handy way to check you've done it right. If you're moving from a smaller unit to a larger unit (e.g. g to kg) the number should get smaller, and vice versa.

# Drawing Conclusions

Congratulations — you're nearly at the end of a gruelling investigation, time to draw conclusions.

## You Can Only Conclude What the Data Shows and **NO MORE**

- 1) Drawing conclusions might seem pretty straightforward — you just look at your data and say what pattern or relationship you see between the dependent and independent variables.

The table on the right shows the rate of a reaction in the presence of two different catalysts:

Catalyst	Rate of reaction ( $\text{cm}^3/\text{s}$ )
A	13.5
B	19.5
No catalyst	5.5

**CONCLUSION:**  
Catalyst **B** makes this reaction go faster than catalyst A.

- 2) But you've got to be really careful that your conclusion matches the data you've got and doesn't go any further.  
 3) You also need to be able to use your results to justify your conclusion (i.e. back up your conclusion with some specific data).  
 4) When writing a conclusion you need to refer back to the original hypothesis and say whether the data supports it or not:

You can't conclude that catalyst B increases the rate of any other reaction more than catalyst A — the results might be completely different.

The rate of this reaction was 6  $\text{cm}^3/\text{s}$  faster using catalyst B compared with catalyst A.

The hypothesis for this experiment might have been that catalyst B would make the reaction go quicker than catalyst A. If so, the data supports the hypothesis.

## Correlation **DOES NOT** Mean Cause

If two things are correlated (i.e. there's a relationship between them) it doesn't necessarily mean a change in one variable is causing the change in the other — this is **REALLY IMPORTANT** — **DON'T FORGET IT**. There are three possible reasons for a correlation:

- 1) **CHANCE**: It might seem strange, but two things can show a correlation purely due to chance.

For example, one study might find a correlation between people's hair colour and how good they are at frisbee. But other scientists don't get a correlation when they investigate it — the results of the first study are just a fluke.

- 2) **LINKED BY A 3RD VARIABLE**: A lot of the time it may look as if a change in one variable is causing a change in the other, but it isn't — a third variable links the two things.

For example, there's a correlation between water temperature and shark attacks. This isn't because warmer water makes sharks crazy. Instead, they're linked by a third variable — the number of people swimming (more people swim when the water's hotter, and with more people in the water you get more shark attacks).

- 3) **CAUSE**: Sometimes a change in one variable does cause a change in the other. You can only conclude that a correlation is due to cause when you've controlled all the variables that could, just could, be affecting the result.

For example, there's a correlation between smoking and lung cancer. This is because chemicals in tobacco smoke cause lung cancer. This conclusion was only made once other variables (such as age and exposure to other things that cause cancer) had been controlled and shown not to affect people's risk of getting lung cancer.



## I conclude that this page is a bit dull...

...although, just because I find it dull doesn't mean that I can conclude it's dull (you might think it's the most interesting thing since that kid got his head stuck in the railings near school). In the exams you could be given a conclusion and asked whether some data supports it — so make sure you understand how far conclusions can go.

# Uncertainties and Evaluations

Hurrah! The end of another investigation. Well, now you have to work out all the things you did **wrong**.

## Uncertainty is the Amount of Error Your Measurements Might Have

- 1) When you **repeat** a measurement, you often get a **slightly different** figure each time you do it due to **random error**. This means that **each result** has some **uncertainty** to it.
- 2) The measurements you make will also have some uncertainty in them due to **limits** in the **resolution** of the equipment you use (see page 6).
- 3) This all means that the **mean** of a set of results will also have some uncertainty to it. You can calculate the uncertainty of a **mean result** using the equation:
- 4) The **larger** the range, the **less precise** your results are and the **more uncertainty** there will be in your results. Uncertainties are shown using the ' $\pm$ ' symbol.

The range is the largest value minus the smallest value (p.7).

$$\text{uncertainty} = \frac{\text{range}}{2}$$

### EXAMPLE:

The table below shows the results of a titration experiment to determine the volume of  $0.5 \text{ mol/dm}^3$  sodium hydroxide solution needed to neutralise  $25 \text{ cm}^3$  of a solution of hydrochloric acid with unknown concentration. Calculate the uncertainty of the mean.

Repeat	1	2	3	mean
Volume of sodium hydroxide ( $\text{cm}^3$ )	20.10	19.80	20.00	19.97

- 1) First work out the range:

$$\text{Range} = 20.10 - 19.80 = 0.300 \text{ cm}^3$$

- 2) Use the range to find the uncertainty:

$$\text{Uncertainty} = \text{range} \div 2 = 0.300 \div 2 = 0.150 \text{ cm}^3 \quad \text{So the uncertainty of the mean} = 19.97 \pm 0.15 \text{ cm}^3$$

- 5) Measuring a **greater amount** of something helps to **reduce uncertainty**. For example, in a rate of reaction experiment, measuring the amount of product formed over a **longer period** compared to a shorter period will **reduce** the **percentage uncertainty** in your results.

## Evaluations — Describe How it Could be Improved

An evaluation is a **critical analysis** of the whole investigation.

- 1) You should comment on the **method** — was it **valid**? Did you control all the other variables to make it a **fair test**?
- 2) Comment on the **quality** of the **results** — was there **enough evidence** to reach a valid **conclusion**? Were the results **repeatable**, **reproducible**, **accurate** and **precise**?
- 3) Were there any **anomalous** results? If there were **none** then **say so**. If there were any, try to **explain** them — were they caused by **errors** in measurement? Were there any other **variables** that could have **affected** the results? You should comment on the level of **uncertainty** in your results too.
- 4) All this analysis will allow you to say how **confident** you are that your conclusion is **right**.
- 5) Then you can suggest any **changes** to the **method** that would **improve** the quality of the results, so that you could have **more confidence** in your conclusion. For example, you might suggest **changing** the way you controlled a variable, or **increasing** the number of **measurements** you took. Taking more measurements at **narrower intervals** could give you a **more accurate result**. For example:

**Enzymes** have an **optimum temperature** (a temperature at which they **work best**). Say you do an experiment to find an enzyme's optimum temperature and take measurements at  $10^\circ\text{C}$ ,  $20^\circ\text{C}$ ,  $30^\circ\text{C}$ ,  $40^\circ\text{C}$  and  $50^\circ\text{C}$ . The results of this experiment tell you the optimum is  $40^\circ\text{C}$ . You could then **repeat** the experiment, taking **more measurements around  $40^\circ\text{C}$**  to get a **more accurate** value for the optimum.

- 6) You could also make more **predictions** based on your conclusion, then **further experiments** could be carried out to test them.

When suggesting improvements to the investigation, always make sure that you say why you think this would make the results better.



## Evaluation — next time, I'll make sure I don't burn the lab down...

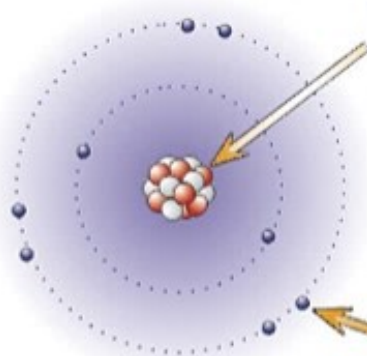
So there you have it — Working Scientifically. Make sure you know this stuff like the back of your hand. It's not just in the lab that you'll need to know how to work scientifically. You can be asked about it in the exams as well.

# Atoms

All substances are made of **atoms**. They're really **tiny** — too small to see, even with your microscope. Atoms are so tiny that a **50p piece** contains about 77 400 000 000 000 000 000 000 of them. Quite a lot then...

## Atoms Contain **Protons**, **Neutrons** and **Electrons**

Atoms have a radius of about **0.1 nanometers** (that's  $1 \times 10^{-10}$  m). There are a few different (and equally useful) modern models of the atom — but chemists tend to like the model below best.



### The Nucleus

- 1) It's in the **middle** of the atom.
- 2) It contains **protons** and **neutrons**.
- 3) The nucleus has a **radius** of around  $1 \times 10^{-14}$  m (that's around 1/10 000 of the radius of an atom)
- 4) It has a **positive charge** because of the protons.
- 5) Almost the **whole** mass of the atom is **concentrated** in the nucleus.

A nanometer (nm) is one billionth of a meter. Shown in standard form, that's  $1 \times 10^{-9}$  m. Standard form is used for showing really large or really small numbers.

Protons are heavy and positively charged.  
Neutrons are heavy and neutral.  
Electrons are tiny and negatively charged.

Particle	Relative Mass	Charge
Proton	1	+1
Neutron	1	0
Electron	Very small	-1

(Electron mass is often taken as zero.)

### The Electrons

- 1) Move **around** the nucleus in electron **shells**.
- 2) They're **negatively charged** and **tiny**, but they cover **a lot of space**.
- 3) The **volume** of their orbits determines the size of the atom.
- 4) Electrons have virtually **no** mass.

## Number of Protons **Equals** Number of Electrons

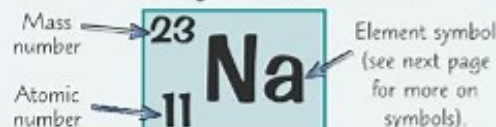
- 1) Atoms are **neutral** — they have **no charge** overall (unlike ions).
- 2) This is because they have the **same number** of **protons** as **electrons**.
- 3) The **charge** on the electrons is the **same** size as the charge on the **protons**, but **opposite** — so the charges **cancel out**.
- 4) In an ion, the number of protons **doesn't equal** the number of **electrons**. This means it has an **overall charge**. For example, an ion with a **2- charge**, has **two more** electrons than protons.

An ion is an atom or group of atoms that has lost or gained electrons.

## Atomic Number and Mass Number Describe an Atom

- 1) The **nuclear symbol** of an atom tells you its **atomic (proton) number** and **mass number**.
- 2) The **atomic number** tells you how many **protons** there are.
- 3) The **mass number** tells you the **total number** of **protons and neutrons** in the atom.
- 4) To get the number of **neutrons**, just subtract the **atomic number** from the **mass number**.

### Nuclear symbol for sodium.



## Let's be positive about this — unless you're an electron of course...

So here we are — the very beginning of GCSE Chemistry. This stuff is super important — if you get to grips with the basic facts then you'll have a better chance understanding the rest of chemistry. Crack on.

Q1 An atom of nitrogen has an atomic number of 7 and a mass number of 14.

Give the number of electrons, protons and neutrons in the atom.

[3 marks]

# Elements

An **element** is a substance made up of atoms that all have the **same** number of **protons** in their nucleus.

## Elements Consist of Atoms With the Same Atomic Number

- 1) Atoms can have different numbers of protons, neutrons and electrons.  
It's the number of **protons** in the nucleus that decides what **type** of atom it is.
- 2) For example, an atom with **one proton** in its nucleus is **hydrogen** and an atom with **two protons** is **helium**.
- 3) If a substance only contains atoms with the **same number** of **protons** it's called an **element**.  
There are about **100 different elements**.
- 4) So **all the atoms** of a particular **element** (e.g. nitrogen) have the **same number** of protons and **different elements** have atoms with **different numbers** of protons.

## Atoms Can be Represented by Symbols

Atoms of each element can be represented by a **one or two letter symbol** — it's a type of **shorthand** that saves you the bother of having to write the full name of the element.

Some make **perfect sense**, e.g.

C = carbon O = oxygen Mg = magnesium

Others less so, e.g.

Na = sodium Fe = iron Pb = lead

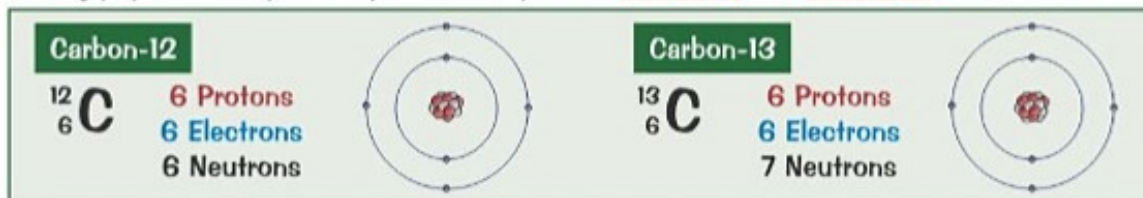
You'll see these symbols on the periodic table (see page 22).

Most of these odd symbols actually come from the Latin names of the elements.

## Isotopes are the Same Except for Extra Neutrons

- 1) **Isotopes** are different forms of the same element, which have the **same number** of **protons** but a **different number** of **neutrons**.
- 2) So isotopes have the **same atomic number** but **different mass numbers**.
- 3) A very popular example of a pair of isotopes are **carbon-12** and **carbon-13**.

The number of neutrons is just the mass number minus the atomic number.



- 4) Because many **elements** can exist as a number of different isotopes, **relative atomic mass** ( $A_r$ ) is used instead of mass number when referring to the element as a whole. This is an **average** mass taking into account the **different masses** and **abundances** (amounts) of all the isotopes that make up the element.
- 5) You can use this **formula** to work out the **relative atomic mass** of an element:
 
$$\text{relative atomic mass } (A_r) = \frac{\text{sum of (isotope abundance} \times \text{isotope mass number)}}{\text{sum of abundances of all the isotopes}}$$

### EXAMPLE:

Copper has two stable isotopes. Cu-63 has an abundance of 69.2% and Cu-65 has an abundance of 30.8%. Calculate the relative atomic mass of copper to 1 decimal place.

$$\text{Relative atomic mass} = \frac{(69.2 \times 63) + (30.8 \times 65)}{69.2 + 30.8} = \frac{4359.6 + 2002}{100} = \frac{6361.6}{100} = 63.616 = 63.6$$

## It's elemental my dear Watson...

Atoms, elements and isotopes — make sure you know what they are and the differences between them.

- Q1 A substance consists of atoms which all have the same number of protons and electrons but different numbers of neutrons. Explain why this substance is an element. [1 mark]
- Q2 An isotope of iron, Fe, has a mass number of 56 and atomic number 26. Give the number of protons and neutrons in an atom of this isotope. [2 marks]

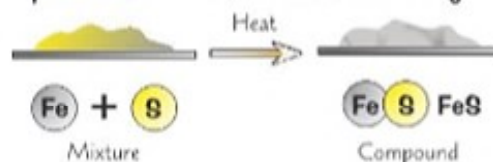
# Compounds

It would be great if we only had to deal with elements. But unluckily for you, elements can mix and match to make lots of new substances called **compounds**. And this makes things a little bit more complicated...

## Atoms Join Together to Make Compounds

- 1) When **elements react**, atoms **combine** with other atoms to form **compounds**.
- 2) Compounds are substances formed from **two or more** elements, the atoms of each are in **fixed proportions** throughout the compound and they're held together by **chemical bonds**.
- 3) **Making bonds** involves atoms giving away, taking or sharing **electrons**. Only the **electrons** are involved — the nuclei of the atoms aren't affected at all when a bond is made.
- 4) It's **usually difficult** to **separate** the original elements of a compound out again — a chemical reaction is needed to do this.
- 5) A compound which is formed from a **metal** and a **non-metal** consists of **ions**. The **metal** atoms **lose** electrons to form **positive ions** and the non-metal atoms **gain** electrons to form **negative ions**. The **opposite charges** (positive and negative) of the ions mean that they're strongly **attracted** to each other. This is called **ionic bonding**. Examples of compounds which are bonded ionically include sodium chloride, magnesium oxide and calcium oxide.
- 6) A compound formed from **non-metals** consists of **molecules**. Each atom **shares** an **electron** with another atom — this is called **covalent bonding**. Examples of compounds that are bonded covalently include hydrogen chloride gas, carbon monoxide, and water.
- 7) The **properties** of a compound are usually **totally different** from the properties of the **original elements**. For example, if iron (a lustrous magnetic metal) and sulfur (a nice yellow powder) react, the compound formed (**iron sulfide**) is a **dull grey solid lump**, and doesn't behave **anything like** either iron or sulfur.

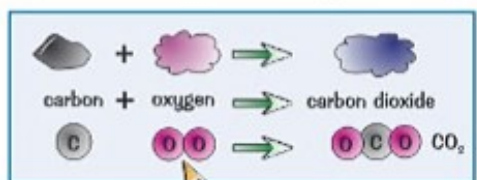
During a chemical reaction, at least one new substance is made. You can usually measure a change in energy, such as a temperature change, as well.



## A Formula Shows What Atoms are in a Compound

Just as elements can be represented by **symbols**, compounds can be represented by **formulas**. The formulas are made up of elemental symbols in the **same proportions** that the elements can be found in the compound.

- 1) For example, carbon dioxide,  $\text{CO}_2$ , is a **compound** formed from a **chemical reaction** between carbon and oxygen. It contains **1 carbon atom** and **2 oxygen atoms**.



Elemental oxygen goes around in pairs of atoms (so it's  $\text{O}_2$ ).

- 2) Here's another example: the formula of **sulfuric acid** is  $\text{H}_2\text{SO}_4$ . So, each molecule contains **2 hydrogen atoms**, **1 sulfur atom** and **4 oxygen atoms**.
- 3) There might be **brackets** in a formula, e.g. calcium hydroxide is  $\text{Ca}(\text{OH})_2$ . The little number outside the bracket applies to **everything** inside the brackets. So in  $\text{Ca}(\text{OH})_2$  there's **1 calcium atom**, **2 oxygen atoms** and **2 hydrogen atoms**.

Here are some examples of formulas which might come in handy:

- |                                   |                                     |  |
|-----------------------------------|-------------------------------------|--|
| 1) Carbon dioxide — $\text{CO}_2$ | 4) Sodium chloride — $\text{NaCl}$  | 7) Calcium chloride — $\text{CaCl}_2$          |
| 2) Ammonia — $\text{NH}_3$        | 5) Carbon monoxide — $\text{CO}$    | 8) Sodium carbonate — $\text{Na}_2\text{CO}_3$ |
| 3) Water — $\text{H}_2\text{O}$   | 6) Hydrochloric acid — $\text{HCl}$ | 9) Sulfuric acid — $\text{H}_2\text{SO}_4$     |

## If you don't learn this stuff it will only compound your problems...

You know when you were little and taught to share things? Turns out atoms have been doing this since the start of the universe. Maybe we could all learn a thing or two from those little guys.

- Q1 How many atoms are in one particle of  $\text{Na}_2\text{CO}_3$ ? [1 mark]
- Q2 A compound has the formula  $\text{CaCl}_2$ . Name the compound and the elements it contains. [2 marks]

# Chemical Equations

**Chemical equations** are fundamental to chemistry. Pretty much like tomato ketchup is to a bacon butty. Mmm... bacon butties... Sorry, I got distracted. Let's do this.

## Chemical Changes are Shown Using Chemical Equations

One way to show a chemical reaction is to write a **word equation**. It's not as **quick** as using chemical symbols and you can't tell straight away **what's happened** to each of the **atoms**, but it's **dead easy**.

Here's an example — you're told that **methane** burns in **oxygen** giving **carbon dioxide** and water:

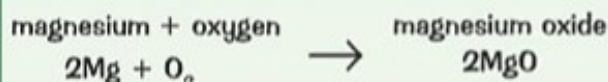
The molecules on the **left-hand side** of the equation are called the **reactants** (because they react with each other).



The molecules on the **right-hand side** are called the **products** (because they've been produced from the reactants).

## Symbol Equations Show the Atoms on Both Sides

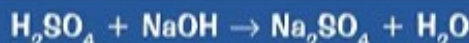
Chemical **changes** can be shown in a kind of **shorthand** using symbol equations. Symbol equations just show the **symbols or formulas** of the **reactants** and **products**...



You'll have spotted that there's a '2' in front of the Mg and the MgO. The reason for this is explained below.

## Symbol Equations Need to be Balanced

- 1) There must always be the **same** number of atoms on **both sides** — they can't just **disappear**.
- 2) You **balance** the equation by putting numbers **in front** of the formulas where needed.  
Take this equation for reacting sulfuric acid with sodium hydroxide:



- 3) The **formulas** are all correct but the numbers of some atoms **don't match up** on both sides.
- 4) You **can't change formulas** like  $\text{H}_2\text{SO}_4$  to  $\text{H}_2\text{SO}_5$ . You can only put numbers **in front of them**.

The more you **practise**, the **quicker** you get, but all you do is this:

- 1) Find an element that **doesn't balance** and **pencil in a number** to try and sort it out.
- 2) **See where it gets you**. It may create **another imbalance**, but if so, pencil in **another number** and see where that gets you.
- 3) Carry on chasing **unbalanced** elements and it'll **sort itself out** pretty quickly.

### EXAMPLE:

In the equation above you'll notice we're short of **H atoms** on the RHS (Right-Hand Side).

- 1) The only thing you can do about that is make it **2H<sub>2</sub>O** instead of just H<sub>2</sub>O:



- 2) But that now gives **too many** H atoms and O atoms on the RHS, so to balance that up you could try putting **2NaOH** on the LHS (Left-Hand Side):



- 3) And suddenly there it is! **Everything balances**. And you'll notice the Na just sorted itself out.



E=mc<sup>2</sup>

## Revision is all about getting the balance right...

Balancing equations is all about practice. Once you have a few goes you'll see it's much less scary than it seemed before you took on, challenged and defeated this page. Go grab some chemistry glory.

Q1 Balance the equation:  $\text{Fe} + \text{Cl}_2 \rightarrow \text{FeCl}_3$  [1 mark]

Q2 Hydrogen and oxygen molecules are formed in a reaction where water splits apart.  
For this reaction: a) State the word equation. b) Give a balanced symbol equation. [3 marks]

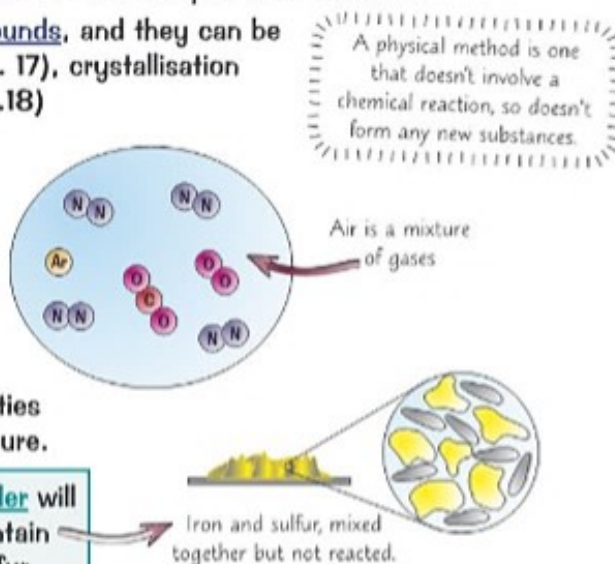
# Mixtures and Chromatography

Mixtures in **chemistry** are just like mixtures in baking, lots of **separate** things all mixed together. But most of the time they're considerably less delicious. And you probably shouldn't eat them. Or put them in an oven.

## Mixtures are Easily Separated — Not Like Compounds

- 1) Unlike in a compound, there's **no chemical bond** between the different parts of a mixture.
- 2) The parts of a mixture can be either **elements** or **compounds**, and they can be separated out by **physical methods** such as filtration (p. 17), crystallisation (p.17), simple distillation (p.18), fractional distillation (p.18) and chromatography (see below).
- 3) **Air** is a **mixture** of gases, mainly nitrogen, oxygen, carbon dioxide and argon. The gases can all be **separated out** fairly easily.
- 4) **Crude oil** is a **mixture** of different length hydrocarbon molecules.
- 5) The **properties** of a mixture are just a **mixture** of the properties of the **separate parts** — the chemical properties of a substance **aren't** affected by it being part of a mixture.

For example, a **mixture** of **iron powder** and **sulfur powder** will show the properties of **both iron and sulfur**. It will contain grey magnetic bits of iron and bright yellow bits of sulfur.



## You Need to Know How to Do Paper Chromatography

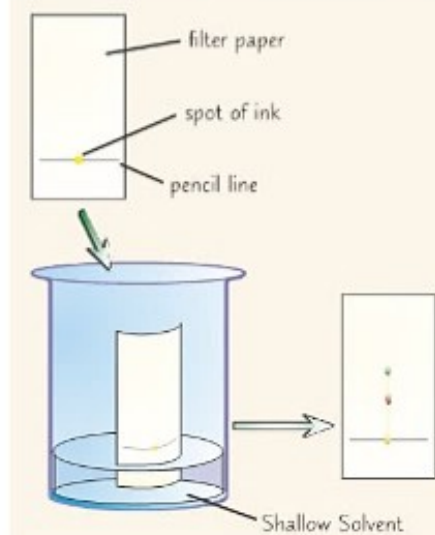
**PRACTICAL**

One method of separating substances in a mixture is through **chromatography**.

This technique can be used to separate different **dyes** in an **ink**. Here's how you can do it:

- 1) Draw a **line** near the bottom of a sheet of **filter paper**.  
(Use a **pencil** to do this — pencil marks are **insoluble** and won't dissolve in the solvent.)
- 2) Add a **spot** of the ink to the line and place the sheet in a beaker of **solvent**, e.g. water.
- 3) The **solvent** used depends on what's being tested. Some compounds **dissolve** well in **water**, but sometimes other solvents, like ethanol, are needed.
- 4) Make sure the ink isn't touching the solvent — you don't want it to **dissolve** into it.
- 5) Place a **lid** on top of the container to stop the solvent **evaporating**.
- 6) The solvent **seeps** up the paper, carrying the ink with it.
- 7) Each different **dye** in the ink will move up the paper at a **different rate** so the dyes will **separate out**. Each dye will form a **spot** in a different place — 1 spot per dye in the ink.
- 8) If any of the dyes in the ink are **insoluble** (won't dissolve) in the solvent you've used, they'll stay on the **baseline**.
- 9) When the **solvent** has nearly reached the **top** of the paper, take the paper out of the beaker and leave it to **dry**.
- 10) The end result is a pattern of spots called a **chromatogram**.

The point the solvent has reached as it moves up the paper is the solvent front.



## Chemistry and fun are a mixture — easily separated...

Chromatography is actually mighty useful in real life. It's used to test athletes' urine samples for performance enhancing drugs, and also to test unknown substances at crime scenes. Eeek...

Q1 Explain why you shouldn't use a pen to draw a line on the filter paper for paper chromatography. [1 mark]

# More Separation Techniques

## PRACTICAL

Filtration and crystallisation are methods of separating mixtures. Chemists use these techniques all the time to separate solids from liquids, so it's worth making sure you know how to do them.

### Filtration Separates Insoluble Solids from Liquids

- 1) Filtration can be used if your product is an insoluble solid that needs to be separated from a liquid reaction mixture.
- 2) It can be used in purification as well. For example, solid impurities in the reaction mixture can be separated out using filtration.

Insoluble means the solid can't be dissolved in the liquid.



Filter paper folded into a cone shape — the solid is left in the filter paper.

### Two Ways to Separate Soluble Solids from Solutions

If a solid can be dissolved it's described as being soluble. There are two methods you can use to separate a soluble salt from a solution — evaporation and crystallisation.

#### Evaporation

- 1) Pour the solution into an evaporating dish.
- 2) Slowly heat the solution. The solvent will evaporate and the solution will get more concentrated. Eventually, crystals will start to form.
- 3) Keep heating the evaporating dish until all you have left are dry crystals.

evaporating dish



You don't have to use a Bunsen burner, you could use a water bath, or an electric heater.

Evaporation is a really quick way of separating a soluble salt from a solution, but you can only use it if the salt doesn't decompose (break down) when it's heated. Otherwise, you'll have to use crystallisation.

#### Crystallisation

- 1) Pour the solution into an evaporating dish and gently heat the solution. Some of the solvent will evaporate and the solution will get more concentrated.
- 2) Once some of the solvent has evaporated, or when you see crystals start to form (the point of crystallisation), remove the dish from the heat and leave the solution to cool.
- 3) The salt should start to form crystals as it becomes insoluble in the cold, highly concentrated solution.
- 4) Filter the crystals out of the solution, and leave them in a warm place to dry. You could also use a drying oven or a desiccator.

You should also use crystallisation if you want to make nice big crystals of your salt.



Salt crystallising out of solution.

### Filtration and Crystallisation can be Used to Separate Rock Salt

- 1) Rock salt is simply a mixture of salt and sand (they spread it on the roads in winter).
- 2) Salt and sand are both compounds — but salt dissolves in water and sand doesn't. This vital difference in their physical properties gives a great way to separate them.

Here's what to do...

- 1) Grind the mixture to make sure the salt crystals are small, so will dissolve easily.
- 2) Put the mixture in water and stir. The salt will dissolve, but the sand won't.
- 3) Filter the mixture. The grains of sand won't fit through the tiny holes in the filter paper, so they collect on the paper instead. The salt passes through the filter paper as it's part of the solution.
- 4) Evaporate the water from the salt so that it forms dry crystals.

You can heat the mixture to help dissolve the salt.

You could also use crystallisation here if you wanted to make nice, big crystals.

### Revise mixtures — just filter out the important bits...

Two out of three pages on separating mixtures done, phew... But before you dash on to the next page (I know, it's just so exciting), make sure you know this page to a T. Talking about Tea, I need a cuppa...

- Q1 A student needs to produce pure crystals of copper sulfate from an aqueous solution of copper sulfate. Describe how the student could use crystallisation for this process.

[4 marks]

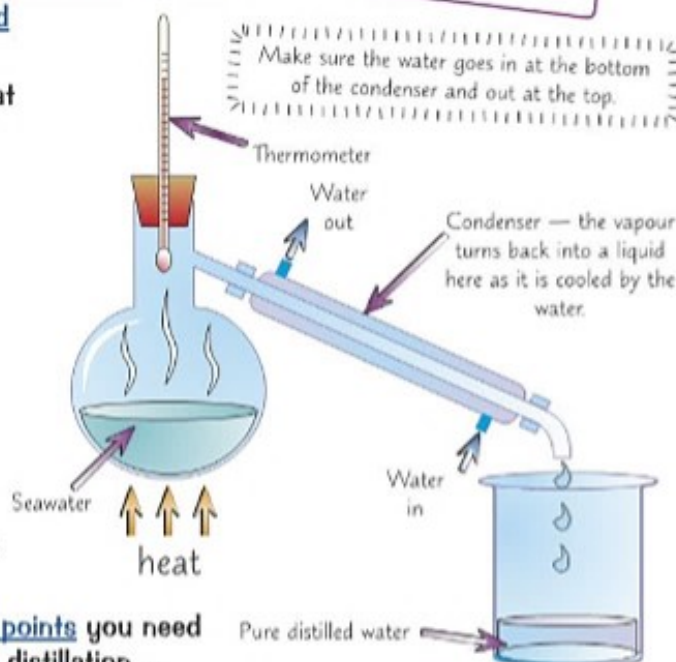
# Distillation

Distillation is used to separate mixtures which contain liquids. There are two types that you should know about — simple and fractional. Hopefully, this page will 'distil' everything you need to know... ho ho.

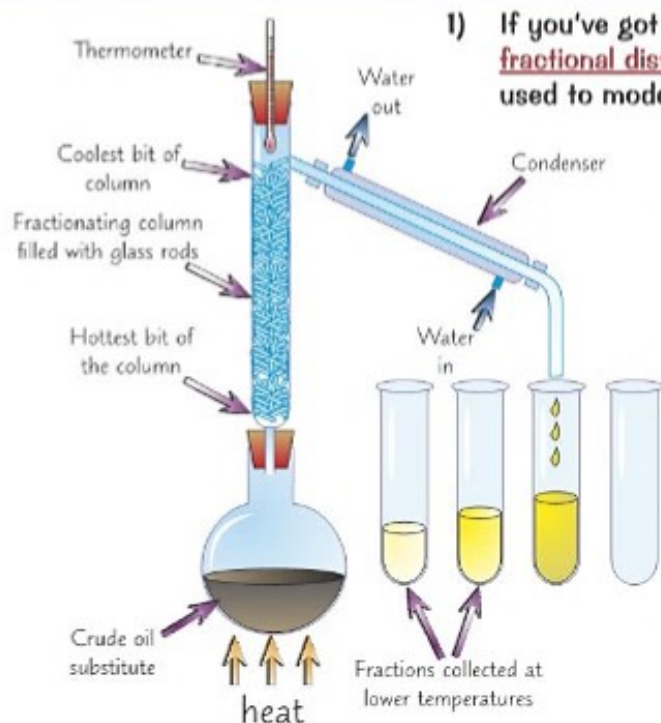
## Simple Distillation is Used to Separate Out Solutions

- 1) Simple distillation is used for separating out a liquid from a solution.
- 2) The solution is heated. The part of the solution that has the lowest boiling point evaporates first.
- 3) The vapour is then cooled, condenses (turns back into a liquid) and is collected.
- 4) The rest of the solution is left behind in the flask.
- 5) You can use simple distillation to get pure water from seawater. The water evaporates and is condensed and collected. Eventually you'll end up with just the salt left in the flask.
- 6) The problem with simple distillation is that you can only use it to separate things with very different boiling points — if the temperature goes higher than the boiling point of the substance with the higher boiling point, they will mix again.
- 7) If you have a mixture of liquids with similar boiling points you need another method to separate them — like fractional distillation...

### PRACTICAL



## Fractional Distillation is Used to Separate a Mixture of Liquids



- 1) If you've got a mixture of liquids you can separate it using fractional distillation. Here is a lab demonstration that can be used to model fractional distillation of crude oil at a refinery.
- 2) You put your mixture in a flask and stick a fractionating column on top. Then you heat it.
- 3) The different liquids will all have different boiling points — so they will evaporate at different temperatures.
- 4) The liquid with the lowest boiling point evaporates first. When the temperature on the thermometer matches the boiling point of this liquid, it will reach the top of the column.
- 5) Liquids with higher boiling points might also start to evaporate. But the column is cooler towards the top. So they will only get part of the way up before condensing and running back down towards the flask.
- 6) When the first liquid has been collected, you raise the temperature until the next one reaches the top.

### Fractionating — sounds a bit too much like maths to me...

You made it to the end of separation techniques. Congratulations. Now all you need to do is learn these techniques. Shouldn't be too tricky. Make sure you scribble all this stuff down — you'd be crazy not to.

- Q1 Propan-1-ol, methanol and ethanol have boiling points of 97 °C, 65 °C and 78 °C respectively. A student uses fractional distillation to separate a mixture of these compounds. State which liquid will be collected in the second fraction and explain why.

[2 marks]

# The History of The Atom

You might have thought you were done with the atom after page 12. Unfortunately amigo, you don't get away that easily — there's more you need to learn. Hold on to your hat, you're going on a journey through **time**...

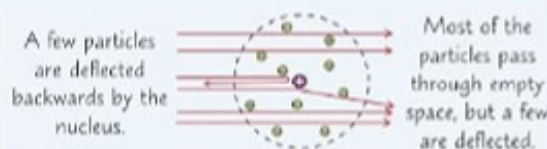
## The Theory of Atomic Structure Has Changed Over Time

- 1) At the start of the 19th century **John Dalton** described atoms as **solid spheres**, and said that different spheres made up the different **elements**.
- 2) In 1897 **J J Thomson** concluded from his experiments that atoms **weren't** solid spheres. His measurements of **charge** and **mass** showed that an atom must contain even smaller, negatively charged particles — **electrons**. The 'solid sphere' idea of atomic structure had to be changed. The new theory was known as the '**plum pudding model**'.
- 3) The plum pudding model showed the atom as a **ball** of **positive charge** with **electrons** stuck in it.



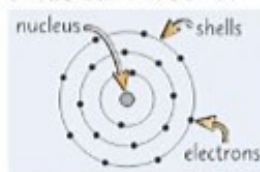
## Rutherford Showed that the Plum Pudding Model Was Wrong

- 1) In 1909 Ernest **Rutherford** and his student Ernest **Marsden** conducted the famous **alpha particle scattering experiments**. They fired positively charged **alpha particles** at an extremely thin sheet of gold.
- 2) From the plum pudding model, they were **expecting** the particles to **pass straight through** the sheet or be **slightly deflected** at most. This was because the positive charge of each atom was thought to be very **spread out** through the 'pudding' of the atom. But, whilst most of the particles **did go straight through** the gold sheet, some were deflected **more than expected**, and a small number were **deflected backwards**. So the plum pudding model **couldn't** be right.
- 3) Rutherford came up with an idea to explain this new evidence — the **nuclear model** of the atom. In this, there's a tiny, positively charged **nucleus** at the centre, where most of the **mass** is concentrated. A 'cloud' of negative electrons surrounds this nucleus — so most of the atom is **empty space**. When alpha particles came near the **concentrated, positive charge** of the **nucleus**, they were **deflected**. If they were fired directly at the nucleus, they were deflected **backwards**. Otherwise, they passed through the empty space.



## Bohr's Nuclear Model Explains a Lot

- 1) Scientists realised that electrons in a 'cloud' around the nucleus of an atom, as Rutherford described, would be attracted to the nucleus, causing the atom to **collapse**. Niels Bohr's nuclear model of the atom suggested that all the electrons were contained in **shells**.
- 2) Bohr proposed that electrons **orbit** the nucleus in **fixed shells** and aren't anywhere in between. Each shell is a fixed distance from the nucleus.
- 3) Bohr's theory of atomic structure was supported by many **experiments** and it helped to explain lots of other scientists' **observations** at the time.



## Further Experiments Showed the Existence of Protons

- 1) Further experiments by Rutherford and others showed that the nucleus can be **divided** into smaller particles, which each have the **same charge** as a **hydrogen nucleus**. These particles were named **protons**.
- 2) About 20 years after scientists had accepted that atoms have nuclei, **James Chadwick** carried out an experiment which provided evidence for **neutral particles** in the nucleus which are now called **neutrons**. The discovery of neutrons resulted in a model of the atom which was **pretty close** to the **modern day** accepted version, known as the **nuclear model** (see page 12).

## I wanted to be a model — but I ate too much plum pudding...

In science, other people's work is constantly being built upon — increasing our understanding of a topic.

- Q1 Describe the 'plum pudding' model of the atom. [1 mark]
- Q2 Rutherford devised an experiment where alpha particles were fired through gold foil. Most of the particles passed through the foil, but some were deflected by different angles, and some were even deflected backwards. Explain why this disproves the plum pudding model. [2 marks]

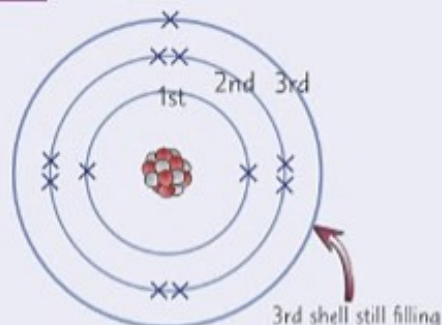
# Electronic Structure

The fact that electrons occupy 'shells' around the nucleus is what causes the whole of chemistry. Remember that, and watch how it applies to each bit of it. It's ace.

## Electron Shell Rules:

- 1) Electrons always occupy **shells** (sometimes called **energy levels**).
- 2) The **lowest** energy levels are **always filled first** — these are the ones closest to the nucleus.
- 3) Only **a certain number** of electrons are allowed in each shell:  
**1st shell: 2**   **2nd shell: 8**   **3rd shell: 8**
- 4) Atoms are much **happier** when they have **full electron shells** — like the **noble gases** in **Group 0**.
- 5) In most atoms, the **outer shell** is **not full** and this makes the atom want to **react** to fill it.

Electron configurations can be shown as **diagrams** like this...



...or as **numbers** like this: **2, 8, 1**

Both of the configurations above are for **sodium**.

## Follow the Rules to Work Out Electronic Structures

You can easily work out the **electronic structures** for the first **20** elements of the periodic table (things get a bit more complicated after that).

### EXAMPLE:

What is the electronic structure of nitrogen?

- 1) Nitrogen's atomic number is 7. This means it has 7 protons... so it must have 7 electrons.
- 2) Follow the 'Electron Shell Rules' above. The first shell can only take 2 electrons and the second shell can take a maximum of 8 electrons.

**So the electronic structure for nitrogen must be 2, 5.**

### EXAMPLE:

What is the electronic structure of magnesium?

- 1) Magnesium's atomic number is 12. This means it has 12 protons... so it must have 12 electrons.
- 2) Follow the 'Electron Shell Rules' above. The first shell can only take 2 electrons and the second shell can take a maximum of 8 electrons, so the third shell must also be partially filled.

**So the electronic structure for magnesium must be 2, 8, 2.**

Here are some more examples of electronic structures:

<b>H</b> Hydrogen  <b>1</b> Proton no. = 1	<b>He</b> Helium  <b>2</b> Proton no. = 2	<b>Li</b> Lithium  <b>2,1</b> Proton no. = 3	<b>C</b> Carbon  <b>2,4</b> Proton no. = 6	<b>Ne</b> Neon  <b>2,8</b> Proton no. = 10	<b>Ca</b> Calcium  <b>2,8,8,2</b> Proton no. = 20
---	--	---	---	---	--

## The electronic structure of the fifth element — it's a bit boron...

Electronic structures may seem a bit complicated at first but once you learn the rules, it's a piece of cake. And just like cake, you'll never regret going back for some more. Better get practising.

- Q1 Give the electronic structure of aluminium (atomic number = 13). [1 mark]
- Q2 Give the electronic structure of argon (atomic number = 18). [1 mark]



# The Modern Periodic Table

So, as you've seen it took a while to get to the [periodic table](#) that you will (soon) know and love. I present to you a chemist's best friend...

## The Periodic Table Helps you to See Patterns in Properties

- 1) There are [100ish elements](#), which all materials are made of.
- 2) In the periodic table the elements are laid out in order of [increasing atomic \(proton\) number](#). Arranging the elements like this means there are [repeating patterns](#) in the [properties](#) of the elements. (The properties are said to occur [periodically](#), hence the name [periodic table](#).)
- 3) If it wasn't for the periodic table [organising everything](#), you'd have a [heck of a job](#) remembering all those properties. It's [ace](#).
- 4) It's a handy tool for working out which elements are [metals](#) and which are [non-metals](#). Metals are found to the [left](#) and non-metals to the [right](#).

relative atomic mass  
atomic number

1 H Hydrogen

alkali metals (see page 24)    transition metals (see page 23)    halogens (see page 25)    noble gases (see page 26)    (pink line separates metals and non-metals)

metals    non-metals

- 5) Elements with [similar properties](#) form [columns](#).
- 6) These [vertical columns](#) are called [groups](#).
- 7) The [group number](#) tells you how many [electrons](#) there are in the [outer shell](#). For example, [Group 1](#) elements all have [one](#) electron in their outer shell and [Group 7](#) all have [seven](#) electrons in their outer shell. The exception to the rule is [group 0](#), for example Helium has two electrons in its outer shell. This is useful as the way atoms react depends upon [the number of electrons](#) in their [outer shell](#). So all elements in the same group are likely to react in a similar way.
- 8) If you know the [properties](#) of [one element](#), you can [predict](#) properties of [other elements](#) in that group — and in the exam, you might be asked to do this. For example the [Group 1](#) elements are Li, Na, K, Rb, Cs and Fr. They're all [metals](#) and they [react in a similar way](#) (see page 24).
- 9) You can also make predictions about trends in [reactivity](#). E.g. in Group 1, the elements react [more vigorously](#) as you go [down](#) the group. And in Group 7, [reactivity decreases](#) as you go down the group.
- 10) The [rows](#) are called [periods](#). Each new period represents another [full shell](#) of electrons.

## I'm in a chemistry band — I play the symbols...

Because the periodic table is organised into groups and periods, it allows us to see trends in both reactivity and properties. And this means we can make predictions on how reactions will occur. How neat is that?

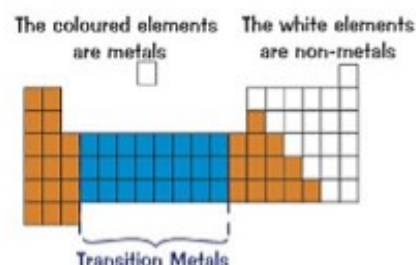
- Q1 Using a periodic table, state how many electrons beryllium has in its outer shell. [1 mark]
- Q2 Chlorine reacts in a similar way to bromine. Suggest a reason why. [1 mark]
- Q3 Sodium readily forms 1+ ions. Suggest what ions potassium forms and explain why. [1 mark]

# Metals and Non-Metals

I can almost guarantee you'll touch something metallic today, that's how important metals are to modern life.

## Most Elements are Metals

- 1) Metals are elements which can form positive ions when they react.
- 2) They're towards the bottom and to the left of the periodic table.
- 3) Most elements in the periodic table are metals.
- 4) Non-metals are at the far right and top of the periodic table.
- 5) Non-metals don't generally form positive ions when they react.



## The Electronic Structure of Atoms Affects How They Will React

- 1) Atoms generally react to form a full outer shell. They do this via losing, gaining or sharing electrons.
- 2) Metals to the left of the periodic table don't have many electrons to remove and metals towards the bottom of the periodic table have outer electrons which are a long way from the nucleus so feel a weaker attraction. Both these effects means that not much energy is needed to remove the electrons so it's feasible for the elements to react to form positive ions with a full outer shell.
- 3) For non-metals, forming positive ions is much more difficult. This is as they are either to the right of the periodic table — where they have lots of electrons to remove to get a full outer shell, or towards the top — where the outer electrons are close to the nucleus so feel a strong attraction. It's far more feasible for them to either share or gain electrons to get a full outer shell.

## Metals and Non-Metals Have Different Physical Properties

- 1) All metals have metallic bonding which causes them to have similar basic physical properties.
  - They're strong (hard to break), but can be bent or hammered into different shapes (malleable).
  - They're great at conducting heat and electricity.
  - They have high boiling and melting points.
- 2) As non-metals don't have metallic bonding, they don't tend to exhibit the same properties as metals. They tend to be dull looking, more brittle, aren't always solids at room temperature, don't generally conduct electricity and often have a lower density.



## Transition Metals can be Found Between Group 2 and Group 3

- 1) Transition metals are in the centre of the periodic table (see above).
- 2) Transition metals are typical metals, and have the properties you would expect of a 'proper' metal — they're good conductors of heat and electricity, and they're very dense, strong and shiny.
- 3) Transition metals also have some pretty special properties...
  - Transition metals can have more than one ion. For example, copper forms  $\text{Cu}^+$  and  $\text{Cu}^{2+}$  ions. Cobalt forms  $\text{Co}^{2+}$  and  $\text{Co}^{3+}$  ions.
  - Transition metal ions are often coloured, and so compounds that contain them are colourful. For example, potassium chromate(VI), which is yellow, potassium(VII) manganate is purple.
  - Transition metal compounds often make good catalysts (things that speed up the rate of a reaction — see p.68). For example, a nickel based catalyst is used in the hydrogenation of alkenes (p.79), and an iron catalyst is used in the Haber process for making ammonia (p.104).

## You can 'rock out' to metal, you can sway gently to non-metal...

Metals and non-metals are like chalk and cheese... Though I hope there's no metal in your cheese.

Q1 State three properties of metals.

[3 marks]

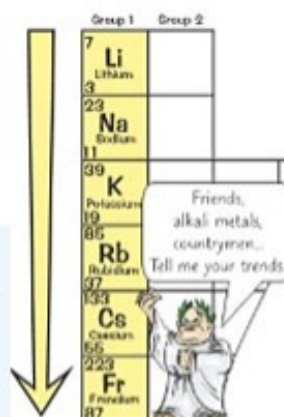
# Group 1 Elements

Group 1 elements are known as the **alkali metals**. As metals go, they're pretty **reactive**.

## The Group 1 Elements are **Reactive, Soft Metals**

- 1) The alkali metals are lithium, sodium, potassium, rubidium, caesium and francium.
- 2) They all have **one electron** in their outer shell which makes them **very reactive** and gives them **similar properties**.
- 3) The alkali metals are all **soft** and have **low density**.
- 4) The **trends** for the alkali metals as you go **down** Group 1 include:

- **Increasing reactivity** — the outer electron is **more easily lost** as the attraction between the nucleus and electron decreases, because the electron is **further away** from the nucleus the further down the group you go.
- **Lower melting** and **boiling** points.
- **Higher relative atomic mass**.



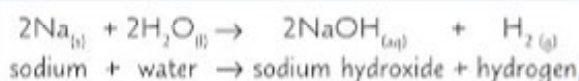
## Alkali Metals Form **Ionic Compounds with Non-Metals**

- 1) The Group 1 elements don't need much energy to lose their one outer electron to form a full outer shell, so they readily form **1+ ions**.
- 2) It's so easy for them to lose their outer electron that they only ever react to form **ionic compounds**. These compounds are generally **white solids** that dissolve in water to form **colourless solutions**.

Don't worry, there's more on ionic compounds on page 30.

### Reaction with water

- When Group 1 metals are put in **water**, they react **vigorously** to produce **hydrogen gas** and **metal hydroxides** — salts that dissolve in water to produce **alkaline solutions**.
- The **more reactive** (lower down in the group) an alkali metal is, the more violent the reaction.
- The amount of **energy** given out by the reaction increases down the group — the reaction with potassium releases enough energy to ignite hydrogen.



All the Group 1 metals react with water in a similar way.

### Reaction with chlorine

- 1) Group 1 metals react **vigorously** when heated in **chlorine gas** to form white **metal chloride salts**.
- 2) As you go down the group, reactivity increases so the reaction with chlorine gets **more vigorous**.



### Reaction with oxygen

The Group 1 metals can react with **oxygen** to form a **metal oxide**. Different types of **oxide** will form depending on the Group 1 metal:

- Lithium reacts to form **lithium oxide** ( $\text{Li}_2\text{O}$ ).
- Sodium reacts to form a mixture of **sodium oxide** ( $\text{Na}_2\text{O}$ ) and **sodium peroxide** ( $\text{Na}_2\text{O}_2$ ).
- Potassium reacts to form a mixture of **potassium peroxide** ( $\text{K}_2\text{O}_2$ ) and **potassium superoxide** ( $\text{KO}_2$ ).

The reactions with oxygen are why Group 1 metals tarnish in the air — the metal reacts with oxygen in the air to form a dull metal oxide layer.

## Group 1 Metals Have **Different Properties to Transition Metals**

- 1) Group 1 metals are much **more reactive** than transition metals — they react more vigorously with **water**, **oxygen** or **Group 7 elements**, for example.
- 2) They're also much less **dense**, **strong** and **hard** than the transition metals, and have much **lower melting points**. E.g. manganese melts at 2000 °C, sodium melts at 98 °C.

## Back to the drawing board with my lithium swim shorts design...

Reactions of alkali metals need safety precautions, but they fizz in water and might explode. Cool.

Q1 Explain the trend in reactivity as you go down Group 1.

[2 marks]

# Group 7 Elements

The Group 7 elements are known as the **halogens**. The whole 'trend thing' happens with the halogens as well — that shouldn't come as a surprise.

## The Halogens are All Non-Metals with Coloured Vapours

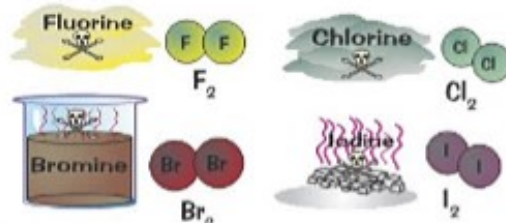
**Fluorine** is a very reactive, poisonous **yellow gas**.

**Chlorine** is a fairly reactive, poisonous **dense green gas**.

**Bromine** is a dense, poisonous, **red-brown volatile liquid**.

**Iodine** is a **dark grey crystalline solid** or a **purple vapour**.

They all exist as molecules which are **pairs of atoms**.



## Learn These Trends:

As you go **DOWN** Group 7, the **halogens**:

- 1) become **LESS REACTIVE** — it's **harder to gain** an extra electron, because the outer shell's further from the nucleus.
- 2) have **HIGHER MELTING AND BOILING POINTS**.
- 3) have **HIGHER RELATIVE ATOMIC MASSES**.

All the Group 7 elements react in **similar ways**. This is because they all have **seven electrons** in their outer shell.

Group 6	Group 7
	19 F Fluorine 9
	35.5 Cl Chlorine 17
	80 Br Bromine 35
	127 I Iodine 53
	210 At Astatine 85

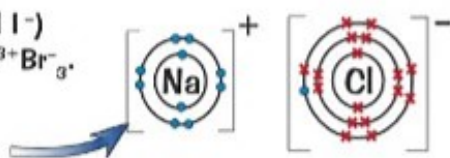
You can use these trends to predict properties of halogens. For example, you know that iodine will have a higher boiling point than chlorine as it's further down the group in the periodic table.

## Halogens can Form Molecular Compounds

Halogen atoms can **share** electrons via covalent bonding (see page 31) with other **non-metals** so as to achieve a **full outer shell**. For example  $\text{HCl}$ ,  $\text{PCl}_5$ ,  $\text{HF}$  and  $\text{CCl}_4$  contain covalent bonds. The compounds that form when halogens react with non-metals all have **simple molecular structures** (see p.32).

## Halogens Form Ionic Bonds with Metals

- 1) The halogens form **1- ions** called **halides** ( $\text{F}^-$ ,  $\text{Cl}^-$ ,  $\text{Br}^-$  and  $\text{I}^-$ ) when they bond with **metals**, for example  $\text{Na}^+\text{Cl}^-$  or  $\text{Fe}^{3+}\text{Br}_3^-$ .
- 2) The compounds that form have **ionic structures**.
- 3) The diagram shows the bonding in sodium chloride,  $\text{NaCl}$ .

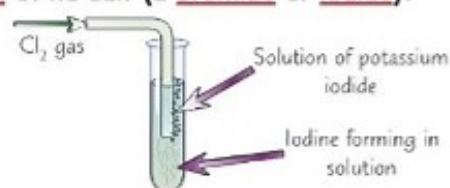
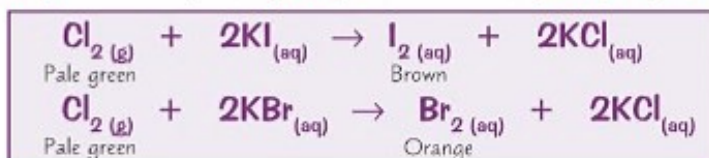


## More Reactive Halogens Will Displace Less Reactive Ones

A **displacement reaction** can occur between a more reactive halogen and the salt of a less reactive one.

E.g. **chlorine** can displace **bromine** and **iodine** from an aqueous **solution** of its salt (a **bromide** or **iodide**).

**Bromine** will also displace **iodine** because of the **trend** in **reactivity**.



## I can see your halo(gen), halo(gen) halo(gen)...

You say 'halo-gen' I say 'ha-logen', let's call the whole thing off... Apart from revision, let's call that whole thing 'on'. Displacement reactions are pretty important in chemistry — better learn the facts.

Q1 Predict whether bromine would displace iodine from sodium iodide and explain why. [2 marks]

Q2 Why do Group 7 elements get less reactive as you go down the group from fluorine to iodine? [3 marks]

# Group 0 Elements

The Group 0 elements are known as **noble gases** — stuffed full of every honourable virtue. They don't react with very much and you can't even see them — making them, well, a bit dull really.

## Group 0 Elements are All Inert, Colourless Gases

- 1) Group 0 elements are called the **noble gases** and include the elements **helium**, **neon** and **argon** (plus a few others).
- 2) They all have **eight electrons** in their outer energy level, apart from helium which has two, giving them a **full outer-shell**. As their outer shell is energetically stable they don't need to **give up** or **gain** electrons to become more stable. This means they are more or less **inert** — they **don't react** with much at all.
- 3) They exist as **monatomic gases** — single atoms **not** bonded to each other.
- 4) All elements in Group 0 are **colourless gases** at room temperature.
- 5) As the noble gases are inert they're **non-flammable** — they won't set on fire.

Helium only has electrons in the first shell, which only needs 2 to be filled.

Group 6	Group 7	Group 0
		2 He Helium
		10 Ne Neon
		18 Ar Argon
		36 Kr Krypton
		54 Xe Xenon
		86 Rn Radon

## There are Patterns in the Properties of the Noble Gases

- 1) The **boiling points** of the noble gases **increase** as you move **down** the group along with increasing relative atomic mass.

Noble Gas
helium
neon
argon
krypton
xenon
radon

Increasing boiling point

- 2) The increase in boiling point is due to an **increase** in the **number of electrons** in each **atom** leading to **greater intermolecular forces** between them which need to be overcome. There's more on intermolecular forces for small molecules on page 32.



Here's another pattern. You don't have to learn this one...

- 3) In the exam you may be given the boiling point of one noble gas and asked to **estimate** the value for **another one**. So make sure you know the **pattern**.

### EXAMPLE:

Neon is a gas at 25 °C. Predict what state helium is at this temperature.

Helium has a lower boiling point than neon as it is further up the group.

So, helium must also be a gas at 25 °C.

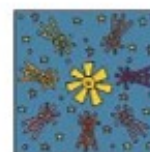
### EXAMPLE:

Radon and krypton have boiling points of -62 °C and -153 °C respectively. Predict the boiling point of xenon.

Xenon comes in between radon and krypton in the group so you can predict that its boiling point would be halfway between their boiling points:  $(-153) + (-62) = -215$   
 $-215 \div 2 = -107.5 \approx -108^\circ\text{C}$

So, xenon should have a boiling point of about -108 °C

The actual boiling point of xenon is -108 °C — just as predicted. Neat!



...or this one.

## Arrrgon — the pirate element...

As noble gases don't really react there isn't too much to learn about them. If you understand why they are unreactive and the trend in boiling points as you go down the group you're sorted.

- Q1 Does xenon or neon have the higher boiling point? [1 mark]
- Q2 Argon is very unreactive. Using your knowledge of its electronic structure, explain why. [2 marks]

# Revision Questions for Topic 1

Well, that wraps up [Topic 1](#) — time to put yourself to the test and find out [how much you really know](#).

- Try these questions and [tick off each one](#) when you [get it right](#).
- When you've done [all the questions](#) under a heading and are [completely happy](#) with it, tick it off.

## Atoms, Elements and Compounds (p.12-15) ☒

- 1) Sketch an atom. Label the nucleus and the electrons. ☒
- 2) What is the charge of a proton? ☒
- 3) True or False? Elements contain more than one type of atom. ☒
- 4) Give the formula for: ☒
  - a) Carbon dioxide
  - b) Sodium carbonate
- 5) Balance these equations: ☒
  - a)  $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$
  - b)  $\text{H}_2\text{SO}_4 + \text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$

## Mixtures and Separation (p.16-18) ☐

- 6) What is the difference between a compound and a mixture? ☐
- 7) What is the name of the pattern formed from carrying out paper chromatography? ☐
- 8) Which method of separation is useful to separate an insoluble solid from a liquid? ☐
- 9) Give the name of a method to separate a soluble solid from a liquid. ☐
- 10) Which method of distillation would you use to separate liquids with similar boiling points? ☐

## Electronic Structure and the History of the Periodic Table (p.19-22) ☒

- 11) Who discovered that the plum pudding model was wrong? ☐
- 12) Who first devised an experiment that proved the existence of the neutron? ☐
- 13) What is the electronic structure of sodium? ☐
- 14) Why did Mendeleev leave gaps in his Table of Elements? ☐

## Groups of the Periodic Table (p.23-26) ☒

- 15) How are the group number and the number of electrons in the outer shell of an element related? ☐
- 16) What kind of ions do metals form? ☐
- 17) Where are non-metals on the periodic table? ☐
- 18) Give three properties which are specific to transition metals. ☐
- 19) State three trends as you go down Group 1. ☐
- 20) State the products of the reaction of sodium and water. ☐
- 21) State the differences between Group 1 and transition metals for the following properties: ☐
  - a) Hardness
  - b) Reactivity
  - c) Melting Points
- 22) How do the boiling points of halogens change as you go down the group from fluorine to astatine? ☐
- 23) What is the charge of the ions that halogens form when they react with metals? ☐
- 24) Predict whether iodine is displaced by chlorine in a solution of potassium iodide. ☐
- 25) What is the trend in boiling point as you go down Group 0? ☐



# Ionic Bonding

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

## Ionic Bonding — Transfer of Electrons

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

## Dot and Cross Diagrams Show How Ionic Compounds are Formed

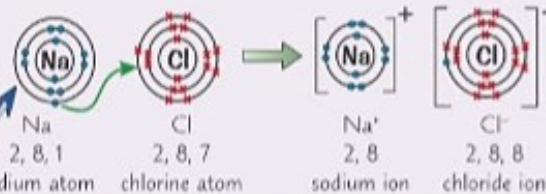
**Dot and cross diagrams** show the **arrangement** of electrons in an atom or ion. Each electron is represented by a **dot** or a **cross**. So these diagrams can show which **atom** the electrons in an **ion** originally came from.

### Sodium Chloride (NaCl)

The **sodium** atom gives up its outer electron, becoming an **Na<sup>+</sup>** ion.

The **chlorine** atom picks up the electron, becoming a **Cl<sup>-</sup>** (**chloride**) ion.

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



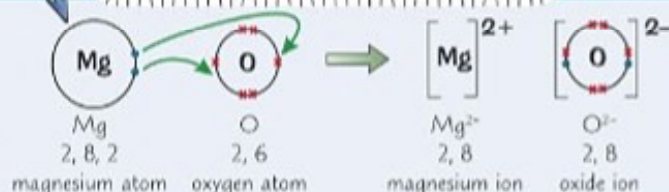
Remember, you can work out how many electrons an atom will gain or lose from its group number.

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

The name's Bond. Ionic Bond.

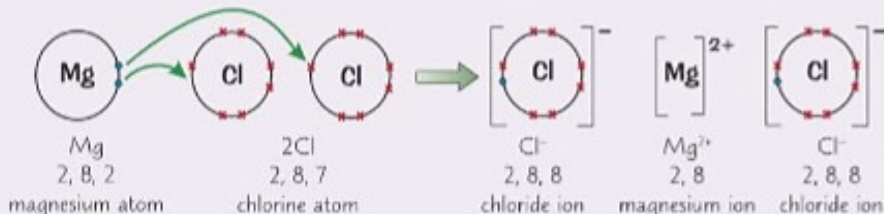
### Magnesium Oxide (MgO)

The **magnesium** atom gives up its **two** outer electrons, becoming an **Mg<sup>2+</sup>** ion. The **oxygen** atom picks up the electrons, becoming an **O<sup>2-</sup>** (**oxide**) ion.



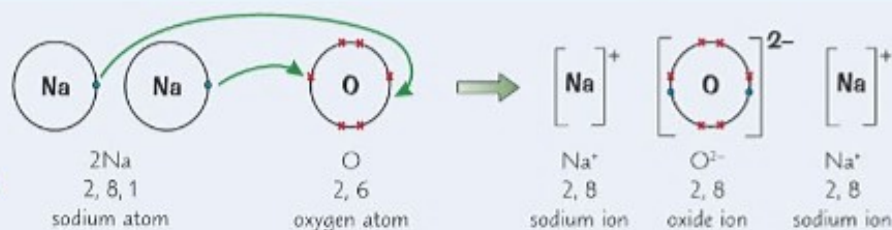
### Magnesium Chloride (MgCl<sub>2</sub>)

The **magnesium** atom gives up its **two** outer electrons, becoming an **Mg<sup>2+</sup>** ion. The two **chlorine** atoms pick up **one electron each**, becoming **two Cl<sup>-</sup>** (**chloride**) ions.



### Sodium Oxide (Na<sub>2</sub>O)

Two **sodium** atoms each give up their single outer electron, becoming **two Na<sup>+</sup>** ions. The **oxygen** atom picks up the **two** electrons, becoming an **O<sup>2-</sup>** ion.



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

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

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

- Q1 Describe, in terms of electron transfer, how sodium (Na) and chlorine (Cl) react to form sodium chloride (NaCl). [3 marks]
- Q2 Draw a dot and cross diagram to show how potassium (a Group 1 metal) and bromine (a Group 7 non metal) form potassium bromide (KBr). [3 marks]

# Ionic Compounds

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

## Ionic Compounds Have A Regular Lattice Structure

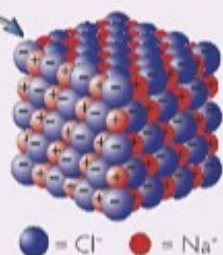
- 1) Ionic compounds have a structure called a giant ionic lattice.
- 2) The ions form a closely packed regular lattice arrangement and there are very strong electrostatic forces of attraction between oppositely charged ions, in all directions in the lattice.

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

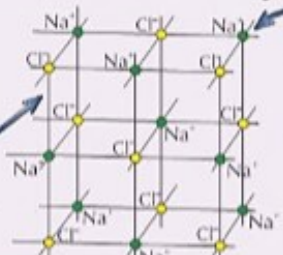
A single crystal of sodium chloride (table salt) is one giant ionic lattice. The  $\text{Na}^+$  and  $\text{Cl}^-$  ions are held together in a regular lattice. The lattice can be represented in different ways...

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

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



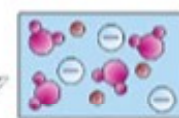
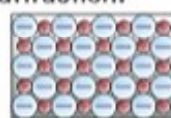
The  $\text{Na}^+$  and  $\text{Cl}^-$  ions alternate.



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

## Ionic Compounds All Have Similar Properties

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



Dissolved in Water



Melted

## Look at Charges to Find the Formula of an Ionic Compound

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

### EXAMPLE:

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

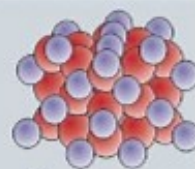
- 1) Look at the diagram to work out what ions are in the compound.
- 2) Work out what charges the ions will form.
- 3) Balance the charges so the charge of the empirical formula is zero.

The compound contains potassium and oxide ions.

Potassium is in Group 1 so forms 1+ ions.

Oxygen is in Group 6 so forms 2- ions.

A potassium ion only has a 1+ charge, so you'll need two of them to balance out the 2- charge of an oxide ion. The empirical formula is  $\text{K}_2\text{O}$ .



● = Potassium ion  
● = Oxide ion

## Giant ionic lattices — all over your chips...

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

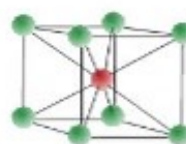
Q1 The structure of an ionic compound is shown on the right.

- a) Predict, with reasoning, whether the compound has a high or a low melting point.
- b) Explain why the compound can conduct electricity when molten.
- c) Use the diagram to find the empirical formula of the compound.

[2 marks]

[1 mark]

[3 marks]



● = chloride ion  
● = caesium ion

# Covalent Bonding

Some elements bond ionically (see page 29) but others form strong **covalent** bonds. This is where atoms **share** electrons with each other so that they've got full outer shells.

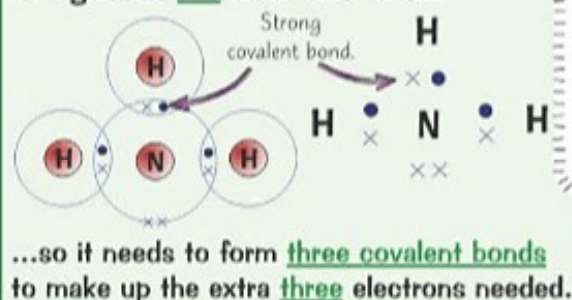
## Covalent Bonds — Sharing Electrons

- 1) When **non-metal** atoms bond together, they **share** pairs of electrons to make **covalent bonds**.
- 2) The positively charged nuclei of the bonded atoms are attracted to the shared pair of electrons by **electrostatic forces**, making covalent bonds very **strong**.
- 3) Atoms only share electrons in their **outer shells** (highest energy levels).
- 4) Each single **covalent bond** provides one **extra** shared electron for each atom.
- 5) Each atom involved generally makes **enough** covalent bonds to **fill up** its outer shell. Having a full outer shell gives them the electronic structure of a **noble gas**, which is very **stable**.
- 6) Covalent bonding happens in **compounds** of **non-metals** (e.g.  $\text{H}_2\text{O}$ ) and in **non-metal elements** (e.g.  $\text{Cl}_2$ ).

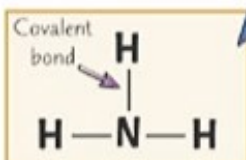
## There are Different Ways of Drawing Covalent Bonds

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

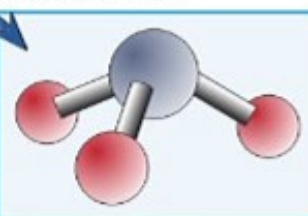
Nitrogen has **five** outer electrons...



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



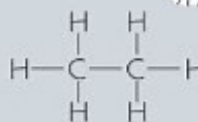
- 4) The **displayed formula** of ammonia ( $\text{NH}_3$ ) shows the covalent bonds as single lines between atoms.
- 5) This is a great way of showing **how** atoms are connected in **large** molecules. However, they **don't** show the **3D structure** of the molecule, or **which atoms** the electrons in the covalent bond have come from.
- 6) The 3D model of ammonia shows the **atoms**, the **covalent bonds** and their **arrangement** in space next to each other. But 3D models can quickly get **confusing** for large molecules where there are lots of atoms to include. They don't show **where** the electrons in the bonds have **come from**, either.
- 7) You can find the **molecular formula** of a simple molecular compound from **any** of these diagrams by **counting up** how many atoms of each element there are.



### EXAMPLE:

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

In the diagram, there are two carbon atoms and six hydrogen atoms. So the molecular formula is  $\text{C}_2\text{H}_6$ .



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

## Sharing is caring...

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

Q1 Draw a dot and cross diagram to show the bonding in a molecule of ammonia ( $\text{NH}_3$ ).

[2 marks]

# Simple Molecular Substances

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

## Learn These Examples of Simple Molecular Substances

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

### Hydrogen, $H_2$

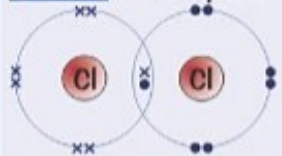
Hydrogen atoms have just one electron. They **only need one more** to complete the first shell...



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

### Chlorine, $Cl_2$

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



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

### Oxygen, $O_2$

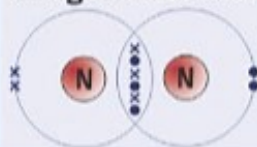
Each oxygen atom needs **two more electrons** to complete its outer shell...



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

### Nitrogen, $N_2$

Nitrogen atoms need **three more electrons**...

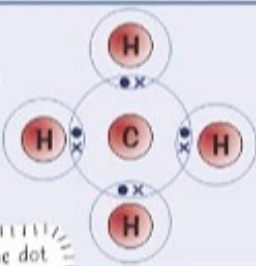


...so **two nitrogen atoms** share **three pairs of electrons** to fill their outer shells. This creates a **triple bond**.

### Methane, $CH_4$

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

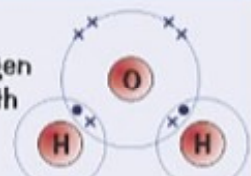
It can form **four covalent bonds** with **hydrogen** atoms to fill up its outer shell.



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

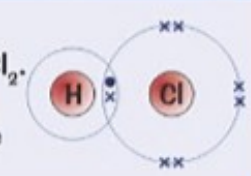
### Water, $H_2O$

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



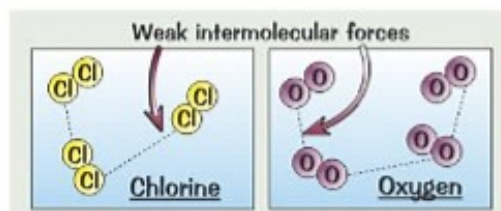
### Hydrogen Chloride, $HCl$

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



## Properties of Simple Molecular Substances

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



## May the intermolecular force be with you...

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

Q1 Explain why oxygen,  $O_2$ , is a gas at room temperature.

[1 mark]

Q2 Explain why nitrogen,  $N_2$ , doesn't conduct electricity.

[1 mark]

# Polymers and Giant Covalent Structures

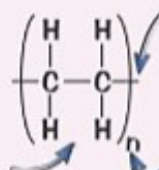
Wouldn't it be simply marvellous if only simple molecular substances had covalent bonds, and it was now time to put your feet up? Well it's not like that. **Polymers** and **giant covalent substances** also have **covalent bonds**.

## Polymers Are Long Chains of Repeating Units

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

This polymer is called 'poly(ethene)'.

The bit in brackets is the repeating unit.



The bonds through the brackets join up to the next repeating unit.

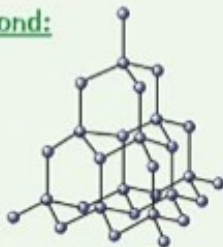
'n' is a large number. It tells you that the unit's repeated lots of times.
- 4) To find the **molecular formula** of a polymer, write down the molecular formula of the **repeating unit** in **brackets**, and put an '**n**' outside.
- 5) So for **poly(ethene)**, the molecular formula of the polymer is **(C<sub>2</sub>H<sub>4</sub>)<sub>n</sub>**.
- 6) The intermolecular forces between polymer molecules are **larger** than between simple covalent molecules, so **more energy** is needed to break them. This means most polymers are **solid** at room temperature.
- 7) The intermolecular forces are still **weaker** than ionic or covalent bonds, so they generally have **lower** boiling points than **ionic** or **giant molecular** compounds.

There's more about how polymers are made on p.80 and p.83 and more about their properties and uses on p.96-97.

## Giant Covalent Structures Are Macromolecules

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

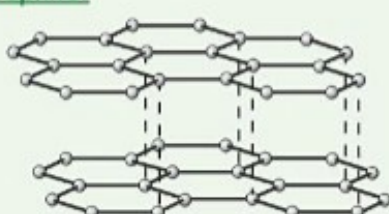
**Diamond:**



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

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

**Graphite:**

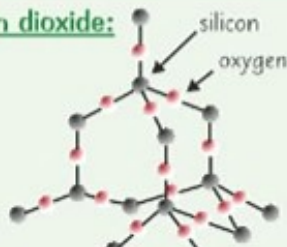


Each carbon atom forms **three covalent bonds** to create **layers of hexagons**. Each carbon atom also has one **delocalised** (free) electron.

A free electron.



**Silicon dioxide:**



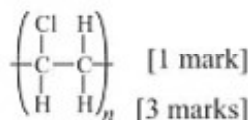
Sometimes called **silica**, this is what **sand** is made of. Each grain of sand is **one giant structure** of silicon and oxygen.

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

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

Q1 The repeating unit of poly(chloroethene) is shown on the right.

What's the molecular formula of poly(chloroethene)?



[1 mark]

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

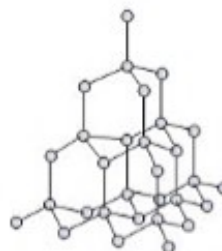
[3 marks]

# Allotropes of Carbon

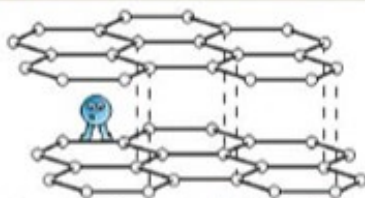
Allotropes are different structural forms of the same element in the same physical state. Carbon's got lots...

## Diamond is Very Hard

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



## Graphite Contains Sheets of Hexagons



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

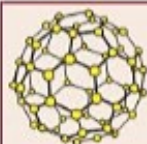
## Graphene is One Layer of Graphite

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



## Fullerenes Form Spheres and Tubes

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



Buckminsterfullerene was the first fullerene to be discovered. It's got the molecular formula C<sub>60</sub> and forms a hollow sphere.

Fullerenes can form nanotubes — tiny carbon cylinders.

The ratio between the length and the diameter of nanotubes is very high.

Nanotubes can conduct both electricity and thermal energy (heat).

They also have a high tensile strength (they don't break when they're stretched).

Technology that uses very small particles such as nanotubes is called nanotechnology. Nanotubes can be used in electronics or to strengthen materials without adding much weight, such as in tennis racket frames.



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

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

Q1 Give three uses of fullerenes.

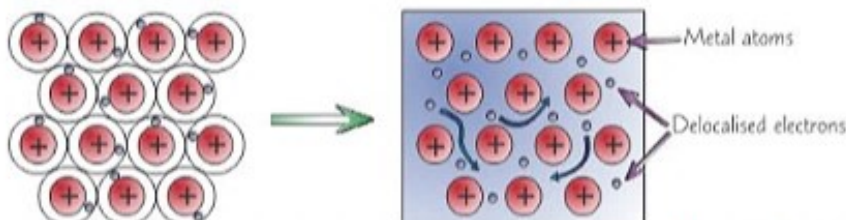
[3 marks]

# Metallic Bonding

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

## Metallic Bonding Involves Delocalised Electrons

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



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

## Most Metals are Solid at Room Temperature

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

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

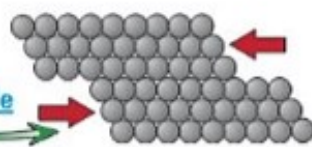


## Metals are Good Conductors of Electricity and Heat

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

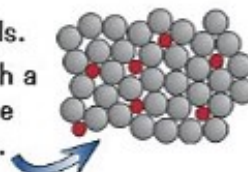
## Most Metals are Malleable

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



## Alloys are Harder Than Pure Metals

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



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

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

- Q1 Copper is a metallic element. Describe and explain what property of copper makes it suitable for using in electrical circuits. [2 marks]
- Q2 Suggest why an alloy of copper, rather than pure copper, is used to make hinges for doors [1 mark]

# States of Matter

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

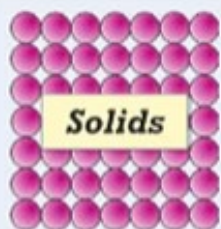
## The Three States of Matter — Solid, Liquid and Gas

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

- the material (the structure of the substance and the type of bonds holding the particles together),
- the temperature,
- the pressure.

The particles could be atoms, ions or molecules.

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



**Solids**

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



**Liquids**

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



**Gases**

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

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

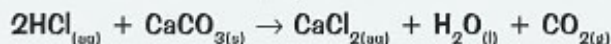
## State Symbols Tell You the State of a Substance in an Equation

You saw on page 15 how a chemical reaction can be shown using a word equation or symbol equation. Symbol equations can also include state symbols next to each substance — they tell you what physical state the reactants and products are in:

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

'Aqueous' means 'dissolved in water'.

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



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

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

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

[1 mark]

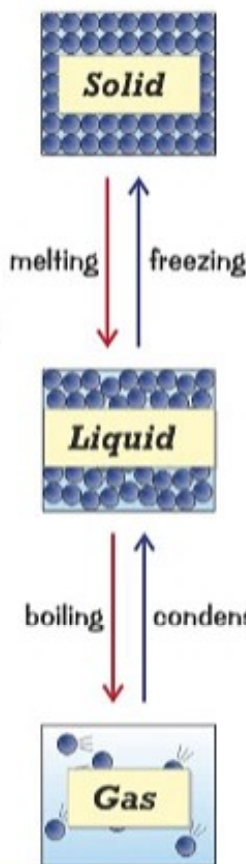
# Changing State

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

## Substances Can Change from One State to Another

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

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



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

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

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

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

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

### EXAMPLE:

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

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

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

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

## Some people are worth melting for...

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

Q1 Ethanol melts at -114 °C and boils at 78 °C. Predict the state that ethanol is in at:

- a) -150 °C      b) 0 °C      c) 25 °C      d) 100 °C

[4 marks]

# Nanoparticles

Just time to squeeze in something really small before the end of the topic...

## Nanoparticles Are Really Really Really Really Tiny

...smaller than that.

Particles are put into categories depending on their diameter. For example:

**Coarse particles** ( $PM_{10}$ ) have a diameter between

2500 nm ( $2.5 \times 10^{-6}$  m) and 10 000 nm ( $1 \times 10^{-5}$  m). They're also called dust.

**Fine particles** ( $PM_{2.5}$ ) have a diameter between 100 nm ( $1 \times 10^{-7}$  m) and 2500 nm ( $2.5 \times 10^{-6}$  m).

**Nanoparticles** have a diameter between 1 nm ( $1 \times 10^{-9}$  m) and 100 nm ( $1 \times 10^{-7}$  m).

These are particles that contain only a few hundred atoms.

There's a whole area of science that investigates the uses and properties of nanoparticles, called nanoscience.

1 nm = 0.000 000 001 m  
(Like I said — really tiny.)

A typical atom has a diameter of about 0.1 nm ( $1 \times 10^{-10}$  m).

## Nanoparticles Have a Large Surface Area to Volume Ratio

- 1) The surface area to volume ratio is an important factor as it can affect the way that a particle behaves.

$$\text{surface area to volume ratio} = \text{surface area} \div \text{volume}$$

- 2) As particles decrease in size, the size of their surface area increases in relation to their volume. This causes the surface area to volume ratio to increase.
- 3) You can see this happening by using two cubes as an example:



### EXAMPLE:

Find the surface area to volume ratio for each of the cubes below.



This cube has sides of length 100 nm.

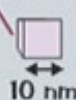
Each face has a surface area of  $100 \text{ nm} \times 100 \text{ nm} = 10\,000 \text{ nm}^2$

The cube has six faces, so the total surface area is  $6 \times 10\,000 \text{ nm}^2 = 60\,000 \text{ nm}^2$

The volume of the cube is  $100 \text{ nm} \times 100 \text{ nm} \times 100 \text{ nm} = 1\,000\,000 \text{ nm}^3$

The surface area to volume ratio = surface area  $\div$  volume =  $60\,000 \div 1\,000\,000 = 0.06 \text{ nm}^{-1}$

The drawings of the cubes aren't to scale. Obviously.



This cube has sides of length 10 nm.

Each face has a surface area of  $10 \text{ nm} \times 10 \text{ nm} = 100 \text{ nm}^2$

The cube has six faces, so the total surface area is  $6 \times 100 \text{ nm}^2 = 600 \text{ nm}^2$

The volume of the cube is  $10 \text{ nm} \times 10 \text{ nm} \times 10 \text{ nm} = 1\,000 \text{ nm}^3$

The surface area to volume ratio = surface area  $\div$  volume =  $600 \div 1\,000 = 0.6 \text{ nm}^{-1}$

As you decrease the side of any cube by a factor of ten, the surface area to volume ratio will always increase by a factor of ten.

- 4) Nanoparticles have a very high surface area to volume ratio — this means the surface area is very large compared to the volume.
- 5) This can cause the properties of a material to be different depending on whether it's a nanoparticle or whether it's in bulk. For example, you'll often need less of a material that's made up of nanoparticles to work as an effective catalyst compared to a material made up of 'normal' sized particles (containing billions of atoms rather than a few hundred).

## Nano nano nano nano nano nano nano — particles...

Nanoparticles are between ten and one hundred times larger than atoms and molecules. That's teeny tiny.

- Q1 Roughly what size are the particles studied in nanoscience? [1 mark]
- Q2 a) The sides of a cube are 50 nm long. What's its surface to volume ratio? [3 marks]
- b) Predict how the surface to volume area would change if the size of the cube was increased so its sides had a length of 500 nm. [2 marks]

# Uses of Nanoparticles

'What's the use of something so small that not even a gnat can see it?' I hear you cry. Well, as you're about to find out, scientists have developed some pretty swanky uses for nanoparticles.

## Nanoparticles Can Be Used in Lots of Things

Finding new ways to use nanoparticles is a really important area of scientific research. Here are some of the uses that have already been developed...

- 1) They have a huge surface area to volume ratio, so they could help make new catalysts.
- 2) Nanomedicine is a hot topic. The idea is that tiny particles (such as fullerenes from page 34) are absorbed more easily by the body than most particles. This means they could deliver drugs right into the cells where they're needed.
- 3) Some nanoparticles conduct electricity, so they can be used in tiny electric circuits for computer chips.
- 4) Silver nanoparticles have antibacterial properties. They can be added to polymer fibres that are then used to make surgical masks and wound dressings and they can also be added to deodorants.
- 5) Nanoparticles are also being used in cosmetics. For example, they're used to improve moisturisers without making them really oily.

You might be asked about why nanoparticles are used for other applications, given their properties. DON'T PANIC. You should be told any extra information that you'll need.

## The Effects of Nanoparticles on Health Aren't Fully Understood

- 1) Although nanoparticles are useful, the way they affect the body isn't fully understood, so it's important that any new products are tested thoroughly to minimise the risks.
- 2) Some people are worried that products containing nanoparticles have been made available before the effects on human health have been investigated properly, and that we don't know what the long-term impacts on health will be.
- 3) As the long-term impacts aren't known, many people believe that products containing nanoscale particles should be clearly labelled, so that consumers can choose whether or not to use them.

- Nanoparticles are being used in sun creams as they have been shown to be better than the materials in traditional sun creams at protecting skin from harmful UV rays.
- They also give better skin coverage than traditional sun creams.
- But it's not yet clear whether the nanoparticles can get into your body, and, if they do, whether they might damage cells.
- It's also possible that when they are washed away they might damage the environment.



## Silver nanoparticles — to make jewellery for nanopeople...

If you thought you'd only ever meet nanoparticles in GCSE Chemistry, then think again. They've got loads of uses out there in the real world, so you'll be sure to hear more about them in the future. If you're still not sure what all the fuss is about, and why something so small might ever be any use, then have a go at the questions below

- Q1 Give one potential application of nanoparticles that can conduct electricity. [1 mark]
- Q2 Nanoparticles may have applications in medicine.  
State one risk of using nanoparticles in this way. [1 mark]
- Q3 Both silver nanoparticles and titanium dioxide nanoparticles can be used to form a transparent layer. Silver nanoparticles conduct heat, whereas titanium dioxide nanoparticles are thermal insulators. State which of the nanoparticles could be used for making car windows that can quickly be defrosted in winter. Explain your answer. [2 marks]

## Revision Questions for Topic 2

And that's the end of [Topic 2](#) — well, nearly the end. You still need to...

- ...try these questions and [tick off each one](#) when you [get it right](#).
- When you've done [all the questions](#) under a heading and are [completely happy](#) with it, tick it off.

### Ions and Ionic Compounds (p.28-30) ☐

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

☐  
☐  
☐  
☐  
☐

### Covalent Substances (p.31-34) ☐

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

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### Metallic Bonding (p.35) ☐

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

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### States of Matter (p.36-37) ☐

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

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### Nanoparticles and their Uses (p.38-39) ☐

- What is nanoscience?
- Give three uses of nanoparticles.
- Explain why people may be wary of using products that contain nanoparticles.

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## Relative Formula Mass

Calculating **relative formula mass** is straight forward enough, but things can get a bit more confusing when you start working out the **percentage compositions** of compounds. Best get cracking, I suppose...

### Compounds Have a Relative Formula Mass, $M_r$

If you have a compound like  $\text{MgCl}_2$  then it has a **relative formula mass**,  $M_r$ , which is just the relative atomic masses of all the atoms in the molecular formula **added together**.

#### EXAMPLE:

Find the relative formula mass of  $\text{MgCl}_2$ .

- 1) Look up the **relative atomic masses** of all the elements in the compound on the periodic table. (In the exams, you might be given the  $A_r$  you need in the question.)  
 $A_r$  of Mg = 24 and the  $A_r$  of Cl = 35.5.

- 2) **Add up** all the relative atomic masses of the atoms in the compound.  
 $\text{Mg} + (2 \times \text{Cl}) = 24 + (2 \times 35.5) = 95$  So  $M_r$  of  $\text{MgCl}_2 = 95$

There are two chlorine atoms in  $\text{MgCl}_2$ , so the relative atomic mass of chlorine needs to be multiplied by 2.

You can find the relative atomic mass ( $A_r$ ) of an element from the periodic table — it's the same as its mass number. See page 13 for more.

### You Can Calculate the % Mass of an Element in a Compound

This is actually **dead easy** — so long as you've learnt this **formula**:

$$\text{Percentage mass of an element in a compound} = \frac{A_r \times \text{number of atoms of that element}}{M_r \text{ of the compound}} \times 100$$

#### EXAMPLE:

Find the percentage mass of sodium in sodium carbonate,  $\text{Na}_2\text{CO}_3$ .

$A_r$  of sodium = 23,  $A_r$  of carbon = 12,  $A_r$  of oxygen = 16

$M_r$  of  $\text{Na}_2\text{CO}_3 = (2 \times 23) + 12 + (3 \times 16) = 106$

Percentage mass of sodium =  $\frac{A_r \times \text{number of atoms of that element}}{M_r \text{ of the compound}} \times 100 = \frac{23 \times 2}{106} \times 100 = 43\%$

You might also come across more complicated questions where you need to work out the percentage mass.

#### EXAMPLE:

A mixture contains 20% iron ions by mass. What mass of iron chloride ( $\text{FeCl}_2$ ) would you need to provide the iron ions in 50 g of the mixture?  $A_r$  of Fe = 56,  $A_r$  of Cl = 35.5.

- 1) Find the **mass** of iron in the mixture.

The mixture contains 20% iron by mass, so in 50 g there will be  $50 \times \frac{20}{100} = 10$  g of iron.

- 2) Calculate the **percentage mass** of iron in **iron chloride**.

Percentage mass of iron =  $\frac{A_r \times \text{number of atoms of that element}}{M_r \text{ of the compound}} \times 100 = \frac{56}{56 + (2 \times 35.5)} \times 100 = 44.09\%$

- 3) Calculate the **mass** of **iron chloride** that contains 10 g of iron.

Iron chloride contains 44.09% iron by mass, so there will be 10 g of iron in  $10 \div \frac{44.09}{100} = 23$  g

So you need 23 g of iron chloride to provide the iron in 50 g of the mixture.

### Relative mass — when you go to church with your parents...

The best way to get to grips with all this stuff is by practising. Start by having a go at these questions...

- Q1 Calculate the relative formula mass ( $M_r$ ) of: a)  $\text{H}_2\text{O}$  b)  $\text{LiOH}$  c)  $\text{H}_2\text{SO}_4$  [3 marks]  
Q2 Calculate the percentage composition by mass of potassium in potassium hydroxide (KOH). [2 marks]

# The Mole

**Moles** can be pretty confusing. It's probably the word that puts people off. It's difficult to see the relevance of the word "mole" to anything but a small burrowing animal.

## "The Mole" is Simply the Name Given to an Amount of a Substance

- Just like "a million" is this many: 1 000 000; or "a billion" is this many: 1 000 000 000, so "**the Avogadro constant**" is this many: **602 000 000 000 000 000 000 000** or  **$6.02 \times 10^{23}$** . And that's **all** it is. Just a **number**.
- One mole** of any substance is just an **amount** of that substance that contains an **Avogadro number of particles** — so  $6.02 \times 10^{23}$  particles. The particles could be atoms, molecules, ions or electrons.
- The burning question, of course, is why is it such a silly long number like that, and with a 6 at the front?
- The answer is that the mass of that number of **atoms** or **molecules** of any substance is exactly the same number of grams as the **relative atomic mass** ( $A_r$ ) or **relative formula mass** ( $M_r$ ) of the element or compound.
- In other words, **one mole** of atoms or molecules of any substance will have **a mass in grams** equal to the **relative formula mass** ( $A_r$  or  $M_r$ ) for that substance. Here are some examples:



**Carbon** has an  $A_r$  of **12**.

So **one mole** of carbon weighs exactly **12 g**.

**Nitrogen gas**,  $N_2$ , has an  $M_r$  of **28** ( $2 \times 14$ ).

So **one mole** of  $N_2$  weighs exactly **28 g**.

**Carbon dioxide**,  $CO_2$ , has an  $M_r$  of **44** ( $12 + [2 \times 16]$ ).

So **one mole** of  $CO_2$  weighs exactly **44 g**.

- This means that 12 g of carbon, or 28 g of  $N_2$ , or 44 g of  $CO_2$ , all contain the **same number of particles**, namely **one mole** or  $6.023 \times 10^{23}$  atoms or molecules.

## Nice Formula to Find the Number of Moles in a Given Mass:

$$\text{Number of moles} = \frac{\text{mass in g (of an element or compound)}}{M_r \text{ (of the element or compound)}}$$

### EXAMPLE:

How many moles are there in 66 g of carbon dioxide ( $CO_2$ )?

- Calculate the  $M_r$  of carbon dioxide.
- Use the formula above to find out how many moles there are.

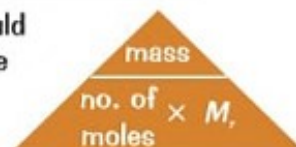
$M_r$  of  $CO_2 = 12 + (16 \times 2) = 44$

No. of moles = Mass (g)  $\div$   $M_r = 66 \div 44 = 1.5$  mol

Easy Peasy.

'mol' is the symbol for the unit 'moles'.

You can **rearrange** the equation above using this handy **formula triangle**. You could use it to find the **mass** of a known number of moles of a substance, or to find the  $M_r$  of a substance from a known mass and number of moles. Just **cover up** the thing you want to find with your finger and write down what's left showing.



### EXAMPLE:

What mass of carbon is there in 4 moles of carbon dioxide?

There are 4 moles of carbon in 4 moles of  $CO_2$ .

Cover up 'mass' in the formula triangle. That leaves you with 'no. of moles  $\times M_r$ '.

So the mass of 4 moles of carbon =  $4 \times 12 = 48$  g

## What do moles have for pudding? Jam moly-poly...

Calculations involving moles can send some people into a spin. Don't be one of those people — there's really no need to freak out about moles. Go back over this page until you've got your head round it all.

Q1 Calculate the number of moles in 90 g of water ( $H_2O$ ).  $A_r(O) = 16$ ,  $A_r(H) = 1$ .

[2 marks]

Q2 Calculate the mass of 0.20 mol of potassium bromide (KBr).  $A_r(K) = 39$ ,  $A_r(Br) = 80$ .

[2 marks]



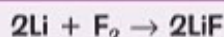
# Conservation of Mass

You've probably realised by now that you can't magic stuff out of thin air, and you can't make it magically disappear, either. This fact is pretty useful for working out the [amounts of substances](#) in chemical reactions.

## In a Chemical Reaction, Mass is Always Conserved

- 1) During a chemical reaction [no atoms are destroyed](#) and [no atoms are created](#).
- 2) This means there are the [same number and types of atoms](#) on each side of a reaction equation.
- 3) Because of this, no mass is lost or gained — we say that mass is [conserved](#) during a reaction.

E.g.



In this reaction, there are [2 lithium atoms](#) and [2 fluorine atoms](#) on [each side](#) of the equation.

- 4) By [adding up](#) the relative formula masses of the substances on each side of a [balanced symbol equation](#), you can see that mass is conserved. The total  $M_r$  of all the reactants [equals](#) the total  $M_r$  of the products.

There's more about balanced symbol equations on p.15.

### EXAMPLE:

Show that mass is conserved in this reaction:  $2\text{Li} + \text{F}_2 \rightarrow 2\text{LiF}$ .

- 1) Add up the relative formula masses on the [left-hand side](#) of the equation.  
 $2 \times M_r(\text{Li}) + 2 \times M_r(\text{F}) = (2 \times 7) + (2 \times 19) = 14 + 38 = 52$
- 2) Add up the relative formula masses on the [right-hand side](#) of the equation.  
 $2 \times M_r(\text{LiF}) = 2 \times (7 + 19) = 2 \times 26 = 52$

The total  $M_r$  on the left hand side of the equation is equal to the total  $M_r$  on the right hand side, so mass is conserved.

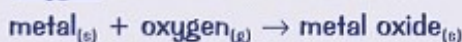
## If the Mass Seems to Change, There's Usually a Gas Involved

In some experiments, you might observe a [change of mass](#) of an [unsealed reaction vessel](#) during a reaction. There are usually two explanations for this:

**Explanation 1:** If the mass [increases](#), it's probably because one of the [reactants](#) is a [gas](#) that's found in air (e.g. oxygen) and all the products are solids, liquids or aqueous.

- [Before](#) the reaction, the gas is floating around in the air. It's there, but it's not contained in the reaction vessel, so you [can't](#) account for its [mass](#).
- When the gas [reacts](#) to form part of the [product](#), it becomes contained inside the reaction vessel — so the [total mass](#) of the stuff [inside](#) the reaction vessel [increases](#).

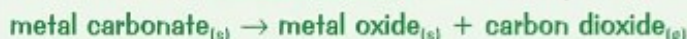
For example, when a [metal](#) reacts with [oxygen](#) in an unsealed container, the mass of the container [increases](#). The mass of the [metal oxide](#) produced [equals](#) the total mass of the [metal](#) and the [oxygen](#) that reacted from the air.



**Explanation 2:** If the mass [decreases](#), it's probably because one of the [products](#) is a [gas](#) and all the reactants are solids, liquids or aqueous.

- [Before](#) the reaction, all the reactants are contained in the reaction vessel.
- If the vessel [isn't enclosed](#), then the gas can [escape](#) from the reaction vessel as it's formed. It's no longer contained in the reaction vessel, so you [can't](#) account for its [mass](#) — the total mass of the stuff [inside](#) the reaction vessel [decreases](#).

For example, when a [metal carbonate](#) thermally decomposes to form a [metal oxide](#) and [carbon dioxide gas](#), the mass of the reaction vessel will [decrease](#) if it isn't sealed. But in reality, the mass of the [metal oxide](#) and the [carbon dioxide](#) produced will [equal](#) the mass of the [metal carbonate](#) that decomposed.

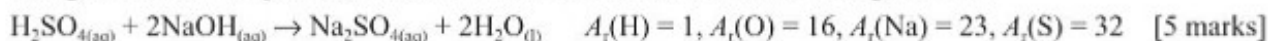


Remember from the particle model on page 36 that a gas will expand to fill any container it's in. So if the reaction vessel isn't sealed, the gas expands out from the vessel, and escapes into the air around.

## Leaving all the potatoes on your plate — that's mass conservation...

Never, ever forget that, in a reaction, the total mass of reactants is the same as the total mass of products.

Q1 Using the balanced equation, show that mass is conserved in the following reaction:



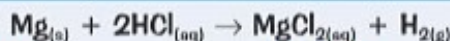
# The Mole and Equations

This is the moment where the 'number of moles = mass ÷ *M<sub>r</sub>*' equation from page 42 comes into it's own.

## You Can Use Moles to Calculate Masses in Reactions

Remember those balanced equations back on page 15? Well, the big numbers in front of the chemical formulas of the reactants and products tell you how many moles of each substance takes part or is formed during the reaction.

For example:



In this reaction, 1 mole of magnesium and 2 moles of hydrochloric acid react together to form 1 mole of magnesium chloride and 1 mole of hydrogen gas.

The little numbers tell you how many atoms of each element there are in each of the substances.

## You Can Balance Equations Using Reacting Masses

If you know the masses of the reactants and products that took part in a reaction, you can work out the balanced symbol equation for the reaction. Here are the steps you should take:

- 1) Divide the mass of each substance by its relative formula mass to find the number of moles.
- 2) Divide the number of moles of each substance by the smallest number of moles in the reaction.
- 3) If any of the numbers aren't whole numbers, multiply all the numbers by the same amount so that they all become whole numbers.
- 4) Write the balanced symbol equation for the reaction by putting these numbers in front of the chemical formulas.

### EXAMPLE:

8.1 g of zinc oxide (ZnO) reacts completely with 0.60 g of carbon to form 2.2 g of carbon dioxide and 6.5 g of zinc. Write a balanced symbol equation for this reaction.  
 $A_r(\text{C}) = 12$ ,  $A_r(\text{O}) = 16$ ,  $A_r(\text{Zn}) = 65$ .

- 1) Work out *M<sub>r</sub>* for each of the substances in the reaction:

$$\text{ZnO: } 65 + 16 = 81$$

$$\text{C: } 12$$

$$\text{CO}_2: 12 + (2 \times 16) = 44$$

$$\text{Zn: } 65$$

- 2) Divide the mass of each substance by its *M<sub>r</sub>* to calculate how many moles of each substance reacted or were produced:

$$\text{ZnO: } \frac{8.1}{81} = 0.10 \text{ mol}$$

$$\text{C: } \frac{0.60}{12} = 0.050 \text{ mol}$$

$$\text{CO}_2: \frac{2.2}{44} = 0.050 \text{ mol}$$

$$\text{Zn: } \frac{6.5}{65} = 0.10$$

- 3) Divide by the smallest number of moles, which is 0.050:

$$\text{ZnO: } \frac{0.10}{0.050} = 2.0$$

$$\text{C: } \frac{0.050}{0.050} = 1.0$$

$$\text{CO}_2: \frac{0.050}{0.050} = 1.0$$

$$\text{Zn: } \frac{0.10}{0.050} = 2.0$$

These numbers give the ratio of the amounts of each substance in the reaction equation.

- 4) The numbers are all whole numbers, so you can write out the balanced symbol equation straight away:



## Where do moles live? Edinburrow...

The calculations on this page have lots of steps, so the best way to learn how to do them is by practising. Luckily, there are some questions below to get you started. Don't say I don't spoil you. Better get cracking...

Q1 84 g of  $\text{N}_2$  reacts completely with 18 g of  $\text{H}_2$  to produce 102 g of  $\text{NH}_3$ .

$$M_r(\text{N}_2) = 28, M_r(\text{H}_2) = 2, M_r(\text{NH}_3) = 17.$$

- a) Calculate how many moles of each substance reacted or was produced.

[3 marks]

- b) Use your answer to part a) to write a balanced symbol equation for this reaction.

[2 marks]

# Limiting Reactants

Reactions don't go on forever — you need stuff in the reaction flask that can react. If one reactant gets **completely used up** in a reaction before the rest, then the reaction will **stop**. That reactant's called **limiting**.

## Reactions Stop When One Reactant is Used Up

When some **magnesium carbonate** ( $\text{MgCO}_3$ ) is placed into a beaker of **hydrochloric acid**, you can tell a **reaction** is taking place because you see lots of **bubbles of gas** being given off. After a while, the amount of fizzing **slows down** and the reaction eventually **stops**...



- 1) The reaction stops when all of one of the reactants is **used up**. Any other reactants are in **excess**. They're usually added in excess to **make sure** that the other reactant is used up.
- 2) The reactant that's **used up** in a reaction is called the **limiting reactant** (because it limits the amount of product that's formed).
- 3) The amount of product formed is **directly proportional** to the amount of **limiting reactant**. For example, if you **halve** the amount of limiting reactant the amount of product formed will also **halve**. If you **double** the amount of limiting reactant the amount of product will **double** (as long as it is still the limiting reactant).
- 4) This is because if you add **more reactant** there will be **more reactant particles** to take part in the reaction, which means **more product particles**.

## The Amount of Product Depends on the Limiting Reactant

You can calculate the **mass of a product** formed in a reaction by using the **mass** of the **limiting reactant** and the **balanced reaction equation**.

You could also use this method to find the mass of a reactant needed to produce a known mass of a product.

- 1) Write out the **balanced equation**.
- 2) **Work out relative formula masses** ( $M_r$ ) of the reactant and product you want.
- 3) Find out **how many moles** there are of the substance you **know** the mass of.
- 4) Use the balanced equation to work out **how many moles** there'll be of the **other** substance. In this case, that's how many moles of **product** will be made of this many moles of **reactant**.
- 5) Use the number of moles to calculate the **mass**.

### EXAMPLE:

Calculate the mass of aluminium oxide formed when 135 g of aluminium is burned in air.

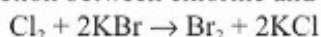
- 1) Write out the **balanced equation**:  $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$
- 2) Calculate the relative formula masses:  $\text{Al: } 27 \quad \text{Al}_2\text{O}_3: (2 \times 27) + (3 \times 16) = 102$
- 3) **Calculate the number of moles** of aluminium in 135 g:  $\text{Moles} = \frac{\text{mass}}{M_r} = \frac{135}{27} = 5$  You don't have to find  $M_r$  of oxygen because it's in excess.
- 4) Look at the **ratio** of moles in the equation:   
 4 moles of Al react to produce 2 moles of  $\text{Al}_2\text{O}_3$  — half the number of moles are produced.   
 So 5 moles of Al will react to produce 2.5 moles of  $\text{Al}_2\text{O}_3$    
If the question asked for the number of moles of aluminium oxide formed, you'd stop here.
- 5) **Calculate the mass** of 2.5 moles of aluminium oxide:  $\text{mass} = \text{moles} \times M_r = 2.5 \times 102 = 255 \text{ g}$

The mass of product (in this case aluminium oxide) is called the **yield** of a reaction. Masses you calculate in this way are called **theoretical yields**. In practice you never get 100% of the yield, so the amount of product you get will be **less** than you calculated.

## 'A Rush of Mud to the Head' — my favourite album by Moldplay...

I've said it before, I'll say it again — practice makes perfect. So before you get distracted by a cute picture of a kitten, or wander off to have a cup of tea, have a go at the question below.

Q1 The balanced equation for the reaction between chlorine and potassium bromide is:



Calculate the mass of potassium chloride produced when 23.8 g of potassium bromide reacts in an excess of chlorine.  $A_r(\text{K}) = 39$ ,  $A_r(\text{Br}) = 80$ ,  $A_r(\text{Cl}) = 35.5$ .

[4 marks]

# Gases and Solutions

Time to find out how to calculate amounts of gases from volumes and amounts of solutes from concentrations.

## One Mole of Any Gas Occupies 24 dm<sup>3</sup> at 20 °C

At the same temperature and pressure, equal numbers of moles of any gas will occupy the same volume.

At room temperature and pressure (r.t.p. = 20 °C and 1 atm) one mole of any gas occupies 24 dm<sup>3</sup>.

You can use this formula to find the volume of a known mass of any gas at r.t.p.:

$$\text{Volume of gas} = \frac{\text{Mass of gas}}{M_r \text{ of gas}} \times 24$$

in g

in dm<sup>3</sup>

### EXAMPLE:

What's the volume of 319.5 g of chlorine at r.t.p?  $\text{Volume} = \frac{\text{Mass}}{M_r} \times 24 = \frac{319.5}{71} \times 24 = 108 \text{ dm}^3$

## You Can Calculate Volumes of Gases in Reactions

For reactions between gases, you can use the volume of one gas to find the volume of another.

### EXAMPLE:

How much carbon dioxide is formed when 30 dm<sup>3</sup> of oxygen reacts with carbon monoxide?  $2\text{CO}_{(g)} + \text{O}_{2(g)} \rightarrow 2\text{CO}_{2(g)}$

1 mole of O<sub>2</sub> → 2 moles of CO<sub>2</sub> so 1 volume of O<sub>2</sub> → 2 volumes of CO<sub>2</sub>

So 30 dm<sup>3</sup> of O<sub>2</sub> → (2 × 30 dm<sup>3</sup>) = 60 dm<sup>3</sup> of CO<sub>2</sub>



## Concentration is a Measure of How Crowded Things Are

- 1) Lots of reactions in chemistry take place between substances that are dissolved in a solution. The amount of a substance (e.g. the mass or the number of moles) in a certain volume of a solution is called its concentration.
- 2) The more solute (the substance that's dissolved) there is in a given volume, the more concentrated the solution.
- 3) One way to measure the concentration of a solution is by calculating the mass of a substance in a given volume of solution. The units will be units of mass/units of volume. Here's how to calculate the concentration of a solution in g/dm<sup>3</sup>:

$$\text{concentration} = \frac{\text{mass of solute}}{\text{volume of solvent}}$$

in g

in g/dm<sup>3</sup>

in dm<sup>3</sup>

4) You can calculate the mass of a solute in solution by rearranging this formula to: mass = conc. × volume.

- 5) Concentration can also be given in mol/dm<sup>3</sup>:

$$\text{concentration} = \frac{\text{number of moles of solute}}{\text{volume of solvent}}$$

in mol

in mol/dm<sup>3</sup>

in dm<sup>3</sup>

### EXAMPLES:

- 1) What's the concentration in g/dm<sup>3</sup> of a solution of sodium chloride where 30 g of sodium chloride is dissolved in 0.2 dm<sup>3</sup> of water?

$$\text{concentration} = \frac{30}{0.2} = 150 \text{ g/dm}^3$$

$$1 \text{ dm}^3 = 1000 \text{ cm}^3$$

- 2) What's the concentration, in mol/dm<sup>3</sup>, of a solution with 2 moles of salt in 500 cm<sup>3</sup>?

• Convert the volume to dm<sup>3</sup> by dividing by 1000:

$$500 \text{ cm}^3 \div 1000 = 0.5 \text{ dm}^3$$

• Now you've got the number of moles and the volume in the right units, just stick them in the formula:

$$\text{Concentration} = \frac{2}{0.5} = 4 \text{ mol/dm}^3$$



Gavin wasn't great at concentration.

## CGP Revision Guides — not from concentrate...

It's pretty handy that any gas, yes I mean ANY gas, occupies the same volume under the same conditions.

Q1 Calculate the volume of 3.5 g of nitrogen (N<sub>2</sub>) gas at room temperature and pressure.

[2 marks]

# Concentration Calculations

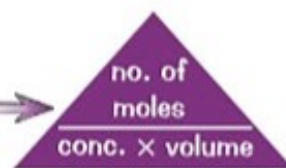
Ever found yourself wondering how to find the **concentration** of a solution from the concentration of a **different** solution and a couple of **volume** values? No? Oh well, here's how to do it anyway.

## You Might Be Asked to Calculate the Concentration

**Titration** experiments let you find the **volumes** needed for two solutions to **react together completely** (see p.52). If you know the **concentration** of one of the solutions, you can use the volumes from the titration experiment, along with the **reaction equation**, to **find** the concentration of the other solution.

### Find the Concentration in $\text{mol/dm}^3$

You might remember the formula for working out the **concentration** of a substance in  $\text{mol/dm}^3$  from page 46. Well, here it is in a handy **formula triangle**. It's dead useful in these calculations.



#### EXAMPLE:

A student started with  $30.0 \text{ cm}^3$  of sulfuric acid ( $\text{H}_2\text{SO}_4$ ) of unknown concentration in a flask. She found that it took an average of  $25.0 \text{ cm}^3$  of  $0.100 \text{ mol/dm}^3$  sodium hydroxide ( $\text{NaOH}$ ) to neutralise the sulfuric acid. Find the concentration of the acid in  $\text{mol/dm}^3$ .

The balanced symbol equation for the reaction is:  $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$

- 1) Work out how many **moles** of the "known" substance you have using the formula:  $\text{no. of moles} = \text{conc.} \times \text{volume}$ .  
 $0.100 \text{ mol/dm}^3 \times (25.0 / 1000) \text{ dm}^3 = 0.00250 \text{ moles of NaOH}$
- 2) Use the reaction equation to work out how many moles of the "unknown" stuff you must have had.  
 Using the equation, you can see that two moles of sodium hydroxide reacts with one mole of sulfuric acid. So  $0.00250 \text{ moles of NaOH}$  must have reacted with  $0.00250 \div 2 = 0.00125 \text{ moles of H}_2\text{SO}_4$ .
- 3) Work out the concentration of the "unknown" stuff.  
 $\text{Concentration} = \text{number of moles} \div \text{volume}$   
 $= 0.00125 \text{ mol} \div (30.0 \div 1000) \text{ dm}^3 = 0.041666... \text{ mol/dm}^3 = 0.0417 \text{ mol/dm}^3$

Make sure the volume is in  $\text{dm}^3$ .

Don't round your answer until right at the end.

Remember — all measurements have some uncertainty to them. Titration experiments will often be repeated, and then the average (mean) of these repeated measurements will be calculated. The range of the results can also be found (see p.7), and can be used to give you an idea of how uncertain the mean value is (see p.11).

### Converting $\text{mol/dm}^3$ to $\text{g/dm}^3$

- 1) To find the concentration in  $\text{g/dm}^3$ , start by finding the concentration in  $\text{mol/dm}^3$  using the steps above.
- 2) Then, **convert** the concentration in  $\text{mol/dm}^3$  to  $\text{g/dm}^3$  using the equation **mass = moles  $\times$   $M_r$** .

You can also convert from  $\text{g/dm}^3$  to  $\text{mol/dm}^3$  by rearranging this equation to:  $\text{moles} = \text{mass} \div M_r$

#### EXAMPLE:

What's the concentration, in  $\text{g/dm}^3$ , of the sulfuric acid solution in the example above?

- 1) Work out the **relative formula mass** for the acid.  
 $M_r(\text{H}_2\text{SO}_4) = (2 \times 1) + 32 + (4 \times 16) = 98$
- 2) Convert the concentration in **moles** (that you've already worked out) into concentration in **grams**. So, in  $1 \text{ dm}^3$ :  
 $\text{Mass in grams} = \text{moles} \times \text{relative formula mass} = 0.041666... \times 98 = 4.08333... \text{ g}$   
 So the concentration in  $\text{g/dm}^3 = 4.08 \text{ g/dm}^3$

## Time to convert my concentration from cat videos to chemistry...

These calculations look pretty scary. But if you get enough practice, then the fear will evaporate and you can tackle them with a smile on your face and a spring in your step. Better get cracking...

- Q1 A student found that  $22.5 \text{ cm}^3$  of  $0.150 \text{ mol/dm}^3$  potassium hydroxide ( $\text{KOH}$ ) was needed to neutralise  $25.0 \text{ cm}^3$  of nitric acid ( $\text{HNO}_3$ ). The balanced equation for the reaction is:  $\text{HNO}_3 + \text{KOH} \rightarrow \text{KNO}_3 + \text{H}_2\text{O}$   
 Calculate the concentration of nitric acid in  $\text{mol/dm}^3$ . [3 marks]

# Atom Economy

It's important in industrial reactions that as much of the **reactants** as possible get turned into **useful products**. This depends on the **atom economy** and the **percentage yield** (see next page) of the reaction.

## "Atom Economy" — % of Reactants Forming Useful Products

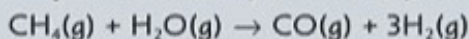
- 1) A lot of reactions make **more than one product**. Some of them will be **useful**, but others will be **waste**.
- 2) The **atom economy** (or atom utilisation) of a reaction tells you how much of the **mass** of the reactants is wasted when manufacturing a chemical and how much ends up as useful products. **Learn** the equation:

$$\text{atom economy} = \frac{\text{relative formula mass of desired products}}{\text{relative formula mass of all reactants}} \times 100$$

- 3) **100%** atom economy means that **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 following reaction to produce hydrogen gas:



- 1) **Identify** the desired product — that's the **hydrogen gas**.

- 2) Work out the ***M<sub>r</sub>*** of **all** the reactants:

$$M_r(\text{CH}_4) = 12 + (4 \times 1) = 16$$

$$M_r(\text{H}_2\text{O}) = (2 \times 1) + 16 = 18$$

$$16 + 18 = 34$$

- 3) Work out the **total *M<sub>r</sub>*** of the **desired product**:

$$3 \times M_r(\text{H}_2) = 3 \times (2 \times 1) = 6$$

- 4) Use the **formula** to calculate the atom economy:

$$\frac{6}{34} \times 100 = 17.6\%$$

So in this reaction, over 80% of the starting materials are wasted.



## High Atom Economy is Better for Profits and the Environment

- 1) Pretty obviously, if you're making **lots of waste**, that's a **problem**.
- 2) Reactions with low atom economy **use up resources** very quickly. At the same time, they make lots of **waste** materials that have to be **disposed** of somehow. That tends to make these reactions **unsustainable** — the raw materials will run out and the waste has to go somewhere.
- 3) For the same reasons, low atom economy reactions aren't usually **profitable**. Raw materials are often **expensive to buy**, and waste products can be expensive to **remove** and dispose of **responsibly**.
- 4) The best way around the problem is to find a **use** for the waste products rather than just **throwing them away**. There's often **more than one way** to make the product you want, so the trick is to come up with a reaction that gives **useful "by-products"** rather than useless ones.
- 5) The reactions with the **highest** atom economy are the ones that only have **one product**. Those reactions have an atom economy of **100%**.  
The more products there are, the lower the atom economy is likely to be.

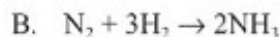
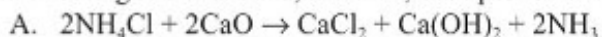
Atom economy isn't the only factor to be considered when choosing which reaction to use to make a certain product. Things such as the **yield** (see next page), the **rate** of the reaction (see page 67) and the **position of equilibrium** for reversible reactions (see page 72) also need to be thought about.

A reaction with a low atom economy, that produces **useful by-products**, might also be used.

## Atom economy class — even less legroom than economy class...

You might get asked about an unfamiliar industrial reaction in the exam. Don't panic — whatever example they give you, the same stuff applies. Have a go at the question below, and you'll see what I mean...

Q1 The following two reactions, A and B, both produce  $\text{NH}_3$ :



$$A_r(\text{N}) = 14, A_r(\text{H}) = 1, A_r(\text{Cl}) = 35.5, A_r(\text{Ca}) = 40, A_r(\text{O}) = 16$$

- a) Which reaction, A or B, has the higher atom economy? Explain your answer.

[1 mark]

- b) Calculate the atom economy for the production of  $\text{NH}_3$  by reaction A.

[5 marks]

# Percentage Yield

Percentage yield tells you about the overall success of an experiment. It compares what you think you should get (theoretical yield) with what you get in practice (actual yield).

## Percentage Yield Compares Actual and Theoretical Yield

The amount of product you get is known as the yield. The more reactants you start with, the higher the actual yield will be — that's pretty obvious. But the percentage yield doesn't depend on the amount of reactants you started with — it's a percentage.

- 1) Percentage yield is given by the formula:

$$\text{percentage yield} = \frac{\text{mass of product actually made} \times 100}{\text{maximum theoretical mass of product}}$$

in g

in g

This is just the maximum theoretical yield. It can be calculated from the balanced reaction equation (see p.45).

- 2) Percentage yield is always somewhere between 0 and 100%. 100% yield means that you got all the product you expected to get. 0% yield means that no reactants were converted into product, i.e. no product at all was made.
- 3) Industrial processes should have as high a percentage yield as possible to reduce waste and reduce costs.

If you know the percentage yield and the maximum theoretical yield of the reaction, you can rearrange this equation to work out what the mass of product actually made would be.

## Yields are Always Less Than 100%

In real life, you never get a 100% yield. Some product or reactant always gets lost along the way — and that goes for big industrial processes as well as school lab experiments. How this happens depends on what sort of reaction it is and what apparatus is being used. Lots of things can go wrong, but three common problems are:

### 1) Not All Reactants React to Make a Product

In reversible reactions (see page 72), the products can turn back into reactants, so the yield will never be 100%.

For example, in the Haber process, at the same time as the reaction  $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$  is taking place, the reverse reaction  $2\text{NH}_3 \rightarrow \text{N}_2 + 3\text{H}_2$  is also happening.

This means the reaction  $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$  never goes to completion (the reactants don't all get used up).



### 2) There Might be Side Reactions

The reactants sometimes react differently to how you expect. They might react with gases in the air, or impurities in the reaction mixture, so they end up forming extra products other than the ones you want.

### 3) You Lose Some Product When You Separate It From the Reaction Mixture

When you filter a liquid to remove solid particles, you nearly always lose a bit of liquid or a bit of solid.

- If you want to keep the liquid, you'll lose the bit that remains with the solid and filter paper (as they always stay a bit wet).
  - If you want to keep the solid, some of it'll get left behind when you scrape it off the filter paper.
- You'll also lose a bit of material when you transfer it from one container to another — even if you manage not to spill it. Some of it always gets left behind on the inside surface of the old container.

## A yield of 100% in a reaction is pretty much impossible...

...but 100% in an exam is not. You just need luck on your side and to have learned your stuff really well.

- Q1 A student carries out the following reaction:  $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$ . She calculates that her maximum theoretical yield of zinc chloride is 2.72 g. The actual yield after purification is 2.31 g. Calculate the percentage yield of zinc chloride. [1 mark]

## Revision Questions for Topic 3

That's almost it for [Topic 3](#) — but before you go there are some questions to test [how much you really know](#).

- Try these questions and [tick off each one](#) when you [get it right](#).
- When you've done [all the questions](#) under a heading and are [completely happy](#) with it, tick it off.

### Moles and Equations (p.41-45) ☐

- 1) How do you calculate the relative formula mass,  $M_r$ , of a substance? ☐
- 2) State the value of the Avogadro constant. ☐
- 3) What is the formula that relates the number of moles of a substance to its mass and  $M_r$ ? ☐
- 4) What does conservation of mass mean? ☐
- 5) Suggest why the mass of a reaction vessel might decrease during a reaction. ☐
- 6) How can you determine the number of moles of each substance that would react together from the balanced reaction equation? ☐
- 7) Explain what is meant by the term 'limiting reactant'. ☐

### Volumes of Gases and Concentrations of Solutions (p.46-47) ☐

- 8) What volume does one mole of any gas occupy at room temperature and pressure? ☐
- 9) State the values of room temperature and pressure. ☐
- 10) What is concentration? ☐
- 11) Give the equation for working out the concentration of a solution in  $\text{g/dm}^3$ . ☐
- 12) Give the equation for working out the concentration of a solution in  $\text{mol/dm}^3$ . ☐
- 13) Describe how you would convert a concentration in  $\text{mol/dm}^3$  to  $\text{g/dm}^3$ . ☐

### Atom Economy and Percentage Yield (p.48-49) ☐

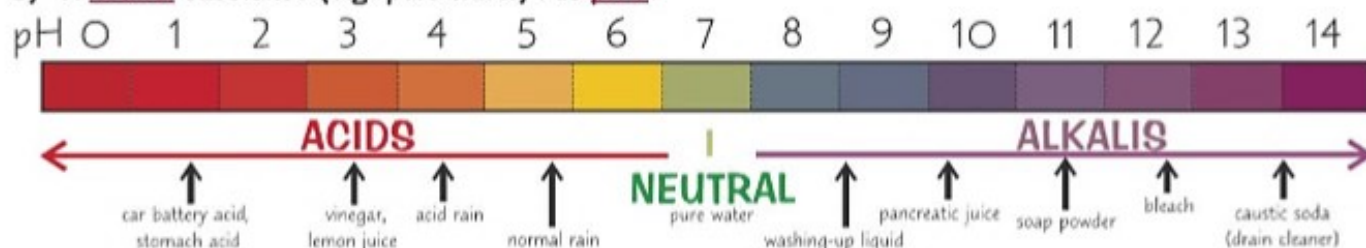
- 14) Give the equation for calculating the atom economy of a reaction. ☐
- 15) Give three reasons why it's better to use reactions that have a high atom economy. ☐
- 16) What's the atom economy of a reaction that only produces one product? ☐
- 17) What information do you need to find the theoretical yield of a reaction? ☐
- 18) Give the equation for calculating the percentage yield of a reaction. ☐
- 19) Give two reasons why it's better to use reactions with high percentage yields. ☐
- 20) Give three reasons why percentage yield is always less than 100%. ☐

# Acids and Bases

Testing the pH of a solution means using an indicator — and that means pretty **colours**...

## The pH Scale Goes From 0 to 14

- 1) The **pH scale** is a measure of how **acidic** or **alkaline** a solution is.
- 2) The **lower** the pH of a solution, the **more acidic** it is.  
The **higher** the pH of a solution, the more **alkaline** it is.
- 3) A **neutral** substance (e.g. pure water) has **pH 7**.



## You Can Measure the pH of a Solution

- 1) An **indicator** is a **dye** that **changes colour** depending on whether it's **above or below a certain pH**. Some indicators contain a **mixture of dyes** that means they **gradually change colour** over a broad range of pH. These are called **wide range indicators** and they're useful for **estimating** the pH of a solution. For example, **universal indicator** gives the colours shown above.
- 2) A **pH probe** attached to a **pH meter** can also be used to measure pH **electronically**. The probe is placed in the solution you are measuring and the pH is given on a digital display as a **numerical value**, meaning it's more accurate than an indicator.



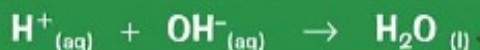
## Acids and Bases Neutralise Each Other

- 1) An **acid** is a substance that forms **aqueous solutions** with a pH of **less than 7**. Acids form **H<sup>+</sup> ions** in **water**.
- 2) A **base** is a substance with a pH **greater than 7**.
- 3) An **alkali** is a base that **dissolves in water** to form a solution with a pH **greater than 7**. Alkalis form **OH<sup>-</sup> ions** in **water**.

The reaction between acids and bases is called **neutralisation**:



Neutralisation between acids and alkalis can be seen in terms of **H<sup>+</sup>** and **OH<sup>-</sup> ions** like this:



Hydrogen (H<sup>+</sup>) ions react with hydroxide (OH<sup>-</sup>) ions to produce water.

When an acid neutralises a base (or vice versa), the **products** are **neutral**, i.e. they have a **pH of 7**. An indicator can be used to show that a neutralisation reaction is over.

**Neutralisation** reactions of **strong** acids and alkalis can be used to calculate the **concentration** of an acid or alkali by **titration** — there is more about this technique on the next page.

## This page should have all bases covered...

pHew, you finished the page... This stuff isn't too bad really, and pH is worth knowing about — it's important to the chemistry in our bodies. For example, here's an interesting(ish) fact — your skin is slightly acidic (pH 5.5).

- Q1 A student uses universal indicator to test the pH of some lemon juice.  
What colour would you expect the indicator to turn? [1 mark]
- Q2 The pH of an unknown solution is found to be 8. Is the solution acidic or alkaline? [1 mark]

## PRACTICAL

## Titrations

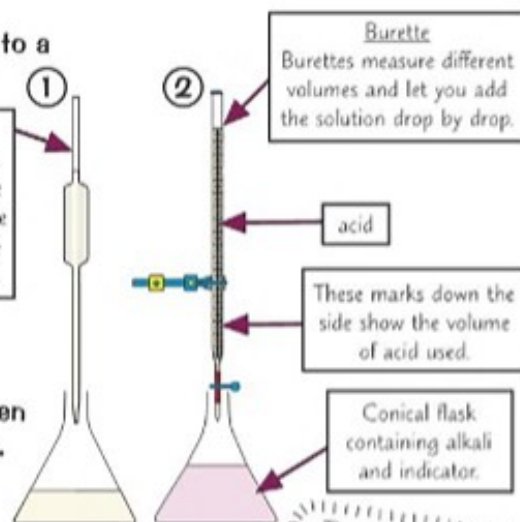
Titrations are a method of analysing the concentrations of solutions. They're pretty important. Some people don't think they're the most exciting game in town. But I secretly enjoy them, now I've got the hang of them.

### Titrations are Used to Find Out Concentrations

**Titrations** allow you to find out exactly how much acid is needed to neutralise a quantity of alkali — or vice versa. You can then use this data to work out the concentration of the acid or alkali (see page 47). For example:

- 1) Say you want to find out the concentration of some alkali. Using a pipette and pipette filler, add a set volume of the alkali to a conical flask. Add two or three drops of indicator too.
- 2) Use a funnel to fill a burette with some acid of known concentration. Make sure you do this BELOW EYE LEVEL — you don't want to be looking up if some acid spills over. (You should wear safety glasses too.) Record the initial volume of the acid in the burette.
- 3) Using the burette, add the acid to the alkali a bit at a time — giving the conical flask a regular swirl. Go especially slowly when you think the end-point (colour change) is about to be reached.
- 4) The indicator changes colour when all the alkali has been neutralised, e.g. phenolphthalein is pink in alkaline conditions, but colourless in acidic conditions.
- 5) Record the final volume of acid in the burette, and use it, along with the initial reading, to calculate the volume of acid used to neutralise the alkali.

**Pipette**  
Pipettes measure only one volume of solution. Fill the pipette to about 3 cm above the line, then carefully drop the level down to the line.



You can also do titrations the other way round — adding alkali to acid.

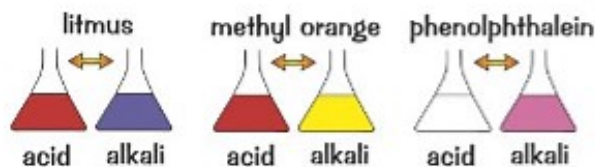
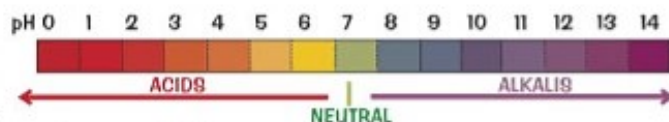
There's more about using burettes and pipettes on page 107.

- To increase the accuracy of your titration and to spot any anomalous results, you need several consistent readings.
- The first titration you do should be a rough titration to get an approximate idea of where the solution changes colour (the end-point).
- You then need to repeat the whole thing a few times, making sure you get (pretty much) the same answer each time (within 0.10 cm<sup>3</sup>).
- Finally, calculate a mean of your results, ignoring any anomalous results.

Anomalous results are ones that don't fit in with the rest (see page 6).

### Use Single Indicators for Titrations

- 1) Universal indicator is used to estimate the pH of a solution because it can turn a variety of colours. Each colour indicates a narrow range of pH values.
- 2) It's made from a mixture of different indicators. The colour gradually changes from red in acidic solutions to violet in alkaline solutions.
- 3) But during a titration between an alkali and an acid, you want to see a sudden colour change, at the end-point.
- 4) So you need to use a single indicator, such as litmus (blue in alkalis and red in acids), phenolphthalein (pink in alkalis and colourless in acids), or methyl orange (yellow in alkalis and red in acids).



### How do you get lean molecules? Feed them tight rations...

Titrations aren't too tricky really — you just need to be as sure as you can be that your results are accurate, which means going slowly near the end-point and then repeating the whole process.

- Q1 How do you determine when the end-point of a titration has been reached? [1 mark]
- Q2 Give two reasons why you should repeat a titration until you have consistent readings. [2 marks]

# Strong Acids and Weak Acids

I like strong acids, and I also like weak acids. But which is better? There's only one way to find out...

## Acids Produce Protons in Water

The thing about acids is that they ionise in aqueous solution — they produce hydrogen ions,  $H^+$ .

For example:

An  $H^+$  ion is just a proton.

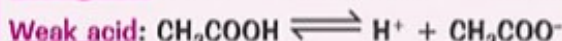
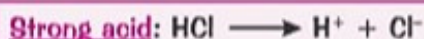


These acids don't produce hydrogen ions until they meet water. So, for example, hydrogen chloride gas isn't acidic.

## Acids Can be Strong or Weak

- 1) Strong acids (e.g. sulfuric, hydrochloric and nitric acids) ionise completely in water. All acid particles dissociate to release  $H^+$  ions.
- 2) Weak acids (e.g. ethanoic, citric and carbonic acids) do not fully ionise in solution. Only a small proportion of acid particles dissociate to release  $H^+$  ions.
- 3) The ionisation of a weak acid is a reversible reaction, which sets up an equilibrium between the undissociated and dissociated acid. Since only a few of the acid particles release  $H^+$  ions, the position of equilibrium lies well to the left.
- 4) Reactions of acids involve the  $H^+$  ions reacting with other substances. If the concentration of  $H^+$  ions is higher, the rate of reaction will be faster, so strong acids will be more reactive than weak acids of the same concentration.

For more on equilibria turn to p.72.



## pH is a Measure of the Concentration of Hydrogen Ions

- 1) The pH of an acid or alkali is a measure of the concentration of  $H^+$  ions in the solution.
- 2) For every decrease of 1 on the pH scale, the concentration of  $H^+$  ions increases by a factor of 10. So, an acid that has a pH of 4 has 10 times the concentration of  $H^+$  ions of an acid that has a pH of 5. For a decrease of 2 on the pH scale, the concentration of  $H^+$  ions increases by a factor of 100. The general rule for this is:

Factor  $H^+$  ion concentration changes by  $= 10^{-x}$

X is the difference in pH. So if pH falls from 7 to 4 the difference is  $-3$ , and the factor the  $H^+$  ion concentration has increased by is  $10^{-(-3)} = 10^3$ .

- 3) So the pH of a strong acid is always less than the pH of a weaker acid if they have the same concentration.

## Don't Confuse Strong Acids with Concentrated Acids

- 1) Acid strength (i.e. strong or weak) tells you what proportion of the acid molecules ionise in water.
- 2) The concentration of an acid is different. Concentration measures how much acid there is in a certain volume of water. Concentration is basically how watered down your acid is.
- 3) The larger the amount of acid there is in a certain volume of liquid, the more concentrated the acid is.
- 4) So you can have a dilute (not very concentrated) but strong acid, or a concentrated but weak acid.
- 5) pH will decrease with increasing acid concentration regardless of whether it's a strong or weak acid.

Concentration describes the total number of dissolved acid molecules — not the number of molecules that are ionised to produce hydrogen ions at any given moment.

## Weak acid or strong acid? I know which goes better with chips...

Acids are acidic because of  $H^+$  ions. And strong acids are strong because they let go of all their  $H^+$  ions at the drop of a hat... Well, at the drop of a drop of water.

Q1 Name a strong acid. [1 mark]

Q2 A student added strong acid to a weakly acidic solution of pH 6. The pH of the new solution was found to be pH 3. By how many times did the concentration of  $H^+$  increase? [1 mark]

# Reactions of Acids

Remember neutralisation from page 51? Well, there's more stuff on [neutralisation reactions](#) coming up...

## Metal Oxides and Metal Hydroxides are Bases

- 1) Some [metal oxides](#) and [metal hydroxides](#) dissolve in [water](#). These soluble compounds are [alkalis](#). As you saw on page 51, alkalis react with acids in [neutralisation reactions](#).
- 2) Even [bases](#) that [won't dissolve](#) in water will still take part in neutralisation reactions with acids.
- 3) So, all [metal oxides](#) and [metal hydroxides](#) react with [acids](#) to form a [salt](#) and [water](#).

### Acid + Metal Oxide → Salt + Water

### Acid + Metal Hydroxide → Salt + Water

The salt that's produced depends upon the [acid](#) and the [metal ion](#) in the [oxide](#) or [hydroxide](#):

hydrochloric acid	+	copper oxide	→	copper chloride	+	water
2HCl	+	CuO	→	CuCl <sub>2</sub>	+	H <sub>2</sub> O
sulfuric acid	+	potassium hydroxide	→	potassium sulfate	+	water
H <sub>2</sub> SO <sub>4</sub>	+	2KOH	→	K <sub>2</sub> SO <sub>4</sub>	+	2H <sub>2</sub> O
nitric acid	+	sodium hydroxide	→	sodium nitrate	+	water
HNO <sub>3</sub>	+	NaOH	→	NaNO <sub>3</sub>	+	H <sub>2</sub> O

To work out the formula of an ionic compound, you need to balance the charges of the positive and negative ions so the overall charge of a compound is neutral. For more on ionic formulas, see p30.

## Acids and Metal Carbonates Produce Carbon Dioxide

Metal carbonates are also [bases](#). They react with acids to produce a salt, water and [carbon dioxide](#).

### Acid + Metal Carbonate → Salt + Water + Carbon Dioxide

hydrochloric acid	+	sodium carbonate	→	sodium chloride	+	water	+	carbon dioxide
2HCl	+	Na <sub>2</sub> CO <sub>3</sub>	→	2NaCl	+	H <sub>2</sub> O	+	CO <sub>2</sub>
sulfuric acid	+	calcium carbonate	→	calcium sulfate	+	water	+	carbon dioxide
H <sub>2</sub> SO <sub>4</sub>	+	CaCO <sub>3</sub>	→	CaSO <sub>4</sub>	+	H <sub>2</sub> O	+	CO <sub>2</sub>

## You can Make Soluble Salts Using an Insoluble Base

- 1) You need to pick the right [acid](#) and [insoluble base](#), such as an [insoluble metal oxide](#), [hydroxide](#) or [carbonate](#). E.g. if you want to make [copper chloride](#), you could mix [hydrochloric acid](#) and [copper oxide](#).  

$$\text{CuO (s)} + 2\text{HCl (aq)} \rightarrow \text{CuCl}_2 \text{ (aq)} + \text{H}_2\text{O (l)}$$
- 2) Gently [warm](#) the dilute acid using a [Bunsen burner](#), then turn off the Bunsen burner.
- 3) Add the [insoluble base](#) to the [acid](#) a bit at a time, until no more reacts (i.e. the base is in [excess](#)). You'll know when all the acid has been neutralised because, even after stirring, the excess solid will just [sink](#) to the bottom of the flask.
- 4) Then [filter](#) out the [excess](#) solid to get the salt solution (see p.17).
- 5) To get [pure, solid](#) crystals of the [salt](#), gently heat the solution using a [water bath](#) or an [electric heater](#) to evaporate some of the water (to make it more concentrated) and then stop heating it and leave the solution to cool. [Crystals](#) of the salt should form, which can be [filtered](#) out of the solution and then [dried](#). This is called [crystallisation](#) (p.17).

### PRACTICAL

You could also react the acid with a metal.



## AHHHHH so many reactions...

In the exam you could get asked to describe how you would go about making a pure, dry sample of a given soluble salt. Make sure you understand the method and what reactants to use.

- Q1 Calcium carbonate is added to hydrochloric acid.  
Write a word equation for the reaction that occurs.

[2 marks]

# The Reactivity Series

You can place [metals](#) in order of reactivity. This can be a lot more useful than it sounds, promise.

## The Reactivity Series — How Well a Metal Reacts

- 1) The [reactivity series](#) lists metals in [order](#) of their [reactivity](#) towards other substances.
- 2) For metals, their reactivity is determined by how [easily](#) they lose electrons — forming positive ions. The [higher](#) up the reactivity series a metal is, the more easily they form [positive ions](#). Make sure you learn this list: →
- 3) When metals [react](#) with [water](#) or [acid](#), they [lose](#) electrons and form positive ions. So, the [higher](#) a metal is in the reactivity series, the more [easily](#) it [reacts](#) with water or acid.
- 4) If you [compare](#) the relative reactivity of different metals with either an [acid](#) or [water](#) and put them in order from [most](#) reactive to the [least](#) reactive, the order you get is the reactivity series.

### The Reactivity Series

Potassium	K	Very Reactive
Sodium	Na	
Lithium	Li	
Calcium	Ca	
Magnesium	Mg	Fairly Reactive
Carbon	C	
Zinc	Zn	
Iron	Fe	
Hydrogen	H	Not very Reactive
Copper	Cu	

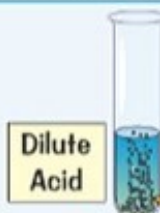
Carbon and hydrogen are non-metals but are often included in the reactivity series.

## How Metals React With Acids Tells You About Their Reactivity

Some metals react with acids to produce a [salt](#) and [hydrogen gas](#).



- 1) The [speed](#) of reaction is indicated by the [rate](#) at which the [bubbles](#) of hydrogen are given off.
- 2) The more [reactive](#) the metal, the [faster](#) the reaction will go. [Very](#) reactive metals like potassium, sodium, lithium and calcium react explosively, but less reactive metals such as magnesium, zinc and iron react less violently. In general, copper [won't](#) react with cold, dilute acids.



Magnesium reacts [vigorously](#) with [cold](#) dilute acids such as  $\text{HCl}_{(\text{aq})}$  or  $\text{H}_2\text{SO}_{4(\text{aq})}$  and produces [loads of bubbles](#).

**Dilute Acid**

**Magnesium**

$$\text{Mg}_{(\text{s})} + 2\text{HCl}_{(\text{aq})} \rightarrow \text{MgCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$$

$$\text{Mg}_{(\text{s})} + \text{H}_2\text{SO}_{4(\text{aq})} \rightarrow \text{MgSO}_{4(\text{aq})} + \text{H}_{2(\text{g})}$$



Both zinc and iron react [slowly](#) with dilute acids but more strongly if you heat them up.

**Dilute Acid**

**Zinc**

$$\text{Zn}_{(\text{s})} + 2\text{HCl}_{(\text{aq})} \rightarrow \text{ZnCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$$

$$\text{Zn}_{(\text{s})} + \text{H}_2\text{SO}_{4(\text{aq})} \rightarrow \text{ZnSO}_{4(\text{aq})} + \text{H}_{2(\text{g})}$$



**Dilute Acid**

**Iron**

$$\text{Fe}_{(\text{s})} + 2\text{HCl}_{(\text{aq})} \rightarrow \text{FeCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$$

$$\text{Fe}_{(\text{s})} + \text{H}_2\text{SO}_{4(\text{aq})} \rightarrow \text{FeSO}_{4(\text{aq})} + \text{H}_{2(\text{g})}$$

HCl reacts to form chloride salts,  $\text{H}_2\text{SO}_4$  reacts to form sulfate salts.

- 3) You can also investigate the reactivity of metals by measuring the [temperature change](#) of the reaction with an acid or water over a set time period. If you use the same [mass](#) and [surface area](#) of metal each time, then the [more reactive](#) the metal, the greater the temperature change should be.

## Metals Also React with Water

The [reactions](#) of metals with [water](#) also show the reactivity of metals.



For example, calcium:  $\text{Ca}_{(\text{s})} + 2\text{H}_2\text{O}_{(\text{l})} \rightarrow \text{Ca}(\text{OH})_{2(\text{aq})} + \text{H}_{2(\text{g})}$

- 1) The metals [potassium](#), [sodium](#), [lithium](#) and [calcium](#) will all react with water.
- 2) Less reactive metals like [zinc](#), [iron](#) and [copper](#) won't react with water.

You can see more on the reactions of Group 1 metals with water on page 24.

## I AM NOT HIGHLY REACTIVE — OK...

See, experiments aren't just for fun — they can give you a thrilling insight into the relative reactivities of elements.

Q1 Give the balanced equation, including state symbols, for the reaction of sodium and water. [3 marks]

# Separating Metals From Metal Oxides

Most metals are not found in the earth as pure lumps. Instead, you have to **extract** them from a compound meaning more work is required. Thanks for nothing nature...

## Metals Often Have to be Separated from their Oxides

- 1) Lots of common metals, like iron and aluminium, react with **oxygen** to form **oxides**. This process is an example of **oxidation**. These oxides are often the **ores** that the metals need to be extracted from.
- 2) A reaction that separates a metal from its oxide is called a **reduction reaction**.

### FORMATION OF METAL ORE:

**Oxidation = Gain of Oxygen**

E.g. magnesium is **oxidised** to make magnesium oxide.



### EXTRACTION OF METAL:

**Reduction = Loss of Oxygen**

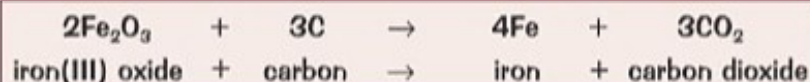
E.g. copper oxide is **reduced** to copper.



An ore is a type of rock that contains metal compounds.

## Some Metals can be Extracted by Reduction with Carbon

- 1) Some metals can be **extracted** from their ores chemically by **reduction** using **carbon**.
- 2) In this reaction, the ore is **reduced** as oxygen is **removed** from it, and carbon **gains** oxygen so is **oxidised**. For example...



- 3) The position of the metal in the **reactivity series** determines whether it can be extracted by **reduction** with **carbon**.

- Metals **higher than carbon** in the reactivity series have to be extracted using **electrolysis** (p.58), which is expensive.
- Metals **below carbon** in the reactivity series can be extracted by **reduction** using **carbon**. For example, **iron oxide** is reduced in a **blast furnace** to make **iron**.
- This is because carbon **can only take the oxygen** away from metals which are **less reactive** than carbon **itself** is.

Extracted using **electrolysis**.

Extracted by **reduction** using **carbon**.

### The Reactivity Series

Potassium	K	More Reactive
Sodium	Na	
Lithium	Li	
Calcium	Ca	
Magnesium	Mg	
<b>Carbon</b>	<b>C</b>	Less Reactive
Zinc	Zn	
Iron	Fe	
Copper	Cu	



Some metals are **so unreactive** they are in the earth as the metal **itself**. For example, **gold** is mined as its elemental form.

Make sure you can explain how and why different metals are extracted in different ways.

## Are you going to revise this page, ore what?

Metals are great aren't they? Loads of uses. Shame extracting them's not always cheap. Make sure you know the difference between reduction and oxidation and why carbon can be used to extract some metals but not others.

Q1 Write a balanced equation for the reduction of zinc oxide, ZnO, by carbon, C. [2 marks]

Q2 A mining company tried to extract calcium from its ore by reduction with carbon. The process did not work. Explain why. [1 mark]

# Redox Reactions

In chemistry, oxidation doesn't just mean gain of oxygen. No, that would be far too easy.


## If Electrons are Transferred, It's a Redox Reaction

- 1) Oxidation can mean the addition of oxygen (or a reaction with it), and reduction can be the removal of oxygen, but on this page we're looking at oxidation and reduction in terms of electrons.
- 2) A loss of electrons is called oxidation. A gain of electrons is called reduction. A handy way to remember this is by the mnemonic OIL RIG — Oxidation Is Loss, Reduction Is Gain.
- 3) REDUction and OXIation happen at the same time — hence the term "REDOX".

- Iron atoms are oxidised to  $\text{Fe}^{2+}$  ions when they react with dilute acid:  $\text{Fe} + 2\text{H}^+ \rightarrow \text{Fe}^{2+} + \text{H}_2$
- The iron atoms lose electrons. They're oxidised by the hydrogen ions:  $\text{Fe} - 2\text{e}^- \rightarrow \text{Fe}^{2+}$
- The hydrogen ions gain electrons. They're reduced by the iron atoms:  $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$

All reactions of metals and acids on p.55 are redox reactions.

## Displacement Reactions are Redox Reactions

- 1) Displacement reactions involve one metal kicking another one out of a compound. Here's the rule: 
- 2) If you put a reactive metal into the solution of a dissolved metal compound, the reactive metal will replace the less reactive metal in the compound (see the reactivity series on p.55).

A MORE REACTIVE metal will displace a LESS REACTIVE metal from its compound.

If you put iron in a solution of copper sulfate ( $\text{CuSO}_4$ ) the more reactive iron will 'kick out' the less reactive copper from the solution. You end up with iron sulfate solution ( $\text{FeSO}_4$ ) and copper metal.

iron + copper sulfate → iron sulfate + copper



In this reaction the iron loses 2 electrons to become a  $2+$  ion — it's oxidised. The copper ion gains these 2 electrons to become a copper atom — it's reduced.

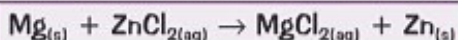
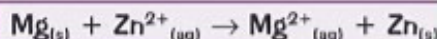


Koalas are more reactive than rucksacks.

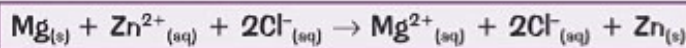
- 3) In displacement reactions, it's always the metal ion that gains electrons and is reduced. The metal atom always loses electrons and is oxidised.
- 4) In the exam you could be asked to write word or symbol equations to show displacement reactions.

## Ionic Equations Show Just the Useful Bits of Reactions

- 1) In an ionic equation only the particles that react and the products they form are shown. For example:
- 2) This just shows the displacement of zinc ions by magnesium metal. Here's what the full equation of the above reaction would be if you'd started off with zinc chloride:



- 3) If you write out the equations so you can see all the ions, you'll see that the chloride ions don't change in the reaction — they're spectator ions. They're of no interest here so can be crossed out.
- 4) Instead, the ionic equation for this displacement reaction just concentrates on the substances which are oxidised or reduced.



## REDOX — great for bubble baths. Oh no, wait...

Ionic equations are hugely important in chemistry. Better practice until you can do them in your sleep.

Q1 The equation for the reaction of zinc and iron sulfate is:  $\text{Zn}_{(s)} + \text{FeSO}_{4(aq)} \rightarrow \text{ZnSO}_{4(aq)} + \text{Fe}_{(s)}$

a) Write an ionic equation for the reaction.

[1 mark]

b) State which species is being reduced and which is being oxidised.

[2 marks]

# Electrolysis

Electrolysis uses an **electrical current** to cause a reaction. It's actually pretty cool. No, really...

## Electrolysis Means 'Splitting Up with Electricity'

- 1) During electrolysis, an electric current is passed through an electrolyte (a **molten** or **dissolved** ionic compound). The ions move towards the electrodes, where they react, and the compound **decomposes**.
- 2) The **positive ions** in the electrolyte will move towards the **cathode** (-ve electrode) and **gain** electrons (they are **reduced**).
- 3) The **negative ions** in the electrolyte will move towards the **anode** (+ve electrode) and **lose** electrons (they are **oxidised**).
- 4) This creates a **flow of charge** through the **electrolyte** as ions travel to the electrodes.
- 5) As ions gain or lose electrons, they form the uncharged element and are **discharged** from the electrolyte.

An electrolyte is just a liquid or solution that can conduct electricity. An electrode is a solid that conducts electricity and is submerged in the electrolyte.

## Electrolysis of Molten Ionic Solids Forms Elements

- 1) An **ionic solid can't** be electrolysed because the ions are in fixed positions and **can't move**.
- 2) **Molten ionic compounds can** be electrolysed because the ions can **move freely** and conduct electricity.
- 3) Molten ionic liquids, e.g. lead bromide, are always broken up into their **elements**.
- 4) Positive **metal** ions are **reduced** to the element at the **cathode**:  $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$
- 5) Negative **non-metal** ions are **oxidised** to the element at the **anode**:  $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$

The electrodes should be inert so they don't react with the electrolyte.

## Metals can be Extracted From Their Ores Using Electrolysis

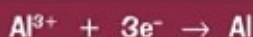
If a metal is **too reactive** to be **reduced** with **carbon** (page 56) or reacts with carbon, then electrolysis can be used to extract it. Extracting metals via this method is very expensive as lots of energy is required to melt the ore and produce the required current.

- 1) Aluminium is extracted from the ore **bauxite** by **electrolysis**. Bauxite contains **aluminium oxide**,  $\text{Al}_2\text{O}_3$ .
- 2) Aluminium oxide has a **very high** melting temperature so it's mixed with **cryolite** to lower the melting point.
- 3) The **molten mixture** contains **free ions** — so it'll **conduct electricity**.
- 4) The **positive  $\text{Al}^{3+}$  ions** are attracted to the **negative electrode** where they **each pick up three electrons** and turn into neutral **aluminium atoms**. These then **sink** to the bottom of the electrolysis tank.
- 5) The **negative  $\text{O}^{2-}$  ions** are attracted to the **positive electrode** where they **each lose two electrons**. The neutral oxygen atoms will then **combine** to form  **$\text{O}_2$  molecules**.

Cryolite is an aluminium based compound with a lower melting point than aluminium oxide.

At the negative electrode:

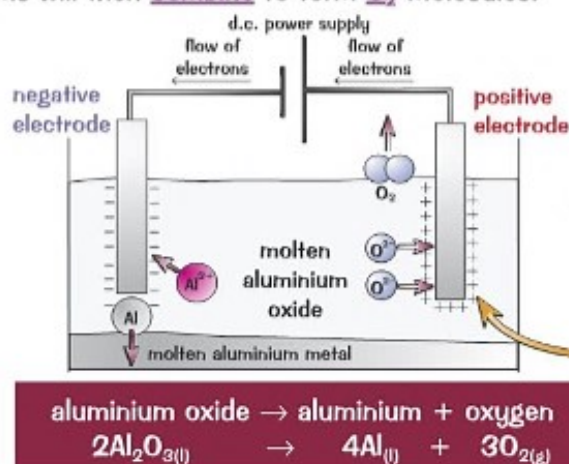
Reduction — a gain of electrons:



Metals form **positive ions**, so they're attracted to the **negative** electrode.

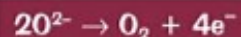
Aluminium is produced at the **negative electrode**.

Overall Equation:



At the positive electrode:

Oxidation — a loss of electrons



Non-metals form **negative ions**, so they're attracted to the **positive** electrode.

Oxygen is produced at the **positive electrode**.

The anode is made of carbon and needs to be replaced regularly as it reacts with oxygen to produce carbon dioxide.

## Faster shopping at the supermarket — use Electrolleys...

It might be jolly useful for your exams to learn the products of electrolysis of molten lead bromide...

Q1 A student carries out electrolysis on molten sodium chloride. What is produced at:

- a) the anode? b) the cathode?

[2 marks]

# Electrolysis of Aqueous Solutions

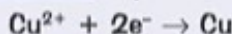
When you electrolyse an aqueous solution, you also have to factor in the ions in the water.

## It May be Easier to Discharge Ions from Water than the Solute

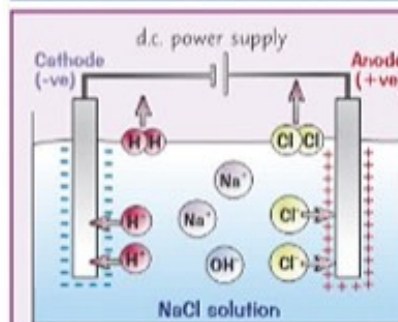
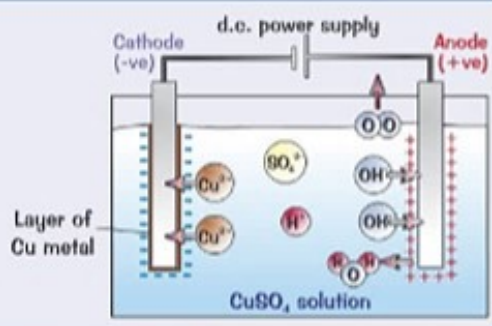
- 1) In aqueous solutions, as well as the ions from the ionic compound, there will be hydrogen ions ( $\text{H}^+$ ) and hydroxide ions ( $\text{OH}^-$ ) from the water:  $\text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}^+_{(aq)} + \text{OH}^-_{(aq)}$
- 2) At the cathode, if  $\text{H}^+$  ions and metal ions are present, hydrogen gas will be produced if the metal ions form an elemental metal that is more reactive than hydrogen (e.g. sodium ions). If the metal ions form an elemental metal that is less reactive than hydrogen (e.g. copper ions), a solid layer of the pure metal will be produced instead.
- 3) At the anode, if  $\text{OH}^-$  and halide ions ( $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$ ) are present, molecules of chlorine, bromine or iodine will be formed. If no halide ions are present, then the  $\text{OH}^-$  ions are discharged and oxygen will be formed.

A solution of copper(II) sulfate ( $\text{CuSO}_4$ ) contains four different ions:  $\text{Cu}^{2+}$ ,  $\text{SO}_4^{2-}$ ,  $\text{H}^+$  and  $\text{OH}^-$ .

- Copper metal is less reactive than hydrogen. So at the cathode, copper metal is produced and coats the electrode.

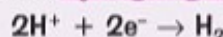


- There aren't any halide ions present. So at the anode oxygen and water are produced. The oxygen can be seen as bubbles.

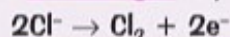


A solution of sodium chloride ( $\text{NaCl}$ ) contains four different ions:  $\text{Na}^+$ ,  $\text{Cl}^-$ ,  $\text{OH}^-$  and  $\text{H}^+$ .

- Sodium metal is more reactive than hydrogen. So at the cathode, hydrogen gas is produced.



- Chloride ions are present in the solution. So at the anode chlorine gas is produced.



If you're drawing the apparatus for an electrolysis experiment, remember to include a d.c. power supply, wires and labels for the anode and the cathode. The anode is the electrode on the same side as the longer line of the d.c. power supply symbol.

**PRACTICAL**

You can set up an electrolysis experiment in the lab like the set-up on page 109.

Once the experiment is finished you can test any gaseous products to work out what was produced.

- Chlorine bleaches damp litmus paper, turning it white.
- Hydrogen makes a "squeaky pop" with a lighted splint.
- Oxygen will relight a glowing splint.

For more on tests for gases, turn to page 88.

## The Half Equations — Make Sure the Electrons Balance

Half equations show the reactions at the electrodes. The important thing to remember when you're combining half equations is that the number of electrons needs to be the same for each half equation. For the electrolysis of aqueous sodium chloride the half equations are:



These combine to form the ionic equation:  $2\text{H}^+ + 2\text{Cl}^- \rightarrow \text{H}_2 + \text{Cl}_2$

The electrons on each side of the half equations balance, so they can be cancelled out in the full ionic equation.

When a halide isn't present in the aqueous solution, the half equation for the anode is:



## I wrote a poem about my tabby — it was a cat ode...

So it's kinda confusing this electrolysis malarkey — you need to take it slow and make sure you get it.

Q1 An aqueous solution of copper chloride,  $\text{CuCl}_2$ , is electrolysed using inert electrodes.

Give the half equations for the anode and the cathode.

[2 marks]

## Revision Questions for Topic 4

Another topic finished. Time to test your knowledge of chemical changes...

- Try these questions and tick off each one when you get it right.
- When you've done all the questions under a heading and are completely happy with it, tick it off.

### Acids and their Reactions (p.51-54) ☐

- 1) State whether the following pH values are acidic, alkaline or neutral.  
a) 9 b) 2 c) 7 d) 6 ☐
- 2) Give the general word equation for the reaction between an acid and a base. ☐
- 3) What type of reagent could be used to show that an acid or base has been completely neutralised? ☐
- 4) Describe the steps involved in carrying out a titration. ☐
- 5) Why should you first carry out a rough titration when carrying out a titration experiment? ☐
- 6) Why shouldn't you use Universal indicator as an indicator in a titration? ☐
- 7) Name two suitable indicators for a titration between an acid and an alkali, and state their colours when in acidic and alkaline solution. ☐
- 8) What is a strong acid? ☐
- 9) By what factor does the  $H^+$  concentration increase for a decrease of 1 on the pH scale? ☐
- 10) Explain the difference between a strong acid and a concentrated one. ☐
- 11) Write a balanced equation for the reaction between hydrochloric acid and sodium carbonate. ☐

### The Reactivity Series (p. 55) ☐

- 12) Is zinc more or less reactive than iron? ☐
- 13) What is the general word equation for the reaction of a metal with an acid? ☐
- 14) Give the balanced equation for the reaction of calcium with water. ☐
- 15) Will copper react with water? ☐

### Reduction and Oxidation (p.56-57) ☐

- 16) What product forms in the oxidation of magnesium by oxygen? ☐
- 17) Explain how you decide whether a metal can be extracted from its oxide by reduction with carbon. ☐
- 18) Why is gold found as itself in the earth? ☐
- 19) In terms of electrons, give the definition of oxidation. ☐
- 20) In a displacement reaction, does the metal atom get reduced or oxidised? ☐

### Electrolysis (p.58-59) ☐

- 21) During electrolysis, which electrode are the positive ions attracted to? ☐
- 22) Why can ionic solids not undergo electrolysis? ☐
- 23) Do ions get reduced or oxidised at the anode? ☐
- 24) During the manufacture of aluminium from bauxite, which electrode is aluminium formed at? ☐
- 25) In what situation will hydrogen gas be given out during the electrolysis of an aqueous solution of an ionic solid? ☐
- 26) If halide ions are present in an aqueous solution of an ionic solid will oxygen gas be released? ☐
- 27) What test could you use to determine if hydrogen gas has been produced in an electrolysis reaction? ☐

# Exothermic and Endothermic Reactions

Whenever chemical reactions occur, there are changes in **energy**. This means that when chemicals get together, things either hot up or cool right off. I'll give you a heads up — this page is a good 'un.

## Energy is *Moved Around* in Chemical Reactions

- 1) Chemicals **store** a certain amount of energy — and **different chemicals** store **different amounts**.
- 2) If the **products** of a reaction store **more** energy than the **original reactants**, then they must have **taken in** the difference in energy between the products and reactants from the **surroundings** during the reaction.
- 3) But if they store **less**, then the **excess** energy was transferred **to the surroundings** during the reaction.
- 4) The **overall** amount of energy doesn't change. This is because energy is **conserved** in reactions — it can't be created or destroyed, only **moved around**. This means the amount of energy in the **universe** always stays the **same**.

## In an *Exothermic* Reaction, Heat is *Given Out*

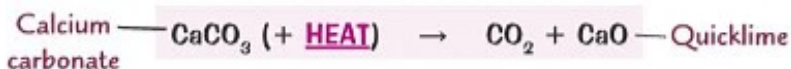
- 1) An **EXOTHERMIC** reaction is one which **transfers** energy to the **surroundings**, usually by **heating**. This is shown by a **rise in temperature**.
- 2) The best example of an exothermic reaction is **burning fuels** — also called **COMBUSTION**. This gives out a lot of energy — it's very exothermic.
- 3) **Neutralisation reactions** (acid + alkali) are also exothermic.
- 4) Many **oxidation reactions** are exothermic. For example, adding sodium to water **releases energy**, so it must be exothermic — see page 24. The reaction **releases energy** and the sodium moves about on the surface of the water as it is oxidised.
- 5) Exothermic reactions have lots of everyday uses. For example:



- Some **hand warmers** use the exothermic oxidation of **iron** in air (with a salt solution catalyst) to **release energy**.
- **Self heating cans** of hot chocolate and coffee also rely on **exothermic reactions** between chemicals in their bases.

## In an *Endothermic* Reaction, Heat is *Taken In*

- 1) An **ENDOTHERMIC** reaction is one which **takes in** energy **from** the surroundings. This is shown by a **fall in temperature**.
- 2) Endothermic reactions are much **less common** than exothermic reactions, but they include:
  - The reaction between **citric acid** and **sodium hydrogencarbonate**.
  - **Thermal decomposition** — e.g. heating calcium carbonate causes it to decompose into calcium oxide (also called quicklime) and carbon dioxide:



Physical processes can also take in or release energy. E.g. freezing is an exothermic process, melting is endothermic.



- 3) Endothermic reactions also have everyday uses. For example:

Endothermic reactions are used in some **sports injury packs** — the chemical reaction allows the pack to become **instantly cooler** without having to put it in the **freezer**.

## Right, so burning gives out heat — really...

Remember, “exo-” = exit, “-thermic” = heat, so an exothermic reaction is one that gives out heat — and endothermic means just the opposite. To make sure you really understand these terms, try this question.

- Q1 A student prepares a flask containing ethanoic acid and measures its temperature as 22.5 °C. He then adds dilute potassium hydroxide solution which is 21 °C. After 2 minutes the temperature of the reaction mixture is 28.5 °C. Is the reaction exothermic or endothermic? [1 mark]

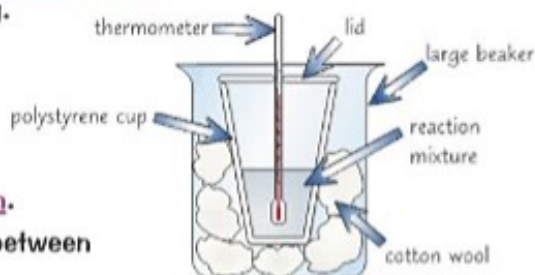
# More Exothermic and Endothermic Reactions

Sometimes it's not enough to just know if a reaction is endothermic or exothermic. You may also need to know how much energy is absorbed or released — you can do experiments to find this out. Fun, fun, fun...

## Energy Transfer can be Measured

**PRACTICAL**

- 1) You can measure the amount of energy released by a chemical reaction (in solution) by taking the temperature of the reagents, mixing them in a polystyrene cup and measuring the temperature of the solution at the end of the reaction. Easy.
- 2) The biggest problem with energy measurements is the amount of energy lost to the surroundings.
- 3) You can reduce it a bit by putting the polystyrene cup into a beaker of cotton wool to give more insulation, and putting a lid on the cup to reduce energy lost by evaporation.
- 4) This method works for neutralisation reactions or reactions between metals and acids, or carbonates and acids.
- 5) You can also use this method to investigate what effect different variables have on the amount of energy transferred — e.g. the mass or concentration of the reactants used.
- 6) Here's how you could test the effect of acid concentration on the energy released in a neutralisation reaction between hydrochloric acid (HCl) and sodium hydroxide (NaOH):



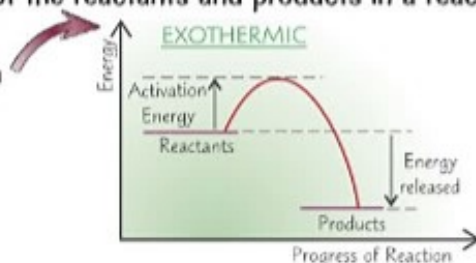
- 1) Put 25 cm<sup>3</sup> of 0.25 mol/dm<sup>3</sup> of hydrochloric acid and sodium hydroxide in separate beakers.
- 2) Place the beakers in a water bath set to 25 °C until they are both at the same temperature (25 °C).
- 3) Add the HCl followed by the NaOH to a polystyrene cup with a lid — as in the diagram above.
- 4) Take the temperature of the mixture every 30 seconds, and record the highest temperature.
- 5) Repeat steps 1-4 using 0.5 mol/dm<sup>3</sup> and then 1 mol/dm<sup>3</sup> of hydrochloric acid.

## Reaction Profiles Show Energy Changes

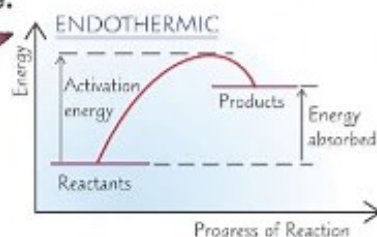
Reaction profiles are diagrams that show the relative energies of the reactants and products in a reaction, and how the energy changes over the course of the reaction.

- 1) This shows an exothermic reaction — the products are at a lower energy than the reactants. The difference in height represents the overall energy change in the reaction (the energy given out) per mole.
- 2) The initial rise in energy represents the energy needed to start the reaction. This is the activation energy (E<sub>a</sub>).
- 3) The activation energy is the minimum amount of energy the reactants need to collide with each other and react. The greater the activation energy, the more energy needed to start the reaction — this has to be supplied, e.g. by heating the reaction mixture.

Reaction profiles are sometimes called energy level diagrams.



- 1) This shows an endothermic reaction because the products are at a higher energy than the reactants.
- 2) The difference in height represents the overall energy change during the reaction (the energy taken in) per mole.



There's more on activation energy and collision theory on pages 67-68.

## Energy transfer — make sure you take it all in...

Don't get confused by these diagrams. In an exothermic reaction the particles release energy to their surroundings — even though the reaction mixture gets warmer, the particles themselves have lost energy.

- Q1 Here is the equation for the combustion of methane in air:  $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$   
Draw a reaction profile for this reaction.

[3 marks]

# Bond Energies

So you know that chemical reactions can take in or release energy — this page is about what **causes** these energy changes. Hint — it's all to do with **making** and **breaking** chemical bonds.

## Energy Must Always be Supplied to Break Bonds

There's more on energy transfer on page 61.

- 1) During a chemical reaction, **old bonds are broken** and **new bonds are formed**.
- 2) Energy must be **supplied** to break existing bonds — so bond breaking is an **endothermic** process.
- 3) Energy is **released** when new bonds are formed — so bond formation is an **exothermic** process.

### BOND BREAKING — ENDOTHERMIC



### BOND FORMING — EXOTHERMIC



- 4) In **exothermic** reactions the energy **released** by forming bonds is **greater** than the energy used to **break** them. In **endothermic** reactions the energy **used** to break bonds is **greater** than the energy released by forming them.

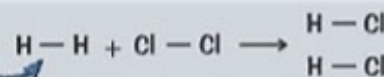
## Bond Energy Calculations — Need to be Practised

Every chemical bond has a particular **bond energy** associated with it. This **bond energy** varies slightly depending on the **compound** the bond occurs in. You can use these **known bond energies** to calculate the **overall energy change** for a reaction. The overall energy change is the **sum** of the energies **needed** to break bonds in the reactants **minus** the energy **released** when the new bonds are formed in the products. You need to **practise** a few of these, but the basic idea is really very simple...



### EXAMPLE:

Using the bond energies given below, calculate the energy change for the reaction between  $\text{H}_2$  and  $\text{Cl}_2$  forming  $\text{HCl}$ :



The bond energies you need are:  $\text{H}-\text{H}$ : +436 kJ/mol;  $\text{Cl}-\text{Cl}$ : +242 kJ/mol;  $\text{H}-\text{Cl}$ : +431 kJ/mol.

- 1) Find the **energy required** to break the original bonds:

$$(1 \times \text{H}-\text{H}) + (1 \times \text{Cl}-\text{Cl}) = 436 \text{ kJ/mol} + 242 \text{ kJ/mol} = 678 \text{ kJ/mol}$$

- 2) Find the **energy released** by forming the new bonds.

$$2 \times \text{H}-\text{Cl} = 2 \times 431 \text{ kJ/mol} = 862 \text{ kJ/mol}$$

- 3) Find the **overall energy change** for the reaction using this equation:

$$\begin{aligned} \text{Overall energy change} &= \text{energy required to break bonds} - \text{energy released by forming bonds} \\ &= 678 \text{ kJ/mol} - 862 \text{ kJ/mol} = -184 \text{ kJ/mol} \end{aligned}$$

You **can't compare** the overall energy changes of reactions unless you know the **numerical differences** in the bond energies.

Chlorine and bromine react with hydrogen in a similar way. **Br-Br** bonds are **weaker** than **Cl-Cl** bonds and **H-Br** bonds are **weaker** than **H-Cl** bonds. So **less energy** is needed to **break** the bonds in the reaction with bromine, but **less energy** is **released** when the new bonds form. So unless you know the **exact difference**, you can't say which reaction releases more energy.

## A student and their bed — a bond that can never be broken...

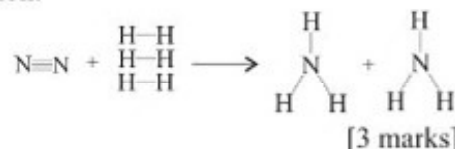
This stuff might look hard at the moment, but with a bit of practice it's dead easy and it'll win you easy marks if you understand all the theory behind it. See how you get on with this question:

Q1  $\text{N}_2$  reacts with  $\text{H}_2$  in the following reaction:  $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$

The bond energies for these molecules are:

$\text{N}\equiv\text{N}$ : 941 kJ/mol;  $\text{H}-\text{H}$ : 436 kJ/mol;  $\text{N}-\text{H}$ : 391 kJ/mol.

Calculate the overall energy change for this reaction



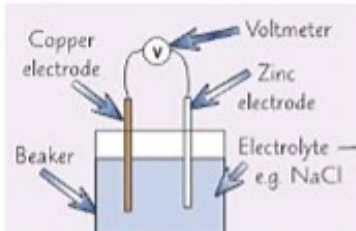
# Cells and Batteries

**Batteries** are dead useful for all sorts of things and, great news — now you can find out [how they work](#).

## Chemical Reactions in a Cell Produce Electricity

An **electrochemical cell** is a basic system made up of **two different electrodes** in contact with an **electrolyte**:

- 1) The two **electrodes** must be able to **conduct electricity** and so are usually metals.
- 2) The **electrolyte** is a liquid that contains **ions** which **react** with the electrodes.
- 3) The chemical reactions between the electrodes and the electrolyte set up a **charge difference** between the electrodes.
- 4) If the electrodes are then **connected** by a wire, the charge is able to flow and **electricity** is produced. A **voltmeter** can also be connected to the circuit to measure the **voltage** of the cell.



## The Voltage of a Cell Depends on Many Factors

- 1) **Different** metals will **react differently** with the same electrolyte — this is what causes the **charge difference**, or the **voltage**, of the cell. So the **type of electrodes** used will affect the voltage of the cell.
- 2) The bigger the **difference** in **reactivity** of the electrodes, the **bigger** the **voltage** of the cell.
- 3) You can **predict** what the voltage of a cell might be from information about the voltages of other cells.

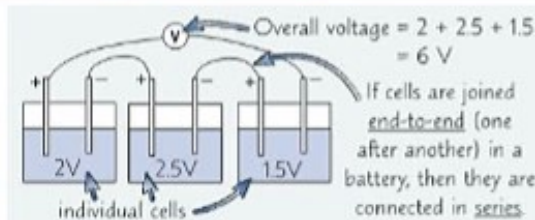
### EXAMPLE:

The voltages of two simple cells with the same electrolyte and different electrodes are shown in the table. If the metal used for Electrode A is less reactive than the other metal electrode, the voltage is positive and vice versa. Calculate the voltage of a cell where Electrode A is lead and Electrode B is tin.

Electrode A	Electrode B	Voltage (V)
Iron	Tin	-0.30
Iron	Lead	-0.31

The voltages of both cells are negative, so iron is more reactive than tin and lead. The voltage for the iron/lead cell is more negative than the iron/tin cell, so lead is less reactive than tin. The difference in voltage of these two cells is 0.01 V. So, for a cell where Electrode A is lead and Electrode B is tin, the voltage is +0.01 V.

- 4) The **electrolyte** used in a cell will also affect the size of the voltage since **different ions** in solution will react differently with the metal electrodes used.
- 5) A **battery** is formed by connecting two or more cells together in **series**. The voltages of the cells in the battery are **combined** so there is a **bigger** voltage overall.



## In Non-Rechargeable Batteries the Reactants Get Used Up

- 1) In some cells, the chemical reactions that happen at the electrodes are **irreversible**.
- 2) Over time the reacting particles — the ions in the electrolyte and the metal ions on the electrode — get **used up** and turned into the **products** of the reaction.
- 3) Once **any one** of the reactants is used up, the reaction **can't happen** and so **no electricity** is produced.
- 4) The products **can't** be turned back into the reactants, so the cell can't be **recharged**.

**Non-rechargeable** batteries, e.g. **alkaline batteries**, contain cells which use irreversible reactions. Once **one** of the reactants is used up, they don't produce any more **charge** and you have to replace them.

- 5) In a **rechargeable** cell, the reaction can be **reversed** by connecting it to an **external electric current**.

## I think I need a new battery for my brain after that...

Most so-called 'batteries' you use are probably cells. Confusing, but remember to use the right word in the exams.

- Q1 The voltages of three cells, where one electrode is zinc and the other electrode changes each time are shown in the table. Zinc is less reactive than the other metals. Write the order of reactivity of the metals used from lowest to highest. Explain your answer. [3 marks]

Metal	1	2	3
Voltage (V)	4	2.5	3

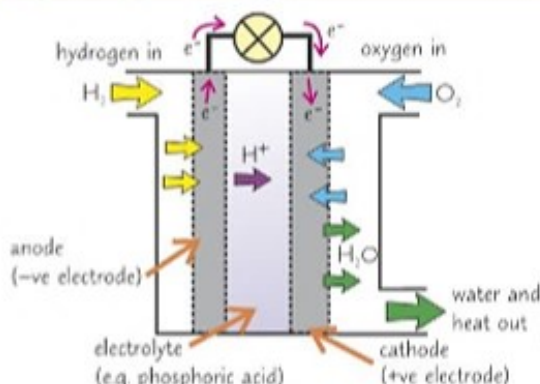
# Fuel Cells

We need fuel to move around — and I don't mean the eggs and bacon to get you out of bed on a Saturday...

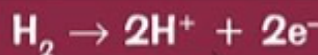
## Fuel Cells Use Fuel and Oxygen to Produce Electrical Energy

- 1) A fuel cell is an electrical cell that's supplied with a fuel and oxygen (or air) and uses energy from the reaction between them to produce electrical energy efficiently.
- 2) When the fuel enters the cell it becomes oxidised and sets up a potential difference within the cell.
- 3) There are a few different types of fuel cells, using different fuels and different electrolytes. One important example is the hydrogen-oxygen fuel cell.
- 4) This fuel cell combines hydrogen and oxygen to produce nice clean water and release energy.

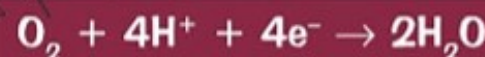
## Hydrogen-Oxygen Fuel Cells Involve a Redox Reaction



- 1) The electrolyte is often a solution of potassium hydroxide. The electrodes are often porous carbon with a catalyst.
- 2) Hydrogen goes into the anode compartment and oxygen goes into the cathode compartment. Watch out — in electrolysis the anode is the positive electrode and the cathode is the negative one. When dealing with fuel cells, they're the other way round.
- 3) At the -ve electrode (the anode), hydrogen loses electrons to produce  $H^+$  ions. This is oxidation.
- 4)  $H^+$  ions in the electrolyte move to the cathode (+ve).

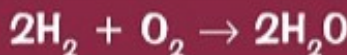


These are the half-equations that show what's happening at each electrode.



There's reduction at the cathode and oxidation at the anode, so it's a redox reaction. See page 57 for more on redox reactions.

- 5) At the +ve electrode (the cathode), oxygen gains electrons from the cathode and reacts with  $H^+$  ions (from the electrolyte) to make water. This is reduction.
- 6) The electrons flow through an external circuit from the anode to the cathode — this is the electric current.
- 7) The overall reaction is hydrogen plus oxygen, which gives water.



## Hydrogen-Oxygen Fuel Cells Could be Used in Vehicles

Conventional fuels for vehicles (such as petrol) have a finite supply — they won't last forever, and they're very polluting. So vehicles that use electrical energy are becoming more and more popular. Batteries are one way of getting cleaner energy but hydrogen-oxygen fuel cells might be even better:

- Fuel cell vehicles don't produce as many pollutants as other fuels — no greenhouse gases, nitrogen oxides, sulfur dioxide or carbon monoxide. The only by-products are water and heat. Electric vehicles don't produce many pollutants either — but their batteries are more polluting to dispose of than fuel cells because they're made from highly toxic metal compounds.
- Batteries in electric vehicles are rechargeable but there's a limit to how many times they can be recharged before they need replacing. And batteries are more expensive to make than fuel cells.
- Batteries also store less energy than fuel cells and so would need to be recharged more often — which can take a long time.

However, there are still some disadvantages of using hydrogen fuel cells:

- Hydrogen is a gas, so it takes up loads more space to store than a rechargeable battery.
- Hydrogen is explosive when mixed with air so it's hard to store safely.
- The hydrogen fuel is often made either from hydrocarbons (from fossil fuels), or by electrolysis of water, which uses electricity (and that electricity has got to be generated somehow — usually from fossil fuels).

## Fuel cells — they're simply electrifying...

Fuel cells could be really important for generating electricity cleanly — but there are some issues to sort out first.

Q1 Write the half equation for the reaction of oxygen in a hydrogen-oxygen fuel cell.

[2 marks]

## Revision Questions for Topic 5

Right, let's see how much you can remember — don't worry though, you can flick back if you get stuck.

- Try these questions and tick off each one when you get it right.
- When you've done all the questions under a heading and are completely happy with it, tick it off.

### Exothermic and Endothermic Reactions (p.61-62) ☐

- 1) In an exothermic reaction is heat transferred to or from the surroundings? ☐
- 2) Name two different types of reaction which are exothermic. ☐
- 3) Define what is meant by an endothermic reaction. ☐
- 4) Write down the equation for the thermal decomposition of calcium carbonate. ☐
- 5) What is the purpose of putting the reaction container into a beaker containing cotton wool during an experiment to investigate the temperature change of an exothermic reaction? ☐
- 6) Sketch an energy level diagram for an endothermic reaction. ☐
- 7) What is the activation energy of a reaction? ☐
- 8) Is the following statement true or false? In an endothermic reaction, the products of the reaction have more energy than the reactants. ☐

### Bond Energies (p.63) ☐

- 9) For the following sentences, use either endothermic or exothermic to fill in the blanks:
  - a) Bond breaking is an \_\_\_\_\_ process.
  - b) Bond forming is an \_\_\_\_\_ process.
  - c) In an \_\_\_\_\_ reaction, the energy released by forming bonds is greater than the energy used to break them. ☐
- 10) What three steps would you use to find the overall energy change in a reaction if you were given the known bond enthalpies for the bonds present in the reactants and products? ☐

### Cells and Batteries (p.64) ☐

- 11) What is an electrolyte in an electrochemical cell? ☐
- 12) Give two factors that affect the voltage produced by an electrochemical cell. ☐
- 13) What is a battery? ☐
- 14) Explain why non-rechargeable batteries eventually lose their charge. ☐

### Fuel Cells (p.65) ☐

- 15) What is a fuel cell? ☐
- 16) Name an electrolyte commonly used in hydrogen-oxygen fuel cells. ☐
- 17) In a hydrogen-oxygen fuel cell, which molecule becomes oxidised and which becomes reduced? ☐
- 18) Write down the overall equation for the reaction in hydrogen-oxygen fuel cells. ☐
- 19) Why is the reaction that takes place in a hydrogen-oxygen fuel cell described as a redox reaction? ☐
- 20) State 3 advantages of hydrogen-oxygen fuel cell powered cars over battery-powered electrical cars. ☐

## Rates of Reaction

Rates of reaction are pretty **important**. In the **chemical industry**, the **faster** you make **chemicals**, the **faster** you make **money** (and the faster everyone gets to go home for tea).

### Reactions Can Go at All Sorts of Different Rates

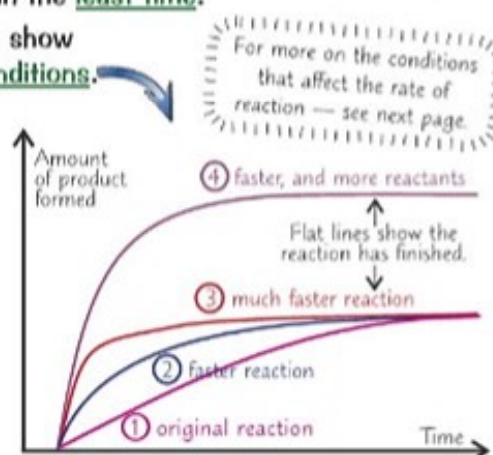
- 1) The rate of a chemical reaction is how fast the **reactants** are **changed** into **products**.
- 2) One of the **slowest** is the rusting of iron (it's not slow enough though — what about my little Mini).
- 3) Other **slow** reactions include **chemical weathering** — like acid rain damage to limestone buildings.
- 4) An example of a **moderate speed** reaction would be the metal **magnesium** reacting with an **acid** to produce a gentle stream of bubbles.
- 5) **Burning** is a **fast** reaction, but **explosions** are even **faster** and release a lot of gas. Explosive reactions are all over in a **fraction of a second**.



### You Need to Understand Graphs for the Rate of Reaction

- 1) You can find the speed of a reaction by recording the amount of **product formed**, or the amount of **reactant used up** over **time** (see page 69).
- 2) The **steeper** the line on the graph, the **faster** the rate of reaction. **Over time** the line becomes **less steep** as the reactants are **used up**.
- 3) The **quickest reactions** have the **steepest** lines and become **flat** in the **least time**.
- 4) The plot below uses the amount of **product formed** over time to show how the **speed** of a particular reaction varies under **different conditions**.

- Graph 1 represents the **original reaction**.
- Graphs 2 and 3 represent the reaction taking place **quicker**, but with the **same initial amounts** of reactants. The slopes of the graphs are **steeper** than for graph 1.
- Graphs 1, 2 and 3 all converge at the **same level**, showing that they all produce the **same amount** of product although they take **different times** to produce it.
- Graph 4 shows **more product** and a **faster reaction**. This can only happen if **more reactant(s)** are added at the start.



### Particles Must Collide with Enough Energy in Order to React

Reaction rates are explained perfectly by **collision theory**. It's simple really. The **rate** of a chemical reaction depends on:

- 1) The **collision frequency** of reacting particles (how **often** they collide). The **more** collisions there are the **faster** the reaction is. E.g. doubling the frequency of collisions doubles the rate.
- 2) The energy **transferred** during a collision. Particles have to collide with **enough energy** for the collision to be successful.

You might remember from page 62 that the **minimum** amount of energy that particles need to react is called the **activation energy**.

Particles need this much energy to **break the bonds** in the reactants and start the reaction.

Factors that **increase** the **number** of collisions (so that a **greater proportion** of reacting particles collide) or the amount of **energy** particles collide with will **increase** the **rate** of the reaction (see next page for more).



A successful collision is a collision that ends in the particles reacting to form products.

### Get a fast, furious reaction — tickle your teacher...

Collision theory's essential for understanding how different factors affect the rate of reaction — so make sure you understand it before moving on to the rest of Topic 6.

Q1 What is meant by the term activation energy?

[1 mark]

# Factors Affecting Rates of Reaction

I'd ask you to guess what this page is about, but the title pretty much says it all really. Read on...

## The Rate of Reaction Depends on Four Things

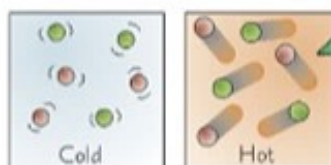
- 1) Temperature.
- 2) The concentration of a solution or the pressure of gas.
- 3) Surface area — this changes depending on the size of the lumps of a solid.
- 4) The presence of a catalyst.



## More Collisions Increases the Rate of Reaction

All four methods of increasing the rate of a reaction can be explained in terms of increasing the number of successful collisions between the reacting particles:

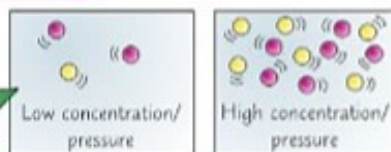
### Increasing the Temperature Increases the Rate



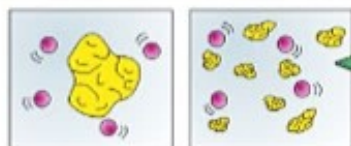
- 1) When the temperature is increased, the particles all move faster.
- 2) If they're moving faster, they're going to collide more frequently.
- 3) Also the faster they move the more energy they have, so more of the collisions will have enough energy to make the reaction happen.

### Increasing the Concentration or Pressure Increases the Rate

- 1) If a solution is made more concentrated, it means there are more particles knocking about in the same volume of water (or other solvent).
- 2) Similarly, when the pressure of a gas is increased, it means that the same number of particles occupies a smaller space.
- 3) This makes collisions between the reactant particles more frequent.



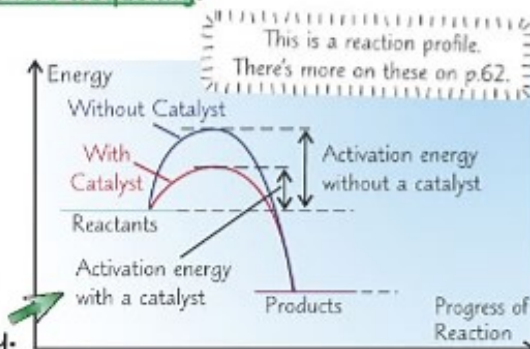
### Increasing the Surface Area Increases the Rate



- 1) If one of the reactants is a solid, then breaking it up into smaller pieces will increase its surface area to volume ratio.
- 2) This means that for the same volume of the solid, the particles around it will have more area to work on — so there will be collisions more frequently.

### Using a Catalyst Increases the Rate

- 1) A catalyst is a substance that speeds up a reaction, without being used up in the reaction itself. This means it's not part of the overall reaction equation.
- 2) Different catalysts are needed for different reactions, but they all work by decreasing the activation energy needed for the reaction to occur. They do this by providing an alternative reaction pathway with a lower activation energy.
- 3) Enzymes are biological catalysts — they catalyse reactions in living things.



## Increase your concentration — burn through that exam paper...

Catalysts are really useful — they don't get used up so you can use them over and over again. Brilliant.

- Q1 For each of the following pairs of reactions, state which one would have the fastest rate (A or B) and why:
- a) A: A 2 g solid strip of magnesium with water. B: 2 g of powdered magnesium with water. [2 marks]
  - b) A: 2 mol/dm<sup>3</sup> KOH with excess ethanoic acid. B: 4 mol/dm<sup>3</sup> KOH with excess ethanoic acid. [2 marks]

# Measuring Rates of Reaction

**PRACTICAL**

All this talk about rates of reactions is fine and dandy, but it's no good if you can't measure it.

## Here Are *Three Ways* to Measure the *Rate* of a Reaction

The rate of a reaction can be observed either by how quickly the reactants are used up or how quickly the products are formed:

$$\text{Rate of Reaction} = \frac{\text{Amount of reactant used or amount of product formed}}{\text{Time}}$$

When the product or reactant is a gas you usually measure the amount in cm<sup>3</sup>. If it's a solid, then you use grams (g). Time is often measured in seconds (s). This means that the units for rate may be in cm<sup>3</sup>/s or in g/s. You can also measure the amount of product or reactant in moles — so the units of rate could also be mol/s. Here are three different ways of measuring the rate of a reaction:

This is the mean rate of reaction. To find the rate of a reaction at a particular time, you'll need to plot a graph and find the gradient at that time (see page 71).

### 1) Precipitation and Colour Change

- 1) You can record the visual change in a reaction if the initial solution is transparent and the product is a precipitate which clouds the solution (it becomes opaque).
- 2) You can observe a mark through the solution and measure how long it takes for it to disappear — the faster the mark disappears, the quicker the reaction.
- 3) If the reactants are coloured and the products are colourless (or vice versa), you can time how long it takes for the solution to lose (or gain) its colour.
- 4) The results are very subjective — different people might not agree over the exact point when the mark 'disappears' or the solution changes colour. Also, if you use this method, you can't plot a rate of reaction graph from the results.

A posh way of saying that the cloudiness of a solution changes is to say that its 'turbidity' changes.



### 2) Change in Mass (Usually Gas Given Off)



- 1) Measuring the speed of a reaction that produces a gas can be carried out using a mass balance.
- 2) As the gas is released, the mass disappearing is measured on the balance.
- 3) The quicker the reading on the balance drops, the faster the reaction.
- 4) If you take measurements at regular intervals, you can plot a rate of reaction graph and find the rate quite easily (see page 71 for more).
- 5) This is the most accurate of the three methods described on this page because the mass balance is very accurate. But it has the disadvantage of releasing the gas straight into the room.

### 3) The Volume of Gas Given Off

- 1) This involves the use of a gas syringe to measure the volume of gas given off.
- 2) The more gas given off during a given time interval, the faster the reaction.
- 3) Gas syringes usually give volumes accurate to the nearest cm<sup>3</sup>, so they're quite accurate. You can take measurements at regular intervals and plot a rate of reaction graph using this method too. You have to be quite careful though — if the reaction is too vigorous, you can easily blow the plunger out of the end of the syringe.



**OK, have you got your stopwatch ready... \*BANG!\* — oh...**

Make sure you've learnt the three different methods on this page, then have a go at this question:

Q1 The reaction between solid Na<sub>2</sub>CO<sub>3</sub> and aqueous HCl releases CO<sub>2</sub> (a gas).

- a) Describe an experiment that would allow you to measure the rate of this reaction.
- b) Suggest units that would be appropriate for expressing the rate of this reaction.

[3 marks]

[1 mark]

## PRACTICAL

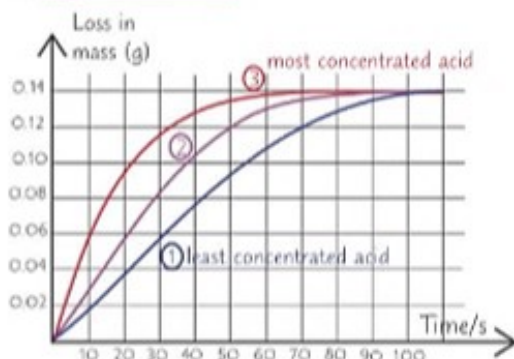
## Two Rates Experiments

Here's a lovely page on [practical investigations](#) into the effect of [concentration](#) on the rate of a reaction. It's particularly lovely because it's got [two](#) methods that you could use. Get your safety goggles on and let's go...

**Magnesium and HCl React to Produce  $H_2$  Gas**

- 1) Start by adding a set volume of [dilute hydrochloric acid](#) to a [conical flask](#) and carefully place on a [mass balance](#).
- 2) Now add some [magnesium ribbon](#) to the acid and quickly [plug](#) the flask with [cotton wool](#).
- 3) Start the [stopwatch](#) and record the [mass](#) on the balance. Take readings of the mass at [regular intervals](#).

Putting cotton wool in the top of the flask lets the gas escape but stops the acid spitting out.



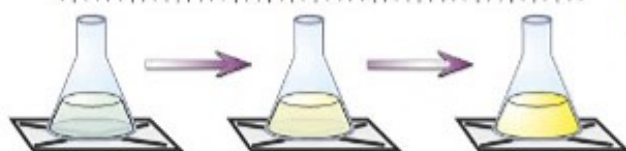
- 4) Plot the results in a table and work out the [mass lost](#) for each reading. Now you can plot a [graph](#) with [time](#) on the x-axis and [loss of mass](#) on the y-axis.
- 5) Repeat with [more concentrated](#) acid solutions. Variables such as the [amount](#) of magnesium ribbon and the [volume](#) of acid used should be kept the same each time — only change the acid's concentration. This is to make your experiment a [fair test](#) — see p.5.
- 6) The three graphs show that a [higher concentration](#) of acid gives a [faster rate of reaction](#).

You could also measure the gas released using a gas syringe, as on the previous page.

**Sodium Thiosulfate and HCl Produce a Cloudy Precipitate**

- 1) These two chemicals are both [clear solutions](#). They react together to form a [yellow precipitate](#) of [sulfur](#).
- 2) Start by adding a set volume of [dilute sodium thiosulfate](#) to a conical flask.
- 3) Place the flask on a piece of paper with a [black cross](#) drawn on it. Add some [dilute HCl](#) to the flask and start the stopwatch.
- 4) Now watch the black cross [disappear](#) through the [cloudy sulfur](#) and [time](#) how long it takes to go.
- 5) The reaction can be [repeated](#) with solutions of either reactant at [different concentrations](#). (Only change the concentration of [one reactant](#) at a time though). The [depth](#) of the liquid must be kept the [same](#) each time.
- 6) These results show the effect of [increasing](#) the [concentration](#) of [HCl](#) on the rate of reaction, when added to an excess of sodium thiosulfate.
- 7) The [higher](#) the concentration, the [quicker](#) the reaction and therefore the [less time](#) it takes for the mark to disappear.
- 8) One sad thing about this reaction is that it [doesn't](#) give a set of [graphs](#). Well I think it's sad. All you get is a set of [readings](#) of how long it took till the mark disappeared for each concentration. Boring.

This reaction releases sulfur dioxide, so the experiment should be carried out in a well-ventilated place.



Concentration of HCl (mol/dm <sup>3</sup> )	0.5	1	1.5	2	2.5
Time taken for mark to disappear (s)	193	184	178	171	164

Although you could draw a graph of concentration against 1/time which will give you an approximate rate.

**Bubbling acid, sulfurous clouds — proper witchcraft this is...**

You should learn the methods involved in these experiments — but remember, other reactions can also be used to investigate the four factors that affect rate. You might see different experiments in your exams, or the same ones but measuring a different factor — so watch out.

- Q1 A student carried out an experiment investigating the effect of changing the HCl concentration on the rate of reaction between HCl and Mg. State two factors that he should have kept constant. [2 marks]

# Finding Reaction Rates from Graphs

You might remember a bit about how to **interpret** graphs on reaction rate from page 67 — well this page shows you how to use them to **calculate** rates.

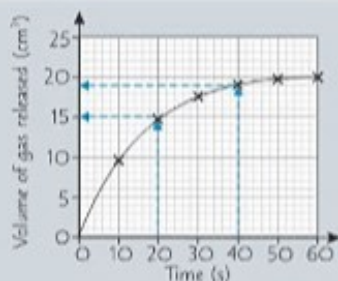
## You can Calculate the Mean Reaction Rate from a Graph

- Remember, a rate of reaction graph shows the amount of **product formed** or amount of **reactant used up** on the **y-axis** and **time** on the **x-axis**.
- So to find the **mean rate** for the **whole reaction**, you just work out the **overall change** in the **y-value** and then **divide this** by the **total time taken** for the reaction.
- You can also use the graph to find the **mean rate** of reaction between **any two points** in time:

### EXAMPLE:

The graph shows the volume of gas released by a reaction, measured at regular intervals. Find the mean rate of reaction between 20 s and 40 s.

$$\begin{aligned}\text{Mean rate of reaction} &= \text{change in } y \div \text{change in } x \\ &= (19 \text{ cm}^3 - 15 \text{ cm}^3) \div 20 \text{ s} \\ &= 0.2 \text{ cm}^3/\text{s}\end{aligned}$$



If you're asked to find the mean rate of reaction for the whole reaction, remember that the reaction finishes as soon as the line on the graph goes flat.

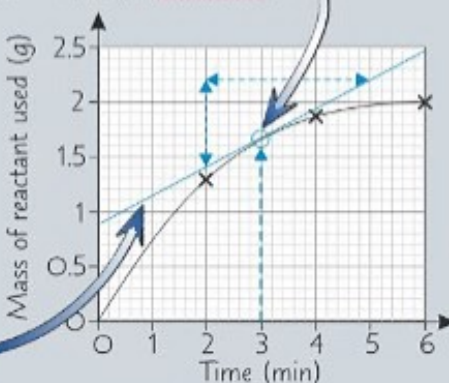
## Draw a Tangent to Find the Reaction Rate at a Particular Point

If you want to find the **rate** of the reaction at a particular point in time, you need to find the **gradient** (slope) of the curve at that point. The easiest way to do this is to draw a **tangent** to the curve — a straight line that touches the curve at one point and doesn't cross it. You then work out the **gradient of the tangent**. It's simpler than it sounds, honest...

### EXAMPLE:

The graph below shows the mass of reactant used up measured at regular intervals during a chemical reaction. What is the rate of reaction at 3 minutes?

- Position a **ruler** on the graph at the point where you want to know the rate — here it's **3 minutes**.
- Adjust the ruler until the **space** between the ruler and the curve is **equal** on **both sides** of the point.
- Draw a line along the ruler to make the **tangent**. Extend the line **right across** the graph.



- Pick **two points** on the line that are easy to read. Use them to calculate the **gradient** of the tangent in order to find the **rate**:

$$\begin{aligned}\text{gradient} &= \text{change in } y \div \text{change in } x \\ &= (2.2 - 1.4) \div (5.0 - 2.0) \\ &= 0.8 \div 3.0 \\ &= 0.27\end{aligned}$$

So, the rate of reaction at 3 minutes was **0.27 g/min**.



## Calculate your reaction to this page. Boredom? How dare you...

There's only one way to learn this stuff properly — practise. So you'd better get going with this question.

Q1 Calcium carbonate powder was added to a conical flask containing dilute HCl.

CO<sub>2</sub> was produced and collected in a gas syringe.

The volume of gas released was recorded at 10 second intervals in the following table:

Time (s)	10	20	30	40	50	60
Volume of CO <sub>2</sub> (cm <sup>3</sup> )	24	32	36	38	39	40

- Plot these results on a graph and draw a line of best fit.

[3 marks]

- Find the rate of the reaction at time = 25 s.

[4 marks]

# Reversible Reactions

Some reactions can go **backwards**. Honestly, that's all you need...

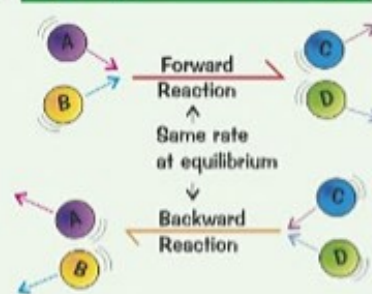
## Reversible Reactions Will Reach Equilibrium

This equation shows a **reversible reaction** — the **products** (C and D) can react to form the **reactants** (A and B) again:



The  $\rightleftharpoons$  shows the reaction goes both ways.

- As the **reactants** react, their concentrations **fall** — so the **forward reaction** will **slow down** (see page 68). But as more and more **products** are made and their concentrations **rise**, the **backward reaction** will **speed up**.
- After a while the forward reaction will be going at **exactly the same rate** as the backward one — the system is at **equilibrium**.
- At equilibrium, **both** reactions are still happening, but there's **no overall effect** (it's a dynamic equilibrium). This means the **concentrations** of reactants and products have reached a balance and **won't change**.
- Equilibrium is only reached if the reversible reaction takes place in a '**closed system**'. A **closed system** just means that **none** of the reactants or products can **escape** and nothing else can get **in**.



## The Position of Equilibrium Can be on the Right or the Left

- When a reaction's at equilibrium it **doesn't** mean the amounts of reactants and products are **equal**.
- If the equilibrium **lies to the right**, the concentration of **products** is **greater** than that of the reactants.
- If the equilibrium **lies to the left**, the concentration of **reactants** is **greater** than that of the products.
- The **position of equilibrium** depends on the following **conditions** (as well as the reaction itself):
  - the **temperature**,
  - the **pressure** (this only affects equilibria involving gases),
  - the **concentration** of the reactants and products.

E.g. ammonium chloride  $\rightleftharpoons$  ammonia + hydrogen chloride  
**Heating** this reaction moves the equilibrium to the **right** (more ammonia and hydrogen chloride) and **cooling** it moves it to the **left** (more ammonium chloride).

The next page tells you why these things affect equilibrium position.

## Reversible Reactions Can Be Endothermic and Exothermic

- In reversible reactions, if the reaction is **endothermic** in one direction, it will be **exothermic** in the other.
- The energy transferred **from** the surroundings by the endothermic reaction is **equal to** the energy transferred **to** the surroundings during the exothermic reaction.
- A good example is the **thermal decomposition** of hydrated copper sulfate:

See page 61 for more on endothermic and exothermic reactions.

'Anhydrous' just means 'without water', and 'hydrated' means 'with water'.



If you **heat blue hydrated** copper(II) sulfate crystals, it drives the water off and leaves **white anhydrous** copper(II) sulfate powder. This is **endothermic**.

If you then **add** a couple of drops of **water** to the **white powder** you get the **blue crystals** back again. This is **exothermic**.



## Dynamic equilibrium — lots of activity, but not to any great effect.\*

Make sure you understand everything on this page before you move on to the next one. Trust me, it'll help.

Q1 What does it mean if a system is at equilibrium?

[1 mark]

# Le Chatelier's Principle

Reversible reactions don't like being messed around — so if you change something, the system will **respond** to undo the change. Sneaky.

## Reversible Reactions Try to Counteract Changes...

- 1) **Le Chatelier's Principle** is the idea that if you change the **conditions** of a reversible reaction at equilibrium, the system will try to **counteract** that change.
- 2) It can be used to **predict** the effect of any changes you make to a reaction system.

## ...Such as Changes to the Temperature...

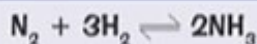
- 1) All reactions are **exothermic** in one direction and **endothermic** in the other (see previous page).
- 2) If you **decrease** the temperature, the equilibrium will move in the **exothermic direction** to produce more heat. This means you'll get **more products** for the **exothermic** reaction and fewer products for the endothermic reaction.
- 3) If you **raise** the temperature, the equilibrium will move in the **endothermic direction** to try and decrease it. You'll now get **more products** for the **endothermic** reaction and fewer products for the exothermic reaction.



Here the forward reaction is exothermic — a decrease in temperature moves equilibrium to the right (more  $\text{NH}_3$ ).

## ...Pressure...

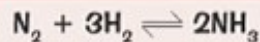
- 1) Changing the pressure only affects an equilibrium involving **gases**.
- 2) If you **increase** the pressure, the equilibrium tries to **reduce** it — it moves in the direction where there are **fewer** molecules of gas.
- 3) If you **decrease** the pressure, the equilibrium tries to **increase** it — it moves in the direction where there are **more** molecules of gas.
- 4) You can use the **balanced symbol equation** for a reaction to see which side has more molecules of gas.



There are 4 moles on the left (1 of  $\text{N}_2$  and 3 of  $\text{H}_2$ ) but only 2 on the right. So, if you increase the pressure, the equilibrium shifts to the right (more  $\text{NH}_3$ ).

## ...or Concentration

- 1) If you change the concentration of **either** the reactants or the products, the system will **no longer** be at equilibrium.
- 2) So the system responds to bring itself **back** to equilibrium again.
- 3) If you increase the concentration of the **reactants** the system tries to decrease it by making more **products**.
- 4) If you decrease the concentration of **products** the system tries to increase it again by reducing the amount of **reactants**.



If more  $\text{N}_2$  or  $\text{H}_2$  is added, the forward reaction increases to produce more  $\text{NH}_3$ .

## An equilibrium is like a particularly stubborn mule...

It's good science this stuff. You do one thing, and the reaction does the other. On the face of it, that sounds like it'd be pretty annoying, but in reality it's what gives you control of what happens in a reversible reaction. And in industry, control is what makes the whole shebang profitable. Mmmm... Money.

Q1 For the following reactions, state the effect of an increase in pressure on the amount of products at equilibrium.

- a)  $\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2\text{NO}_{2(g)}$
- b)  $\text{ClNO}_{2(g)} + \text{NO}_{(g)} \rightleftharpoons \text{NO}_{2(g)} + \text{ClNO}_{(g)}$
- c)  $2\text{CO}_{(g)} + \text{O}_{2(g)} \rightleftharpoons 2\text{CO}_{2(g)}$

[1 mark]

[1 mark]

[1 mark]

## Revision Questions for Topic 6

We'll you've almost made it — you're just one more page away from a lovely cup of tea and a biscuit...

- Try these questions and tick off each one when you get it right.
- When you've done all the questions under a heading and are completely happy with it, tick it off.

### Rates of Reaction and Factors Affecting Them (p.67-68) ☐

- 1) On a rate of reaction graph, what does the line getting steeper show? ☐
- 2) What does a flat line on a graph of amount of products against time show? ☐
- 3) What two factors relating to the collisions between particles influence the rate of a reaction? ☐
- 4) What are the four factors that affect the rate of a chemical reaction? ☐
- 5) Why does increasing the temperature of a reaction mixture increase the rate of a reaction? ☐
- 6) Other than increasing the temperature, describe two ways of increasing the rate of reaction between a solution and a solid. ☐
- 7) What is a catalyst? ☐
- 8) How does a catalyst increase the rate of a reaction? ☐

### Measuring and Calculating Rates of Reaction (p.69-71) ☐

- 9) State the equation that could be used to calculate the mean rate of a reaction. ☐
- 10) Give three possible units for the rate of a chemical reaction. ☐
- 11) How would you measure the rate of a reaction between two clear solutions, in which the product formed was a precipitate? ☐
- 12) Explain why measuring a mass change during a reaction is an accurate method of measuring rate. ☐
- 13) Describe how you could investigate the effect of increasing HCl concentration on the rate of reaction between HCl and Mg. ☐
- 14) Describe how you could use a graph to find the mean rate of a reaction between two points in time. ☐
- 15) What is a tangent? ☐
- 16) How would you use a tangent to find the gradient of a curve at a particular point? ☐

### Reversible Reactions and Le Chatelier's Principle (p.72-73) ☐

- 17) Which one of the following statements is true? ☐
  - a) In a reaction at equilibrium, there is the same amount of products as reactants.
  - b) If the forward reaction in a reversible reaction is exothermic, then the reverse reaction is endothermic.
  - c) If the equilibrium of a system lies to the right, then the concentration of products is less than the concentration of reactants.
- 18) What effect will decreasing the temperature have on a reversible reaction in which the forward reaction is exothermic? ☐
- 19) How can you predict the effect of changing the pressure of a gaseous reaction? ☐
- 20) According to Le Chatelier's Principle, what will the effect of decreasing the concentration of the products for the forward reaction have on a reversible reaction? ☐

# Hydrocarbons

Organic chemistry is about compounds that contain **carbon**. **Hydrocarbons** are the simplest organic compounds. As you're about to discover, the **properties** of hydrocarbons make them really useful.

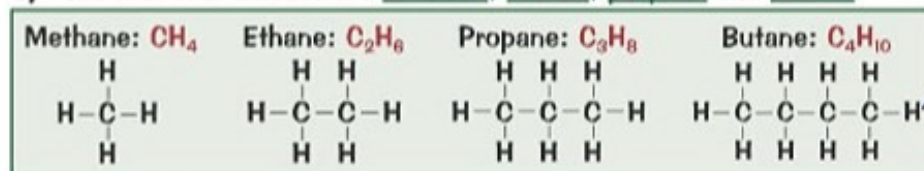
## Hydrocarbons Only Contain Hydrogen and Carbon Atoms

A hydrocarbon is any compound that is formed from **carbon and hydrogen atoms only**.

So  $C_{10}H_{22}$  (decane, an alkane) is a hydrocarbon, but  $CH_3COOC_3H_7$  (an ester) is **not** — it contains oxygen.

## Alkanes Have All C–C Single Bonds

- 1) **Alkanes** are the simplest type of hydrocarbon you can get. They have the general formula  $C_nH_{2n+2}$ .
- 2) The alkanes are a **homologous series** — a group of organic compounds that react in a similar way.
- 3) Alkanes are **saturated compounds** — each carbon atom forms four single covalent bonds.
- 4) The first four alkanes are **methane**, **ethane**, **propane** and **butane**.



A drawing showing all the atoms and bonds in a molecule is called a displayed formula.

Alkane, Al saw, Al conquered.

Give it a rest, Alan!



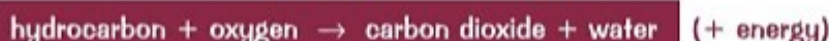
## Hydrocarbon Properties Change as the Chain Gets Longer

As the **length** of the carbon chain changes, the **properties** of the hydrocarbon change.

- 1) The **shorter** the carbon chain, the **more runny** a hydrocarbon is — that is, the **less viscous** (gooey) it is.
- 2) Hydrocarbons with shorter carbon chains are also **more volatile**, i.e. they have lower boiling points.
- 3) Also, the **shorter** the carbon chain, the more **flammable** (easier to ignite) the hydrocarbon is.
- 4) The **properties** of hydrocarbons affect how they're used for fuels. E.g. **short chain** hydrocarbons with **lower** boiling points are used as 'bottled gases' — stored **under pressure** as **liquids** in bottles.

## Complete Combustion Occurs When There's Plenty of Oxygen

- 1) The **complete combustion** of any hydrocarbon in oxygen releases lots of energy. The only waste products are **carbon dioxide** and **water** vapour.

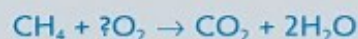


- 2) During combustion, both carbon and hydrogen from the hydrocarbon are **oxidised**.
- 3) Hydrocarbons are used as **fuels** due to the **amount of energy** released when they combust completely.
- 4) You need to be able to give a **balanced symbol equation** for the **complete combustion** of a simple hydrocarbon fuel when you're given its **molecular formula**. It's pretty easy — here's an example:

### EXAMPLE:

Write a balanced equation for the complete combustion of methane ( $CH_4$ ).

- 1) On the **left hand side**, there's **one** carbon atom, so only **one** molecule of  $CO_2$  is needed to balance this.
- 2) On the **left hand side**, there are **four** hydrogen atoms, so **two** water molecules are needed to balance them.
- 3) There are **four** oxygen atoms on the **right hand side** of the equation. **Two** oxygen molecules are needed on the left to balance them.



Oxidation can be defined as the gain of oxygen.

See p.15 for more on balancing equations.

## The name's bond — single covalent bond...

So hydrocarbons only contain two ingredients — carbon and hydrogen. Jamie Oliver would not be happy.

Q1 A student has two alkanes,  $C_5H_{12}$  and  $C_{10}H_{22}$ . Compare the following properties of the alkanes:

- a) viscosity b) boiling point c) flammability

[3 marks]

Q2 Write a balanced symbol equation for the complete combustion of propane,  $C_3H_8$ .

[2 marks]

# Fractional Distillation

Crude oil can be used to make loads of useful things, such as fuels. But you can't just put crude oil in your car. First, the different hydrocarbons have to be separated. That's where fractional distillation comes in.

## Crude Oil is Made Over a Long Period of Time

- 1) Crude oil is a fossil fuel. It's formed from the remains of plants and animals, mainly plankton, that died millions of years ago and were buried in mud. Over millions of years, with high temperature and pressure, the remains turn to crude oil, which can be drilled up from the rocks where it's found.
- 2) Fossil fuels like coal, oil and gas are called non-renewable fuels as they take so long to make that they're being used up much faster than they're being formed. They're finite resources (see p.99) — one day they'll run out.

## Fractional Distillation can be Used to Separate Hydrocarbon Fractions

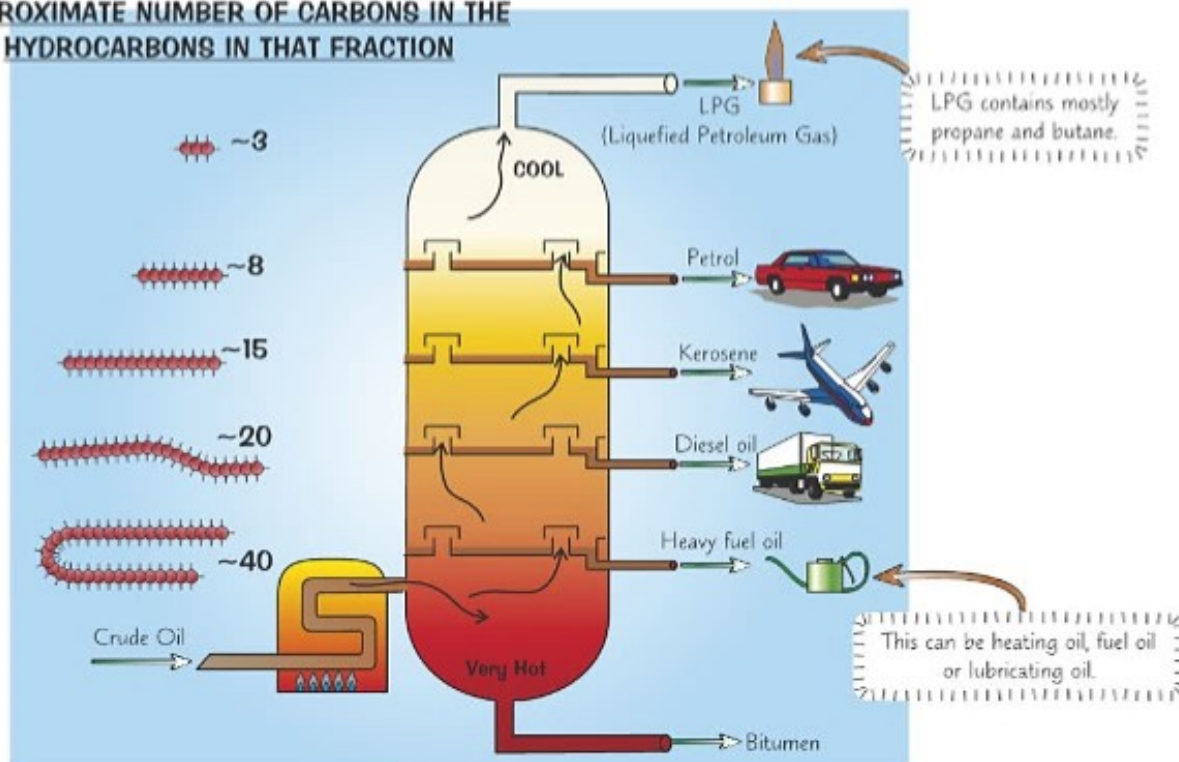
Crude oil is a mixture of lots of different hydrocarbons, most of which are alkanes. The different compounds in crude oil are separated by fractional distillation.

Here's how it works:

- 1) The oil is heated until most of it has turned into gas. The gases enter a fractionating column (and the liquid bit is drained off).
- 2) In the column there's a temperature gradient (it's hot at the bottom and gets cooler as you go up).
- 3) The longer hydrocarbons have high boiling points. They condense back into liquids and drain out of the column early on, when they're near the bottom. The shorter hydrocarbons have lower boiling points. They condense and drain out much later on, near to the top of the column where it's cooler.
- 4) You end up with the crude oil mixture separated out into different fractions. Each fraction contains a mixture of hydrocarbons that all contain a similar number of carbon atoms, so have similar boiling points.

Hydrocarbons are molecules containing only hydrogen and carbon.

### APPROXIMATE NUMBER OF CARBONS IN THE HYDROCARBONS IN THAT FRACTION



## How much petrol is there in crude oil? Just a fraction...

Make sure you understand how fractional distillation works — it might just save your life... OK, maybe not.

- Q1 Petrol drains further up a fractionating column than diesel. What does this suggest about the boiling points of the hydrocarbons which make up petrol compared to those in diesel? [1 mark]
- Q2 Describe the temperature gradient in a fractionating column used for fractional distillation. [1 mark]

# Uses and Cracking of Crude Oil

Crude oil has fuelled modern civilisation — it would be a very different world if we hadn't discovered oil.

## Crude Oil has Various Uses Important in Modern Life

- 1) Oil provides the fuel for most modern transport — cars, trains, planes, the lot. Diesel oil, kerosene, heavy fuel oil and LPG (liquid petroleum gas) all come from crude oil.
- 2) The petrochemical industry uses some of the hydrocarbons from crude oil as a feedstock to make new compounds for use in things like polymers, solvents, lubricants, and detergents.
- 3) All the products you get from crude oil are examples of organic compounds (compounds containing carbon atoms). The reason you get such a large variety of products is because carbon atoms can bond together to form different groups called homologous series. These groups contain similar compounds with many properties in common. Alkanes, alkenes, as well as the other families you'll meet in this topic, are all examples of different homologous series.

I'm more than just fuel,  
I'm going to be a star.



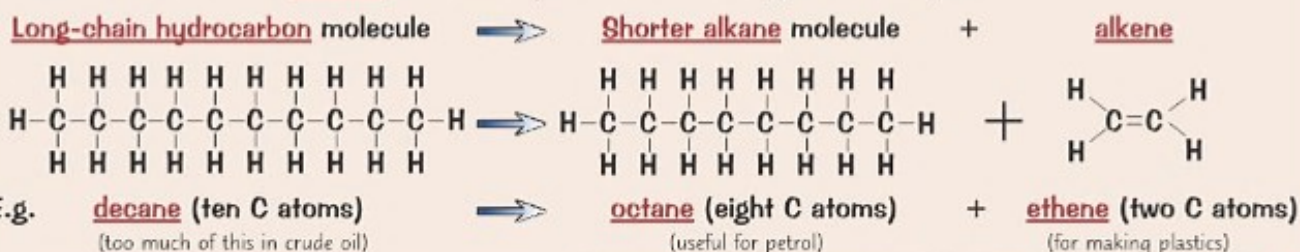
## Cracking Means Splitting Up Long-Chain Hydrocarbons

- 1) Short-chain hydrocarbons are flammable so make good fuels and are in high demand. However, long-chain hydrocarbons form thick gloopy liquids like tar which aren't all that useful, so...
- 2) ...a lot of the longer alkane molecules produced from fractional distillation are turned into smaller, more useful ones by a process called cracking.
- 3) As well as alkanes, cracking also produces another type of hydrocarbon called alkenes. Alkenes are used as a starting material when making lots of other compounds and can be used to make polymers (see p.80).
- 4) Some of the products of cracking are useful as fuels, e.g. petrol for cars and paraffin for jet fuel.

## There are Different Methods of Cracking

- 1) Cracking is a thermal decomposition reaction — breaking molecules down by heating them.
- 2) The first step is to heat long-chain hydrocarbons to vaporise them (turn them into a gas).
- 3) Then the vapour is passed over a hot powdered aluminium oxide catalyst.
- 4) The long-chain molecules split apart on the surface of the specks of catalyst — this is catalytic cracking.
- 5) You can also crack hydrocarbons if you vaporise them, mix them with steam and then heat them to a very high temperature. This is known as steam cracking.

You need to be able to balance chemical equations for cracking. For example:



Make sure that when writing equations for cracking, there are the same number of carbon and hydrogen atoms on both sides of the equation.

## This page is tough — better get cracking...

We use lots of oil — we're dependent on it for loads of things. So we could be in a proper pickle when it runs out. Just like you could be in a pickle if you don't revise this page. See what I did there? Ahaha...

- Q1 Pentane,  $\text{C}_5\text{H}_{12}$ , can be cracked into ethene,  $\text{C}_2\text{H}_4$ , and one other hydrocarbon.  
Give the formula of the other hydrocarbon.

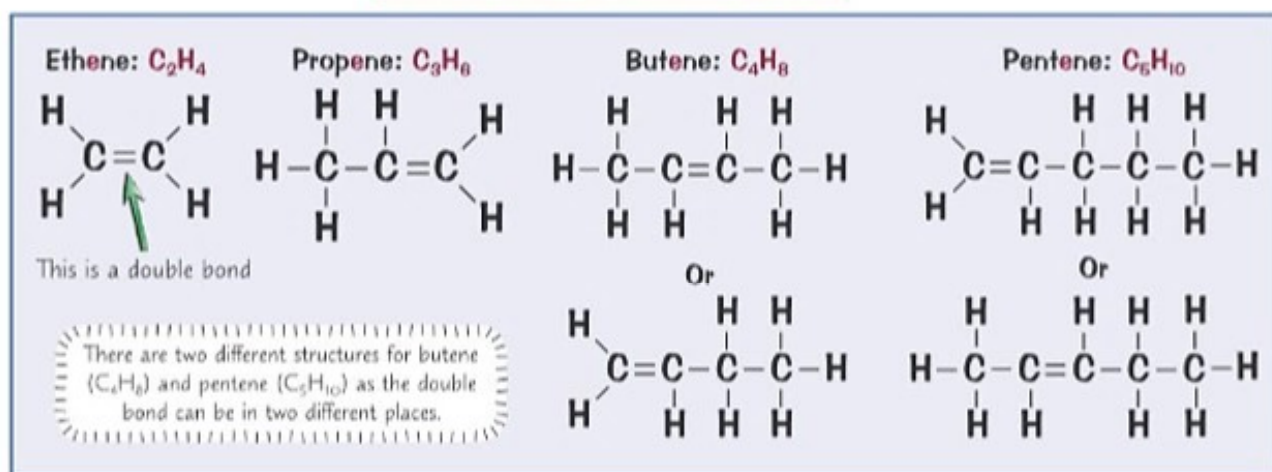
[1 mark]

# Alkenes

**Alkenes** are **unsaturated** because they have a **double** carbon-carbon bond. They're **hydrocarbons**, and their names sound like the alkanes on p.75. Don't mix them up, or you'll be the joke of the examiners' tea room...

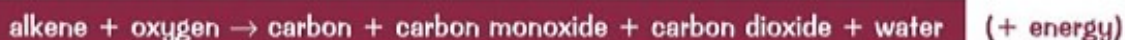
## Alkenes Have a C=C Double Bond

- 1) Alkenes are hydrocarbons which have a **double bond** between two of the **carbon** atoms in their chain.
- 2) The C=C double bond means that alkenes have **two fewer** hydrogens compared with alkanes containing the **same number** of carbon atoms. This makes them **unsaturated**.
- 3) The C=C double bond can open up to make a **single bond**, allowing the two carbon atoms to bond with **other atoms** (see the next page). This makes alkenes **reactive** — far more reactive than **alkanes**.
- 4) The first four alkenes are **ethene** (with two carbon atoms), **propene** (three Cs), **butene** (four Cs) and **pentene** (five Cs).
- 5) Straight-chain **alkenes** have twice as many hydrogen atoms as carbon:

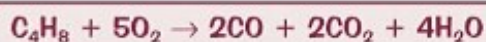


## Alkenes Burn With a Smoky Flame

- 1) In a large amount of oxygen, alkenes **combust completely** to produce only water and carbon dioxide (see p.75).
- 2) However, there isn't enough oxygen in the air for this, so when you burn them they tend to undergo **incomplete combustion**. Carbon dioxide and water are still produced, but you can also get **carbon** and **carbon monoxide** (CO) which is a poisonous gas.



- 3) Incomplete combustion results in a **smoky yellow flame**, and **less energy** being released compared to complete combustion of the same compound.
- 4) Here's an example of an **equation** for incomplete combustion:



This is just **one possibility**. The products depend on how much oxygen is present. E.g. you could also have:  $C_4H_8 + 3O_2 \rightarrow 2C + 2CO + 4H_2O$  — but the equation has to be **balanced**.

## I hope you alkene to learn this page...

Don't get alkenes confused with alkanes — that letter 'e' makes all the difference. Alkenes have a C=C bond, alkanes don't. The first part of both their names tells you how many carbon atoms they have.

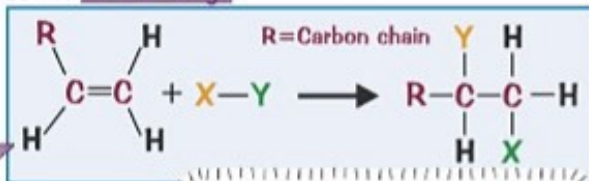
- Q1 Give the chemical formula for the alkene that contains eight carbon atoms. [1 mark]
- Q2 Name two products which are produced during incomplete combustion which aren't produced during complete combustion. [2 marks]

# Reactions of Alkenes

Alkenes **react** with lots of compounds, which is great. Unless you have to learn the reactions that is...

## Alkenes React via Addition Reactions

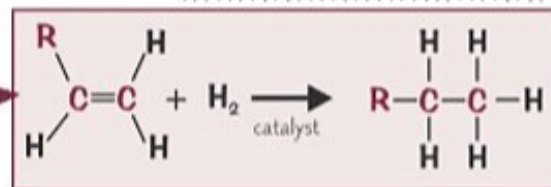
- 1) A **functional group** is a **group of atoms** in a molecule that determine **how** that molecule typically **reacts**.
- 2) All **alkenes** have the functional group 'C=C', so they all react in **similar ways**.  
So you can suggest the products of a reaction based on your knowledge of how alkenes react **in general**.
- 3) Most of the time, alkenes react via **addition reactions**.  
The carbon-carbon double bond will **open** up to leave a **single bond** and a new atom is added to each carbon:



Alkenes are a homologous series because they all have the same functional group and react in similar ways.

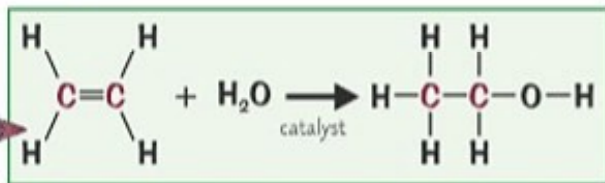
## Addition of Hydrogen is Known as Hydrogenation

**Hydrogen** can **react** with the **double-bonded** carbons to **open up** the double bond and form the equivalent, saturated, alkane.  
The alkene is reacted with hydrogen in the presence of a **catalyst**:



## Steam can React with Alkenes to Form Alcohols

- 1) When alkenes react with **steam**, **water** is added across the double bond and an **alcohol** is formed.
- 2) For example, **ethanol** can be made by mixing **ethene** with steam and then passing it over a **catalyst**:
- 3) The conversion of ethene to ethanol is one way of making ethanol **industrially**. After the reaction has taken place, the reaction mixture is passed from the reactor into a condenser. Ethanol and water have a **higher** boiling point than ethene, so both **condense** whilst any **unreacted ethene** gas is **recycled** back into the reactor. The alcohol can then be **purified** from the mixture by **fractional distillation**.

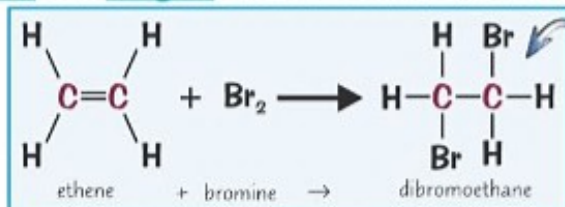


Have a look at p.81 for more on alcohols

## Halogens can React with Alkenes

- 1) Alkenes will also react in **addition reactions** with **halogens** such as bromine, chlorine and iodine. The molecules formed are saturated, with the C=C carbons each becoming bonded to a halogen atom.
- 2) For example, bromine and ethene react together to form **dibromoethane**:

There are two bromine atoms so it's called **dibromoethane**.



This reaction is the same for Cl<sub>2</sub> or I<sub>2</sub> except the halogen added is different.

The **addition** of bromine to a double bond can be used to test for **alkenes**:

- 1) When orange **bromine water** is added to a **saturated compound**, like an **alkane**, no reaction will happen and it'll stay **bright orange**.
- 2) If it's added to an **alkene** the **bromine** will add **across** the double bond, making a **colourless** dibromo-compound — so the bromine water is decolourised.



## Double the carbon bonds, double the fun...

These reactions aren't too bad to get your head around. They're as easy as saying 'diiodoethane'...

Q1 Propene reacts with hydrogen in the presence of a nickel catalyst at 60 °C.

Give the displayed formula of the product of this reaction.

[1 mark]

Q2 State what colour change occurs when butene, C<sub>4</sub>H<sub>8</sub>, reacts with bromine water.

[1 mark]

# Addition Polymers

**Polymers** are made up of **lots** of the same molecule **joined together** in one long chain. They're what make up plastics. They have lots of weird and wonderful properties which make them darn useful to modern society.

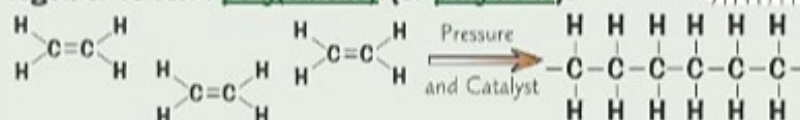
## Plastics are Made Up of Long-Chain Molecules Called Polymers

- 1) **Polymers** are long molecules formed when lots of small molecules called **monomers** join together. This reaction is called **polymerisation** — and it usually needs high pressure and a catalyst.
- 2) Plastics are made up of polymers. They're usually **carbon based** and their monomers are often **alkenes** (see page 78).

## Addition Polymers are Made From Unsaturated Monomers

- 1) The monomers that make up **addition** polymers have a **double covalent bond**.
- 2) Lots of **unsaturated monomer molecules** (**alkenes**) can open up their **double bonds** and join together to form **polymer chains**. This is called **addition polymerisation**.

For example, lots of ethene molecules can react together to form **poly(ethene)** (or **polythene**):



The reaction can also be shown like this:

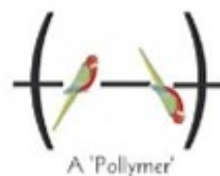


- 3) When the **monomers** react in addition polymerisation reactions, the **only** product is the polymer, so an addition polymer contains exactly the **same type** and **number** of atoms as the monomers that formed it.

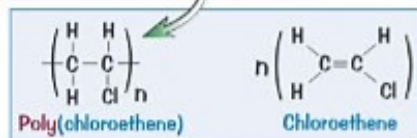
## You May Need to Draw the Repeating Unit of a Polymer

- 1) Drawing the **displayed formula** of an addition polymer from the displayed formula of its **monomer** is easy.

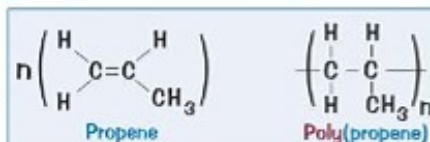
- Start by drawing the two **alkene carbons**, replace the **double bond** with a **single bond** and add an extra single bond to each of the carbons.
- Then fill in the rest of the groups in the same way that they surrounded the double bond in the monomer. Finally, stick a pair of **brackets** around the repeating bit, and put an '**n**' after it (to show that there are lots of monomers).



- 2) To get from the **displayed formula** of the **polymer** to the displayed formula of the **monomer**, just do the reverse. Draw out the **repeating bit** of the polymer, get rid of the two bonds going out through the brackets and put a **double bond** between the **carbons**.



- 3) The **name** of the polymer comes from the **type of monomer** it's made from — you just stick the word '**poly**' in front of it and put the monomer name in brackets. So **propene** becomes **poly(propene)**.



## Which polymer is good for making a cuppa? Poly(putthekettleon)...

Which monomer a polymer is made from affects the properties of the plastic — as there are so many monomers, chemists can make all sorts of things. Make sure you can recognise and draw polymers and their monomers.

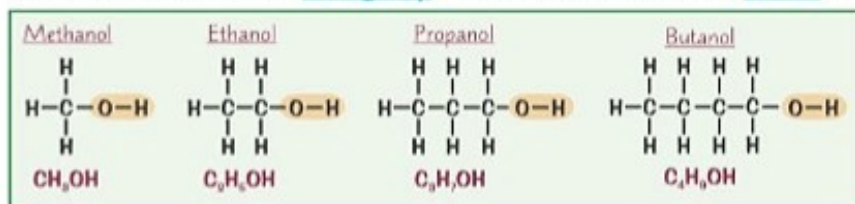
- Q1 What is a polymer? [1 mark]
- Q2 Two different polymers can be formed from butene. This is dependent on whether the double bond of the butene monomer is located in the middle or at the end of the carbon chain. Draw the two possible repeating units of the polymer made from butene molecules. [2 marks]

# Alcohols

**Alcohol** isn't just the thing your Nan has for special occasions. In fact, **alcohols** are yet another **homologous series** of organic compounds. They're handy for lots of things, and you need to know how they're used.

## Alcohols Have an '-OH' Functional Group and End in '-ol'

- 1) The **general formula** of an alcohol is  $C_nH_{2n+1}OH$ . So an alcohol with 2 carbons has the formula  $C_2H_5OH$ .
- 2) All alcohols contain an **-OH group**. You need to know the **first 4** in the homologous series:

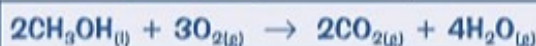


Remember — a homologous series is a group of chemicals that react in a similar way because they have the same functional group (in alcohols it's the -OH group).

- 3) The basic **naming** system is the same as for alkanes — but replace the final '-e' with '-ol'.
- 4) Don't write  $CH_4O$  instead of  $CH_3OH$  — it doesn't show the **-OH functional group**.

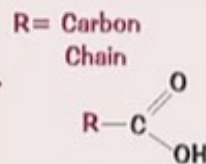
## The First Four Alcohols Have Similar Properties

- 1) Alcohols are **flammable**. They undergo complete combustion in air to produce **carbon dioxide** and **water**. For example:



Make sure you can write balanced equations for the combustion of alcohols.

- 2) Methanol, ethanol, propanol and butanol are all soluble in water. Their solutions have a **neutral pH**.
- 3) They also react with **sodium**. One of the products of this reaction is **hydrogen**.
- 4) **Alcohols** can be **oxidised** by reacting with oxygen (e.g. from the air) to produce a **carboxylic acid** (see the next page):
- 5) Different alcohols form different carboxylic acids. For example, **methanol** is oxidised to **methanoic acid**, while **ethanol** is oxidised to **ethanoic acid**.



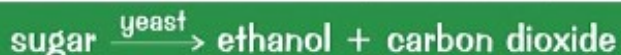
## Alcohols are Used as Solvents and Fuels

- 1) Alcohols such as methanol and ethanol are used as solvents in industry. This is because they can **dissolve** most things water can dissolve, but they can also dissolve substances that **water can't dissolve** — e.g. hydrocarbons, oils and fats.
- 2) The first four alcohols are used as fuels. For example, **ethanol** is used as a fuel in **spirit burners** — it burns fairly **cleanly** and it's **non-smelly**.

## Ethanol can be Made by Fermentation

**Ethanol** is the alcohol found in **alcoholic drinks** such as wine or beer. It's usually made using **fermentation**.

- 1) **Fermentation** uses an enzyme in **yeast** to convert **sugars** into **ethanol**. Carbon dioxide is also produced. The reaction occurs in solution so the ethanol produced is aqueous.



Ethanol can also be produced from ethene. See p.79.

- 2) Fermentation happens fastest at a temperature of around **37 °C**, in a **slightly acidic** solution and under **anaerobic conditions** (no oxygen).
- 3) Under these conditions the enzyme in yeast works best to convert the sugar to alcohol. If the conditions were different, for example a **lower pH/higher temperature** or **higher pH/lower temperature**, the enzyme could be denatured (destroyed) or could work at a much slower rate.

## Alcohol — drink too much and it's likely tequila...

For the exams, it might be jolly useful to learn the structures of alcohols... In fact, it might be essential...

Q1 Give the balanced equation for the complete combustion of ethanol,  $C_2H_5OH$ .

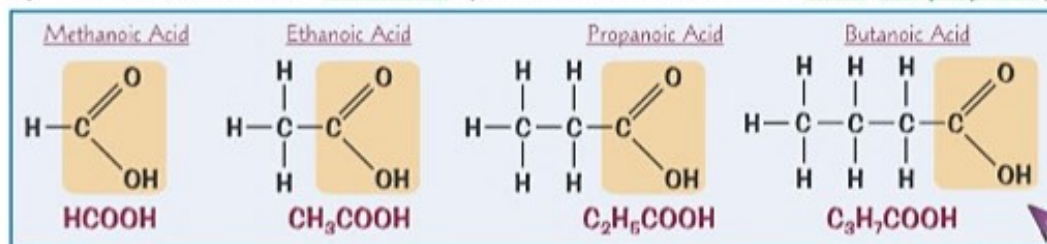
[2 marks]

# Carboxylic Acids

So you would have seen a little bit about [carboxylic acids](#) on the previous page. Now it's time to dive deeper into their exciting and wondrous world... (Disclaimer: may not be exciting or wondrous.)

## Carboxylic Acids Have the Functional Group -COOH

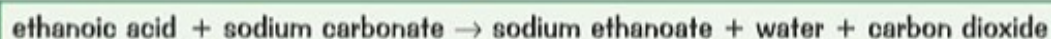
- 1) [Carboxylic acids](#) are a homologous series of compounds that all have '-COOH' as a [functional group](#).
- 2) Their names end in '-anoic acid' (and start with the normal 'meth/eth/prop/but')



Make sure you know the names and the structures of these four carboxylic acids. You could be asked about them in the exams.

## Carboxylic Acids React Like Other Acids

- 1) They react (like [any other acid](#)) with [carbonates](#) to produce a [salt](#), [water](#) and [carbon dioxide](#).
- 2) The [salts](#) formed in these reactions end in '-anoate' — e.g. methanoic acid will form a [methanoate](#), ethanoic acid an [ethanoate](#), etc. For example:



- 3) Carboxylic acids can [dissolve in water](#). When they dissolve, they [ionise](#) and [release H<sup>+</sup> ions](#) resulting in an acidic solution. But, because they [don't ionise completely](#) (not all the acid molecules release their H<sup>+</sup> ions), they just form [weak](#) acidic solutions. This means that they have a [higher pH](#) (are less acidic) than aqueous solutions of [strong acids](#) with the [same concentration](#). There's more about strong and weak acids on page 53.

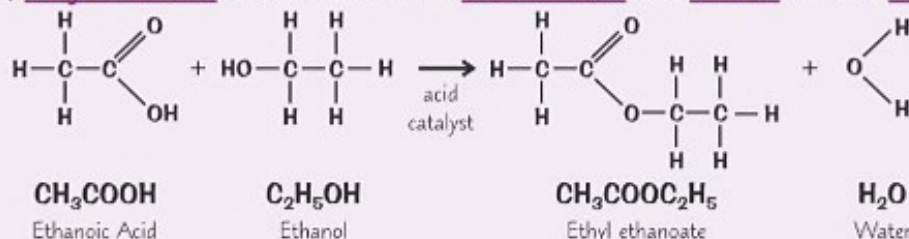
## Esters can be Made from Carboxylic Acids

- 1) [Esters](#) have the functional group '-COO-'.
- 2) [Esters](#) are formed from an [alcohol](#) and a [carboxylic acid](#).
- 3) An [acid catalyst](#) is usually used (e.g. concentrated [sulfuric acid](#)).

The names of esters can be a bit complicated and are often tongue twisters. The one you need to make sure you learn is ethyl ethanoate — the examiners could ask you about this one.



For example, [ethyl ethanoate](#) can be made from [ethanoic acid](#) and [ethanol](#) with an [acid catalyst](#):



## What's a chemist's favourite chocolate? Ester eggs...

So carboxylic acids react in the same way as normal acids, however they don't ionise fully in water which means they're weak acids. But let's not judge them on that basis, they can make esters after all.

Q1 Sodium carbonate is added to an aqueous solution of methanoic acid.

- a) Describe and explain what you would see during this reaction.
- b) Give the word equation for this reaction.

[2 marks]

[1 mark]

# Condensation Polymers

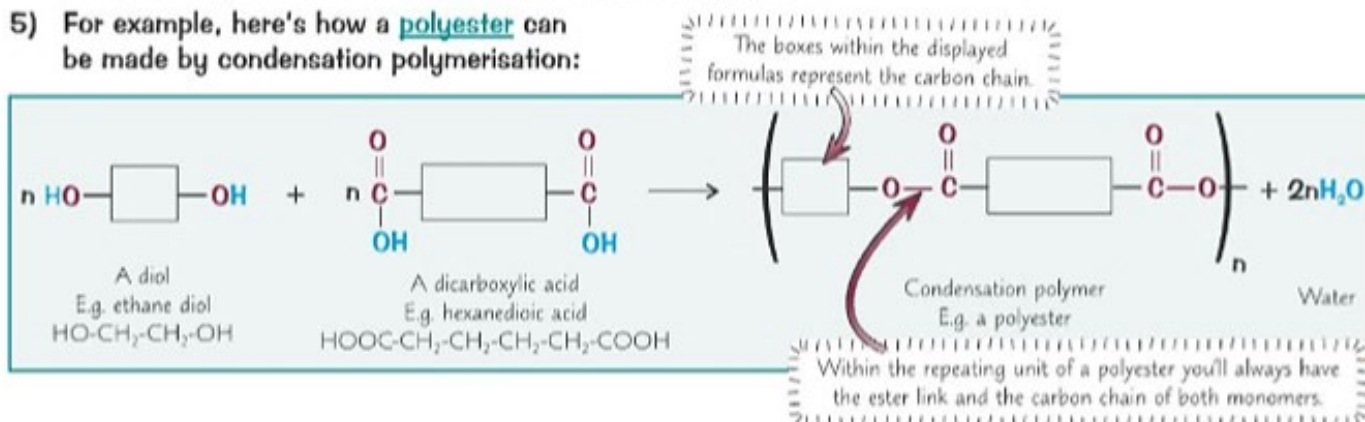
You might remember [addition polymerisation](#) from back on page 80, but there is [another](#) type of polymerisation that you could be asked about. This time the monomers need to have [TWO](#) functional groups — gasp.

## Polymers can be Made by Condensation Polymerisation

- 1) Condensation polymerisation involves monomers which contain [different](#) functional groups.
- 2) The monomers react together and [bonds](#) form between them, making polymer chains.
- 3) For [each new bond](#) that forms, a [small molecule](#) (for example, water) is [lost](#). This is why it's called [condensation](#) polymerisation.
- 4) The simplest types of condensation polymers contain [two different types](#) of monomer, each with two of the same functional groups.
- 5) For example, here's how a [polyester](#) can be made by condensation polymerisation:



I'm a bear on a bicycle and I endorse this message.



## Addition and Condensation Polymerisation are Different

Both addition and condensation polymerisation produce [polymers](#). However, the [products](#) and [reactants](#) are very different:

	Addition Polymerisation	Condensation Polymerisation
<u>Number of types of monomers</u>	Only one monomer type containing a C=C bond.	Two monomer types each containing two of the same functional groups. <b>or</b> One monomer type with two different functional groups (see next page).
<u>Number of products</u>	Only one product formed.	Two types of product — the polymer and a small molecule (e.g. water).
<u>Functional groups involved in polymerisation</u>	Carbon-carbon double bond in monomer.	Two reactive groups on each monomer.

## Revision's like polymers — it's all about stringing facts together...

So, addition polymers are formed when identical monomers containing C=C bonds are joined together in a polymer chain. Condensation polymers are made when different monomers, or those with two different reactive functional groups, react together. Got it? It's a bit confusing but that table should help.

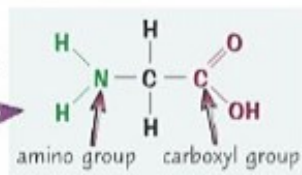
- Q1 Suggest what small molecule can be produced during a polymerisation reaction to form a polyester. [1 mark]
- Q2 How many products are formed by an addition polymerisation reaction? [1 mark]

# Naturally Occurring Polymers

Polymers might sound futuristic and artificial but they're actually found throughout the natural world...

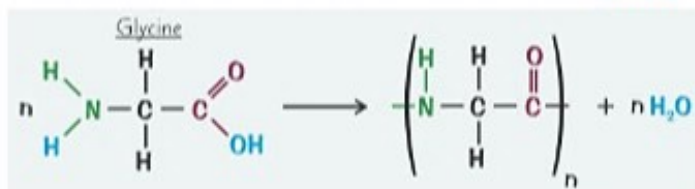
## Amino Acids have an Amino Group and a Carboxyl Group

- 1) An amino acid contains two different functional groups — a basic amino group ( $\text{NH}_2$ ) and an acidic carboxyl group ( $\text{COOH}$ ).
- 2) An example of an amino acid is glycine — the smallest and simplest amino acid possible.



## Proteins are Polymers of Amino Acids

- 1) Amino acids can form polymers known as polypeptides via condensation polymerisation.
- 2) The amino group of an amino acid can react with the acid group of another, and so on, to form a polymer chain. For every new bond that is formed a molecule of water is lost.



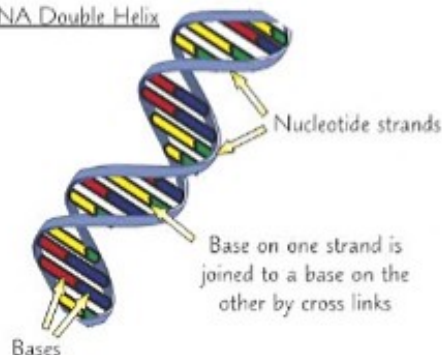
- 3) One or more long-chains of polypeptides are known as proteins. Proteins have loads of important uses in the human body. For example, enzymes work as catalysts, haemoglobin transports oxygen, antibodies form part of the immune system, and the majority of body tissue is made from proteins.
- 4) Polypeptides and proteins can contain different amino acids in their polymer chains. The order of the amino acids is what gives proteins their different properties and shapes.

## DNA Molecules are Made From Nucleotide Polymers

DNA (deoxyribonucleic acid) is found in every living thing and many viruses. It contains genetic instructions that allow the organism to develop and operate. It's a large molecule that takes a double helix structure.

- 1) DNA is made of two polymer chains of monomers called 'nucleotides'. The nucleotides each contain a small molecule known as a 'base'. There are four different bases, known by their initials — A, C, G, T.
- 2) The bases on the different polymer chains pair up with each other and form cross links keeping the two strands of nucleotides together and giving the double helix structure.
- 3) The order of the bases acts as a code for an organism's genes.

A DNA Double Helix



## Simple Sugars Can Form Polymers

- 1) Sugars are small molecules that contain carbon, oxygen and hydrogen.
- 2) Sugars can react together through polymerisation reactions to form larger carbohydrate polymers, e.g. starch, which living things use to store energy, and cellulose, which is found in plant cell walls.

## Protein — not just for bodybuilders...

Polymers are found everywhere in nature. You can't get away from them — trust me I've tried. Even the paper on this page is made of cellulose fibres which are polymers. Like revision, there is no escape.

Q1 Amino acids can bond together by condensation polymerisation.

State the type of polymer formed from this reaction.

[1 mark]

Q2 Give three examples of naturally occurring polymers.

[3 marks]

# Revision Questions for Topic 7

Well, that wraps up [Topic 7](#) — I think this was my favourite topic so far, along with 1, 2, 3, 4, 5, and 6...

- Try these questions and [tick off each one](#) when you [get it right](#).
- When you've done [all the questions](#) under a heading and are [completely happy](#) with it, tick it off.

## Hydrocarbons (p.75) ☐

- 1) What two elements do hydrocarbons contain? ☐
- 2) What is the general formula for alkanes? ☐
- 3) Draw the displayed formula of butane. ☐
- 4) What two waste products form from the complete combustion of hydrocarbons? ☐

## Crude Oil (p.76-77) ☐

- 5) How is crude oil formed? ☐
- 6) Where are the shortest carbon chains found in the fractional distillation column? ☐
- 7) Give three products that can be made from crude oil. ☐
- 8) Why is cracking used? ☐
- 9) Give a product of cracking that is used for making plastics. ☐

## Alkenes and Addition Polymerisation (p.78-80) ☐

- 10) What is an alkene? ☐
- 11) Give the general formula for an alkene. ☐
- 12) Why do alkenes burn in air with a smoky flame? ☐
- 13) Name the molecule that ethene reacts with when it is hydrogenated. ☐
- 14) Draw the product of the reaction of ethene with bromine water. ☐
- 15) What is used to test for alkenes? ☐
- 16) What type of compounds are used as the monomers to make addition polymers? ☐
- 17) Draw the displayed formula for the repeat unit of poly(ethene). ☐

## Alcohols and Carboxylic Acids (p.81-82) ☐

- 18) Give the names and formulas of the first four alcohols. ☐
- 19) Give the balanced equation for the combustion of methanol. ☐
- 20) Name the gas that is produced when ethanol reacts with sodium. ☐
- 21) Name the process that converts sugar to ethanol using yeast. ☐
- 22) Draw the displayed formula of  $C_2H_5COOH$ . ☐
- 23) Why are carboxylic acids considered weak acids? ☐
- 24) Give the general word equation for the reaction of an alcohol and a carboxylic acid. ☐

## Condensation Polymerisation and Natural Polymers (p.83-84) ☐

- 25) How many different products are there in a condensation polymerisation reaction? ☐
- 26) What are polypeptides? ☐
- 27) What small molecule is lost when two amino acids react together to form a new bond? ☐

## Purity and Formulations

In an ideal world, every compound a chemist made would be **100% pure**. Unfortunately, in the real world it **doesn't** always work out like that — but luckily, there are ways to find out **how pure** a substance is.

### Purity is Defined Differently in Chemistry to Everyday

- 1) **Usually** when you refer to a **substance** as being **pure** you mean that **nothing** has been **added** to it, so it's in its **natural state**. For example: pure milk or beeswax.
- 2) In **chemistry**, a pure substance is something that only contains **one compound** or **element** throughout — not mixed with anything else.

### The Boiling or Melting Point Tells You How Pure a Substance Is

- 1) A chemically pure substance will **melt** or **boil** at a **specific** temperature.
- 2) You can test the purity of a sample by measuring its **melting** or **boiling point** and comparing it with the melting or boiling point of the **pure substance** (which you can find from a **data book**).
- 3) The **closer** your measured value is to the actual melting or boiling point, the **purer** your sample is.
- 4) Impurities in your sample will **lower** the **melting point** and **increase** the **melting range** of your substance.
- 5) Impurities in your sample will also **increase** the **boiling point** and may result in your sample boiling at a **range** of temperatures.

### Formulations are Mixtures with Exact Amounts of Components

- 1) **Formulations** are useful mixtures with a **precise purpose** that are made by following a 'formula' (a recipe). Each component in a formulation is present in a **measured quantity**, and **contributes** to the properties of the formulation so that it meets its **required function**.

Take a look at p.16 for more on mixtures.

For example, paints are formulations composed of:

- **Pigment** — gives the paint colour, for example titanium oxide is used as a pigment in white paints.
- **Solvent** — used to dissolve the other components and alter the viscosity.
- **Binder** (resin) — forms a film that holds the pigment in place after it's been painted on.
- **Additives** — added to further change the physical and chemical properties of the paint.

Depending on the **purpose** of the paint, the **chemicals** used and their **amounts** will be changed so the paint produced is right for the job.

- 2) Formulations are really important in the **pharmaceutical industry**. For example, by altering the formulation of a pill, chemists can make sure it delivers the drug to the correct **part of the body** at the right **concentration**, that it's **consumable** and has a long enough **shelf life**.
- 3) In **everyday life**, formulations can be found in cleaning products, fuels, cosmetics, fertilisers (see p.105), metal alloys and even food and drink.
- 4) When you buy a product, you might find that it has **information** about its composition on the packaging. For example, the **ratio** or **percentage** of each component. This tells you the product's a **formulation**. It also lets you choose a formulation with the **right composition** for your particular use.



### Cake and tea are key to the revision success formula...

Knowing how pure a product is can be vital in industries such as pharmaceuticals and the food industry. Luckily for us, chemists have lots of different ways to make sure they're making exactly what they want.

- Q1 The melting point of a sample of aspirin made by a student is measured as being between 128–132 °C. The melting point and boiling points of pure aspirin are 136 °C and 140 °C respectively.
- a) Give two reasons why the melting point measured shows that the sample is not pure. [2 marks]
  - b) Suggest a value for the boiling point of the sample. [1 mark]

# Paper Chromatography

You met chromatography on page 16. Now it's time to see how it works. Careful — things might get crazy...

## Chromatography uses Two Phases

Chromatography is an analytical method used to separate the substances in a mixture. You can then use it to identify the substances. There are different types of chromatography, but they all have two 'phases':

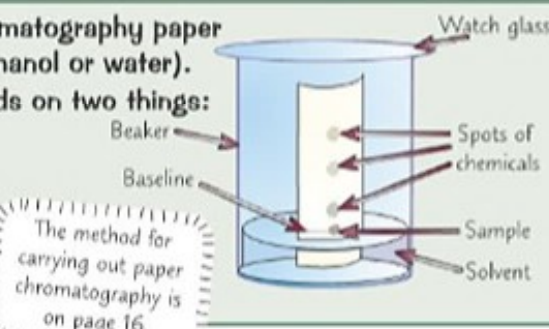
- A mobile phase — where the molecules can move. This is always a liquid or a gas.
  - A stationary phase — where the molecules can't move. This can be a solid or a really thick liquid.
- 1) During a chromatography experiment, the substances in the sample constantly move between the mobile and the stationary phases — an equilibrium is formed between the two phases.
  - 2) The mobile phase moves through the stationary phase, and anything dissolved in the mobile phase moves with it. How quickly a chemical moves depends on how it's 'distributed' between the two phases — whether it spends more time in the mobile phase or the stationary phase.
  - 3) The chemicals that spend more time in the mobile phase than the stationary phase will move further through the stationary phase.
  - 4) The components in a mixture will normally separate through the stationary phase, so long as all the components spend different amounts of time in the mobile phase. The number of spots may change in different solvents as the distribution of the chemical will change depending on the solvent. A pure substance will only ever form one spot in any solvent as there is only one substance in the sample.

During paper chromatography the stationary phase is the chromatography paper (often filter paper) and the mobile phase is the solvent (e.g ethanol or water).

The amount of time the molecules spend in each phase depends on two things:

- How soluble they are in the solvent.
- How attracted they are to the paper.

Molecules with a higher solubility in the solvent, and which are less attracted to the paper, will spend more time in the mobile phase — and they'll be carried further up the paper.

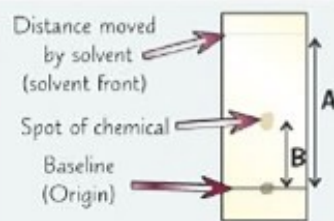


## You can Calculate the $R_f$ Value for Each Chemical

- 1) The result of chromatography analysis is called a chromatogram.
- 2) An  $R_f$  value is the ratio between the distance travelled by the dissolved substance (the solute) and the distance travelled by the solvent. The further through the stationary phase a substance moves, the larger the  $R_f$  value. You can calculate  $R_f$  values using the formula:

$$R_f = \frac{\text{distance travelled by substance (B)}}{\text{distance travelled by solvent (A)}}$$

This is the distance from the baseline to the centre of the spot.



- 3) Chromatography is often carried out to see if a certain substance is present in a mixture. To do this, you run a pure sample of that substance (a reference) alongside the unknown mixture. If the  $R_f$  values of the reference and one of the spots in the mixture match, the substance may be present (although you haven't yet proved they're the same).
- 4) The  $R_f$  value is dependent on the solvent — if you change the solvent the  $R_f$  value for a substance will change. You can test both the mixture and the reference in a number of different solvents. If the  $R_f$  value of the reference compound matches the  $R_f$  value of one of the spots in the mixture in all the solvents, then it's likely the reference compound is present in the mixture. If the spots in the mixture and the spot in the reference only have the same  $R_f$  value in some of the solvents, then the reference compound isn't present in the mixture.

## Chromatography revision — it's a phase you have to get through...

You can't see the chemicals moving between the two phases, but it does happen. You'll just have to trust me.

Q1 Explain how paper chromatography separates mixtures.

[4 marks]

## PRACTICAL

## Tests for Gases and Anions

Ahh... tests, glorious tests. Luckily, these aren't the kind of tests you have to revise for, but you should probably revise these tests for your exam — it's swings and roundabouts really...

There are **Tests for 4 Common Gases**1) **Chlorine**

Chlorine bleaches damp litmus paper, turning it white. (It may turn red for a moment first though — that's because a solution of chlorine is acidic.)

2) **Oxygen**

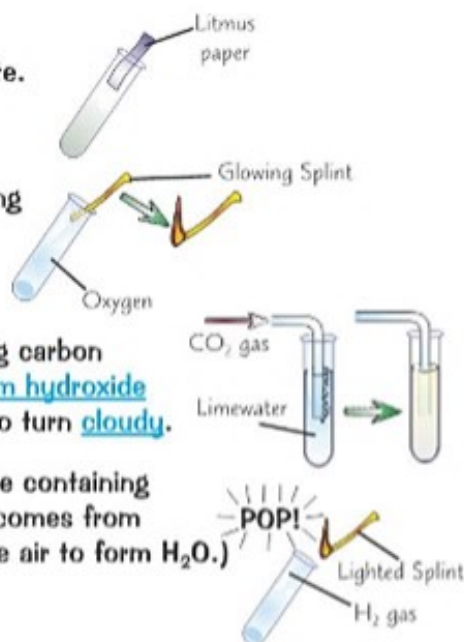
If you put a glowing splint inside a test tube containing oxygen, the oxygen will relight the glowing splint.

3) **Carbon Dioxide**

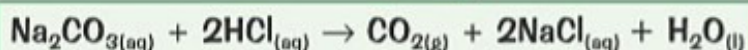
Bubbling carbon dioxide through (or shaking carbon dioxide with) an aqueous solution of calcium hydroxide (known as limewater) causes the solution to turn cloudy.

4) **Hydrogen**

If you hold a lit splint at the open end of a test tube containing hydrogen, you'll get a "squeaky pop". (The noise comes from the hydrogen burning quickly with the oxygen in the air to form  $H_2O$ .)

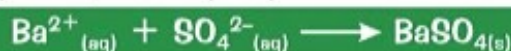
Tests for Anions Often **Give Precipitates****Dilute Acid Can Help Detect Carbonates**

Carbonates are substance that contain  $CO_3^{2-}$  ions. You can test whether a mystery solution contains carbonate ions by putting a sample in a test tube and then using a dropping pipette to add a couple of drops of dilute acid. You should then connect the test tube to a test tube of limewater. If carbonate ions are present, carbon dioxide will be released that will turn the limewater cloudy when it bubbles through it (as shown above). For example:

**Test for Sulfates with HCl and Barium Chloride**

To identify sulfate ions ( $SO_4^{2-}$ ), use a dropping pipette to add a couple of drops of dilute hydrochloric acid (HCl) followed by a couple of drops of barium chloride solution ( $BaCl_2$ ) to a test tube containing your mystery solution.

If sulfate ions are present, a white precipitate of barium sulfate will form:



Hydrochloric acid is added to get rid of any traces of carbonate ions before you do the test. These would also produce a precipitate, so they'd confuse the results.

**Test for Halides ( $Cl^-$ ,  $Br^-$ ,  $I^-$ ) with Nitric Acid and Silver Nitrate**

To identify a halide ion, add a couple of drops of dilute nitric acid ( $HNO_3$ ), followed by a couple of drops of silver nitrate solution,  $AgNO_3$ , to your mystery solution.

$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$  A chloride gives a white precipitate of silver chloride.

$Ag^+_{(aq)} + Br^-_{(aq)} \rightarrow AgBr_{(s)}$  A bromide gives a cream precipitate of silver bromide.

$Ag^+_{(aq)} + I^-_{(aq)} \rightarrow AgI_{(s)}$  An iodide gives a yellow precipitate of silver iodide.

**Hopefully this page won't be too testing for you...**

How to learn this page — don't stare at the whole thing till your eyes swim and you don't want to see the word 'precipitate' ever again. Each test has been given its own section, so take them one by one and learn it that way.

Q1 A student adds dilute nitric acid and silver nitrate solution to an aqueous solution of halide ions.

A white precipitate is formed. Suggest what halide may be present in the solution.

[1 mark]

# Tests for Cations

**PRACTICAL**

I haven't met a person who doesn't like flame tests for positive ions — everybody loves 'em...

## Flame Tests Identify Metal Ions

- Compounds of some metals burn with a characteristic colour.
- So you can test for various metal ions by heating your substance and seeing whether it burns with a distinctive colour flame:

Lithium ions,  $\text{Li}^+$ , burn with a crimson flame.

Sodium ions,  $\text{Na}^+$ , burn with a yellow flame.

Potassium ions,  $\text{K}^+$ , burn with a lilac flame.

Calcium ions,  $\text{Ca}^{2+}$ , burn with an orange-red flame.

Copper ions,  $\text{Cu}^{2+}$ , burn with a green flame.



Flame emission spectroscopy uses the light emitted by metal ions when they enter a flame to analyse and identify them. There's more on this on the next page.

- To do the test, you first need to clean a platinum wire loop by dipping it in some dilute HCl and then holding it in a blue flame from a Bunsen burner until it burns without any colour. Then, dip the loop into the sample you want to test and put it back in the flame. Record the colour of the flame.
- You can use these colours to detect and identify different ions. However it only works for samples that contain a single metal ion. If the sample tested contains a mixture of metal ions, the flame colours of some ions may be hidden by the colours of others.

## Some Metals Form a Coloured Precipitate with NaOH

- Many metal hydroxides are insoluble and precipitate out of solution when formed. Some of these hydroxides have a characteristic colour.
- So in this test you add a few drops of sodium hydroxide solution to a solution of your mystery compound — all in the hope of forming an insoluble hydroxide.
- If you get a coloured insoluble hydroxide you can often tell which metal was in the compound.



Metal Ions	Colour of Precipitate	Ionic Equation for Precipitate Formation
Calcium, $\text{Ca}^{2+}$	White	$\text{Ca}^{2+}_{(\text{aq})} + 2\text{OH}^{-}_{(\text{aq})} \rightarrow \text{Ca}(\text{OH})_{2(\text{s})}$
Copper(II), $\text{Cu}^{2+}$	Blue	$\text{Cu}^{2+}_{(\text{aq})} + 2\text{OH}^{-}_{(\text{aq})} \rightarrow \text{Cu}(\text{OH})_{2(\text{s})}$
Iron(II), $\text{Fe}^{2+}$	Green	$\text{Fe}^{2+}_{(\text{aq})} + 2\text{OH}^{-}_{(\text{aq})} \rightarrow \text{Fe}(\text{OH})_{2(\text{s})}$
Iron(III), $\text{Fe}^{3+}$	Brown	$\text{Fe}^{3+}_{(\text{aq})} + 3\text{OH}^{-}_{(\text{aq})} \rightarrow \text{Fe}(\text{OH})_{3(\text{s})}$
Aluminium, $\text{Al}^{3+}$	White at first. But then redissolves in excess NaOH to form a colourless solution.	$\text{Al}^{3+}_{(\text{aq})} + 3\text{OH}^{-}_{(\text{aq})} \rightarrow \text{Al}(\text{OH})_{3(\text{s})}$
Magnesium, $\text{Mg}^{2+}$	White	$\text{Mg}^{2+}_{(\text{aq})} + 2\text{OH}^{-}_{(\text{aq})} \rightarrow \text{Mg}(\text{OH})_{2(\text{s})}$

You don't need to know the equation for how the aluminium hydroxide precipitate dissolves.

## Cations — ions with a pawsitive charge...

Did you know that compounds containing metal cations are put into fireworks? And that when fireworks explode it's these metal cations that give fireworks their pretty colours? November 5th will never seem the same again...

- Q1 Why can you use a flame test to identify the metal ion in a compound? [1 mark]
- Q2 Sodium hydroxide is added to a solution of an unknown metal ion.  
A white precipitate is formed which redissolves following the addition of excess NaOH.
- Suggest what the unknown metal ion is. [1 mark]
  - Give the ionic equation for the formation of the precipitate. [1 mark]

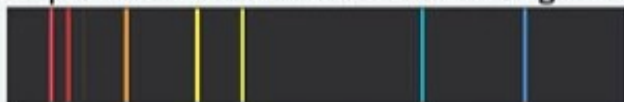
# Flame Emission Spectroscopy

As well as producing pretty colours, the science behind flame tests can be used to identify different metal ions in solution accurately and find their concentrations. Phwoar, colours and science. This is my kind of page...

## Every Metal Ion Gives a Characteristic Line Spectrum

- 1) During flame emission spectroscopy a sample is placed in a flame. As the ions heat up their electrons become excited. When the electrons drop back to their original energy levels, they transfer energy as light.
- 2) The light passes through a spectroscope, which can detect different wavelengths of light to produce a line spectrum.
- 3) The combination of wavelengths emitted by an ion depends on its charge and its electron arrangement (see page 20). Since no two ions have the same charge and the same electron arrangement, different ions emit different wavelengths of light. So each ion produces a different pattern of wavelengths, and has a different line spectrum.
- 4) The intensity of the spectrum indicates the concentration of that ion in solution.
- 5) This means that line spectra can be used to identify ions in solution and calculate their concentrations.

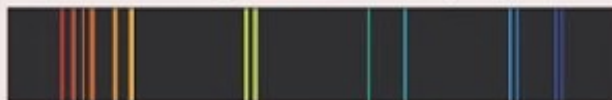
A line spectrum for an ion could look something like this:



## Flame Emission Spectroscopy Works for Mixtures

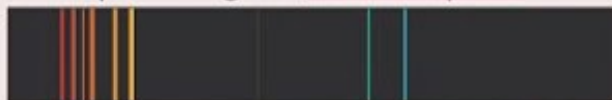
Flame emission spectroscopy can also be used to identify different ions in mixtures. This makes it more useful than flame tests, which only work for substances that contain a single metal ion.

For example, a solution containing different ions might give the following spectrum:

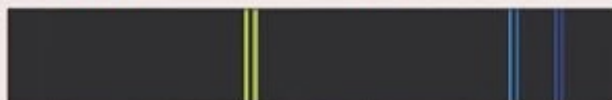


You can compare this spectrum against reference spectra:

Spectrum for ion X:



Spectrum for ion Y:



The combination of the spectra for ion X and for ion Y gives the spectra for the sample tested — this suggests the mixture contains ion X and ion Y.

## Machines can Analyse Unknown Substances

Chemists often use instrumental analysis (i.e. tests that use machines), such as flame emission spectroscopy, instead of conducting tests.

Advantages of Using Machines:

- Very sensitive — they can detect even the  tiniest amounts  of substances.
- Very fast and tests can be automated.
- Very accurate.

Must identify ions!



## Spectroscopy — it's a flaming useful technique...

As you can see, there are nifty ways of identifying substances as well as things like your bog standard flame test.

Q1 Why might it be necessary to use flame emission spectroscopy instead of a flame test to identify metal ions?

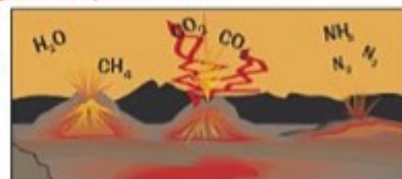
[2 marks]

# The Evolution of the Atmosphere

Theories for how the Earth's atmosphere evolved have changed a lot over the years — it's hard to gather evidence from such a long time period and from so long ago (4.6 billion years). Here's one idea we've got:

## Phase 1 — Volcanoes Gave Out Gases

- 1) The first billion years of Earth's history were pretty explosive — the surface was covered in volcanoes that erupted and released lots of gases. We think this was how the early atmosphere was formed.
- 2) The early atmosphere was probably mostly carbon dioxide, with virtually no oxygen. This is quite like the atmospheres of Mars and Venus today.
- 3) Volcanic activity also released nitrogen, which built up in the atmosphere over time, as well as water vapour and small amounts of methane and ammonia.

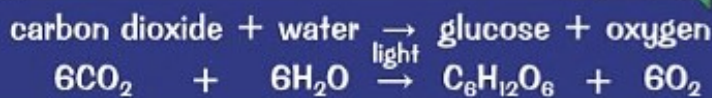


## Phase 2 — Oceans, Algae and Green Plants Absorbed Carbon Dioxide

- 1) When the water vapour in the atmosphere condensed, it formed the oceans.
- 2) Lots of carbon dioxide was removed from the early atmosphere as it dissolved in the oceans. This dissolved carbon dioxide then went through a series of reactions to form carbonate precipitates that formed sediments on the seabed.
- 3) Green plants and algae evolved and absorbed some of the carbon dioxide so that they could carry out photosynthesis (see below). Later, marine animals evolved. Their shells and skeletons contained carbonates from the oceans.
- 4) Some of the carbon these organisms took in from the atmosphere and oceans became locked up in rocks and fossil fuels after the organisms died.
  - When plants, plankton and marine animals die, they fall to the seabed and get buried by layers of sediment. Over millions of years, they become compressed and form sedimentary rocks, oil and gas — trapping the carbon within them and helping to keep carbon dioxide levels in the atmosphere reduced.
  - Things like coal, crude oil and natural gas that are made by this process are called 'fossil fuels'.
  - Crude oil and natural gas are formed from deposits of plankton. These fossil fuels form reservoirs under the seabed when they get trapped in rocks.
  - Coal is a sedimentary rock made from thick plant deposits.
  - Limestone is also a sedimentary rock. It's mostly made of calcium carbonate deposits from the shells and skeletons of marine organisms.

## Phase 3 — Green Plants and Algae Produced Oxygen

- 1) As well as absorbing the carbon dioxide in the atmosphere, green plants and algae produced oxygen by photosynthesis — this is when plants use light to convert carbon dioxide and water into sugars:
- 2) Algae evolved first — about 2.7 billion years ago. Then over the next billion years or so, green plants also evolved.
- 3) As oxygen levels built up in the atmosphere over time, more complex life (like animals) could evolve.
- 4) Eventually, about 200 million years ago, the atmosphere reached a composition similar to what it is today: approximately 80% nitrogen, 20% oxygen and small amounts of other gases (each only makes up less than 1% of the atmosphere), mainly carbon dioxide, noble gases and water vapour.



## The atmosphere's evolving — shut the window will you...

We've learnt about the atmosphere from Antarctic ice cores. Each year, a layer of ice forms with bubbles of air trapped in it. The deeper the ice, the older the air, so examining air in different layers shows us how it's changed.

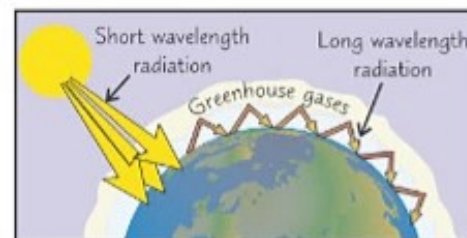
Q1 Describe how sedimentary rocks are formed.

[2 marks]

# Greenhouse Gases and Climate Change

Greenhouse gases are important but can also cause problems — it's all about keeping a delicate balance.

## Carbon Dioxide is a Greenhouse Gas



- 1) Greenhouse gases like carbon dioxide, methane and water vapour act like an insulating layer in the Earth's atmosphere — this, amongst other factors, allows the Earth to be warm enough to support life.
- 2) All particles absorb certain frequencies of radiation. Greenhouse gases don't absorb the incoming short wavelength radiation from the sun — but they do absorb the long wavelength radiation that gets reflected back off the Earth. Then they re-radiate it in all directions — including back towards the Earth. The longwave radiation is thermal radiation, so it results in warming of the surface of the Earth. This is the greenhouse effect.
- 3) Some forms of human activity affect the amount of greenhouse gases in the atmosphere. E.g:

- Deforestation: fewer trees means less  $\text{CO}_2$  is removed from the atmosphere via photosynthesis.
- Burning fossil fuels: carbon that was 'locked up' in these fuels is released as  $\text{CO}_2$ .
- Agriculture: more farm animals produce more methane through their digestive processes.
- Creating waste: more landfill sites and more waste from agriculture means more  $\text{CO}_2$  and methane released by decomposition of waste.

## Increasing Carbon Dioxide is Linked to Climate Change

- 1) The Earth's temperature varies naturally, but recently the average temperature of the Earth's surface has been increasing. Most scientists agree that the extra carbon dioxide from human activity is causing this increase and that this will lead to climate change.
- 2) Evidence for this has been peer-reviewed (see page 2) — so you know that the information out there is reliable.
- 3) Unfortunately, it's hard to fully understand the Earth's climate — this is because it's so complex, and there are so many variables, that it's very hard to make a model that isn't oversimplified.
- 4) This has led to speculation, particularly in the media — where stories may be biased or only some of the information given.

See page 3 for more on science in the media.

## Climate Change Could Have Dangerous Consequences

The Earth's climate is complex, but it's still important to make predictions about the consequences of climate change so that policy-makers can make decisions now. For example:

- 1) An increase in global temperature could lead to polar ice caps melting — causing a rise in sea levels, increased flooding in coastal areas and coastal erosion.
- 2) Changes in rainfall patterns (the amount, timing and distribution) may cause some regions to get too much or too little water. This, along with changes in temperature, may affect the ability of certain regions to produce food.
- 3) The frequency and severity of storms may also increase.
- 4) Changes in temperature and the amount of water available in a habitat may affect wild species, leading to differences in their distribution.

## Eee, problems, problems — there's always summat goin' wrong...

Everyone's talking about climate change these days — it's pretty scary stuff, so make sure you get it.

Q1 Describe three potential consequences of climate change.

[3 marks]



# Carbon Footprints

It's generally accepted that greenhouse gas emissions from human activities is causing climate change. Knowing what leads to a lot of emissions of carbon dioxide could be useful for stopping it happening.

## Carbon Footprints are Tricky to Measure

- 1) Carbon footprints are basically a measure of the amount of carbon dioxide and other greenhouse gases released over the full life cycle of something. That can be a service (e.g. the school bus), an event (e.g. the Olympics), a product (e.g. a toastie maker) — almost anything.
- 2) Measuring the total carbon footprint of something can be really hard, though — or even impossible.
- 3) That's because there are so many different factors to consider — for example, you would have to count the emissions released as a result of sourcing all the parts of your toastie maker and in making it, not to mention the emissions produced when you actually use it and finally dispose of it. Eugh, complicated...
- 4) Still, a rough calculation can give a good idea of what the worst emitters are — so that people can avoid them in the future.

## There are Ways of Reducing Carbon Footprints

You can't always measure a carbon footprint exactly, but there are always methods to try and reduce it. Anything that reduces the amount of greenhouse gases (e.g. carbon dioxide or methane) given out by a process will also reduce its carbon footprint. Here are some things that can be done:

- Renewable energy sources or nuclear energy could be used instead of fossil fuels.
- Using more efficient processes could conserve energy and cut waste. Lots of waste decomposes to release methane, so this will reduce methane emissions.
- Governments could tax companies or individuals based on the amount of greenhouse gases they emit — e.g. taxing cars based on the amount of carbon dioxide they emit over a set distance could mean that people choose to buy ones that are more fuel-efficient and so less polluting.
- Governments can also put a cap on emissions of all greenhouse gases that companies make — then sell licences for emissions up to that cap.
- There's also technology that captures the CO<sub>2</sub> produced by burning fossil fuels before it's released into the atmosphere — it can then be stored deep underground in cracks in the rock such as old oil wells.

## But Making Reductions is Still Difficult

- 1) It's easy enough saying that we should cut emissions, but actually doing it — that's a different story.
- 2) For a start, there's still a lot of work to be done on alternative technologies that result in lower CO<sub>2</sub> emissions.
- 3) A lot of governments are also worried that making these changes will impact on the economic growth of communities — which could be bad for people's well-being. This is particularly important for countries that are still developing.
- 4) Because not everyone is on board, it's hard to make international agreements to reduce emissions. Most countries don't want to sacrifice their economic development if they think that others won't do the same.
- 5) It's not just governments, though — individuals in developed countries need to make changes to their lifestyles. But it might be hard to get people to make changes if they don't want to and if there isn't enough education provided about why the changes are necessary and how to make them.

## Who has the biggest carbon footprint then? Clowns of course...

Carbon footprints are a game of 'fortunately/unfortunately'. Unfortunately, carbon emissions can lead to global warming. Fortunately, there are steps we can take to cut our carbon dioxide emissions. Unfortunately, not everyone's on board. Fortunately, as time goes on, people are doing more to reduce their emissions. And so on...

Q1 State two things governments can do to try to reduce the greenhouse gas emissions of businesses. [2 marks]

# Air Pollution

Increasing carbon dioxide is causing climate change. But  $\text{CO}_2$  isn't the only gas released when fossil fuels burn — you also get other nasties like oxides of nitrogen, sulfur dioxide and carbon monoxide.

## Combustion of Fossil Fuels Releases Gases and Particles

- 1) Fossil fuels, such as crude oil and coal, contain hydrocarbons. During combustion, the carbon and hydrogen in these compounds are oxidised so that carbon dioxide and water vapour are released into the atmosphere.
- 2) When there's plenty of oxygen, all the fuel burns — this is called complete combustion.
- 3) If there's not enough oxygen, some of the fuel doesn't burn — this is called incomplete combustion. Under these conditions, solid particles (called particulates) of soot (carbon) and unburnt fuel are released and carbon monoxide can be produced as well as carbon dioxide.

Hydrocarbons are compounds that only contain hydrogen and carbon (see page 75).

- 4) Particulates in the air can cause all sorts of problems:

- If particulates are inhaled, they can get stuck in the lungs and cause damage. This can then lead to respiratory problems.
- They're also bad for the environment — they themselves, or the clouds they help to produce, reflect sunlight back into space. This means that less light reaches the Earth — causing global dimming.

There's more about complete combustion on p.75 and incomplete combustion on p.78.

- 5) It's not just particulates from incomplete combustion that cause problems. Carbon monoxide is pretty nasty too.

- Carbon monoxide ( $\text{CO}$ ) is really dangerous because it can stop your blood from doing its proper job of carrying oxygen around the body.
- It does this by binding to the haemoglobin in your blood that normally carries  $\text{O}_2$  — so less oxygen is able to be transported round your body.
- A lack of oxygen in the blood can lead to fainting, a coma or even death.
- Carbon monoxide doesn't have any colour or smell, so it's very hard to detect. This makes it even more dangerous.

## Sulfur Dioxide and Oxides of Nitrogen Can be Released

- 1) Sulfur dioxide ( $\text{SO}_2$ ) is released during the combustion of fossil fuels, such as coal, that contain sulfur impurities — the sulfur in the fuel becomes oxidised.
- 2) Nitrogen oxides are created from a reaction between the nitrogen and oxygen in the air, caused by the heat of the burning. (This can happen in the internal combustion engines of cars.)
- 3) When these gases mix with clouds they form dilute sulfuric acid or dilute nitric acid. This then falls as acid rain.
- 4) Acid rain kills plants and damages buildings and statues. It also makes metal corrode. It's shocking.
- 5) Not only that, but sulfur dioxide and nitrogen oxides can also be bad for human health — they cause respiratory problems if they're breathed in.

$\text{SO}_2$  gas reacts with water to form sulfuric acid. You can test for sulfur impurities in a fuel by bubbling the gases from combustion through a solution containing Universal indicator — if the fuel contains sulfur, the gases will contain  $\text{SO}_2$  which will form sulfuric acid and turn the Universal indicator red.



## Revision and pollution — the two bugbears of modern life...

Eeee.... cars and fossil fuels — they're nowt but trouble. But at least this topic is kind of interesting, what with its relevance to everyday life and all. Just think... you could see this kind of stuff on TV.

- Q1 Name three potential pollutants that could be released as a result of incomplete combustion of hydrocarbons, that wouldn't be released as a result of complete combustion.

[3 marks]

## Revision Questions for Topics 8 & 9

That's all for [Topics 8 and 9](#), but before you breathe a sigh of relief, there are some questions to get through.

- Try these questions and [tick off each one](#) when you [get it right](#).
- When you've done [all the questions](#) under a heading and are [completely happy](#) with it, tick it off.

### Purity, Formulations and Paper Chromatography (p.86-87) ☐

- 1) What is a formulation? ☐
- 2) What are the two phases called in chromatography? ☐
- 3) In paper chromatography, how many spots will a pure substance form on the paper? ☐
- 4) Give the formula for working out the  $R_f$  value of a substance. ☐
- 5) Would you expect the  $R_f$  value of a substance to change if you changed the solvent used in the chromatography experiment? ☐

### Chemical Tests (p.88-89) ☐

- 6) What colour does litmus paper turn in the presence of chlorine? ☐
- 7) Describe a test you could use to test for the presence of carbonate ions. ☐
- 8) How do you conduct a flame test? ☐
- 9) What colour precipitate do iron(II) compounds form with sodium hydroxide? ☐
- 10) What occurs on addition of excess sodium hydroxide to a solution containing  $Al^{3+}$  ions? ☐
- 11) Give the ionic equation for the formation of a precipitate when a solution containing magnesium ions is reacted with sodium hydroxide. ☐

### Flame Emission Spectroscopy (p.90) ☐

- 12) Describe how flame emission spectroscopy works. ☐
- 13) What does the intensity of a flame emission spectrum of a solution relate to? ☐
- 14) Give three advantages of using machines to conduct chemical analysis. ☐

### The Evolution of the Atmosphere (p.91) ☐

- 15) How do scientists think the atmosphere was formed during the first billion years or so of Earth's history? ☐
- 16) Name five gases that scientists think were present in the early atmosphere. ☐
- 17) Describe how the levels of carbon dioxide in the atmosphere were reduced. ☐
- 18) Write the balanced chemical equation for photosynthesis. ☐
- 19) State the approximate composition of the atmosphere today. ☐

### Pollution and Climate Change (p.92-94) ☐

- 20) Name three greenhouse gases. ☐
- 21) Explain how the greenhouse effect works to keep the Earth warm. ☐
- 22) State three ways in which human activity is leading to an increase in carbon dioxide in the atmosphere. ☐
- 23) What is a carbon footprint? ☐
- 24) Explain why reducing carbon dioxide emissions can be a difficult issue. ☐
- 25) Describe how the following air pollutants are produced:
  - a) particulates,
  - b) carbon monoxide,
  - c) sulfur dioxide,
  - d) nitrogen oxides.☐
- 26) Why is carbon monoxide dangerous? ☐
- 27) Explain how acid rain forms. ☐
- 28) State two problems caused by acid rain. ☐

# Ceramics, Composites and Polymers

We've got very good at using **materials** from the earth to produce things we can use to make our lives **easier**.

## Ceramics Come in Many Different Forms

**Ceramics** are **non-metal** solids with high **melting points** that aren't made from carbon-based compounds.

- 1) Some ceramics can be made from **clay**.
- 2) **Clay** is a **soft** material when it's dug up out of the ground, so can be **moulded** into different **shapes**.
- 3) When it's **fired** at high temperatures, it **hardens** to form a clay ceramic.
- 4) Its ability to be moulded when wet and then hardened makes clay **ideal** for making **pottery** and **bricks**.
- 5) Another example of a ceramic is **glass**. Glass is generally **transparent**, can be **moulded** when hot and can be **brittle** when thin.
- 6) Most glass made is **soda-lime glass**, which is made by heating a mixture of **limestone**, **sand** and **sodium carbonate** (soda) until it melts. When the mixture cools it comes out as **glass**.
- 7) **Borosilicate glass** has a higher **melting point** than soda-lime glass. It's made in the same way as soda-lime glass, using a mixture of **sand** and **boron trioxide**.

## Composites are Generally Made of Two Different Materials

**Composites** are made of one material **embedded** in another. **Fibres** or **fragments** of a material (known as the **reinforcement**) are surrounded by a **matrix** acting as a **binder**. The **properties** of a composite depend on the properties of the materials it is **made from**. For example:

- 1) **Fibreglass** consists of fibres of **glass** embedded in a matrix made of **polymer** (plastic). It has a **low density** (like plastic) but is **very strong** (like glass). It's used for things like skis, boats and surfboards.
- 2) **Carbon fibre** composites also have a polymer matrix. The reinforcement is either made from long chains of carbon atoms bonded together (carbon fibres) or from carbon nanotubes (see p.34). These composites are very **strong** and **light** so are used in aerospace and sports car manufacturing.
- 3) **Concrete** is made from **aggregate** (a mixture of sand and gravel) embedded in **cement**. It's very **strong**. This makes it ideal for use as a **building material**, e.g. in skate parks.
- 4) **Wood** is a natural composite of **cellulose fibres** held together by an organic polymer matrix.

## Polymers Can Have Very Different Properties

Two important things can **influence** the **properties** of a polymer — **how** it's made and **what** it's made from.

For example, the properties of poly(ethene) depend on the **catalyst** that was used and the **reaction conditions** (the temperature and pressure) that it was made under.

- **Low density** (LD) poly(ethene) is made from ethene at a **moderate temperature** under a **high pressure** and with a catalyst. It's **flexible** and is used for bags and bottles.
- **High density** (HD) poly(ethene) is also made from ethene but at a **lower** temperature and pressure with a different catalyst. It's **more rigid** and is used for water tanks and drainpipes.

The **monomers** that a polymer is made from determine the type of bonds that form between the polymer chains. These weak bonds **between** the different molecule chains determine the **properties** of the polymer:

- **Thermosoftening polymers** contain individual polymer chains entwined together with weak forces between the chains. You can melt these plastics and remould them.
- **Thermosetting polymers** contain **monomers** that can form **cross-links** between the polymer chains, holding the chains together in a **solid structure**. Unlike thermosoftening polymers, these polymers **don't soften** when they're **heated**. Thermosetting polymers are strong, hard and rigid.

## People in soda-lime glass houses shouldn't throw heated objects...

Modern day composites are amazing. I reckon one will end up running the country — David Carbonron.

- Q1 Give one difference in properties between thermosetting and thermosoftening polymers. [1 mark]

# Properties of Materials

It's all very well making something, but the material it's made from needs to be fit for purpose.

## Different Materials are Suited to Different Jobs

What materials are used for depends on their properties. You need to be able to interpret information about the properties of materials and decide how suitable these materials would be for different purposes.

**Ceramics** include glass and clay ceramics such as porcelain and bricks. They're insulators of heat and electricity, brittle (they aren't very flexible and break easily) and stiff.

**Polymers** are insulators of heat and electricity, they can be flexible (they can be bent without breaking) and are easily moulded. Polymers have many applications including in clothing and insulators in electrical items.

The properties of composites depend on the matrix/binder and the reinforcement used to make them, so they have many different uses.

**Metals** are malleable, good conductors of heat and electricity, ductile (they can be drawn into wires), shiny and stiff (see p.23). Metals have many uses, including in electrical wires, car body-work, and cutlery.

### EXAMPLE:

Suggest which material would be best for making the hull of a large ship.

Material	Brittleness	Water resistance
Clay ceramic	brittle	porous (absorbs water)
Rigid PVC (a polymer)	brittle	non-porous, corrosion resistant
Low carbon steel	tough	non-porous, corrodes in water, but cheap to protect against corrosion.

Low carbon steel is non-porous, strong and can be easily protected against corrosion.

- 1) A ship's hull needs to be non-porous so it doesn't absorb water, and shouldn't corrode. This rules out clay ceramic — it absorbs water so the ship could sink. Low carbon steel corrodes when unprotected, but it's cheap to protect so could still be suitable.
- 2) If the hull's brittle, it could suddenly break. Rigid PVC is brittle, so not suitable.

Low carbon steel looks like the best choice.

## Pure Metals Don't Always Have the Properties Needed

- 1) As you saw on page 35, the regular structure of pure metals makes them soft — often too soft for use in everyday life.
- 2) Alloys are made by adding another element to the metal. This disrupts the structure of the metal, making alloys harder than pure metals.
- 3) For example, alloys of iron called steels are often used instead of pure iron. Steels are made by adding small amounts of carbon and sometimes other metals to iron.
- 4) Many other alloys are used in everyday life.

TYPE OF STEEL	PROPERTIES	USE
Low carbon steel (0.1–0.3% carbon)	easily shaped	car bodies
High carbon steel (0.22–2.5% carbon)	very strong, inflexible, brittle	bridges
Stainless steel (chromium added, and sometimes nickel)	corrosion-resistant, hard	cutlery

An unsuitable material.



- **Bronze = Copper + Tin:** Bronze is harder than copper. It's used to make medals, decorative ornaments and statues.
- **Brass = Copper + Zinc:** Brass is more malleable than bronze and is used in situations where lower friction is required, such as in water taps and door fittings.
- **Gold alloys are used to make jewellery:** Pure gold is very soft. Metals such as zinc, copper, and silver are used to harden the gold. Pure gold is described as 24 carat, so 18 carats means that 18 out of 24 parts of the alloy are pure gold. In other words, 18 carat gold is 75% gold.
- **Aluminium alloys are used to make aircraft:** Aluminium has a low density which is an important property in aircraft manufacture. But pure aluminium is too soft for making aeroplanes so it's alloyed with small amounts of other metals to make it stronger.

For more on the structure of metals have a look at p.35.

## A brass band — harder than Iron Maiden...

There's a perfect purpose for every material out there — makes me believe in love again.

Q1 To the nearest percent, what percentage of pure gold is in 14 carat gold?

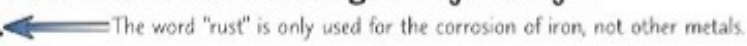
[1 mark]

# Corrosion

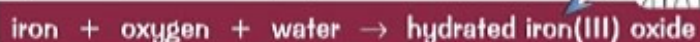
Some metals react with substances in their environment, and this can weaken them.

## Iron and Steel Corrode Much More than Aluminium

Corrosion is where metals react with substances in their environment and are gradually destroyed.

- 1) Iron corrodes easily. In other words, it rusts.  The word "rust" is only used for the corrosion of iron, not other metals.
- 2) In order to rust, iron needs to be in contact with both oxygen and water, which are present in air.
- 3) The stuff we call rust is actually the compound hydrated iron(III) oxide.

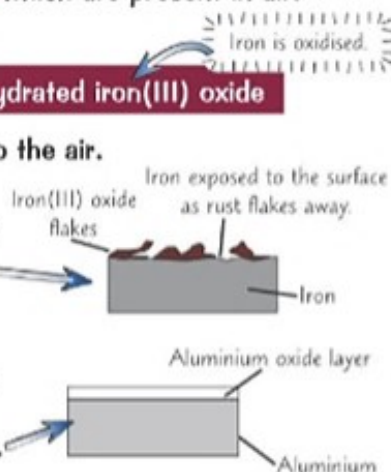
Here's the equation for it's formation:



- 4) Corrosion only happens on the surface of a material, where it's exposed to the air.

- 5) Unfortunately, rust is a soft crumbly solid that soon flakes off to leave more iron available to rust again. This means that, eventually, all the iron in an object corrodes away even if it wasn't initially at the surface.

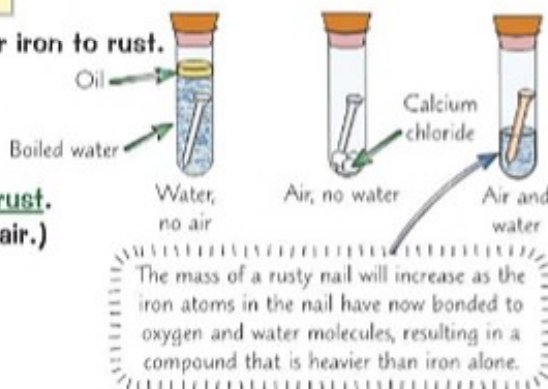
- 6) Aluminium also corrodes when exposed to air. Unlike iron objects, things made from aluminium aren't completely destroyed by corrosion. This is because the aluminium oxide that forms when aluminium corrodes doesn't flake away. In fact, it forms a nice protective layer that sticks firmly to the aluminium below and stops any further reaction taking place.



## Both Air and Water are Needed for Iron to Rust

Experiments can show that both oxygen and water are needed for iron to rust.

- 1) If you put an iron nail in a boiling tube with just water, it won't rust. (The water is boiled to remove oxygen and oil is used to stop air getting in.)
- 2) If you put an iron nail in a boiling tube with just air, it won't rust. (Calcium chloride can be used to absorb any water from the air.)
- 3) However, if you put an iron nail in a boiling tube with air and water, it will rust.



## There are Two Main Ways to Prevent Rusting

- 1) The obvious way to prevent rusting is to coat the iron with a barrier to keep out the water and oxygen. This can be done by:
  - Painting/Coating with plastic — ideal for big and small structures alike. It can be decorative too.
  - Electroplating — this uses electrolysis to reduce metal ions onto an iron electrode. It can be used to coat the iron with a layer of a different metal that won't be corroded away.
  - Oiling/Greasing — this has to be used when moving parts are involved, like on bike chains.
- 2) Another method is the sacrificial method. This involves placing a more reactive metal such as zinc or magnesium with the iron. Water and oxygen then react with the sacrificial metal instead of with the iron.
- 3) Some protection techniques employ both the methods above. For example:

An object can be galvanised by spraying it with a coating of zinc. The zinc layer is firstly protective, but if it's scratched, the zinc around the site of the scratch works as a sacrificial metal.



## The sacrificial method — chemistry sure is bloodthirsty...

If you've wondered how Iron Man avoids rusting, he greases himself up and sends out Zinc Man if it's raining.

- Q1 Two nails are left in a test tube containing water and air for a period of time. One nail is galvanised, the other is pure iron. Predict which nail will rust and give a reason for your answer. [2 marks]

# Finite and Renewable Resources

There are lots of different resources that humans use to provide **energy** for things like **heating** or **travelling**, as well as for **building materials** and **food**. Some of these resources get replaced, some don't.

## Natural Resources Come From the Earth, Sea and Air

- 1) Natural resources form without **human input**. They include anything that comes from the earth, sea or air. For example, cotton for clothing or oil for fuel.
- 2) Some of these natural products can be **replaced** by synthetic products or **improved upon** by man-made processes. For example, **rubber** is a natural product that can be extracted from the sap of a tree, however man-made **polymers** have now been made which can **replace** rubber in uses such as tyres.
- 3) **Agriculture** provides **conditions** where **natural resources** can be enhanced for our needs. E.g. the development of fertilisers have meant we can produce a high yield of crops.

## Some Natural Resources will Run Out

- 1) **Renewable resources** reform at a similar rate to, or faster than, we use them.
- 2) For example, timber is a renewable resource as trees can be planted following a harvest and only take a few years to regrow. Other examples of renewable resources include fresh water and food.
- 3) **Finite (non-renewable) resources**, aren't formed quickly enough to be considered replaceable.
- 4) Finite resources include **fossil fuels** and **nuclear fuels** such as **uranium** and **plutonium**. **Minerals** and **metals** found in **ores** in the earth are also non-renewable materials.
- 5) After they've been **extracted**, many finite resources undergo **man-made processes** to provide fuels and materials necessary for modern life. For example, **fractional distillation** (see p.76) is used to produce usable products such as petrol from crude oil and metal ores are **reduced** to produce a pure metal (see p.56).



## Tables, Charts and Graphs can Give You an Insight Into Different Resources

You may be asked to **interpret** information about resources in the exam.

### EXAMPLE:

The table below shows information for two resources, coal and timber. Identify which resource is which.

	Energy Density (MJ/m <sup>3</sup> )	Time it takes to form
Resource 1	7600-11400	10 years
Resource 2	23000-26000	10 <sup>6</sup> years

The time it takes for Resource 1 to reform is 10<sup>5</sup> times shorter than Resource 2 suggesting it is a renewable resource. Resource 1 is also a far less energetic fuel than Resource 2, so is more likely to be timber than coal.

Resource 1 is timber and Resource 2 is coal.

10<sup>6</sup> is a shorthand way of showing 1 000 000. This is because 10<sup>6</sup> = 10 × 10 × 10 × 10 × 10 × 10 = 1 000 000.

## Extracting Finite Resources has Risks

- 1) Many modern materials are made from **raw, finite resources**, for example most plastics, metals and building materials.
- 2) People have to balance the **social**, **economic** and **environmental** effects of extracting finite resources.
- 3) For example, mining metal ores is **good** because **useful products** can be made. It also provides local people with **jobs** and brings **money** into the area. However, mining ores is **bad for the environment** as it uses loads of energy, scars the landscape, produces lots of waste and destroys habitats.

## This book is a renewable resource — a gift that keeps on giving...

Unfortunately we can't just run around using every resource we get our hands on — we have to consider the impacts of our actions. If you ever start a major mining project think... What would David Attenborough do?

Q1 Using examples, state the difference between a finite and renewable resource.

[2 marks]

# Reuse and Recycling

Many materials used in the modern world are **limited**. Once they're finished with, it's usually far better to **recycle** them than to use new finite resources which will eventually run out.

## Chemistry is Improving Sustainability

- 1) **Sustainable development** is an approach to development that takes account of the needs of **present society** while not damaging the lives of **future generations**.
- 2) As you saw on the last page, not all resources are **renewable** so it's **unsustainable** to keep using them.
- 3) As well as using resources, **extracting** resources can be unsustainable due to the amount of **energy** used and **waste** produced. **Processing** the resources into useful materials, such as **glass** or **bricks**, can be unsustainable too, as the processes often use **energy** that's made from **finite resources**.
- 4) One way of reducing the use of finite resources is for people to use **less**. This doesn't just reduce the use of that resource but also anything needed to produce it.
- 5) We can't stop using finite resources altogether, but chemists can **develop** and **adapt** processes that use **lower amounts** of **finite resources** and **reduce** damage to the environment. For example, chemists have developed **catalysts** that **reduce** the amount of **energy** required for certain industrial processes.

## Copper-Rich Ores are in Short Supply

- 1) Copper is a finite resource. One way to improve its sustainability is by extracting it from **low-grade ores** (ores without much copper in). Scientists are looking into new ways of doing this:
  - **Bioleaching** — **bacteria** are used to convert copper compounds in the ore into soluble copper compounds, separating out the copper from the ore in the process. The **leachate** (the solution produced by the process) contains copper ions, which can be extracted, e.g. by electrolysis (see p.58) or displacement (see p.57) with a more reactive metal, e.g. scrap iron.
  - **Phytomining** — this involves growing **plants** in **soil** that **contains copper**. The plants **can't use** or **get rid** of the copper so it gradually **builds up** in the **leaves**. The plants can be **harvested**, **dried** and **burned** in a furnace. The ash contains soluble copper compounds from which copper can be extracted by electrolysis or displacement using scrap iron.
- 2) **Traditional methods** of copper mining are pretty **damaging** to the **environment**. These new methods of extraction have a much **smaller impact**, but the disadvantage is that they're **slow**.

These methods can be used to extract other metals too.

## Recycling Metals is Important

- 1) **Mining** and **extracting** metals takes lots of **energy**, most of which comes from burning **fossil fuels**.
- 2) Recycling metals often uses much **less energy** than is needed to mine and extract new metal, **conserves** the finite amount of each metal in the earth and cuts down on the amount of **waste** getting sent to **landfill**.
- 3) Metals are usually recycled by **melting** them and then **casting** them into the shape of the new product.
- 4) Depending on what the metal will be used for after recycling, the amount of **separation** required for recyclable metals can change. For example, waste steel and iron can be kept together as they can both be added to iron in a **blast furnace** to reduce the amount of iron ore required.

Recycling is way to reduce our need for copper rich ores.

A blast furnace is used to extract iron from its ore at a high temperature using carbon.

## Glass can Also be Recycled

**Glass recycling** can help **sustainability** by reducing the amount of energy needed to make new glass products, and also the amount of waste created when used glass is thrown away.

- 1) **Glass bottles** can often be **reused** without reshaping.
- 2) Other forms of glass can't be reused so they're **recycled** instead. Usually the glass is separated by **colour** and **chemical composition** before being recycled.
- 3) The glass is crushed and then melted to be reshaped for use in glass products such as **bottles** or jars. It might also be used for a **different** purpose such as **insulating** glass wool for wall insulation in homes.

## CGP Jokes — 85% recycled since 1996...

Recycling is really handy — as well as saving limited finite materials it also saves energy.

Q1 Give three positive effects of recycling metals.

[3 marks]

# Life Cycle Assessments

If a company wants to manufacture a new product, they carry out a life cycle assessment (LCA).

## Life Cycle Assessments Show Total Environmental Costs

A life cycle assessment (LCA) looks at every stage of a product's life to assess the impact it would have on the environment.

### 1 Getting the Raw Materials:

- 1) Extracting raw materials needed for a product can damage the local environment, e.g. mining metals. Extraction can also result in pollution due to the amount of energy needed.
- 2) Raw materials often need to be processed to extract the desired materials and this often needs large amounts of energy. E.g. extracting metals from ores or fractional distillation of crude oil.

### 2 Manufacture and Packaging:

- 1) Manufacturing products and their packaging can use a lot of energy resources and can also cause a lot of pollution, e.g. harmful fumes such as carbon monoxide or hydrogen chloride.
- 2) You also need to think about any waste products and how to dispose of them. The chemical reactions used to make compounds from their raw materials can produce waste products. Some waste can be turned into other useful chemicals, reducing the amount that ends up polluting the environment.

### 3 Using the Product:

- 1) The use of a product can damage the environment. For example, burning fuels releases greenhouse gases and other harmful substances. Fertilisers can leach into streams and rivers causing damage to ecosystems.
- 2) How long a product is used for or how many uses it gets is also a factor — products that need lots of energy to produce but are used for ages mean less waste in the long run.

### 4 Product Disposal:

- 1) Products are often disposed of in landfill sites. This takes up space and pollutes land and water, e.g. if paint washes off a product and gets into rivers.
- 2) Energy is used to transport waste to landfill, which causes pollutants to be released into the atmosphere.
- 3) Products might be incinerated (burnt), which causes air pollution.

## You Can Compare Life Cycle Assessments for Plastic and Paper Bags

Life Cycle Assessment Stage	Plastic Bag	Paper Bag
Raw Materials	Crude oil	Timber
Manufacturing and Packaging	The compounds needed to make the plastic are extracted from crude oil by fractional distillation, followed by cracking and then polymerisation. Waste is reduced as the other fractions of crude oil have other uses.	Pulped timber is processed using lots of energy. Lots of waste is made.
Using the Product	Can be reused. Can be used for other things as well as shopping, for example bin liners.	Usually only used once.
Product Disposal	Recyclable but not biodegradable and will take up space in landfill and pollute land.	Biodegradable, non-toxic and can be recycled.

Life cycle assessments have shown that even though plastic bags aren't biodegradable, they take less energy to make and have a longer lifespan than paper bags, so may be less harmful to the environment.

## There are Problems with Life Cycle Assessments

- 1) The use of energy, some natural resources and the amount of certain types of waste produced by a product over its lifetime can be easily quantified. But the effect of some pollutants is harder to give a numerical value to. E.g. it's difficult to apply a value to the negative visual effects of plastic bags in the environment compared to paper ones.
- 2) So, producing an LCA is not an objective method as it takes into account the values of the person carrying out the assessment. This means LCAs can be biased.
- 3) Selective LCAs, which only show some of the impacts of a product on the environment can also be biased as they can be written to deliberately support the claims of a company, in order to give them positive advertising.

## Need exercise? Go life-cycling then...

In the exam you may be asked to evaluate the use of different materials for a particular product, using an LCA.

Q1 What are the four stages that need to be considered to conduct a life cycle assessment? [4 marks]

# Potable Water

We all need safe drinking water. The way that water's made safe depends on local conditions.

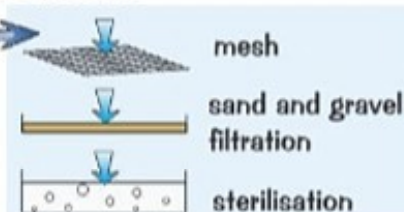
## Potable Water is Water You Can Drink

- 1) Potable water is water that's been treated or is naturally safe for humans to drink — it's essential for life.
- 2) Chemists wouldn't call it pure, though. Pure water only contains  $H_2O$  molecules whereas potable water can contain lots of other dissolved substances.
- 3) The important thing is that the levels of dissolved salts aren't too high, that it has a pH between 6.5 and 8.5 and also that there aren't any nasties (like bacteria or other microbes) swimming around in it.

## The Way that Potable Water is Produced Depends on Where You Are

- 1) Rainwater is a type of fresh water. Fresh water is water that doesn't have much dissolved in it.
- 2) When it rains, water can either collect as surface water (in lakes, rivers and reservoirs) or as groundwater (in rocks called aquifers that trap water underground).
- 3) In the UK, the source of fresh water used depends on location. Surface water tends to dry up first, so in warm areas, e.g. the south-east, most of the domestic water supply comes from groundwater.
- 4) Even though it only has low levels of dissolved substances, water from these fresh water sources still needs to be treated to make it safe before it can be used. This process includes:

- Filtration — a wire mesh screens out large twigs etc, and then gravel and sand beds filter out any other solid bits.
- Sterilisation — the water is sterilised to kill any harmful bacteria or microbes. This can be done by bubbling chlorine gas through it or by using ozone or ultraviolet light.



- 5) In some very dry countries, e.g. Kuwait, there's not enough surface or groundwater and instead sea water must be treated by desalination to provide potable water.

- 6) Distillation can be used to desalinate sea water.

Here's how you can test and distil water in the lab:

Chemicals can also be added to the water supply, such as fluoride (which is good for teeth). This is controversial, because people aren't given any choice over whether they consume them or not.

### PRACTICAL

- First, test the pH of the water using a pH meter. If the pH is too high or too low, you'll need to neutralise it. You can do this with a titration (see p.52), but use a pH meter to tell you when the solution's neutral, rather than an indicator, as this won't contaminate the water.
- You should also test the water for the presence of sodium chloride (the main salt in seawater). To test for sodium ions, do a flame test on a small sample (see p.89). If sodium ions are present the flame will turn yellow. To test for chloride ions, take another sample of your water and add a few drops of dilute nitric acid, followed by a few drops of silver nitrate solution. If chloride ions are present, a white precipitate will form.
- To distil the water, pour the salty water into a distillation apparatus, like the one on p.18. Heat the flask from below. The water will boil and form steam, leaving any dissolved salts in the flask. The steam will condense back to liquid water in the condenser and can be collected as it runs out.
- Then, retest the distilled water for sodium chloride to check that it has been removed. Also retest the pH of the water with a pH meter to check that it's neutral (has a pH of 7).

- 7) Sea water can also be treated by processes that use membranes — like reverse osmosis. The salty water is passed through a membrane that only allows water molecules to pass through. Ions and larger molecules are trapped by the membrane so separated from the water.
- 8) Both of distillation and reverse osmosis need loads of energy, so they're really expensive and not practical for producing large quantities of fresh water.

## Potable water — nothing to do with gardening...

Distilling salty water could be useful if you ever end up stranded on a desert island. Oh, and for your exams too.

Q1 Describe the steps you could take to treat water from a groundwater source to make it potable. [2 marks]

# Waste Water Treatment

It might not be pretty, but dealing with our waste is really important to ensure that we don't **pollute** the natural environment — it also means that we can access **nice clean water**. Super.

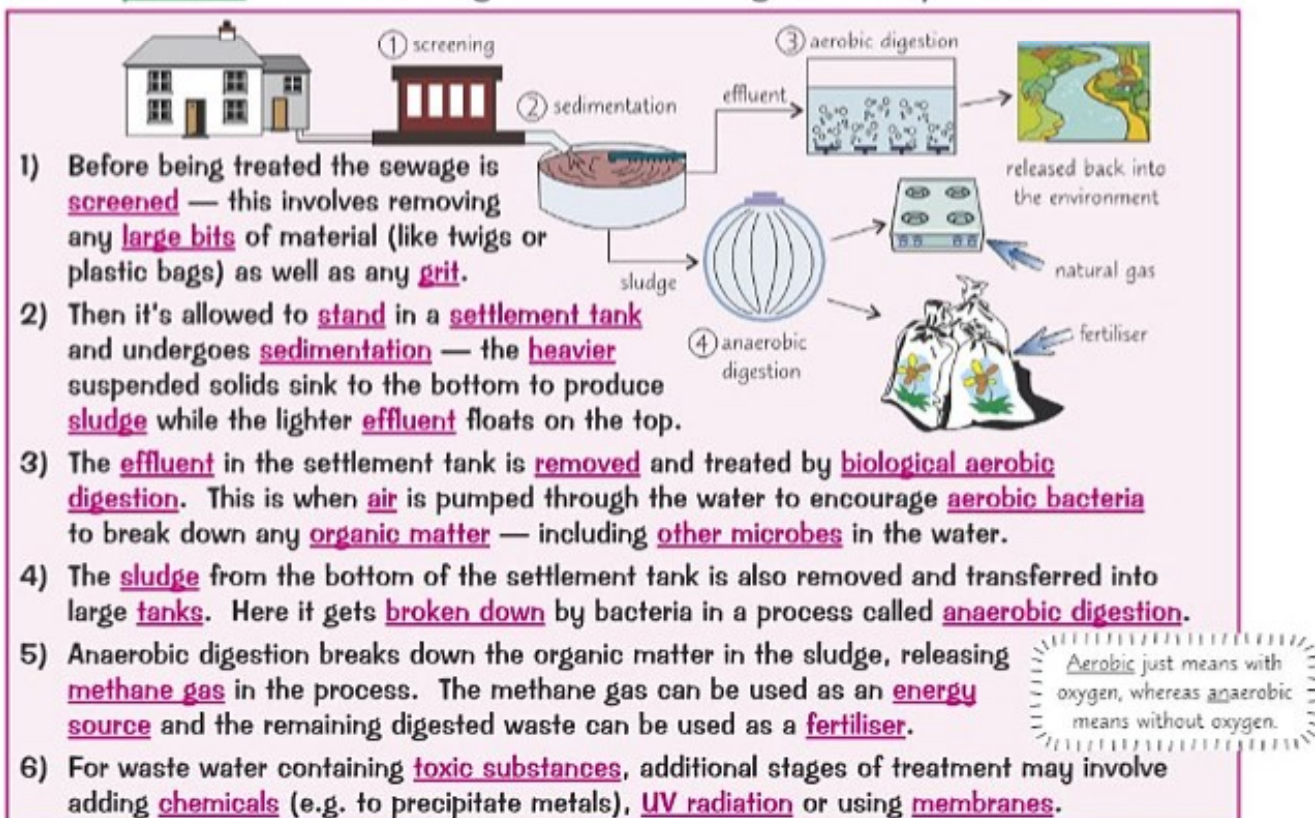
## Waste Water Comes from Lots of Different Sources

- 1) We use water for lots of things at home — like **having a bath**, going to the **toilet**, doing the **washing-up**, etc. When you flush this water down the drain, it goes into the **sewers** and towards **sewage treatment plants**.
- 2) **Agricultural systems** also produce a lot of waste water including **nutrient run-off** from fields and **slurry** from animal farms.
- 3) Sewage from **domestic** or **agricultural** sources has to be **treated** to remove any **organic matter** and **harmful microbes** before it can be put **back** into freshwater sources like **rivers** or **lakes**. Otherwise it would make them very **polluted** and would pose **health risks**.
- 4) **Industrial processes** such as the Haber Process (see the next page) also produce a lot of waste water that has to be **collected** and **treated**.
- 5) As well as **organic matter**, industrial waste water can also contain **harmful chemicals** — so it has to undergo **additional stages** of treatment before it is safe to release it into the environment.



## Sewage Treatment Happens in Several Stages

Some of the **processes** involved in treating waste water at sewage treatment plants include:



Sewage treatment requires **more processes** than treating **fresh water** but uses **less energy** than the **desalination** of **salt water**, so could be used as an alternative in areas where there's not much fresh water. For example, **Singapore** is treating waste water and recycling it back into drinking supplies. However, people don't like the idea of drinking water that used to be sewage.

## Is it just me, or does this page stink? Phew...

Modern sewage systems have done wonders to make life in developed countries much less... well... smelly.

Q1 List three stages of treatment that domestic sewage undergoes at a sewage treatment plant. [3 marks]

# The Haber Process

This is an important industrial process. It produces ammonia ( $\text{NH}_3$ ), which is used to make fertilisers.

## Nitrogen and Hydrogen are Needed to Make Ammonia

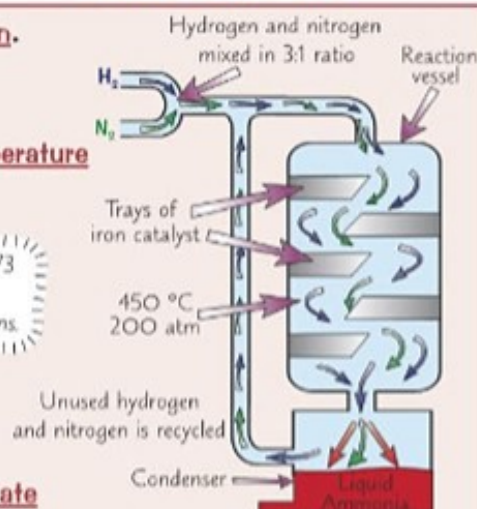
The Haber process is used to make ammonia from hydrogen and nitrogen using the following reaction:



This reaction is well suited for an industrial scale as the reactants aren't too difficult or expensive to obtain.

- 1) The nitrogen is obtained easily from the air, which is 78% nitrogen.
- 2) The hydrogen mainly comes from reacting methane (from natural gas) with steam to form hydrogen and carbon dioxide.
- 3) The reactant gases are passed over an iron catalyst. A high temperature ( $450^\circ\text{C}$ ) and a high pressure (200 atmospheres) are used.
- 4) Because the reaction is reversible (it occurs in both directions), some of the ammonia produced converts back into hydrogen and nitrogen again. It eventually reaches a dynamic equilibrium.
- 5) The ammonia is formed as a gas, but as it cools in the condenser it liquefies and is removed. The unused hydrogen ( $\text{H}_2$ ), and nitrogen, ( $\text{N}_2$ ), are recycled, so nothing is wasted.
- 6) The ammonia produced can then be used to make ammonium nitrate — a very nitrogen-rich fertiliser (see next page).

See pages 72-73  
for more on  
reversible reactions.

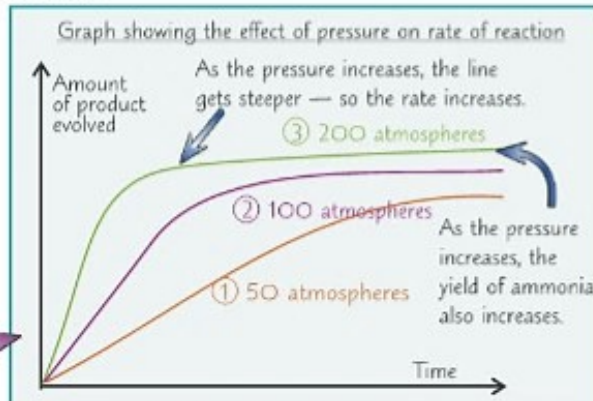


## The Reaction is Reversible, So There's a Compromise to be Made

- 1) You might remember from page 68 that certain conditions, such as the temperature and pressure, can affect the rate of a reaction.
- 2) These factors can also affect the position of equilibrium for a reversible reaction — and sometimes there is a trade-off between increasing the rate and maximising the yield (see page 49) of a reaction.
- 3) For example, the forward reaction in the Haber process is exothermic. That means that increasing the temperature will move the equilibrium the wrong way — away from ammonia and towards nitrogen and hydrogen. So the yield of ammonia would be greater at lower temperatures.

For more on factors that  
affect the position of  
equilibrium, see page 73.

- 4) The trouble is, lower temperatures mean a slower rate of reaction (and so equilibrium is reached more slowly).
- 5) The  $450^\circ\text{C}$  is a compromise between maximum yield and speed of reaction. It's better to wait just 20 seconds for a 10% yield than to have to wait 60 seconds for a 20% yield.
- 6) Higher pressures move the position of equilibrium towards the products since there are four molecules of gas on the left-hand side for every two molecules on the right. So increasing pressure maximises the percentage yield. It also increases the rate of reaction.



- 7) So the pressure is set as high as possible, without making the process too expensive to or dangerous to build and maintain. Hence the 200 atmospheres operating pressure.
- 8) And finally, the iron catalyst makes the reaction go faster, but doesn't affect the yield.

## You need to learn this stuff — go on, Haber go at it...

The trickiest bit of this page is remembering that temperature is raised not for a better equilibrium, but for speed. Remember that industries have to consider costs — so a higher pressure may give a better yield, but it's not cheap.

Q1 What are the industrial conditions used in the Haber process and why?

[3 marks]

# NPK Fertilisers

Fertilisers allow us to keep growing crops on the same land every year — thanks to our old friend [ammonia](#).

## NPK Fertilisers Provide Plants with the Essential Elements for Growth

- 1) Farmers used to use [manure](#) to fertilise fields. Formulated fertilisers are better as they're more [widely available](#), [easier](#) to use, don't [smell](#) and have just enough of each [nutrient](#) so more crops can be grown.
- 2) The [three](#) main essential elements in fertilisers are [nitrogen](#), [phosphorus](#) and [potassium](#). If plants don't get enough of these elements, their growth and life processes are affected. These elements may be [missing](#) from the soil if they've been used up by a [previous crop](#).
- 3) Fertilisers [replace](#) these missing elements or [provide more](#) of them. This helps to increase the [crop yield](#), as the crops can grow [faster](#) and [bigger](#). For example, fertilisers add more [nitrogen](#) to plant proteins, which makes the plants grow [faster](#) — increasing [productivity](#).
- 4) [NPK fertilisers](#) are formulations containing salts of nitrogen (N), phosphorus (P) and potassium (K) (hence, "NPK") in the [right percentages](#) of the elements.



There's more on formulations on p.86.

## Ammonia is Used to Produce Nitrogen-Containing Compounds

- 1) Ammonia can be reacted with oxygen and water in a series of reactions to make [nitric acid](#).
- 2) You can also react ammonia with acids, [including](#) nitric acid, to get [ammonium salts](#).
- 3) Ammonia and nitric acid react together to produce [ammonium nitrate](#) — this is an especially good compound to use in a fertiliser because it has nitrogen from [two sources](#).



- 4) This reaction is carried out [differently](#) in [industry](#) to how it is done in the [lab](#). You may be asked to compare differences between the different methods in your exam:

### In Industry

The reaction is carried out in [giant vats](#), at [high concentrations](#) resulting in a [very exothermic](#) reaction. The heat released is used to [evaporate water](#) from the mixture to make a very concentrated ammonium nitrate product.

If you get asked to compare reactions in the exams, you'll be given all the details you need.

### In the Lab

The reaction is carried out on a much [smaller scale](#) by [titration](#) and [crystallisation](#). The reactants are at a much [lower concentration](#) than in industry, so less heat is produced by the reaction and it's [safer](#) for a person to carry it out. After the titration, the mixture then needs to be [crystallised](#) to give pure ammonium nitrate [crystals](#). Crystallisation isn't used in industry because it's very [slow](#).

For how to carry out a titration or crystallisation have a look at p.52 and p.17.

## Phosphate and Potassium are Sourced from Mined Compounds

- 1) [Potassium chloride](#) and [potassium sulphate](#) can be [mined](#) and [used](#) as a source of potassium.
- 2) [Phosphate rock](#) is also mined. However, because the phosphate salts in the rock are [insoluble](#), plants can't use them as nutrients.
- 3) Reacting phosphate rock with a number of different types of [acids](#) produces [soluble](#) phosphates:
  - Reaction with [nitric acid](#) produces [phosphoric acid](#) and [calcium nitrate](#).
  - Reaction with [sulfuric acid](#) produces [calcium sulfate](#) and [calcium phosphate](#) (this mixture is known as [single superphosphate](#)).
  - Reaction with [phosphoric acid](#) only produces [calcium phosphate](#) (the product of this reaction can be called [triple superphosphate](#)).

## There's nowt wrong wi' just spreadin' muck on it...

Not the most thrilling of pages, I'm afraid. Just loads of reactions for you to learn... It's important stuff, though.

Q1 Suggest a source of potassium in an NPK fertiliser.

[1 mark]

## Revision Questions for Topic 10

**Topic 10** — done. Now is as good a time as any to have another read back over the section.

- Try these questions and tick off each one when you get it right.
- When you've done all the questions under a heading and are completely happy with it, tick it off.

### Properties and Uses of Materials (p.96-97) ☐

- 1) How is soda-lime glass made? ☐
- 2) Give three examples of composite materials. ☐
- 3) Describe the differences in the properties of low density poly(ethene) and high density poly(ethene). ☐
- 4) Name two elements that can be added to iron to make stainless steel. ☐
- 5) What two metals is bronze made of? ☐
- 6) Give two uses of brass. ☐

### Corrosion (p.98) ☐

- 7) What two substances does iron react with when it rusts? ☐
- 8) Aluminium objects can corrode. Why is aluminium not completely destroyed by corrosion? ☐
- 9) Suggest three ways of making a barrier that can be used to protect an iron object from rusting. ☐
- 10) Suggest one metal that can be used in the sacrificial method to prevent rusting. ☐

### Chemistry and Sustainability (p.99-101) ☐

- 11) Give two examples of renewable resources. ☐
- 12) Suggest one way in which chemistry is improving sustainability. ☐
- 13) State two methods that can be used to extract copper from low grade ores. ☐
- 14) Other than recycling, how can glass bottles be used in a sustainable way? ☐
- 15) What do life cycle assessments (LCA) do? ☐

### Potable Water and Waste Water Treatment (p.102-103) ☐

- 16) What is potable water? ☐
- 17) Suggest a source of fresh water. ☐
- 18) Describe how you could distil seawater in the lab. ☐
- 19) What other method could you use for making sea water potable? ☐
- 20) Name three different sources of waste water. ☐
- 21) Why is it important to treat waste water before releasing it into the environment? ☐
- 22) What type of waste water could contain harmful chemicals? ☐
- 23) What two products can be obtained by the anaerobic digestion of sewage sludge? ☐

### The Haber Process (p.104) ☐

- 24) What is the balanced symbol equation for the Haber process? ☐
- 25) From what source is the hydrogen used in the Haber process obtained? ☐
- 26) What happens to the unused hydrogen and nitrogen in the Haber process? ☐

### NPK Fertilisers (p.105) ☐

- 27) What three elements are present in NPK fertilisers? ☐
- 28) What substances can be reacted with each other to produce the following:
  - a) ammonium nitrate,    b) phosphoric acid and calcium nitrate,    c) calcium phosphate only? ☐

# Practical Techniques

Safety specs out and lab coats on, it's time to find out about the skills you'll need in **experiments**. Finally time to look like a real scientist... hurrah! But you also need to know about this stuff in your exams... boooo...

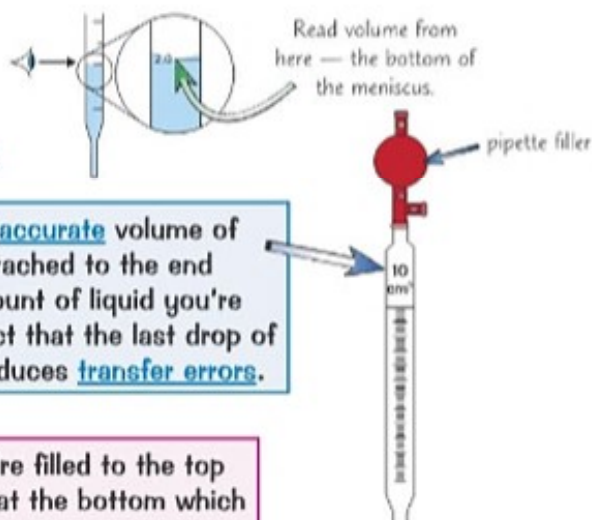
## Solids Should Be Measured Using a Balance

- 1) To weigh a solid, start by putting the **container** you're weighing your substance **into** on the **balance**.
- 2) Set the balance to exactly **zero** and then start weighing out your substance.
- 3) It's **no good** carefully weighing out your solid if it's not all transferred to your reaction vessel — the amount in the **reaction vessel** won't be the same as your measurement. Here are a couple of methods you can use to make sure that none gets left in your weighing container...

- If you're **dissolving** a solid in a solvent to make a **solution**, you could **wash** any remaining solid into the new container using the **solvent**.
- You could set the balance to zero **before** you put your **weighing container** on the balance. Then **reweigh** the weighing container **after** you've transferred the solid. Use the **difference in mass** to work out **exactly** how much solid you added to your experiment.

## Three Ways to Measure Liquids

There are a few methods you might use to measure the volume of a liquid. Whichever method you use, always read the volume from the **bottom of the meniscus** (the curved upper surface of the liquid) when it's at **eye level**.



**Pipettes** are long, narrow tubes that are used to suck up an **accurate** volume of liquid and **transfer** it to another container. A **pipette filler** attached to the end of the pipette is used so that you can **safely control** the amount of liquid you're drawing up. Pipettes are often **calibrated** to allow for the fact that the last drop of liquid stays in the pipette when the liquid is ejected. This reduces **transfer errors**.



**Burettes** measure from top to bottom (so when they're filled to the top of the scale, the scale reads zero). They have a tap at the bottom which you can use to release the liquid into another container (you can even release it drop by drop). To use a burette, take an **initial reading**, and once you've released as much liquid as you want, take a **final reading**. The **difference** between the readings tells you **how much** liquid you used.

Burettes are used a lot for titrations. There's loads more about titrations on page 52.

**Measuring cylinders** are the most common way to measure out a liquid. They come in all different **sizes**. Make sure you choose one that's the right size for the measurement you want to make. It's no good using a huge 1000 cm<sup>3</sup> cylinder to measure out 2 cm<sup>3</sup> of a liquid — the graduations will be too big, and you'll end up with **massive errors**. It'd be much better to use one that measures up to 10 cm<sup>3</sup>.



If you only want a couple of drops of liquid, and don't need it to be accurately measured, you can use a dropping pipette to transfer it. For example, this is how you'd add a couple of drops of indicator into a mixture.

## Gas Syringes Measure Gas Volumes

Gases can be measured with a gas syringe. They should be measured at **room temperature and pressure** as the **volume** of a gas **changes** with temperature and pressure. You should also use a gas syringe that's the **right size** for the measurement you're making. Before you use the syringe, you should make sure it's completely sealed and that the plunger moves smoothly.

# Practical Techniques

## Measure *Temperature* Accurately

You can use a **thermometer** to measure the temperature of a substance:

- 1) Make sure the **bulb** of your thermometer is **completely submerged** in any mixture you're measuring.
- 2) If you're taking an initial reading, you should wait for the temperature to **stabilise** first.
- 3) Read your measurement off the **scale** on a thermometer at **eye level** to make sure it's correct.

## You May Have to Measure the *Time Taken* for a Change

- 1) You should use a **stopwatch** to **time** experiments. These measure to the nearest **0.1 s** so are **sensitive**.
- 2) Always make sure you **start** and **stop** the stopwatch at exactly the right time. For example, if you're investigating the rate of an experiment, you should start timing at the **exact moment** you mix the reagents and start the reaction. If you're measuring the time taken for a precipitate to form, you should watch the reaction like a hawk so you can **stop** timing the moment it goes cloudy.

## Measure *pH* to Find Out How *Acidic* or *Alkaline* a Solution Is

You need to be able to decide the best method for measuring pH, depending on what your experiment is.

- 1) **Indicators** are dyes that **change colour** depending on whether they're in an **acid** or an **alkali**. You use them by adding a couple of drops of the indicator to the solution you're interested in. They're useful for titration reactions, when you want to find the point at which a solution is neutralised.
- 2) **Universal indicator** is a **mixture** of indicators that changes colour **gradually** as pH changes. It doesn't show a **sudden** colour change. It's useful for **estimating** the pH of a solution based on its colour.
- 3) Indicators can be soaked into **paper** and strips of this paper can be used for testing pH. If you use a dropping pipette to spot a small amount of a solution onto some indicator paper, it will **change colour** depending on the pH of the solution.
- 4) Indicator paper is useful when you **don't** want to change the colour of **all** of the substance, or if the substance is **already** coloured so might **obscure** the colour of the indicator. You can also hold a piece of **damp indicator paper** in a **gas sample** to test its pH.
- 5) **pH probes** are attached to pH meters which have a **digital display** that gives a **numerical** value for the pH of a solution. They're used to give an **accurate value** of pH.

**Litmus paper** turns **red** in acidic conditions and **blue** in basic conditions. **Universal indicator paper** can be used to **estimate** the pH based on its colour.

There's loads more about pH on page 51

## Be Careful When You *Handle* or *Mix* Substances

- 1) There are lots of hazards in chemistry experiments, so **before** you start any experiment, you should read any **safety precautions** to do with your method or the chemicals you're using.
- 2) The substances used in chemical reactions are often **hazardous**. For example, they might catch fire easily (they're flammable), or they might irritate or burn your skin if you come into contact with them.
- 3) Whenever you're doing an experiment, you should wear a **lab coat**, **safety goggles** and **gloves**.
- 4) Always be careful that the chemicals you're using aren't flammable before you go lighting any Bunsen burners, and make sure you're working in an area that's **well ventilated**.
- 5) If you're doing an experiment that might produce nasty **gases** (such as chlorine), you should carry out the experiment in a **fume hood** so that the gas can't escape out into the room you're working in.
- 6) Never directly touch any chemicals (even if you're wearing gloves). Use a **spatula** to transfer **solids** between containers. Carefully **pour** liquids between different containers, using a **funnel** to avoid spillages.
- 7) Be careful when you're **mixing** chemicals, as a reaction might occur. If you're **diluting** a liquid, add the **concentrated substance** to the **water** (not the other way around) or the mixture could get very **hot**.

## The Tempipettes and MacBurette — Shakespeare for chemists...

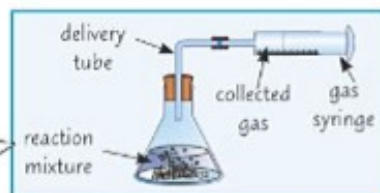
It's no good throwing chemicals around willy nilly and calling it an experiment. To make sure your results are reproducible, so can be trusted by other scientists, you have to make sure all your measurements are accurate.

# Setting Up Equipment

Setting up the equipment for an experiment correctly is **just as important** as making accurate measurements.

## To Collect Gases, the System Needs to be Sealed

- 1) There are times when you might want to **collect** the gas produced by a reaction. For example, to investigate the **rate** of reaction.
- 2) The most accurate way to measure the volume of a gas that's been produced is to collect it in a **gas syringe** (see page 107).
- 3) You could also collect it by **displacing water** from a measuring cylinder. Here's how you do it...



- Fill a **measuring cylinder** with **water**, and carefully place it **upside down** in a container of water. Record the **initial level** of the water in the measuring cylinder.
- Position a **delivery tube** coming **from** the reaction vessel so that it's **inside** the measuring cylinder, pointing upwards. Any gas that's produced will pass **through** the delivery tube and **into** the **measuring cylinder**. As the gas enters the measuring cylinder, the **water** is **pushed out**.
- Record the **level of water** in the measuring cylinder and use this value, along with your **initial value**, to calculate the **volume** of gas produced.

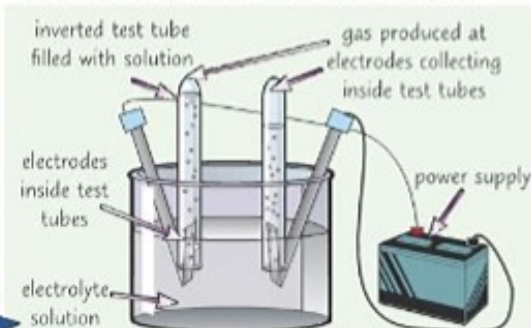


If the delivery tube is underneath the measuring cylinder rather than inside it then some of the gas might escape out into the air.

- 4) This method is **less accurate** than using a gas syringe to measure the volume of gas produced. This is because some gases can **dissolve** in water, so less gas ends up in the measuring cylinder than is **actually produced**.
- 5) If you just want to **collect** a sample to test (and don't need to know the volume), you can collect it over water, as above, using a **test tube**. Once the test tube is full of gas, stopper it to store the gas for later.

## You May Have to Identify the Products of Electrolysis

- 1) When you electrolyse an **aqueous solution**, the products of electrolysis will depend on how reactive the ions in the solution are compared to the  $H^+$  and  $OH^-$  ions that come from water.
- 2) At the **cathode** you'll either get a **pure metal** coating the electrode or bubbles of **hydrogen gas**.
- 3) At the **anode**, you'll get bubbles of **oxygen gas** unless a **halide ion** is present, when you'll get the **halogen**.
- 4) You may have to predict and identify what's been made in an electrolysis experiment. To do this, you need to be able to **set up the equipment** so that you can **collect** any gas that's produced. The easiest way to collect the gas is in a **test tube**.
- 5) Here's how to set up the equipment...

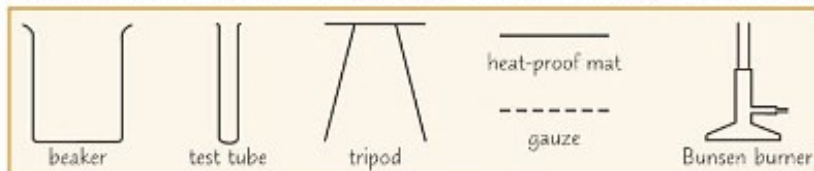


There's more about electrolysis on p58-59.

The tests for gases are described on page 88.

## Make Sure You Can Draw Diagrams of Your Equipment

When you're writing out a **method** for your experiment, it's always a good idea to draw a **labelled diagram** showing how your apparatus will be **set up**. The easiest way to do this is to use a scientific drawing, where each piece of apparatus is drawn as if you're looking at its **cross-section**. For example:



The pieces of glassware are drawn without tops so they aren't sealed. If you want to draw a closed system, remember to draw a bung in the top.

## I set up my equipment — they had a blind date at the cinema...

Being a dab hand at setting up experiments won't just make your investigations more reliable. You might also be asked to comment on how an experiment's been set up in the exam. So best get learning. You'll thank me for it...

## Heating Substances

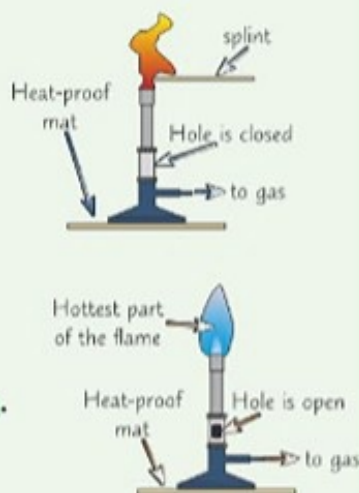
Heating a reaction isn't as simple as wrapping it up in a lumpy wool jumper and a stripy scarf. There's more than one way to do it, and you need to be able to decide on the **best**, and the **safest**, method.

### Bunsen Burners Have a Naked Flame

Bunsen burners are good for heating things quickly. You can easily adjust how strongly they're heating. But you need to be careful not to use them if you're heating **flammable** compounds as the flame means the substance would be at risk of **catching fire**.

#### Here's how to use a Bunsen burner...

- Connect the Bunsen burner to a gas tap, and check that the hole is **closed**. Place it on a **heat-proof mat**.
- Light a **splint** and hold it over the Bunsen burner. Now, turn on the gas. The Bunsen burner should light with a **yellow flame**.
- The **more open** the hole is, the **more strongly** the Bunsen burner will heat your substance. Open the hole to the amount you want. As you open the hole more, the flame should turn more **blue**.
- The **hottest** part of the flame is just above the **blue cone**, so you should heat things here.
- If your Bunsen burner is alight but not heating anything, make sure you **close** the hole so that the flame becomes **yellow** and **clearly visible**.
- If you're heating something so that the container (e.g. a test tube) is **in** the flame, you should hold the vessel at the **top**, furthest away from the substance (and so the flame) using a pair of **tongs**.
- If you're heating something **over** the flame (e.g. an evaporating dish), you should put a **tripod and gauze** over the Bunsen burner before you light it, and place the vessel on this.

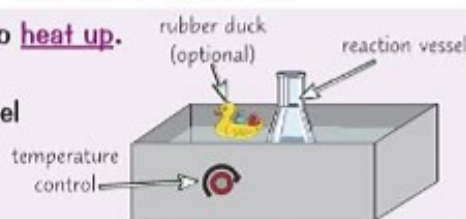


You'd use a Bunsen burner to carry out **flame tests** to identify metal ions in a compound (see page 89). A sample of the compound is placed on a metal wire that you hold just above the cone of a Bunsen burner with a blue flame. The flame should then **change colour** depending on what metal ion is in the sample.

### The Temperature of Water Baths & Electric Heaters Can Be Set

- 1) A **water bath** is a container filled with water that can be heated to a **specific temperature**.

- **Set** the temperature on the water bath, and allow the water to **heat up**.
- Place the vessel containing your substance in the water bath using a pair of tongs. The level of the water outside the vessel should be **just above** the level of the substance inside the vessel. The substance will then be warmed to the **same temperature** as the water.



As the substance in the vessel is surrounded by water, the heating is very **even**. Water boils at **100 °C** though, so you **can't** use a water bath to heat something to a higher temperature than this — the water **won't** get **hot** enough.

Handle any glassware you've heated with tongs until you're sure it's cooled down.

- 2) **Electric heaters** are often made up of a metal **plate** that can be heated to a certain temperature. The vessel containing the substance you want to heat is placed on top of the hot plate. You can heat substances to **higher temperatures** than you can in a water bath but, as the vessel is only heated from **below**, you'll usually have to **stir** the substance inside to make sure it's **heated evenly**.

### A bath and an electric heater — how I spend my January nights...

You know, I used to have a chemistry teacher who'd play power ballads when the Bunsen burners were alight and sway at the front of the class like he was at a gig. You think I made that up, but it's true.

# Answers

## p.12 — Atoms

- Q1 protons = atomic number = 7 [1 mark]  
electrons = protons = 7 [1 mark]  
neutrons = mass number – atomic number  
= 14 – 7 = 7 [1 mark]

## p.13 — Elements

- Q1 E.g. it's the number of protons in an atom that determines what type of atom it is, so if all the atoms have the same number of protons then the substance is an element [1 mark].

- Q2 protons = 26 [1 mark]  
neutrons = 56 – 26 = 30 [1 mark]

## p.14 — Compounds

- Q1  $(2 \times \text{Na}) + (1 \times \text{C}) + (3 \times \text{O}) = 6$  [1 mark]  
Q2 calcium chloride [1 mark]  
Calcium and chlorine [1 mark].

## p.15 — Chemical Equations

- Q1  $2\text{Fe} + 3\text{Cl}_2 \rightarrow 2\text{FeCl}_3$  [1 mark]  
Q2 a) water  $\rightarrow$  hydrogen + oxygen [1 mark]  
b)  $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$   
[1 mark for correct products and reactants,  
1 mark for being balanced correctly]

## p.16 — Mixtures and Chromatography

- Q1 The pen ink might dissolve in the solvent and rise up the filter paper [1 mark].

## p.17 — More Separation Techniques

- Q1 Pour the copper sulfate solution into an evaporating dish and slowly heat the solution until crystals start to form or some of the solvent has evaporated [1 mark]. Leave the dish to cool until crystals form [1 mark]. Filter [1 mark] and then dry the crystals using a desiccator/drying oven [1 mark].



## p.18 — Distillation

- Q1 Ethanol [1 mark]. Ethanol has the second lowest boiling point and will be collected once all the methanol has been distilled off and the temperature increased [1 mark].

## p.19 — The History of the Atom

- Q1 In the plum pudding model, the atom is a ball of positive charge with electrons spread throughout it [1 mark].  
Q2 If the plum pudding model was correct you would expect most of the alpha particles to have passed through the foil or only to have been deflected slightly [1 mark]. The actual deflections of the particles suggests that atoms contain a small nucleus where the positive charge is concentrated [1 mark].

## p.20 — Electronic Structure

- Q1 2,8,3 or  [1 mark]  
Q2 2,8,8 or  [1 mark]

## p.21 — Development of the Periodic Table

- Q1 By relative atomic mass [1 mark].  
Q2 E.g. he left gaps in order to keep elements with similar properties in the same group [1 mark]. He switched the order of elements based on their properties, even if their atomic masses were no longer in order. [1 mark].

## p.22 — The Modern Periodic Table

- Q1 2 [1 mark]  
Q2 Both chlorine and bromine are in Group 7 and so have the same number of electrons in their outer shell [1 mark].  
Q3 E.g. potassium forms 1+ ions as it's in the same group as sodium so will react in a similar way [1 mark].

## p.23 — Metals and Non-Metals

- Q1 Any three from: e.g. metals tend to be strong / good conductors of heat / good at conducting electricity / malleable / high melting/boiling temperatures [1 mark for each].

## p.24 — Group 1 Elements

- Q1 As you go further down the group the outer electron is further away from the nucleus [1 mark]. This means the attraction between the nucleus and the electron decreases so is more easily removed resulting in an increase in reactivity [1 mark].

## p.25 — Group 7 Elements

- Q1 Bromine would displace iodine in potassium iodide [1 mark] because bromine is more reactive than iodine [1 mark].  
Q2 Halogens react by gaining electrons to make a full outer-shell [1 mark]. As you go further down the group the outer electrons are further away so there is less attraction between them and the nucleus [1 mark]. This means electrons are harder to gain so they become less reactive [1 mark].

## p.26 — Group 0 Elements

- Q1 Xenon has a higher boiling point than neon [1 mark].  
Q2 Argon has a full outer-shell [1 mark] so is electronically stable and does not readily lose or gain electrons [1 mark].

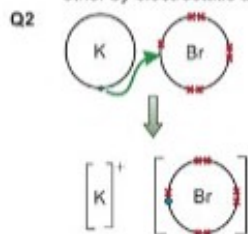
## p.28 — Formation of Ions

- Q1 Noble gas electronic structures have a full shell of outer electrons [1 mark], which is a very stable structure [1 mark].

- Q2 a) 1- [1 mark]  
b) 2+ [1 mark]  
c) 1+ [1 mark]

## p.29 — Ionic Bonding

- Q1 Each sodium atom loses an electron to form an  $\text{Na}^+$  ion [1 mark]. Each chlorine atom gains an electron to form a  $\text{Cl}^-$  ion [1 mark]. The oppositely charged ions are attracted to each other by electrostatic attraction [1 mark].

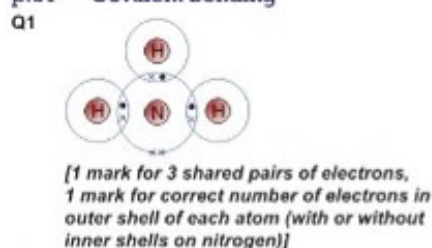


[1 mark for arrow showing electron transferred from potassium to bromine,  
1 mark for correct outer shell electron configurations (with or without inner shells),  
1 mark for correct charges]

## p.30 — Ionic Compounds

- Q1 a) It will have a high melting point [1 mark] because a lot of energy is needed to break the strong attraction between the ions/the strong ionic bonds [1 mark].  
b) When melted, the ions are free to move, so they can carry an electric current [1 mark].  
c) The compound contains caesium and chloride ions. Caesium is in Group 1 so forms 1+ ions [1 mark], and chlorine is in Group 7 so forms 1- ions [1 mark]. The charges balance with one of each ion, so the empirical formula is  $\text{CsCl}$  [1 mark].

## p.31 — Covalent Bonding



## p.32 — Simple Molecular Substances

- Q1 The intermolecular forces between molecules of  $\text{O}_2$  are weak and don't need much energy to break [1 mark].  
Q2  $\text{N}_2$  molecules aren't charged/There aren't any free electrons or ions [1 mark].

## p.33 — Polymers and Giant Covalent Structures

- Q1  $(\text{C}_2\text{H}_3\text{Cl})_n$  [1 mark]  
Q2 To melt diamond you have to break the covalent bonds between atoms which are very strong [1 mark] but to melt poly(ethene) you only have to break the weaker intermolecular forces which needs less energy [1 mark]. So diamond has a higher melting point [1 mark].

## p.34 — Allotropes of Carbon

- Q1 Any three from, e.g.: lubricants / catalysts / drug delivery / to strengthen materials [1 mark for each, up to 3 marks]

## p.35 — Metallic Bonding

- Q1 Copper is a good electrical conductor [1 mark] as it contains delocalised electrons which are able to carry an electrical current [1 mark].  
Q2 Pure copper would be too soft to use for a door hinge, but alloys are harder than pure metals, so are suitable [1 mark].

## p.36 — States of Matter

- Q1 The gaseous state [1 mark].

## p.37 — Changing State

- Q1 a) solid [1 mark]  
b) liquid [1 mark]  
c) liquid [1 mark]  
d) gas [1 mark]

## p.38 — Nanoparticles

- Q1 Between 1 nm and 100 nm [1 mark].  
Q2 a) Surface area of one face  
=  $50 \text{ nm} \times 50 \text{ nm} = 2500 \text{ nm}^2$   
Total surface area =  $6 \times 2500 \text{ nm}^2 = 15\,000 \text{ nm}^2$  [1 mark]  
Volume =  $50 \text{ nm} \times 50 \text{ nm} \times 50 \text{ nm}$   
=  $125\,000 \text{ nm}^3$  [1 mark]  
Surface area to volume ratio  
=  $15\,000 \text{ nm}^2 \div 125\,000 \text{ nm}^3$   
=  $0.12 \text{ nm}^{-1}$  [1 mark]  
b) It would decrease [1 mark] by a factor of ten [1 mark]. (As the length of the sides has increased by a factor of ten).

## p.39 — Uses of Nanoparticles

- Q1 E.g. in electrical circuits for computer chips [1 mark].  
Q2 E.g. they might have long-term impacts on health [1 mark].  
Q3 Silver nanoparticles. They form a transparent layer, so the window will still be see-through [1 mark] and they conduct heat, so will be able to warm the window and melt the frost [1 mark].

## p.41 — Relative Formula Mass

- Q1 a)  $A_r$  of H = 1 and  $A_r$  of O = 16  
 $M_r$  of  $H_2O = (2 \times 1) + 16 = 18$  [1 mark]  
 b)  $A_r$  of Li = 7,  $A_r$  of O = 16 and  $A_r$  of H = 1  
 So  $M_r$  of LiOH =  $7 + 16 + 1 = 24$  [1 mark]  
 c)  $A_r$  of H = 1,  $A_r$  of S = 32 and  $A_r$  of O = 16  
 $M_r$  of  $H_2SO_4 = (2 \times 1) + 32 + (4 \times 16) = 98$  [1 mark]  
 Q2  $A_r$  of K = 39,  $A_r$  of O = 16 and  $A_r$  of H = 1  
 $M_r$  of KOH =  $39 + 16 + 1 = 56$  [1 mark]  
 $\frac{39}{56} \times 100 = 70\%$  [1 mark]

## p.42 — The Mole

- Q1  $M_r$  of  $H_2O = 16 + (2 \times 1) = 18$  [1 mark]  
 number of moles = mass  $\div M_r$   
 number of moles =  $90 \text{ g} \div 18 = 5$  moles [1 mark]  
 Q2  $M_r$  of KBr =  $39 + 80 = 119$  [1 mark]  
 mass = number of moles  $\times M_r$   
 mass =  $0.20 \times 119 = 24 \text{ g}$  [1 mark]

## p.43 — Conservation of Mass

- Q1 Total mass on the left hand side  
 =  $M_r(H_2SO_4) + 2 \times M_r(NaOH)$   
 $M_r$  of  $H_2SO_4 = (2 \times 1) + 32 + (4 \times 16) = 98$   
 $2 \times M_r$  of NaOH =  $2 \times (23 + 16 + 1) = 80$   
 So total mass on the left hand side =  $98 + 80 = 178$  [2 marks for 178, 1 mark for either 98 or 80]  
 Total mass on right hand side  
 =  $M_r(Na_2SO_4) + 2 \times M_r(H_2O)$   
 $M_r$  of  $Na_2SO_4 = (2 \times 23) + 32 + (4 \times 16) = 142$   
 $2 \times M_r$  of  $H_2O = 2 \times (2 \times 1 + 16) = 36$   
 $142 + 36 = 178$  [2 marks for 178, 1 mark for either 142 or 36]  
 The total  $M_r$  on the left-hand side is equal to the total  $M_r$  on the right-hand side, so mass is conserved [1 mark].

## p.44 — The Mole and Equations

- Q1 a)  $N_2: \frac{84}{28} = 3.0 \text{ mol}$  [1 mark]  
 $H_2: \frac{18}{2} = 9 \text{ mol}$  [1 mark]  
 $NH_3: \frac{102}{17} = 6.0 \text{ mol}$  [1 mark]  
 b) Divide by the smallest number of moles (3.0):  
 $N_2: \frac{3.0}{3.0} = 1$   $H_2: \frac{9}{3.0} = 3$   $NH_3: \frac{6.0}{3.0} = 2.0$  [1 mark]  
 The balanced symbol equation is:  
 $N_2 + 3H_2 \rightarrow 2NH_3$  [1 mark]

## p.45 — Limiting Reactants

- Q1  $M_r(KBr) = 119$ ,  $M_r(KCl) = 74.5$  [1 mark]  
 No. of moles of KBr =  $23.8 \div 119 = 0.200 \text{ mol}$  [1 mark]  
 From the reaction equation, 2 moles of KBr react to form 2 moles of KCl. So 0.200 moles of KBr reacts to form 0.200 moles of KCl [1 mark].  
 Mass KCl =  $74.5 \times 0.200 = 14.9 \text{ g}$  [1 mark]

## p.46 — Gases and Solutions

- Q1  $M_r(N_2) = 2 \times 14 = 28$  [1 mark]  
 Volume =  $\frac{3.5}{28} \times 24 = 3.0 \text{ dm}^3$  [1 mark]

## p.47 — Concentration Calculations

- Q1 Moles of KOH =  $0.150 \times (22.5 + 1000) = 0.003375$  [1 mark]  
 From the reaction equation, 1 mole of KOH reacts with 1 mole of  $HNO_3$ , so 0.003375 mol of KOH reacts with 0.003375 mol of  $HNO_3$  [1 mark].  
 Concentration of  $HNO_3$ :  
 $(0.003375) \div (25.0 + 1000) = 0.135 \text{ mol/dm}^3$  [1 mark]

## p.48 — Atom Economy

- Q1 a) Reaction B, because it only has one product / the atom economy is 100% [1 mark].  
 b) Total mass of all reactants:  
 $M_r(NH_4Cl) = 14 + (4 \times 1) + 35.5 = 53.5$  [1 mark]  
 $M_r(CaO) = 40 + 16 = 56$  [1 mark]  
 $(2 \times 53.5) + (2 \times 56) = 219$  [1 mark]  
 Total mass of desired product  
 =  $2 \times M_r(NH_3) = 34$  [1 mark]  
 Atom economy =  $\frac{34}{219} \times 100 = 16\%$  [1 mark]

## p.49 — Percentage Yield

- Q1  $\frac{2.31}{2.72} \times 100 = 84.9\%$  [1 mark]

## p.51 — Acids and Bases

- Q1 red/orange [1 mark]  
 Q2 alkaline [1 mark]

## p.52 — Titrations

- Q1 The indicator will have just changed colour [1 mark].  
 Q2 E.g. to increase the accuracy of your result [1 mark] and to spot any anomalous results [1 mark].

## p.53 — Strong Acids and Weak Acids

- Q1 Any one of, e.g. sulfuric acid / nitric acid / hydrochloric acid [1 mark].  
 Q2 Change in pH =  $3 - 6 = -3$   
 Change in concentration of  $H^+ = 10^{-(-3)} = 10^3 = 1000$   
 So, the concentration of  $H^+$  is 1000 times greater at pH = 3 than at pH = 6 [1 mark].

## p.54 — Reactions of Acids

- Q1 calcium carbonate + hydrochloric acid  $\rightarrow$  calcium chloride + carbon dioxide + water [1 mark for calcium chloride, 1 mark for carbon dioxide and water].

## p.55 — The Reactivity Series

- Q1  $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$  [1 mark for correct reactants and products, 1 mark for balancing, 1 mark for state symbols]

## p.56 — Separating Metals From Metal Oxides

- Q1  $2ZnO + C \rightarrow 2Zn + CO_2$  [1 mark for the correct products, 1 mark for the correctly balanced equation]  
 Q2 Carbon is less reactive than calcium and therefore will not reduce calcium oxide / Calcium is more reactive than carbon, so calcium oxide won't be reduced by carbon [1 mark].

## p.57 — Redox Reactions

- Q1 a)  $Zn_{(s)} + Fe^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$  [1 mark]  
 b)  $Zn_{(s)}$  is being oxidised [1 mark].  $Fe^{2+}_{(aq)}$  is being reduced [1 mark].

## p.58 — Electrolysis

- Q1 a) chlorine gas/ $Cl_2$  [1 mark]  
 b) sodium atoms/Na [1 mark]

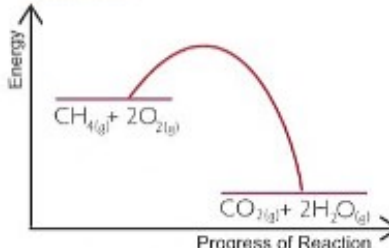
## p.59 — Electrolysis of Aqueous Solutions

- Q1 anode:  $2Cl^- \rightarrow Cl_2 + 2e^-$  [1 mark]  
 cathode:  $Cu^{2+} + 2e^- \rightarrow Cu$  [1 mark]

## p.61 — Exothermic and Endothermic Reactions

- Q1 exothermic [1 mark]

## p.62 — More Exothermic and Endothermic Reactions

- Q1  [1 mark for correct axes, 1 mark for correct energy levels of reactants and products, 1 mark for correct shape of curve linking the reactants to the products]

## p.63 — Bond Energies

- Q1 Energy required to break original bonds:  
 $(1 \times N=N) + (3 \times H-H)$   
 =  $941 \text{ kJ/mol} + 1308 \text{ kJ/mol} = 2249 \text{ kJ/mol}$  [1 mark]  
 Energy released by forming new bonds:  
 $(6 \times N-H)$   
 =  $2346 \text{ kJ/mol}$  [1 mark]  
 Overall energy change:  
 =  $2249 \text{ kJ/mol} - 2346 \text{ kJ/mol} = -97 \text{ kJ/mol}$  [1 mark]

## p.64 — Cells and Batteries

- Q1 Metal 2, Metal 3, Metal 1 [1 mark]. Zinc is less reactive than all 3 metals, so the most reactive metal will make the cell with the greatest voltage, since there will be the biggest difference in reactivity [1 mark]. So the order of reactivity is simply the order of the size of voltage [1 mark].

## p.65 — Fuel Cells

- Q1  $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$  [1 mark for correct reactants and products, 1 mark for balancing].

## p.67 — Rates of Reaction

- Q1 The activation energy for a reaction is the minimum amount of energy that particles need to react [1 mark].

## p.68 — Factors Affecting Rates of Reaction

- Q1 a) B [1 mark], because the powder has a higher surface area to volume ratio than the solid strip [1 mark].  
 b) B [1 mark], because the  $4 \text{ mol/dm}^3$  KOH solution is more concentrated than the  $2 \text{ mol/dm}^3$  solution [1 mark].

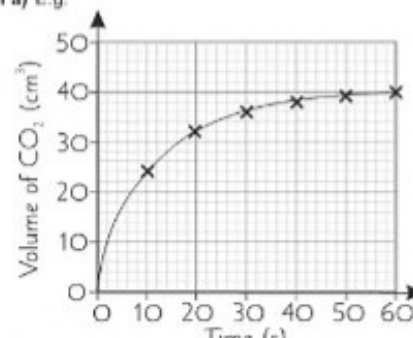
## p.69 — Measuring Rates of Reaction

- Q1 a) Add  $Na_2CO_3$  to a flask containing HCl [1 mark] and take readings of the mass of the flask at regular intervals [1 mark] using a mass balance [1 mark]. / Add  $Na_2CO_3$  to a flask containing HCl [1 mark] and take regular readings of the volume of gas released [1 mark] using a gas syringe [1 mark]. Maximum of 3 marks available.  
 b) E.g.  $\text{cm}^3/\text{s}$  [1 mark]

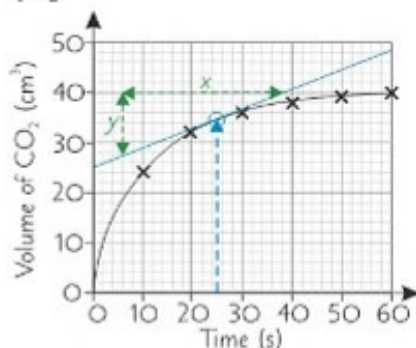
## p.70 — Two Rates Experiments

- Q1 E.g. volume of HCl added [1 mark], mass of magnesium used [1 mark].

## p.71 — Finding Reaction Rates from Graphs

- Q1 a) E.g.  [1 mark for correctly marking on all 6 points, 1 mark for choosing a sensible scale for the axes, 1 mark for drawing a line of best fit.]

b) E.g.



$$y = 40 - 27 = 13,$$

$$x = 38 - 6 = 32,$$

$$\text{rate} = y/x = 13/32 = 0.4 \text{ cm}^3/\text{s}$$

[1 mark for drawing a tangent at 25 s, 1 mark for correctly calculating a change in y from the tangent, 1 mark for correctly calculating a change in x from the tangent and 1 mark for a rate between 0.3 cm<sup>3</sup>/s and 0.5 cm<sup>3</sup>/s]

#### p.72 — Reversible Reactions

Q1 A system is at equilibrium when both the forward and reverse reactions are happening at the same rate [1 mark].

#### p.73 — Le Chatelier's Principle

Q1 a) More N<sub>2</sub>O<sub>4(g)</sub> produced [1 mark].

b) No effect [1 mark].

c) More CO<sub>2</sub> produced [1 mark].

#### p.75 — Hydrocarbons

Q1 a) C<sub>10</sub>H<sub>22</sub> is more viscous than C<sub>6</sub>H<sub>14</sub> [1 mark].

b) C<sub>10</sub>H<sub>22</sub> has a higher boiling point than C<sub>6</sub>H<sub>14</sub> [1 mark].

c) C<sub>10</sub>H<sub>22</sub> is less flammable than C<sub>6</sub>H<sub>14</sub> [1 mark].

Q2 C<sub>2</sub>H<sub>6</sub> + 5O<sub>2</sub> → 3CO<sub>2</sub> + 4H<sub>2</sub>O [1 mark for correct reactants and products, 1 mark for correctly balancing]

#### p.76 — Fractional Distillation

Q1 It suggests that the hydrocarbons in petrol have a lower boiling point than those in diesel [1 mark].

Q2 It's hot at the bottom and cooler at the top [1 mark].

#### p.77 — Uses and Cracking of Crude Oil

Q1 Number of C atoms = 5 - 2 = 3

Number of H atoms = 12 - 4 = 8

Formula = C<sub>3</sub>H<sub>8</sub> [1 mark]

#### p.78 — Alkenes

Q1 C<sub>4</sub>H<sub>6</sub> [1 mark]

Q2 Carbon monoxide [1 mark], carbon [1 mark].

#### p.79 — Reactions of Alkenes

Q1 
$$\begin{array}{c} \text{H} & \text{H} & \text{H} \\ | & | & | \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{H} \\ | & | & | \\ \text{H} & \text{H} & \text{H} \end{array}$$
 [1 mark]

Q2 The solution turns from orange to colourless [1 mark].

#### p.80 — Addition Polymers

Q1 A polymer is a large molecule formed when lots of monomers join together [1 mark].

Q2 
$$\left( \begin{array}{c} \text{CH}_3\text{H} \\ | \\ \text{C}-\text{C} \\ | \quad | \\ \text{H} \quad \text{CH}_3 \end{array} \right)_n$$
 [1 mark]

$$\left( \begin{array}{c} \text{H} & \text{H} \\ | & | \\ \text{C} & -\text{C} \\ | & | \\ \text{H} & \text{C}_2\text{H}_5 \end{array} \right)_n$$
 [1 mark]

#### p.81 — Alcohols

Q1 C<sub>2</sub>H<sub>5</sub>OH + 3O<sub>2</sub> → 2CO<sub>2</sub> + 3H<sub>2</sub>O [1 mark for correct products and reactants, 1 mark for correctly balancing]

#### p.82 — Carboxylic Acids

Q1 a) You would see bubbles of gas form [1 mark] as carbon dioxide is produced [1 mark].

b) sodium carbonate + methanoic acid → sodium methanoate + water + carbon dioxide [1 mark]

#### p.83 — Condensation Polymers

Q1 E.g. water [1 mark]

Q2 one [1 mark]

#### p.84 — Naturally Occurring Polymers

Q1 polypeptides [1 mark]

Q2 E.g. proteins/polypeptides [1 mark], starch/cellulose [1 mark], DNA [1 mark].

#### p.86 — Purity and Formulations

Q1 a) The sample melts over a range of temperatures [1 mark]. The melting point is lower than that of pure aspirin [1 mark].

b) Any range or single value within the range: 141-200 °C [1 mark]

#### p.87 — Paper Chromatography

Q1 During paper chromatography, the molecules of each chemical in the sample move between the stationary phase and the mobile phase [1 mark]. The mobile phase moves through the stationary phase over the course of the experiment, and anything that's dissolved in it will move with it [1 mark]. The distance a compound moves through the stationary phase depends on how long it spends dissolved in the mobile phase compared to on the stationary phase [1 mark]. Since different compounds will interact differently with the mobile phase and the stationary phase, they'll move different amounts through the stationary phase, and so be separated from each other [1 mark].

#### p.88 — Tests for Gases and Anions

Q1 Chloride [1 mark]

#### p.89 — Tests for Cations

Q1 The metal ions burn with a distinctive colour [1 mark].

Q2 a) Al<sup>3+</sup> [1 mark]

b) Al<sup>3+</sup> + 3OH<sup>-</sup> → Al(OH)<sub>3</sub> [1 mark]

#### p.90 — Flame Emission Spectroscopy

Q1 Flame emission spectroscopy can be used to identify the metal ions in mixtures [1 mark], whereas flame test can only be used for substances that contain a single metal ion [1 mark].

#### p.91 — The Evolution of the Atmosphere

Q1 Sedimentary rocks are formed when organic matter such as plant deposits or the shells and skeletons dead of marine animals fall to the seabed [1 mark]. These become buried and compressed by sediments over millions of years forming rocks [1 mark].

#### p.92 — Greenhouse Gases and Climate Change

Q1 Any three from, e.g. polar ice caps melting / sea levels rising / coastal erosion / changing rainfall patterns / some regions may have too much or too little water / it may be difficult to produce food / there may be an increase in the frequency and severity of storms. / the distribution of wild species may be affected [1 mark for each, up to a maximum of 3 marks].

#### p.93 — Carbon Footprints

Q1 E.g. governments can put a cap on the amount of greenhouse gases that a business can emit and issue licences for set amounts of emissions up to this point [1 mark]. They can also impose taxes on companies according to the amount of greenhouse gases that they emit to encourage them to cut down on emissions [1 mark].

#### p.94 — Air Pollution

Q1 E.g. particulates (soot) [1 mark], unburnt fuels [1 mark] and carbon monoxide [1 mark].

#### p.96 — Ceramics, Composites and Polymers

Q1 E.g. thermosetting polymers don't soften when heated whereas thermosoftening polymers do [1 mark].

#### p.97 — Properties of Materials

Q1 (14 + 24) × 100 = 58.33 = 58% [1 mark]

#### p.98 — Corrosion

Q1 The pure iron nail will rust [1 mark]. E.g. the galvanised nail won't rust due to the zinc forming a protective layer [1 mark].

#### p.99 — Finite and Renewable Resources

Q1 A finite resource, such as e.g. crude oil, will take a long time to replenish [1 mark]. On the other hand, a renewable resource, such as e.g. timber, can be replaced within a relatively short time scale [1 mark].

#### p.100 — Reuse and Recycling

Q1 E.g. saves energy needed to extract metals from the earth / conserves limited supplies of metals from the earth / cuts down on the amount of waste going to landfill [1 mark for each].

#### p.101 — Life Cycle Assessments

Q1 Choice of material / manufacturing and packaging / using the product / product disposal [1 mark for each].

#### p.102 — Potable Water

Q1 E.g. filter the water first, using a wire mesh followed by sand and gravel beds [1 mark]. Then sterilise the filtered water using chlorine / ozone / UV radiation [1 mark].

#### p.103 — Waste Water Treatment

Q1 Screening [1 mark], sedimentation [1 mark], digestion (aerobic / anaerobic) [1 mark].

#### p.104 — The Haber Process

Q1 A temperature of 450 °C is used because this is high enough to ensure a fast rate of reaction without decreasing the yield of ammonia by too much [1 mark]. A pressure of 200 atm is used in order to maximise the rate and the yield of the reaction. Higher pressure would be too expensive [1 mark]. An iron catalyst is used in order to speed up the rate of reaction [1 mark].

#### p.105 — NPK Fertilisers

Q1 E.g. potassium chloride / potassium sulphate [1 mark].

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# The Periodic Table

Periods		Groups																Group 0	
1	Group 1	Group 2											Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	
1	7 <b>Li</b> Lithium 3	9 <b>Be</b> Beryllium 4											11 <b>B</b> Boron 5	12 <b>C</b> Carbon 6	14 <b>N</b> Nitrogen 7	16 <b>O</b> Oxygen 8	19 <b>F</b> Fluorine 9	20 <b>Ne</b> Neon 10	4 <b>He</b> Helium 2
2	23 <b>Na</b> Sodium 11	24 <b>Mg</b> Magnesium 12											27 <b>Al</b> Aluminium 13	28 <b>Si</b> Silicon 14	31 <b>P</b> Phosphorus 15	32 <b>S</b> Sulfur 16	35.5 <b>Cl</b> Chlorine 17	40 <b>Ar</b> Argon 18	
3	39 <b>K</b> Potassium 19	40 <b>Ca</b> Calcium 20	45 <b>Sc</b> Scandium 21	48 <b>Ti</b> Titanium 22	51 <b>V</b> Vanadium 23	52 <b>Cr</b> Chromium 24	55 <b>Mn</b> Manganese 25	56 <b>Fe</b> Iron 26	59 <b>Co</b> Cobalt 27	59 <b>Ni</b> Nickel 28	63.5 <b>Cu</b> Copper 29	65 <b>Zn</b> Zinc 30	70 <b>Ga</b> Gallium 31	73 <b>Ge</b> Germanium 32	75 <b>As</b> Arsenic 33	79 <b>Se</b> Selenium 34	80 <b>Br</b> Bromine 35	84 <b>Kr</b> Krypton 36	
4	85 <b>Rb</b> Rubidium 37	88 <b>Sr</b> Strontium 38	89 <b>Y</b> Yttrium 39	91 <b>Zr</b> Zirconium 40	93 <b>Nb</b> Niobium 41	96 <b>Mo</b> Molybdenum 42	98 <b>Tc</b> Technetium 43	101 <b>Ru</b> Ruthenium 44	103 <b>Rh</b> Rhodium 45	106 <b>Pd</b> Palladium 46	108 <b>Ag</b> Silver 47	112 <b>Cd</b> Cadmium 48	115 <b>In</b> Indium 49	119 <b>Sn</b> Tin 50	122 <b>Sb</b> Antimony 51	128 <b>Te</b> Tellurium 52	127 <b>I</b> Iodine 53	131 <b>Xe</b> Xenon 54	
5	133 <b>Cs</b> Caesium 55	137 <b>Ba</b> Barium 56	139 <b>La</b> Lanthanum 57	178 <b>Hf</b> Hafnium 72	181 <b>Ta</b> Tantalum 73	184 <b>W</b> Tungsten 74	186 <b>Re</b> Rhenium 75	190 <b>Os</b> Osmium 76	192 <b>Ir</b> Iridium 77	195 <b>Pt</b> Platinum 78	197 <b>Au</b> Gold 79	201 <b>Hg</b> Mercury 80	204 <b>Tl</b> Thallium 81	207 <b>Pb</b> Lead 82	209 <b>Bi</b> Bismuth 83	209 <b>Po</b> Polonium 84	210 <b>At</b> Astatine 85	[222] <b>Rn</b> Radon 86	
6	[223] <b>Fr</b> Francium 87	[226] <b>Ra</b> Radium 88	[227] <b>Ac</b> Actinium 89	[261] <b>Rf</b> Rutherfordium 104	[262] <b>Db</b> Dubnium 105	[266] <b>Sg</b> Seaborgium 106	[264] <b>Bh</b> Bohrium 107	[277] <b>Hs</b> Hassium 108	[268] <b>Mt</b> Meitnerium 109	[271] <b>Ds</b> Darmstadtium 110	[272] <b>Rg</b> Roentgenium 111	[285] <b>Cn</b> Copernicium 112	[286] <b>Uut</b> Ununtrium 113	[289] <b>Fl</b> Flerovium 114	[289] <b>Uup</b> Ununpentium 115	[293] <b>Lv</b> Livermorium 116	[294] <b>Uus</b> Ununseptium 117	[294] <b>Uuo</b> Ununoctium 118	

Relative atomic mass

Atomic (proton) number

The Lanthanides (atomic numbers 58-71) and the Actinides (atomic numbers 90-103) are not shown in this table.

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