GCSE
Chemistry

GCSE Chemistry

Complete Revision & Practice

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Relative atomic mass → (top number); Atomic number → (bottom number)

1 H Hydrogen 1

Periods	Group 1	Group 2											Group 3	Group 4	Group 5	Group 6	Group 7	Group 0
1																		4 He Helium 2
2	7 Li Lithium 3	9 Be Beryllium 4											11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10
3	23 Na Sodium 11	24 Mg Magnesium 12											27 Al Aluminium 13	28 Si Silicon 14	31 P Phosphorus 15	32 S Sulfur 16	35.5 Cl Chlorine 17	40 Ar Argon 18
4	39 K Potassium 19	40 Ca Calcium 20	45 Sc Scandium 21	48 Ti Titanium 22	51 V Vanadium 23	52 Cr Chromium 24	55 Mn Manganese 25	56 Fe Iron 26	59 Co Cobalt 27	59 Ni Nickel 28	63.5 Cu Copper 29	65 Zn Zinc 30	70 Ga Gallium 31	73 Ge Germanium 32	75 As Arsenic 33	79 Se Selenium 34	80 Br Bromine 35	84 Kr Krypton 36
5	85.5 Rb Rubidium 37	88 Sr Strontium 38	89 Y Yttrium 39	91 Zr Zirconium 40	93 Nb Niobium 41	96 Mo Molybdenum 42	98 Tc Technetium 43	101 Ru Ruthenium 44	103 Rh Rhodium 45	106 Pd Palladium 46	108 Ag Silver 47	112 Cd Cadmium 48	115 In Indium 49	119 Sn Tin 50	122 Sb Antimony 51	128 Te Tellurium 52	127 I Iodine 53	131 Xe Xenon 54
6	133 Cs Caesium 55	137 Ba Barium 56	139 La Lanthanum 57	178.5 Hf Hafnium 72	181 Ta Tantalum 73	184 W Tungsten 74	186 Re Rhenium 75	190 Os Osmium 76	192 Ir Iridium 77	195 Pt Platinum 78	197 Au Gold 79	201 Hg Mercury 80	204 Tl Thallium 81	207 Pb Lead 82	209 Bi Bismuth 83	210 Po Polonium 84	210 At Astatine 85	222 Rn Radon 86
7	223 Fr Francium 87	226 Ra Radium 88	227 Ac Actinium															

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Theories Come, Theories Go

SCIENTISTS ARE ALWAYS RIGHT — OR ARE THEY?

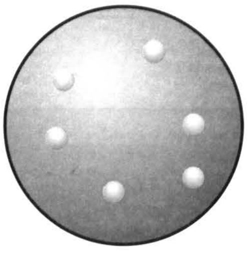
Well it'd be nice if that were so, but it just ain't — never has been and never will be. Increasing scientific knowledge involves making mistakes along the way. Let me explain...

Scientists come up with ***hypotheses*** *— then* ***test*** *them*

1) Scientists try and explain things. Everything.

2) They start by observing or thinking about something they don't understand — it could be anything, e.g. planets in the sky, a person suffering from an illness, what matter is made of... anything.

3) Then, using what they already know (plus a bit of insight), they come up with a hypothesis (a theory) that could explain what they've observed.



About 100 years ago, we thought atoms looked like this.

> Remember, a hypothesis is just a theory, a belief. And believing something is true doesn't make it true — not even if you're a scientist.

4) So the next step is to try and convince other scientists that the hypothesis is right — which involves using evidence. First, the hypothesis has to fit the evidence already available — if it doesn't, it'll convince no one.

5) Next, the scientist might use the hypothesis to make a prediction — a crucial step. If the hypothesis predicts something, and then evidence from experiments backs that up, that's pretty convincing.

> This doesn't mean the hypothesis is true (the 2nd prediction, or the 3rd, 4th or 25th one might turn out to be wrong) — but a hypothesis that correctly predicts something in the future deserves respect.

A hypothesis is a good place to start

You might have thought that science was all about facts... well, it's not as cut and dried as that — you also need to know about the process that theories go through to become accepted, and how those theories change over time. Remember, nothing is set in stone...

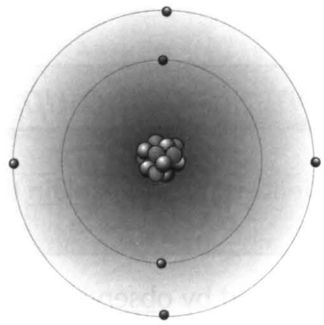
Theories Come, Theories Go

*Other scientists will **test** the hypotheses too*

1) Now then... other scientists will want to use the hypothesis to make their own predictions, and they'll carry out their own experiments. (They'll also try to reproduce earlier results.) And if all the experiments in all the world back up the hypothesis, then scientists start to have a lot of faith in it.

2) However, if a scientist somewhere in the world does an experiment that doesn't fit with the hypothesis (and other scientists can reproduce these results), then the hypothesis is in trouble. When this happens, scientists have to come up with a new hypothesis (maybe a modification of the old theory, or maybe a completely new one).

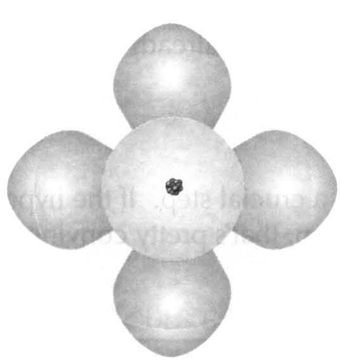
3) This process of testing a hypothesis to destruction is a vital part of the scientific process. Without the 'healthy scepticism' of scientists everywhere, we'd still believe the first theories that people came up with — like thunder being the belchings of an angered god (or whatever).



Then we thought they looked like this.

*If **evidence** supports a hypothesis, it's **accepted** — for now*

1) If pretty much every scientist in the world believes a hypothesis to be true because experiments back it up, then it usually goes in the textbooks for students to learn.



Now we think it's more like this.

2) Our currently accepted theories are the ones that have survived this 'trial by evidence' — they've been tested many, many times over the years and survived (while the less good ones have been ditched).

3) However... they never, never become hard and fast, totally indisputable fact.

You can never know... it'd only take one odd, totally inexplicable result, and the hypothesising and testing would start all over again.

You expect me to believe that — then show me the evidence...

If scientists think something is true, they need to produce evidence to convince others — it's all part of testing a hypothesis. One hypothesis might survive these tests, while others won't — it's how things progress. And along the way some hypotheses will be disproved — i.e. shown not to be true. So, you see... not everything scientists say is true. It's how science works.

Your Data's Got to Be Good

Evidence is the key to science — but not all evidence is equally good.
The way that evidence is gathered can have a big effect on how trustworthy it is.

Lab experiments *are better than* ***rumour*** *or* ***small samples***

1) Results from controlled experiments in laboratories are great. A lab is the easiest place to control variables so that they're all kept constant (except for the one you're investigating).

   This makes it easier to carry out a fair test.

   It's also the easiest way for different scientists around the world to carry out the same experiments. (There are things you can't study in a lab though, like climate.)

2) Old wives' tales, rumours, hearsay, 'what someone said', and so on, should be taken with a pinch of salt. They'd need to be tested in controlled conditions to be genuinely scientific.

3) Data based on samples that are too small don't have much more credibility that rumours do.

   A sample should be representative of the whole population (i.e. it should share as many of the various characteristics in the whole population as possible) — a small sample just can't do that.

Evidence is only ***reliable*** *if* ***other people can repeat it***

Scientific evidence needs to be reliable (or reproducible). If it isn't, then it doesn't really help.

RELIABLE means that the data can be reproduced by others.

Example: Cold fusion

In 1989, two scientists claimed that they'd produced 'cold fusion' (the energy source of the Sun — but without the enormous temperatures).

It was huge news — if true, this could have meant energy from sea water — the ideal energy solution for the world... forever.

However, other scientists just couldn't get the same results — i.e. the results weren't reliable. And until they are, 'cold fusion' isn't going to be generally accepted as fact.

Reliability is really important in science

The scientific community won't accept someone's data if it can't be repeated by anyone else. It may sound like a really fantastic new theory, but if there's no other support for it, it just isn't reliable.

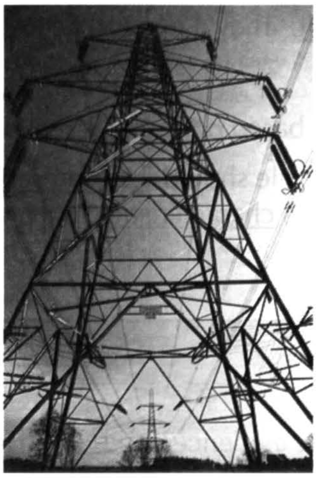
Your Data's Got to Be Good

*Evidence also needs to be **valid***

To answer scientific questions scientists often try to link changes in one variable with changes in another. This is useful evidence, as long as it's valid.

VALID means that the data is reliable AND answers the original question.

Example: Do power lines cause cancer?



Some studies have found that children who live near overhead power lines are more likely to develop cancer. What they'd actually found was a correlation between the variables "presence of power lines" and "incidence of cancer" — they found that as one changed, so did the other.

But this evidence is not enough to say that the power lines cause cancer, as other explanations might be possible.

For example, power lines are often near busy roads, so the areas tested could contain different levels of pollution from traffic. Also, you need to look at types of neighbourhoods and lifestyles of people living in the tested areas (could diet be a factor... or something else you hadn't thought of...).

So these studies don't show a definite link and so don't answer the original question.

Controlling** all the variables is **really hard

In reality, it's very hard to control all the variables that might (just might) be having an effect.

You can do things to help — e.g. choose two groups of people (those near power lines and those far away) who are as similar as possible (same mix of ages, same mix of diets etc).
But you can't easily rule out every possibility.

If you could do a properly controlled lab experiment, that'd be better — but you just can't do it without cloning people and exposing them to things that might cause cancer... hardly ethical.

Does the data really say that?

If it's so hard to be definite about anything, how does anybody ever get convinced about anything?
Well, what usually happens is that you get a load of evidence that all points the same way.
If one study can't rule out a particular possibility, then maybe another one can. So you gradually build up a whole body of evidence, and it's this (rather than any single study) that convinces people.

Bias and How to Spot it

Scientific results are often used to make a point, but results are sometimes presented in a biased way.

You don't need to lie to make things biased

1) For something to be misleading, it doesn't have to be untrue. We tend to read scientific facts and assume that they're the 'truth', but there are many different sides to the truth. Look at this headline...

1 in 2 people are of above average weight ← *Sounds like we're a nation of fatties.*

2) But an average is a kind of 'middle value' of all your data. Some readings are higher than average (about half of them, usually). Others will be lower than average (the other half).

So the above headline could just as accurately say: → **1 in 2 people are of below average weight**

3) The point is... both headlines sound quite worrying, even though they're not. That's the thing... you can easily make something sound really good or really bad — even if it isn't. You can...

(1) ...use only some of the data, rather than all of it:	"Many people lost weight using the new SlimAway diet. Buy it now!!"	*"Many" could mean anything — e.g. 50 out of 5000 (i.e. 1%). But that could be ignoring most of the data.*
(2) ...phrase things in a 'leading' way:	90% fat free!	*Would you buy it if it were "90% cyanide free"? That 10% is the important bit, probably.*
(3) ...use a statistic that supports your point of view:	The amount of energy wasted is increasing. / Energy wasted per person is decreasing. / The rate at which energy waste is increasing is slowing down.	*These describe the same data. But two sound positive and one negative.*

Think about why things might be biased

1) People who want to make a point can sometimes present data in a biased way to suit their own purposes (sometimes without knowing they're doing it).

2) And there are all sorts of reasons why people might want to do this — for example...

- Governments might want to persuade voters, other governments, journalists, etc. Evidence might be ignored if it could create political problems, or emphasised if it helps their cause.
- Companies might want to 'big up' their products. Or make impressive safety claims, maybe.
- Environmental campaigners might want to persuade people to behave differently.

3) People do it all the time. This is why any scientific evidence has to be looked at carefully. Are there any reasons for thinking the evidence is biased in some way?

- Does the experimenter (or the person writing about it) stand to gain (or lose) anything?
- Might someone have ignored some of the data for political or commercial reasons?
- Is someone using their reputation rather than evidence to help make their case?

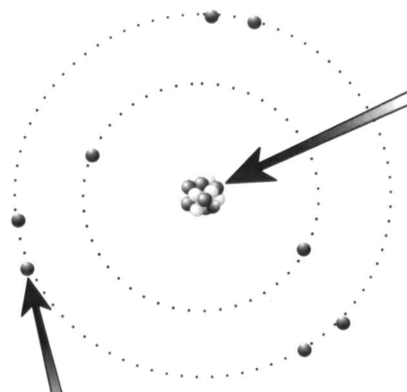
Scientific data's not always misleading, you just need to be careful. The most credible argument will be the one that describes all the data that was found, and gives the most balanced view of it.

Atoms

Hello, good evening and welcome to Chemistry. This section covers all of Chemistry's essential gory details — about atoms, their innards, and what they get up to with each other when no one's looking.

Structure of the atom — there's nothing to it

The structure of atoms is quite simple. Just learn and enjoy, my friend.



The Nucleus

1) It's in the middle of the atom. It contains protons and neutrons. (It's the number of protons in an atom that decides what element it is.)
2) The nucleus has an overall positive charge because protons are positively charged while neutrons have no charge.
3) Almost the whole mass of the atom is concentrated in the nucleus. But size-wise it's tiny compared to the atom as a whole.

The Electrons

1) They move around the nucleus in energy levels called shells. (Each shell is only allowed a certain number of electrons.)
2) They have a negative charge (electrons and protons have equal but opposite charges).
3) They're tiny compared to the nucleus (they have virtually no mass), but as they move around they cover a lot of space. (The size of their orbits determines how big the atom is.)

Number of protons equals number of electrons

1) Neutral atoms have no charge overall.
2) This is because the number of protons always equals the number of electrons in a neutral atom, and the charge on the electrons is the same size as the charge on the protons, but opposite.
3) The number of neutrons isn't fixed but is usually about the same as the number of protons.

Know your particles

1) Protons are heavy and positively charged.
2) Neutrons are heavy and neutral (no charge).
3) Electrons are tiny and negatively charged.

Particle	Relative mass	Relative charge
Proton	1	+1
Neutron	1	0
Electron	1/2000	-1

Each element has an atomic number and a mass number

1) The atomic number says how many protons there are in an atom, and is unique to that element.
2) The atomic number also tells you the number of electrons.
3) The mass number is the total number of protons and neutrons in the atom. So if you want to find the number of neutrons in an atom, just subtract the atomic number from the mass number.

MASS NUMBER
Total number of protons and neutrons. → 16

ATOMIC NUMBER (OR PROTON NUMBER)
Number of protons, which is equal to the number of electrons. → 8

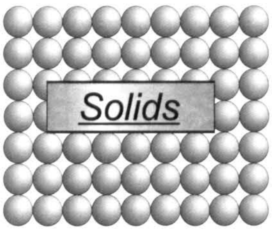
$^{16}_{8}O$

Solids, Liquids and Gases

You can explain a lot of things (including perfumes) if you get your head round this lot.

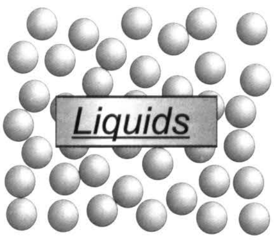
States of matter — depend on the forces between particles

All stuff is made of particles (molecules, ions or atoms) that are constantly moving, and the forces between these particles can be weak or strong, depending on whether it's a solid, liquid or a gas.



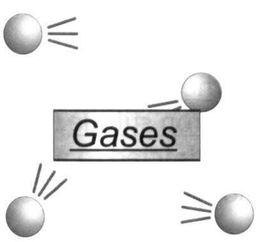
1) There are strong forces of attraction between particles, which holds them in fixed positions in a very regular lattice arrangement.
2) The particles don't move from their positions, so all solids keep a definite shape and volume, and don't flow like liquids.
3) The particles vibrate about their positions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

If you heat the solid (give the particles more energy), eventually the solid will melt and become liquid.



1) There is some force of attraction between the particles. They're free to move past each other, but they do tend to stick together.
2) Liquids don't keep a definite shape and will flow to fill the bottom of a container.
3) 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.

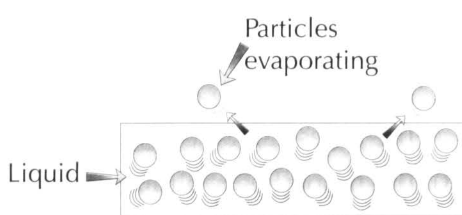
If you now heat the liquid, eventually it will boil and become gas.



1) There's next to no force of attraction between the particles — they're free to move. They travel in straight lines and only interact when they collide.
2) Gases don't keep a definite shape or volume and will always fill any container. When particles bounce off the walls of a container they exert a pressure on the walls.
3) 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.

Some liquids are more volatile than others

1) When a liquid is heated, the heat energy is transferred to the particles, which makes them move faster.
2) Some particles move faster than others.
3) Fast-moving particles at the surface will overcome the forces of attraction from the other particles and escape. This is evaporation.
4) How easily a liquid evaporates is called its volatility.



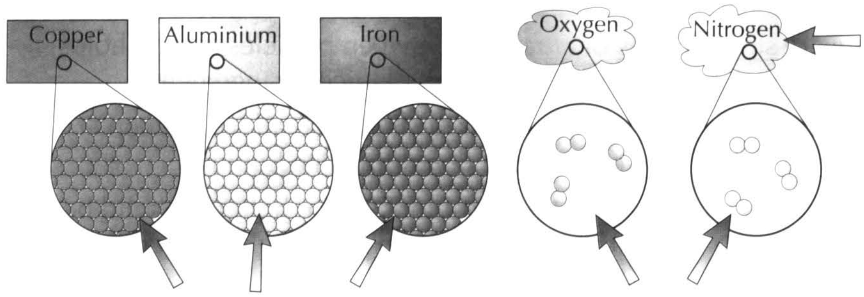
Evaporation is why you can smell stuff, even solids and liquids. A few particles have enough energy to evaporate and the smell receptors in your nose detect them — and hey presto — you smell the substance. Perfumes, air fresheners, etc. are usually volatile liquids so they evaporate enough for you to smell them.

Elements, Compounds and Mixtures

There are only about 100 or so different kinds of atoms, which doesn't sound too bad.
But they can join together in loads of different combinations, which makes life more complicated.

Elements consist of one type of atom only

Quite a lot of everyday substances are elements:



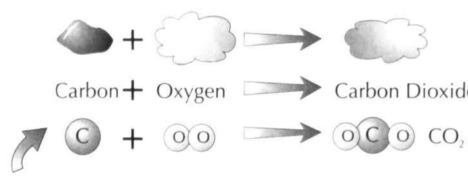
Nitrogen is the most common element in the air (about 78%).

The atoms in solids are tightly packed.

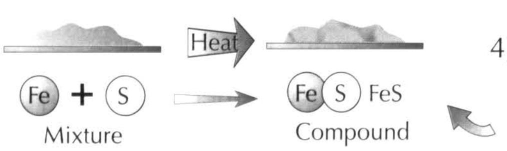
Atoms in gases often go round in pairs. A molecule with two atoms in it is called a diatomic molecule.

Compounds are chemically bonded

A compound is a substance that is made of two or more different elements which are chemically joined (bonded) together.

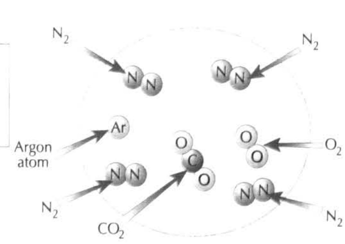


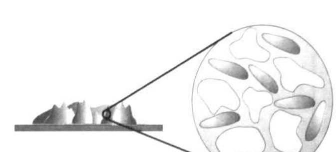
1) For example, carbon dioxide is a compound formed from a chemical reaction. One carbon atom reacts with two oxygen atoms to form a molecule of carbon dioxide, with the formula CO_2.
2) It's very difficult to separate the two original elements out again.
3) The properties of a compound are often totally different from the properties of the original elements.
4) For example, if a mixture of iron and sulfur is heated, the iron and sulfur atoms react to form the compound iron sulfide (FeS). Iron sulfide is not much like iron (e.g. it's not attracted to a magnet), nor is it much like sulfur (e.g. it's not yellow in colour).



Mixtures are easily separated — not like compounds

1) Unlike in a compound, there's no chemical bond between the different parts of a mixture. The parts can be separated out by physical methods such as distillation (see page 32).
2) Air is a mixture of gases, mainly nitrogen, oxygen, carbon dioxide and argon. The gases can all be separated out fairly easily.
3) The properties of a mixture are just a mixture of the properties of the separate parts.
4) 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.
5) Crude oil is a mixture of different length hydrocarbon molecules — see page 32.





Iron and sulfur mixed together, but unreacted.

The Periodic Table

In the 1800s chemists were keen to try and find patterns in the elements they knew about. And the more elements that were identified, the clearer those patterns became...

*Dmitri **Mendeleev** arranged the **elements in groups***

Mendeleev's Table of the Elements

H																
Li	Be											B	C	N	O	F
Na	Mg											Al	Si	P	S	Cl
K	Ca	*	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	*	*	As	Se	Br
Rb	Sr	Y	Zr	Nb	Mo	*	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I
Cs	Ba	*	*	Ta	W	*	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi		

1) In 1869, a Russian scientist called Dmitri Mendeleev arranged the 50 or so known elements in order of atomic mass to make a Table of Elements.

2) Mendeleev's table placed elements with similar chemical properties in the same vertical groups — but he found that he had to leave gaps in his table to make this work.

3) The gaps in Mendeleev's table of elements were really clever because they predicted the properties of undiscovered elements.

4) Since then new elements have been found which fit into the gaps in Mendeleev's table. Over the last hundred years or so the table has been refined to produce the periodic table we know (and love) today...

Elementary my dear Mendeleev

Even though its not the periodic table we use today, its important to know how much of an influence Mendeleev's periodic table has been on our modern periodic table. Make sure you know how Mendeleev arranged his table and how it came to look like the one we're used to using today.

The Periodic Table

The periodic table is really important. You can try to do chemistry without it, but it's likely to all end in disaster. Learn the rules, know the trends and practise safe chemistry.

*The **periodic table** puts **elements** with **similar properties** together*

1) The periodic table is laid out so that elements with similar properties form columns.
2) These vertical columns are called groups and Roman numerals are often (but not always) used for them.
3) If you know the properties of one element, you can predict properties of other elements in that group.
4) For example the Group 1 elements are Li, Na, K, Rb, Cs and Fr. They're all metals and they react the same way. E.g. they all react with water to form an alkaline solution and hydrogen gas.
5) You can also make predictions about 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.
6) There are 100ish elements, which all materials are made of. 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.

	Group 1	Group 2											Group 3	Group 4	Group 5	Group 6	Group 7	Group 0
							1 H Hydrogen 1											4 He Helium 2
2	7 Li Lithium 3	9 Be Beryllium 4											11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10
3	23 Na Sodium 11	24 Mg Magnesium 12											27 Al Aluminium 13	28 Si Silicon 14	31 P Phosphorus 15	32 S Sulfur 16	35.5 Cl Chlorine 17	40 Ar Argon 18
4	39 K Potassium 19	40 Ca Calcium 20	45 Sc Scandium 21	48 Ti Titanium 22	51 V Vanadium 23	52 Cr Chromium 24	55 Mn Manganese 25	56 Fe Iron 26	59 Co Cobalt 27	59 Ni Nickel 28	63.5 Cu Copper 29	65 Zn Zinc 30	70 Ga Gallium 31	73 Ge Germanium 32	75 As Arsenic 33	79 Se Selenium 34	80 Br Bromine 35	84 Kr Krypton 36
5	85 Rb Rubidium 37	88 Sr Strontium 38	89 Y Yttrium 39	91 Zr Zirconium 40	93 Nb Niobium 41	96 Mo Molybdenum 42	99 Tc Technetium 43	101 Ru Ruthenium 44	103 Rh Rhodium 45	106 Pd Palladium 46	108 Ag Silver 47	112 Cd Cadmium 48	115 In Indium 49	119 Sn Tin 50	122 Sb Antimony 51	128 Te Tellurium 52	127 I Iodine 53	131 Xe Xenon 54
6	133 Cs Caesium 55	137 Ba Barium 56	57-71 Lanthanides	178 Hf Hafnium 72	181 Ta Tantalum 73	184 W Tungsten 74	186 Re Rhenium 75	190 Os Osmium 76	192 Ir Iridium 77	195 Pt Platinum 78	197 Au Gold 79	201 Hg Mercury 80	204 Tl Thallium 81	207 Pb Lead 82	209 Bi Bismuth 83	210 Po Polonium 84	210 At Astatine 85	222 Rn Radon 86
7	223 Fr Francium 87	226 Ra Radium 88	89-103 Actinides															

reactive metals · transition metals · other metals · non-metals · noble gases · separates metals from non-metals

The periodic table gives you the mass number and atomic number of all known elements.

mass number (total number of protons & neutrons) → 4
He Helium
atomic number (number of protons) → 2

The periodic table is all you need

One important thing to understand on these pages is that a scientific theory (such as "elements can be grouped in a table according to their properties") can be used to make predictions (such as "there are gaps in the table so there must be some undiscovered elements to fill those gaps"). Got it... good.

Balancing Equations

All chemical reactions can be shown using an equation. Unfortunately, getting equations right takes a bit of practice. So make sure you get a bit of practice — don't just skate over them.

Atoms *aren't lost or made* in chemical reactions

1) During chemical reactions, things don't appear out of nowhere and things don't just disappear.
2) You still have the same atoms at the end of a chemical reaction as you had at the start. They're just arranged in different ways.
3) Balanced symbol equations show the atoms at the start (the reactant atoms) and the atoms at the end (the product atoms) and how they're arranged. For example:

Word equation:	Magnesium	+	Oxygen	→	Magnesium oxide
Balanced symbol equation:	$2Mg$	+	O_2	→	$2MgO$

4) Because atoms aren't gained or lost, the mass of the reactants equals the mass of the products. So, if you react 6 g of magnesium with 4 g of oxygen, you'd end up with 10 g of magnesium oxide.

Balancing the equation — match them up one by one

1) There must always be the same number of atoms of each element 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 (H_2SO_4) with sodium hydroxide ($NaOH$) to get sodium sulfate (Na_2SO_4) and water (H_2O):

$$H_2SO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$$

The formulas are all correct but the numbers of some atoms don't match up on both sides. E.g. there are 3 H's on the left, but only 2 on the right. You can't change formulas like H_2O to H_3O. You can only put numbers in front of them:

Method: balance just *ONE* type of atom at a time

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 — if so, just pencil in another number and see where that gets you.
3) Carry on chasing unbalanced elements and it'll sort itself out pretty quickly.

I'll show you. In the equation above you soon notice we're short of H atoms on the right-hand side.

1) The only thing you can do about that is make it $2H_2O$ instead of just H_2O:

$$H_2SO_4 + NaOH \rightarrow Na_2SO_4 + 2H_2O$$

2) But that now causes too many H atoms and O atoms on the right-hand side, so to balance that up you could try putting $2NaOH$ on the left-hand side:

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

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

Warm-Up and Exam Questions

It's easy to think you've learnt everything in the section until you try the warm-up questions. Don't panic if there are bits you've forgotten. Just go back over those bits until they're firmly fixed in your brain.

Warm-Up Questions

1) In a neutral atom, which particles are always equal in number?
2) Explain the difference between mass number and atomic number.
3) What happens to a gas when it is heated?
4) Name an element in which the atoms are tightly packed at room temperature.
5) Balance this equation for the reaction of glucose ($C_6H_{12}O_6$) and oxygen:
   $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
6) In this equation: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
   explain the difference between the meaning of the $_2$ in H_2SO_4 and the 2 in 2NaOH.

Exam Questions

1 The proton has a relative mass of 1. What is the relative mass of the neutron?

A 2000

B 1

C 1/2000

D 2

(1 mark)

2 Which of the following statements about Mendeleev's periodic table is **not** true?

A Elements with similar chemical properties were placed in vertical groups.

B Gaps were left which helped in predicting the properties of undiscovered elements.

C The elements were arranged in order of atomic mass.

D Mendeleev's periodic table contained over 100 elements.

(1 mark)

3 (a) Using the ideas of kinetic theory, explain:

(i) why liquids flow.

(2 marks)

(ii) why gases fill their containers.

(2 marks)

(iii) why solids have a fixed shape.

(2 marks)

(b) (i) Explain the process of evaporation, in terms of particles.

(1 mark)

(ii) Explain why perfumes are made from substances which evaporate easily.

(1 mark)

Exam Questions

4 Air is a mixture of gases, mainly nitrogen, oxygen, carbon dioxide and argon.

(a) How is a compound different from a mixture?

(1 mark)

(b) (i) What group is argon in?

(1 mark)

(ii) Give the mass number of argon.

(1 mark)

(c) Name one gas found in air which is a compound.

(1 mark)

(d) Oxygen gas, O_2, reacts with magnesium to form magnesium oxide, MgO.
Write a balanced symbol equation for this reaction.

(2 marks)

5 The modern periodic table can be divided into metals and non-metals.
The non-metals are

A on the left of the periodic table

B on the right of the periodic table

C in the middle of the periodic table

D in Group 2

(1 mark)

6 Which of these statements about chemical reactions is **not** true?

A The mass of the reactants is always equal to the mass of the products.

B Atoms are neither created nor destroyed in a reaction.

C The mass of the products is always less than the mass of the reactants.

D In a written equation, the mass of all the atoms on the left of the arrow is equal to the mass of all the atoms on the right of the arrow.

(1 mark)

7 Sulfuric acid, H_2SO_4, reacts with ammonia, NH_3, to form ammonium sulfate, $(NH_4)_2SO_4$.

(a) Write the word equation for this reaction.

(1 mark)

(b) Write a balanced symbol equation for this reaction.

(2 marks)

(c) In the balanced equation, how many atoms are there in the reactants?

(1 mark)

Revision Summary for Section One

There wasn't anything too ghastly in this section, and a few bits were even quite interesting, I reckon. But you've got to make sure the facts are all firmly embedded in your brain and that you really understand them. These questions will let you see what you know and what you don't. If you get stuck on any, you need to look at that stuff again. Keep going till you can do them all without coming up for air.

1) Sketch an atom. Label the nucleus and the electrons.

2) Name the three types of particle in an atom. State the relative mass and charge of each particle.

3) The element boron is written as $^{11}_{5}B$. How many neutrons does an atom of this element contain? How many electrons does a neutral boron atom have?

4) A substance keeps the same volume, but changes its shape according to the container it's held in. Is it a solid, a liquid or a gas?

5) Are the forces of attraction between the particles in a liquid stronger or weaker than those in a gas?

6) What does it mean if a liquid is said to be very volatile?

7) Explain the difference between a mixture and a compound.

8) Compounds and mixtures are both equally difficult to separate out — true or false?

9)* Which elements make up a molecule of Na_2CO_3? How many atoms are there of each element?

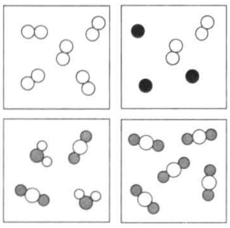
10)* Say which of the diagrams on the right show:

a) a mixture of compounds,

b) a mixture of elements,

c) an element,

d) a compound.



11) Explain how Mendeleev arranged the known elements in a table. How did he predict new elements?

12)* Which element's properties are more similar to magnesium's: calcium or iron?

13) Name one element that you would find in Group 1 of the periodic table.

14)* Balance these equations:

a) $CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$

b) $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$

15)* Write a balanced equation for the reaction between potassium and water to form potassium hydroxide (KOH) and hydrogen.

*Answers on page 253.

Using Limestone

The Mendip Hills and the Yorkshire Dales are mainly made of a rock called limestone. When limestone is dug out of the ground it's great for building stuff like houses and churches from. You need limestone to make mortar, cement, concrete and glass too. In fact, it's blooming marvellous.

Limestone is used as a **building material**

1) Limestone is a bit of a boring grey/white colour. It's often formed from sea shells and, although the original shells are mostly crushed, there are still quite a few fossilised shells remaining.
2) It's quarried out of the ground. This causes some environmental problems though — see next page.

St Paul's Cathedral is made from limestone.

3) It's great for making into blocks for building with. Fine old buildings like cathedrals are often made purely from limestone blocks. It's also used for statues and fancy carved bits on nice buildings too.
4) Limestone's virtually insoluble in plain water. But acid rain is a big problem. The acid reacts with the limestone and dissolves it away.
5) Limestone can also be crushed up into chippings and used in road surfacing.

Limestone is mainly **calcium carbonate**

1) Limestone is mainly calcium carbonate — $CaCO_3$. When it's heated it thermally decomposes (breaks down) to make calcium oxide (quicklime) and carbon dioxide.

calcium carbonate → calcium oxide + carbon dioxide

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

When other carbonates are heated, they decompose in the same way (e.g. $Na_2CO_3 \rightarrow Na_2O + CO_2$).

2) When you add water to quicklime you get slaked lime. Slaked lime is actually calcium hydroxide.

quicklime + water → slaked lime or $CaO + H_2O \rightarrow Ca(OH)_2$

3) Slaked lime is an alkali which can be used to neutralise acid soils in fields. Powdered limestone can be used for this too, but the advantage of slaked lime is that it works much faster.

Using Limestone

Limestone is really very handy. However, digging huge amounts of limestone out of the ground can have a quite a significant negative effect on the environment.

*Limestone is used to make other **building materials** too*

1) Powdered limestone is heated in a kiln with powdered clay to make cement.
2) Cement can be mixed with sand and water to make mortar. Mortar is the stuff you stick bricks together with. You can also use it to cover outside walls.

3) Or you can mix cement with sand, water and gravel to make concrete. And by including steel rods, you get reinforced concrete — a composite material with the hardness of concrete and the strength of steel.

4) And believe it or not — limestone is also used to make glass. You just heat it with sand and sodium carbonate until it melts.

*Extracting rocks can cause **environmental damage***

1) Quarrying uses up land and destroys habitats. It costs money to make quarry sites look pretty again. And the waste materials from mines and quarries produce unsightly tips.
2) Transporting rock can cause noise and pollution, and the quarrying process itself produces dust and makes a lot of noise — they often use dynamite to blast the rock out of the ground.
3) Disused sites can be dangerous. Disused mines have been known to collapse. And quarries are sometimes turned into (very deep) lakes — people drown in them every year.

Limestone's amazingly useful

It sounds like you can achieve pretty much anything with limestone. Fred Flintstone even managed to make his car wheels and bowling balls out of rock (although I'm not 100% certain it was limestone).

Metals from Rocks

You don't tend to find big lumps of pure metal in the ground — the metal atoms tend to be joined to other atoms in compounds. It can be a bit of a tricky, expensive process to separate the metal out.

Ores contain enough metal to make extraction worthwhile

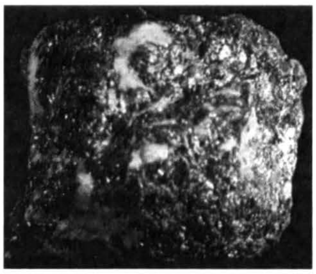
1) Rocks are made of minerals. (Minerals are just solid elements and compounds.)
2) A metal ore is a mineral which contains enough metal to make it worthwhile extracting the metal from it.
3) In many cases the ore is an oxide of the metal. Here are a few examples:

As technology improves, it becomes possible to extract more metal from a sample of rock than previously. So it might now be worth extracting metal that wasn't worth extracting in the past.

a) A type of iron ore is called haematite. This is iron(III) oxide (Fe_2O_3).

b) The main aluminium ore is called bauxite. This is aluminium oxide (Al_2O_3).

c) A type of copper ore is called chalcopyrite. This is copper iron sulfide ($CuFeS_2$).



Chalcopyrite — a copper ore

4) There's a limited amount of ores — they're "finite resources".
5) People have to balance the social, economic and environmental effects of mining the ores. Most of the issues are exactly the same as those to do with quarrying limestone on page 16.

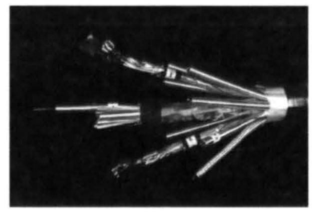
6) So mining metal ores is good because it means that useful products can be made. It also provides local people with jobs and brings money into the area. This means services such as transport and health can be improved.
7) But mining ores is bad for the environment as it causes noise, scarring of the landscape and loss of habitats. Deep mine shafts can also be dangerous for a long time after the mine has been abandoned.

Metals from Rocks

Copper has some important uses. So it's useful to know how to extract it as efficiently as possible.

Copper is purified by electrolysis

1) Copper is a transition metal. It is hard, strong and has a high melting point.
2) It is a good conductor of electricity, so it's ideal for drawing out into electrical wires.



3) It can also be made into pipes, and as it's below hydrogen in the reactivity series (see p.20), it doesn't react with water. This makes it great for using in plumbing.

4) If you look on page 20 you'll see that copper can be easily extracted by reduction with carbon. But the copper produced this way is impure — and impure copper doesn't conduct electricity very well.
5) So electrolysis is used to purify it (see p.144), even though it's expensive. This produces very pure copper, which is a much better conductor.

Electrolysis uses electricity to separate the metal from the ore. Aluminium is another metal separated from its ore in this way.

Copper-rich ores are in short supply

1) The supply of copper-rich ores is limited, so it's important to recycle as much copper as possible.
2) The demand for copper is growing and this may lead to shortages in the future.
3) Scientists are looking into new ways of extracting copper from low-grade ores (ores that only contain small amounts of copper) or from the waste that is currently produced when copper is extracted.
4) One way is to use bacteria to separate copper from copper sulfide. The bacteria get energy from the bond between copper and sulfur, separating out the copper from the ore in the process.
5) This process is slow but it's more environmentally friendly than other methods, which need lots of energy and release sulfur dioxide gas which causes acid rain.

Copper is a really useful metal

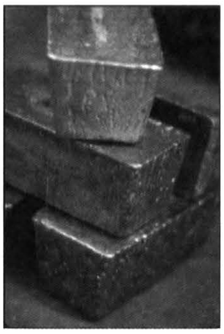
The skin of the Statue of Liberty is made of copper — about 80 tonnes of it in fact. Its surface reacts with gases in the air to form copper carbonate — which is why it's that pretty shade of green. It was a present from France to the United States — I wonder if they found any wrapping paper big enough?

Extraction of Metals

How easy it is to get a metal out of its ore all comes down to the metal's position in the reactivity series.

More reactive metals are harder to get

1) A few unreactive metals like gold are found in the Earth as the metal itself, rather than as a compound.



2) But most metals need to be extracted from their ores using a chemical reaction.

3) More reactive metals, like sodium, are harder to extract — that's why it took longer to discover them.

A more reactive metal displaces a less reactive metal

1) More reactive metals react more strongly than less reactive metals.
2) This means that a metal can be extracted from its oxide by any more reactive metal. The more reactive metal bonds more strongly to the oxygen and pushes out the less reactive metal.

   E.g. tin could be extracted from tin oxide by more reactive iron.

   tin oxide + iron → iron oxide + tin

3) And if you put a more reactive metal into the solution of a dissolved metal compound, the more reactive metal will replace the less reactive metal in the compound.

   E.g. put an iron nail in a solution of copper sulfate and the more reactive ireon will "kick out" the less reactive copper from the solution. You end up with iron sulfate solution and copper metal.

   copper sulfate + iron → iron sulfate + copper

4) But if a piece of silver metal is put into a solution of copper sulfate, nothing happens. The more reactive metal (copper) is already in the solution.

Extraction of Metals

*Some metals can be **extracted** by **reduction** with **carbon***

1) Electrolysis (splitting with electricity) is one way of extracting a metal from its ore.
2) The other common way is chemical reduction using carbon or carbon monoxide.
3) When an ore is reduced, oxygen is removed from it, e.g.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

iron(III) oxide + carbon monoxide → iron + carbon dioxide

*Position in the **reactivity series** is important for extraction*

1) Metals higher than carbon in the reactivity series have to be extracted using electrolysis, which is expensive.
2) 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.
3) This is because carbon can only take the oxygen away from metals which are less reactive than carbon itself is.

	The Reactivity Series		
Extracted using electrolysis	Potassium	K	*more reactive*
	Sodium	Na	
	Calcium	Ca	
	Magnesium	Mg	
	Aluminium	Al	
	CARBON	C	
Extracted by reduction using carbon	Zinc	Zn	
	Iron	Fe	
	Tin	Sn	
	Copper	Cu	*less reactive*

Extraction of metals is difficult

Extracting metals isn't cheap. You have to pay for special equipment, energy and labour. Then there's the cost of getting the ore to the extraction plant. If there's a choice of extraction methods, a company always picks the cheapest, unless there's a good reason not to — they're not extracting it for fun.

Warm-Up and Exam Questions

You've arrived at the next set of warm-up and exam questions. It's really important to find out what you know (as well as what you think you know but actually don't). So give them a go.

Warm-Up Questions

1) Give three major uses of limestone.
2) Give an example of environmental damage caused by quarrying.
3) Describe the difference between cement and mortar.
4) Name a metal ore.
5) Explain why copper used in electrical wires needs to be purified by electrolysis.
6) Name a metal which can be extracted from its ore by reduction with carbon.

Exam Questions

1 Limestone is mainly calcium carbonate, $CaCO_3$. When heated, it thermally decomposes to produce calcium oxide and carbon dioxide.

(a) Write a balanced symbol equation for this reaction.

(1 mark)

(b) Calcium oxide is also known as quicklime. When water is added to quicklime, slaked lime is produced.

(i) Write the chemical name and formula of slaked lime.

(2 marks)

(ii) Give **one** use of slaked lime.

(1 mark)

2 Which of the following statements does **not** describe copper?

A A hard, strong transition metal.

B The main product extracted from the ore bauxite.

C A metal that is less reactive than iron.

D A material used to make electrical wires.

(1 mark)

3 Mining metal ores has social, economic and environmental effects.

(a) Give **two** positive effects of mining metal ores.

(2 marks)

(b) Give **two** negative effects of mining metal ores.

(2 marks)

Exam Questions

4 The diagram shows part of the reactivity series of metals, together with carbon.

Potassium	K	*more*
Sodium	Na	*reactive*
Calcium	Ca	
Magnesium	Mg	
Aluminium	Al	
<u>CARBON</u>	<u>C</u>	
Zinc	Zn	
Iron	Fe	
Tin	Sn	*less*
Copper	Cu	*reactive*

(a) Name one metal which is extracted from its ore using electrolysis.

(1 mark)

(b) Some metals can be extracted from their ores by reduction with carbon, producing the metal and carbon dioxide.

(i) Explain the meaning of reduction.

(1 mark)

(ii) Write a word equation for the reduction of zinc oxide by carbon.

(1 mark)

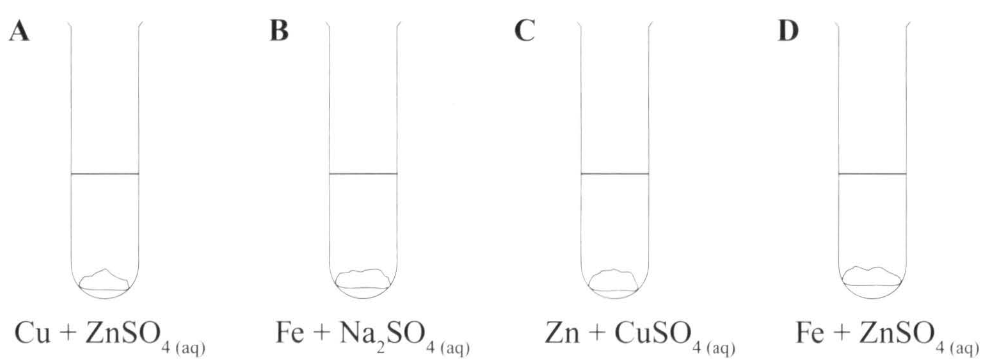
(c) Iron can be extracted by the reduction of iron(III) oxide (Fe_2O_3) with carbon monoxide (CO), to produce iron and carbon dioxide.

Write a balanced symbol equation for this reaction, including state symbols.

(3 marks)

(d) In which of these test tubes will a reaction occur?

(1 mark)



5 Copper needs to be extracted from its ore before it can be used.

(a) Give **two** uses of copper.

(2 marks)

(b) Why are scientists trying to find new ways to extract copper from low-grade ores?

(1 mark)

(c) It is possible to obtain copper from copper sulfide using bacteria.

(i) Give **one** advantage of using this method over other methods for extracting copper from copper sulfide.

(1 mark)

(ii) Give **one** disadvantage of using this method rather than other methods.

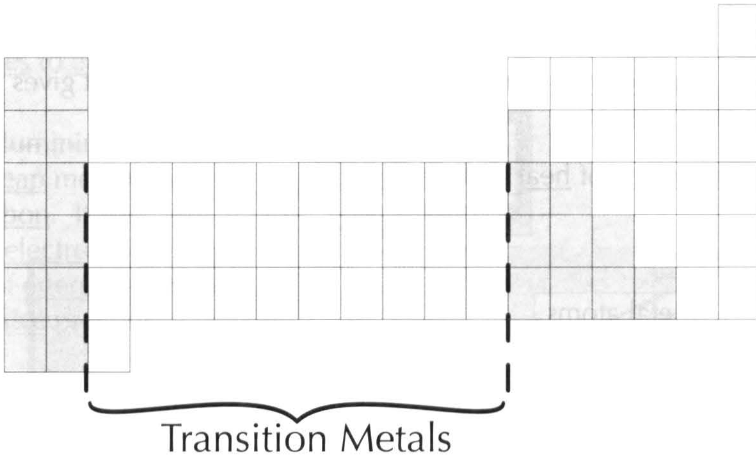
(1 mark)

Properties of Metals

Metals are all similar but slightly different. They have some basic properties in common, but each has its own specific combination of properties, which means you use different ones for different purposes.

Metals are on the left and middle of the periodic table

Most of the elements are metals — so they cover most of the periodic table.
In fact, only the elements on the far right are non-metals.
The so-called transition metals are found in the centre block of the periodic table.
Many of the metals in everyday use are transition metals — such as titanium, iron and nickel.



Metals are strong and bendy, and they're great conductors

All metals have some fairly similar basic properties.

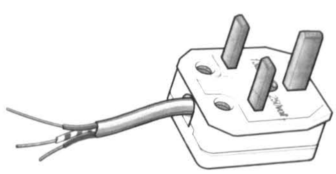
1) Metals are strong (hard to break), but they can be bent or hammered into different shapes.
2) They're great at conducting heat.
3) They conduct electricity well.

Metals (and especially transition metals) have loads of everyday uses because of these properties...

- Their strength and 'bendability' makes them handy for making into things like bridges and car bodies.

- Metals are ideal if you want to make something that heat needs to travel through, like a saucepan base.

- And their conductivity makes them great for making things like electrical wires.



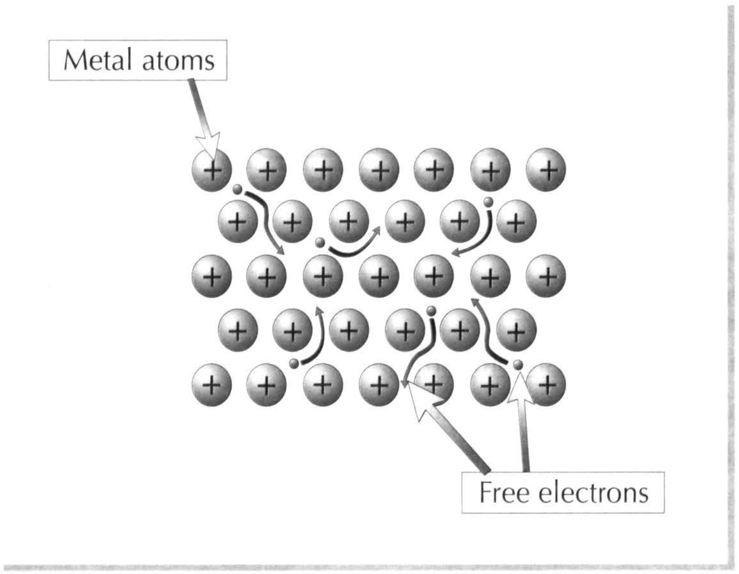
Transition metals have loads of everyday uses — partly because they're not crazily reactive like, say, potassium (which would catch fire if it got rained on).

Properties of Metals

The reason all metals have the same basic properties is because of the bonding in metals.
It's their exact properties which are used to match metals to their uses.

*It's the **structure of metals** that gives them their **properties***

1) All metals have the same basic properties. These are due to the special type of bonding in metals.
2) Metals consist of a giant structure of atoms held together with metallic bonds.
3) These special bonds allow the outer electron(s) of each atom to move freely.
4) This creates a "sea" of free electrons throughout the metal, which is what gives rise to many of the properties of metals.
5) This includes their conduction of heat and electricity.



*A metal's **exact properties** decide how it's best **used***

1) The properties above are typical properties of metals. Not all metals are the same though — their exact properties determine how they're used.
2) If you wanted to make an aeroplane, you'd probably use metal as it's strong and can be bent into shape, but you'd also need it to have a low density — so aluminium would be a good choice.
3) And if you were making replacement hips, you'd pick a metal that won't corrode when it comes in contact with water — it'd also have to have a low density too, and not too bendy. Titanium has all of these properties so it's used for this.

Properties of metals are all due to the "sea" of free electrons

So, all metals conduct electricity and heat and can be bent into shape. But lots of them have special properties too. You have to decide what properties you need and use a metal with those properties.

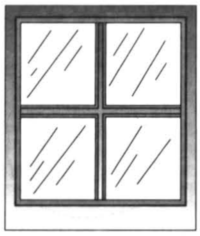
More Metals

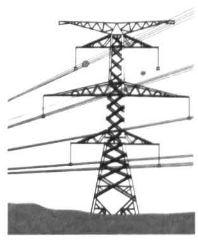
There are loads of metals. But if none of them have quite the properties you need, you could try an alloy.

Aluminium is useful, but expensive to extract

1) Aluminium has a low density and is corrosion-resistant.
2) These properties make aluminium a very useful structural material. It can be used for loads of things from window frames to electricity cables and aircraft.
3) You can't extract aluminium from its oxide by the cheap method of reduction with carbon. It has to be extracted by electrolysis. This requires lots of energy, which makes it an expensive process.

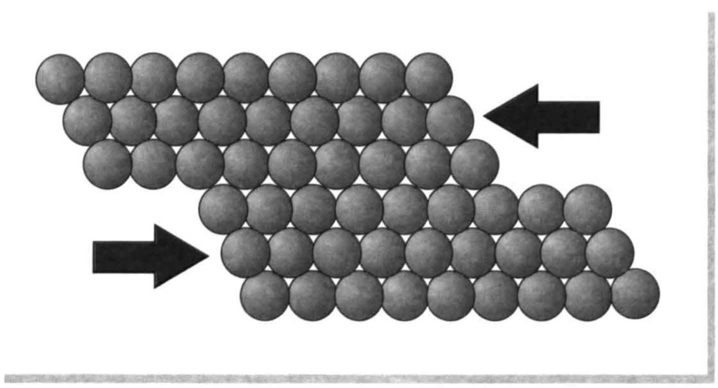
The aluminium reacts with oxygen in the air to form aluminium oxide. This sticks firmly to the aluminium below and stops any further reaction taking place.





Pure iron tends to be a bit too bendy

1) 'Iron' straight from the blast furnace is only 96% iron. The other 4% is impurities.
2) This impure iron is brittle. It's used for ornamental railings but it doesn't have many other uses. So all the impurities are removed from most blast furnace iron.
3) This pure iron has a regular arrangement of identical atoms. The layers of atoms can slide over each other, which makes the iron soft and easily shaped. This iron is far too bendy for most uses.



4) Most of the pure iron is changed into alloys called steels. Steels are made by adding small amounts of carbon (plus maybe other metals) to the iron.

An alloy is a mixture of two or more metals, or a mixture of a metal and a non-metal.

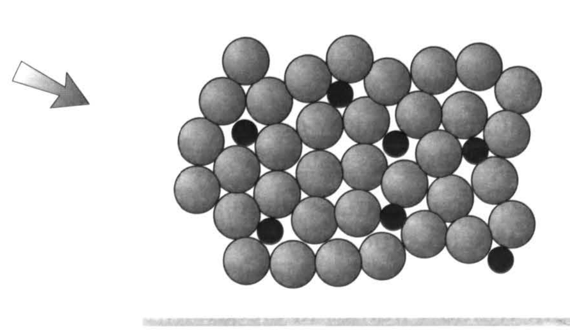
Most iron is changed into steel, otherwise it's too bendy or too brittle

The Eiffel Tower is made of iron — but the problem with iron is that it goes rusty if air and water get to it. So the Eiffel Tower has to be painted every seven years to make sure that it doesn't rust. This is quite a job and takes an entire year for a team of 25 painters. Too bad they didn't use stainless steel.

More Metals

Alloys are harder than pure metals

1) Different elements have <u>different sized atoms</u>. So when an element such as carbon is added to pure iron, the <u>smaller</u> carbon atom will <u>upset</u> the layers of pure iron atoms, making it more difficult for them to slide over each other. So alloys are <u>harder</u>.



2) Many metals in use today are actually <u>alloys</u>. For example:

BRONZE = COPPER + TIN

Bronze is <u>harder</u> than copper.
It's good for making medals and statues from.

CUPRONICKEL = COPPER + NICKEL

This is <u>hard</u> and <u>corrosion resistant</u>.
It's used to make "silver" coins.

GOLD ALLOYS ARE USED TO MAKE JEWELLERY

Pure gold is <u>too soft</u>. Metals such as zinc, copper, silver, palladium and nickel are used to harden the "gold".

ALUMINIUM ALLOYS ARE USED TO MAKE AIRCRAFT

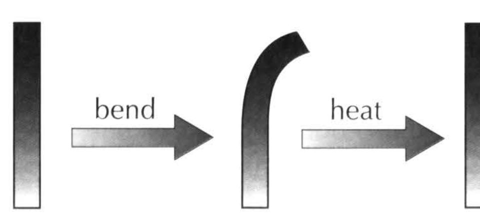
Aluminium has a <u>low density</u>, but it's <u>alloyed</u> with small amounts of other metals to make it <u>stronger</u>.

3) In the past, the development of alloys was by <u>trial and error</u>. But nowadays we understand much more about the properties of metals, so alloys can be <u>designed</u> for specific uses.

Smart alloys return to their original shape

1) <u>Nitinol</u> is a "<u>shape memory alloy</u>" — it has a <u>shape memory</u> property.

2) If you <u>bend</u> a wire made of this smart alloy, it'll go back to its <u>original shape</u> when it's <u>heated</u>. You can get specs with frames made from a smart alloy — you can sit on them and not destroy them.



3) At the moment, <u>metal fatigue</u> in smart alloys is a lot <u>worse</u> than in normal alloys. Smart alloys are also <u>more expensive</u> than steel or aluminium.

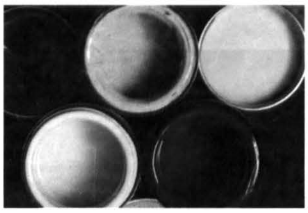
Alloys are really important in industry

If the properties of a metal aren't quite suited to a job, an alloy is often used instead. To make an alloy you mix one metal with another metal or non-metal. The finished alloy can be a lot harder, or less brittle — the properties can be varied and they can be made to suit a particular job really well.

Paints and Pigments

So, rocks are useful themselves (like limestone) and you can also extract useful metals from them. But they're not finished there — they also have a hand in art and interior design. Minerals come in lots of weird and wonderful colours, and they've been used as the pigments in paints for thousands of years.

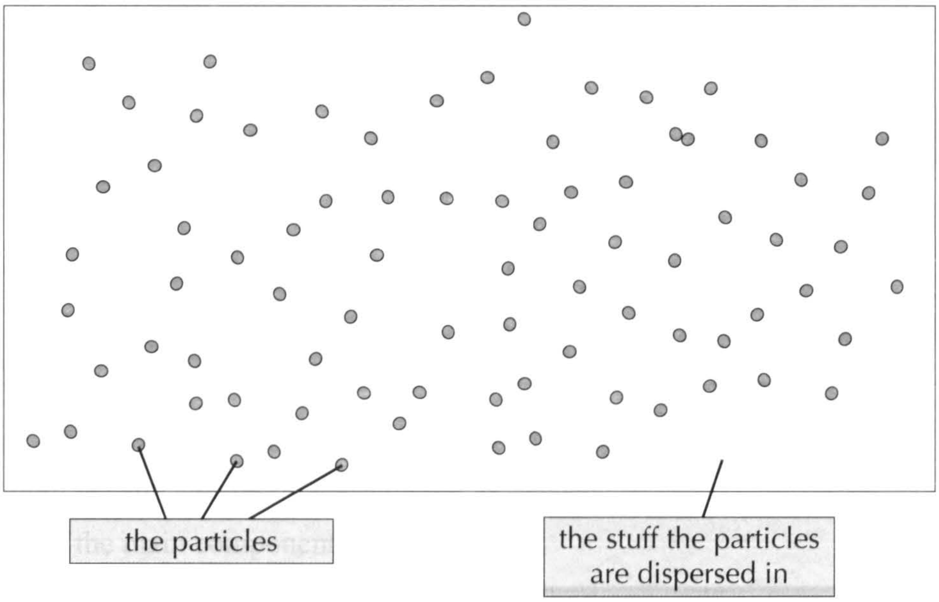
Pigments give paints their colours



1) Paint usually contains the following bits: solvent, binding medium and pigment.
2) The pigment gives the paint its colour.
3) The binding medium is a substance that carries the pigment bits and holds them together. When the binding medium turns to solid, it sticks the pigments to the surface you've painted.
4) The solvent is the liquid used to dissolve the binding medium and pigment bits and keep them in liquid form until they're used. Then the solvent evaporates, leaving the binding medium and pigment behind.

Paints are colloids

1) A colloid consists of really tiny particles of one kind of stuff dispersed in (mixed in with) another kind of stuff. They're mixed in, but not dissolved.
2) The particles can be bits of solid, droplets of liquid or bubbles of gas.
3) Colloids don't separate out because the particles are so small. They don't settle out at the bottom.
4) In an oil paint, the pigment is in really tiny bits dispersed in the oil. And then the solvent (if there is one — there isn't always) dissolves the oil to keep it all runny.



Paints and Pigments

Some paints are **water-based** and some are **oil-based**

1) Emulsion paints are water-based. The solvent in the paint is water, and the binding medium is usually a polymer such as polyurethane, acrylic or latex.

2) Traditional gloss paint and artists' oil paints are oil-based. This time, the binding material is oil, and the solvent in the paint is an organic compound that'll dissolve oil.

Some modern gloss paints are water-based.

3) Turpentine is used as a solvent for artists' oil paints. Some solvents in oil-based paints produce fumes which can be harmful — it's best to make sure there's plenty of ventilation when using oil-based gloss.

4) Whether you're creating a masterpiece in oils or painting your bedroom wall, you normally brush on the paint as a thin layer. The paint dries as the solvent evaporates. (A thin layer dries a heck of a lot quicker than a thick layer.)

5) With a water-based emulsion, the solvent evaporates, leaving behind the binder and pigment as a thin solid film. A thin layer of emulsion paint dries quite quickly.

6) Oil-based paints take rather longer to dry because the oil has to be oxidised by oxygen in the air before it turns solid.

In the exam, you might be asked to choose the best kind of paint for a job, given some info about paints.

- For example, to paint the outside part of a door you'd want a waterproof, hard-wearing paint.
- Oil-based paints are more hard-wearing than water-based paints, so you'd probably go for an oil-based gloss.
- To paint bedroom walls you'd want a paint that goes on easily, dries quickly, and doesn't produce harmful fumes. Water-based emulsion fits the bill here.

Paints are made of three parts

Paint is just solvent, binding medium, and pigment. Pigment gives paint its colour, the binding medium holds the pigment together and the solvent temporarily keeps the paint in liquid form. Some paints use water as a solvent and some use oil, but the idea is essentially the same.

Warm-Up and Exam Questions

The warm-up questions run quickly over the basic facts you'll need in the exam.
Unless you've learnt the facts first you'll find the exam questions pretty difficult.

Warm-Up Questions

1) Name two transition metals.
2) Give three useful physical properties of most metals.
3) What type of bonding do metals contain?
4) What is an alloy? Give two examples, with a use for each.
5) Which three components are found in most paints?
6) What is the difference between water-based and oil-based paints?

Exam Questions

1 Many paints are colloids.

(a) What is a colloid?

(1 mark)

(b) Why don't colloids separate out?

(1 mark)

(c) Read the following sentences about different paints and answer the questions that follow.

> Modern emulsion paints use water as a solvent. These paints produce very low levels of harmful fumes and are fast drying.
>
> Most gloss paints use organic compounds as solvents. Gloss paint should only be used in well ventilated areas because it can produce harmful fumes. It is harder-wearing than emulsion but takes longer to dry.

(i) Which type of paint would you use to paint the front door of a house — emulsion or gloss? Explain your answer

(1 mark)

(ii) Which type of paint would you use to paint the walls of a bedroom? Explain your answer

(1 mark)

2 Which of these statements best describes aluminium?

A A high density, corrosion resistant metal.

B One of the main components in steel.

C A very tough, completely unreactive, dense metal.

D A low density, versatile metal that can't be extracted by reduction with carbon.

(1 mark)

Exam Questions

3 Alloys are often used instead of pure metals because

A they are more plentiful.

B their properties make them more suitable for the application.

C their melting points are higher.

D they are completely inert.

(1 mark)

4 Nitinol is an example of a smart alloy.

(a) What property of nitinol makes it a 'smart' alloy?

(1 mark)

(b) Suggest a use for nitinol.

(1 mark)

(c) Give **two** disadvantages that smart alloys have compared to ordinary alloys.

(2 marks)

5 Match words **A**, **B**, **C** and **D** with the numbers **1 - 4** in the sentences below.

A titanium

B tin

C iron

D gold

Pure ...**1**... is a soft metal, so metals such as nickel are alloyed with it to make it harder.
...**2**... is used for making replacement hips as it has a low density and does not corrode.
...**3**... is added to copper to make bronze.
Carbon is added to ...**4**... to make steel.

(4 marks)

6 All metals have the same basic properties.

(a) Describe how the structure of a metal allows it to carry an electric current.

(1 mark)

(b) Give two other properties of metals.

(1 mark)

Revision Summary for Section Two

Okay, if you were just about to turn the page without doing these revision summary questions, then stop. What kind of an attitude is that... Is that really the way you want to live your life... running, playing and having fun... Of course not. That's right. Do the questions. It's for the best all round.

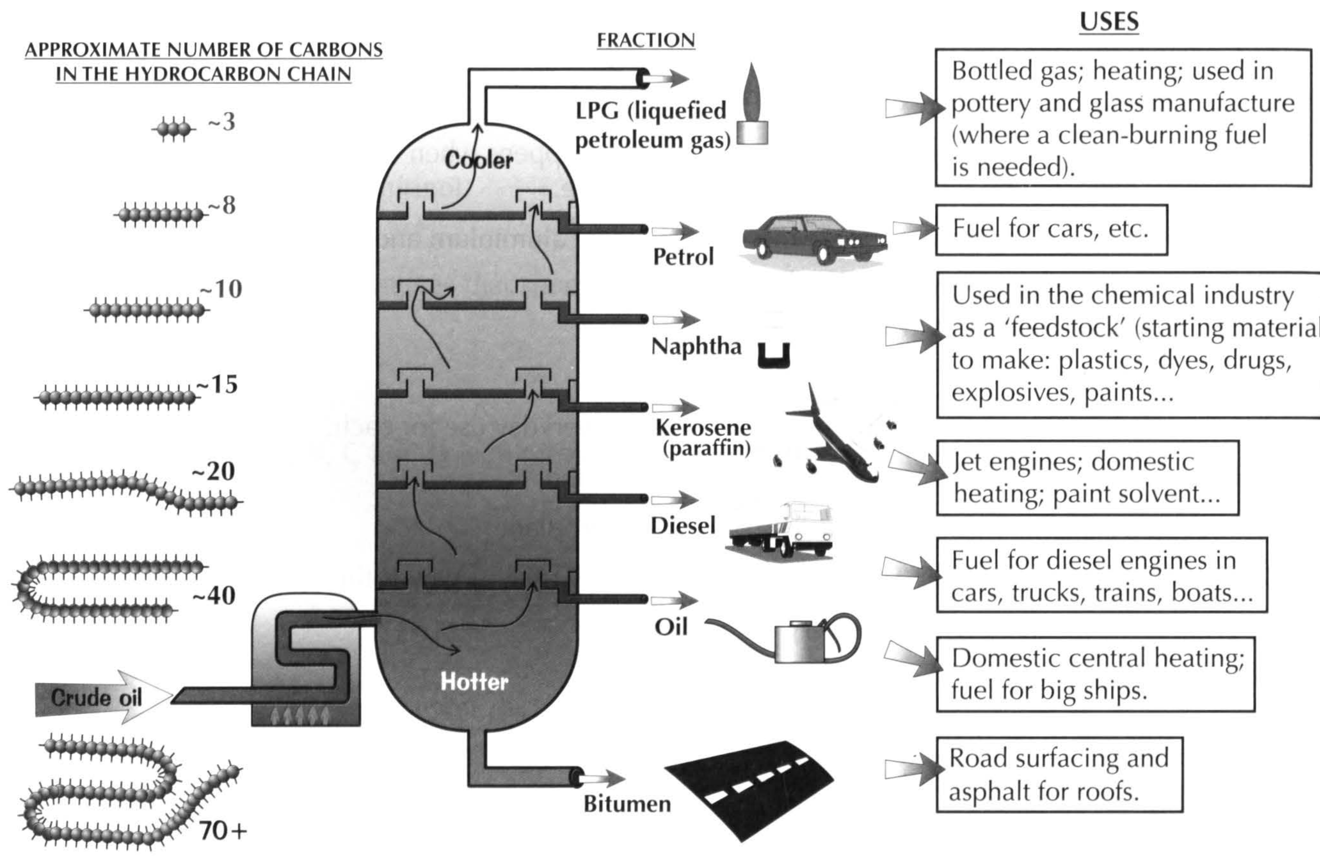
1) Explain how acid rain could pose a threat to buildings like St. Paul's Cathedral.
2) Give the word and symbol equations for the reaction that happens when calcium carbonate is heated.
3) How is glass made? Cement? Concrete?
4) List three environmental impacts of extracting rocks from the Earth.
5) What is the definition of an ore?
6) How would you extract the copper from its ore?
7) Give two reasons why it is important to recycle copper.
8) Explain why zinc can be extracted by reduction with carbon but magnesium can't.
9) Write down the word equation for the reaction that happens when the iron in the ore iron(III) oxide is extracted by reduction with carbon monoxide.
10) Explain why gold and copper were discovered before aluminium and magnesium.
11) What happens if you put:
    a) a piece of magnesium in a solution of zinc sulfate?
    b) a copper bracelet in a solution of iron chloride?
12) Name three different transition metals and give an everyday use for each.
13) Give three properties that most metals have in common.
14) Suggest a suitable metal to use for the body of an aeroplane.
15) Iron straight from the blast furnace doesn't have many uses — why is this?
16) What is the problem with using very pure iron?
17) Explain why alloys tend to be harder than pure metals.
18) Give two examples of alloys and say what's in them.
19) Give an example of a smart alloy.
20) What is the name for the substance that gives a paint its colour?
21) What other two components do paints usually have?
22) Suggest why a water-based paint might not be the best choice for painting an outside door.

Fractional Distillation of Crude Oil

Crude oil is formed from the buried remains of plants and animals — it's a fossil fuel. Over millions of years, with high temperature and pressure, the remains turn to crude oil, which can be drilled up.

Crude oil can be *split* into *separate hydrocarbons*

1) Crude oil is a mixture of hydrocarbons — molecules which are made of just carbon and hydrogen.
2) Fractional distillation splits crude oil into fractions (groups of compounds with carbon chains of similar length).
3) Heated crude oil is piped in at the bottom of a fractionating column. The various fractions are constantly tapped off at the different levels where they condense.



Fractional distillation is an example of a physical process — there are no chemical reactions.

The *properties* of *hydrocarbon molecules* depend on their *size*

The big hydrocarbon molecules are the first to condense, because they have higher boiling points. As the molecules get smaller, they condense higher up the fractionating column. The smaller the molecule...

1) The lower the boiling point — the substance stays as a gas at lower temperatures.
2) The more flammable it is — it sets fire more easily.
3) The less viscous it is — it's less 'gloopy' and flows more easily.
4) The more volatile it is — it evaporates more readily.

The vapours of the more volatile hydrocarbons are very flammable and pose a serious fire risk. So don't smoke at the petrol station. (In fact, don't smoke at all, it's stupid.)

Burning Hydrocarbons

A fuel is a substance that reacts with oxygen to release useful energy. Remember that.

Complete combustion happens when there's plenty of oxygen

The complete combustion of any hydrocarbon in oxygen will produce only carbon dioxide and water as waste products, which are both quite clean and non-poisonous.

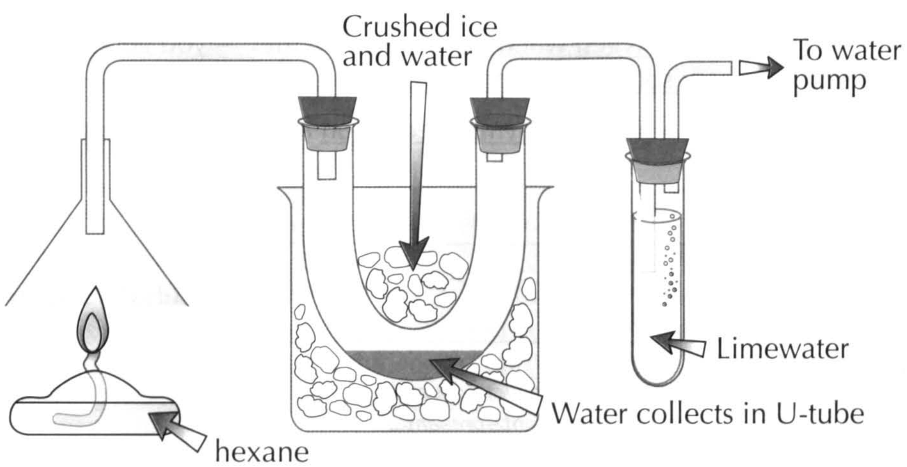
hydrocarbon + oxygen → carbon dioxide + water (+ energy)

1) Many gas heaters release these waste gases into the room, which is perfectly OK. As long as the gas heater is working properly and the room is well ventilated there's no problem.
2) This reaction, when there's plenty of oxygen, is known as complete combustion. It releases lots of energy and only produces those two harmless waste products (lots of CO_2 isn't ideal, but the alternatives are worse — see next page). When there's plenty of oxygen and combustion is complete, the gas burns with a clean blue flame.
3) 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:

$$CH_4 + 2O_2 \rightarrow 2H_2O + CO_2 \quad \text{(+ energy)}$$

You've just got to make sure you end up with the same number of Cs, Hs and Os on either side of the arrow.

You can show a fuel burns to give CO_2 and H_2O...



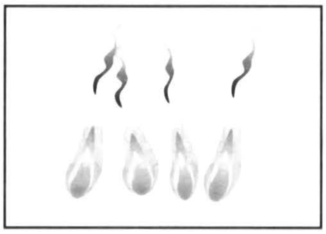
- The water pump draws gases from the burning hexane through the apparatus.
- Water collects inside the cooled U-tube and you can show that it's water by checking its boiling point.
- The limewater turns milky, showing that carbon dioxide is present.

Burning Hydrocarbons

Incomplete combustion also releases useful energy from fuel — but a lot less energy than is released from complete combustion.

Incomplete combustion of hydrocarbons is NOT safe

1) If there isn't enough oxygen the combustion will be incomplete.
   This gives carbon monoxide and carbon as waste products too, and produces a smoky yellow flame.



   Incomplete combustion produces less energy than complete combustion does.

hydrocarbon + oxygen →
carbon dioxide + water + carbon monoxide + carbon (+ energy)

2) The carbon monoxide is a colourless, odourless and poisonous gas and it's very dangerous (see p.79). Every year people are killed while they sleep due to faulty gas fires and boilers filling the room with deadly carbon monoxide (CO) and nobody realising — so it's important to regularly service gas appliances. The black carbon given off produces sooty marks — a clue the fuel's not burning fully.
3) So basically, you want lots of oxygen when you're burning fuel — you get more energy given out, and you don't get any messy soot or poisonous gases.
4) You need to be able to write a balanced symbol equation for incomplete combustion too, e.g.

$$4CH_4 + 6O_2 \rightarrow C + 2CO + CO_2 + 8H_2O \quad \text{(+ energy)}$$

This is just one possibility. The products depend on the exact quantity of the reactants present...

...E.g. you could also have: $2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$ — the important thing is that the equation is balanced.

Blue flame good, yellow flame bad

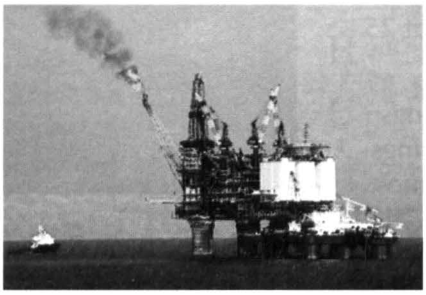
This is why people should get their gas appliances serviced every year, and get carbon monoxide detectors fitted. Carbon monoxide really can kill people in their sleep — scary stuff. Don't let that scare you off learning everything that's on this page — any of it could come up in the exam.

Using Crude Oil as a Fuel

Nothing as amazingly useful as crude oil would be without its problems.
No, that'd be too good to be true.

Crude oil provides an important **fuel** for modern life

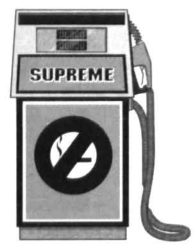
1) Crude oil fractions burn cleanly so they make good fuels. Most modern transport is fuelled by a crude oil fraction, e.g. cars, boats, trains and planes. Parts of crude oil are also burned in central heating systems in homes and in power stations to generate electricity.



2) There's a massive industry with scientists working to find oil reserves, take it out of the ground, and turn it into useful products. As well as fuels, crude oil also provides the raw materials for making various chemicals, including plastics. There's more on this on page 42.

3) Often, alternatives to using crude oil fractions as fuel are possible. E.g. electricity can be generated by nuclear power or wind power, solar energy can be used to heat water, and there are hydrogen-powered cars (see page 83).

4) But things tend to be set up for using oil fractions. For example, cars are designed for petrol or diesel and it's readily available. There are filling stations all over the country, with storage facilities and pumps specifically designed for these crude oil fractions. So crude oil fractions are often the easiest and cheapest thing to use.



5) Crude oil fractions are often more reliable too — e.g. solar and wind power won't work without the right weather conditions. Nuclear energy is reliable, but there are lots of concerns about its safety and the storage of radioactive waste.

Using Crude Oil as a Fuel

Crude oil is really useful fuel that we use every day — there is a possibility that it might run out.

Crude oil might **run out** one day... eeek

1) Most scientists think that oil will run out. But no one knows exactly when.
2) There have been heaps of different predictions — e.g. about 40 years ago, scientists predicted that it'd all be gone by the year 2000.
3) New oil reserves are discovered from time to time — e.g. a major new oil field was found in southern Oman in the Middle East in 2002. No one knows how much oil will be discovered in the future though.
4) Also, technology is constantly improving, so it's now possible to extract oil that was once too difficult or expensive to extract. It's likely that technology will improve further — but who knows how much?
5) In the worst-case scenario, oil may be pretty much gone in about 25 years — and that's not far off.
6) Some people think we should immediately stop using oil for things like transport, for which there are alternatives, and keep it for things that it's absolutely essential for, like some chemicals and medicines.

7) It will take time to develop alternative fuels that will satisfy all our energy needs (see page 83 for more info). It'll also take time to adapt things so that the fuels can be used on a wide scale. E.g. we might need different kinds of car engines, or special storage tanks built.
8) So however long oil does last for, it's a good idea to start conserving it and finding alternatives now.

Crude oil is **not** the **environment's** best friend

1) Oil spills can happen as the oil is being transported by tanker — this spells disaster for the local environment. Birds get covered in the stuff and are poisoned as they try to clean themselves. Other creatures, like sea otters and whales, are poisoned too.

2) You have to burn oil to release the energy from it. But burning oil is thought to be a major cause of global warming (p.75), acid rain (p.78) and global dimming (p.77).

If oil alternatives aren't developed, we might get caught short

Crude oil is really important to our lives. Take petrol for instance — at the first whisper of a shortage, there's mayhem. Loads of people dash to the petrol station and start filling up their tanks. This causes a queue, which starts everyone else panicking. I don't know what they'll do when it runs out totally.

Alkanes

Crude oil contains lots of alkanes and some alkenes (see the next page).
They have different properties, and it's all down to their structure.

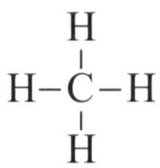
Alkanes have all C–C single bonds

1) Alkanes are made up of chains of carbon atoms joined by single covalent bonds, and surrounded by hydrogen atoms.

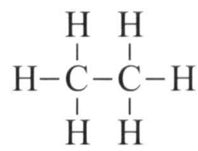
Covalent bonds are when atoms share electrons. Carbon atoms like to make 4 bonds altogether. Hydrogen atoms like to make 1.

2) Different alkanes have chains of different lengths.
The first four alkanes are methane (natural gas), ethane, propane and butane.

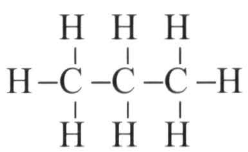
1) **Methane**: CH_4
(natural gas)



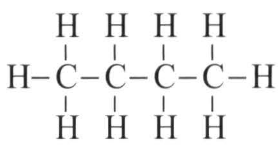
2) **Ethane**: C_2H_6



3) **Propane**: C_3H_8



4) **Butane**: C_4H_{10}



3) All alkanes have the formula: C_nH_{2n+2}

4) They're called saturated hydrocarbons because they have no spare bonds left (i.e. no double bonds that can open up and have things join onto them — see the next page).

5) You can tell the difference between an alkane and an alkene by adding the substance to bromine water.
An alkane won't decolourise the bromine water.
This is because it has no spare bonds, so it can't react with the bromine.

6) Alkanes won't form polymers — again, no spare bonds.

7) They burn cleanly, producing carbon dioxide and water.

Alkenes

Alkenes have very different properties to alkanes. Due to their double bonds they can do lots of clever things — like forming polymers.

Alkenes have a C=C double bond

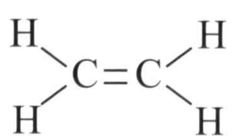
1) Alkenes have chains of carbon atoms with one or more double covalent bonds.
2) They're called unsaturated hydrocarbons because double bonds can open up and let things join on.
3) This is why they will decolourise bromine water. They form bonds with the bromine.

4) Alkenes are more reactive — due to the double bond all poised and ready to just pop open. They can form polymers by opening up their double bonds to 'hold hands' in a long chain. (See page 42 for more info on polymers.)

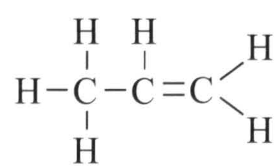
A double bond means that atoms are sharing two pairs of electrons. A double bond counts as two of a carbon atom's four bonds.

5) The first three alkenes are ethene, propene and butene...

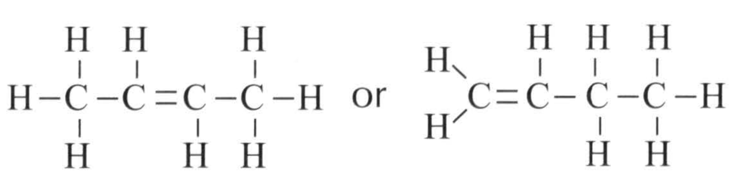
1) Ethene: C_2H_4



2) Propene: C_3H_6



3) Butene: C_4H_8



There are two different forms of butene — the double bond can be in different places.

6) All alkenes containing one double bond have the formula: C_nH_{2n}
7) They tend to burn with a smoky flame, producing soot (carbon).

Alkane anybody who doesn't learn this lot properly

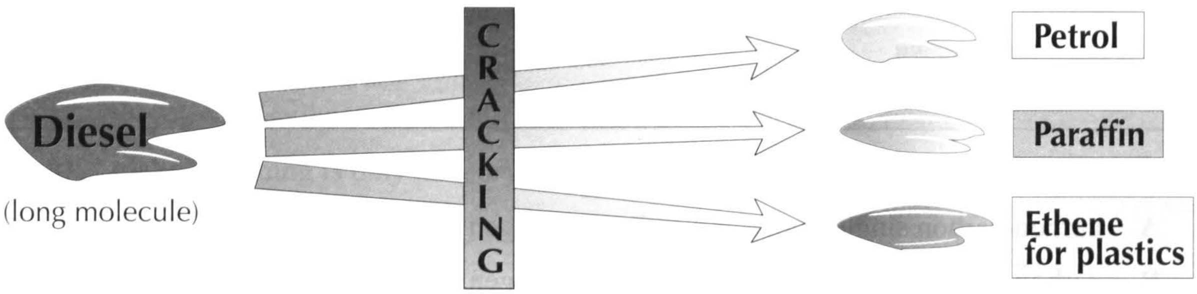
Don't get alkenes confused with alkanes — that one letter makes all the difference. Alkenes have a C=C bond, alkanes don't. The first part of their names is the same though. "Meth-" means "one carbon atom", "eth-" means "two C atoms", "prop-" means "three C atoms", "but-" means "four C atoms", etc.

Cracking Crude Oil

After the distillation of crude oil, you've still got both short and long hydrocarbons, just not all mixed together. But there's more demand for some products, like petrol, than for others.

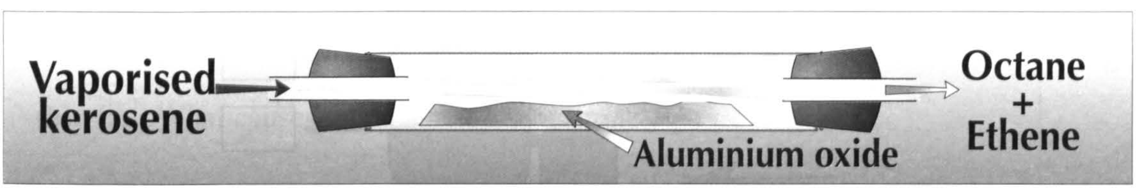
Cracking means splitting up long-chain hydrocarbons...

1) Long-chain hydrocarbons form thick gloopy liquids like tar which aren't all that useful, so...
2) ... a lot of the longer molecules produced from fractional distillation are turned into smaller ones by a process called cracking.
3) Some of the products of cracking are useful as fuels, e.g. petrol for cars and paraffin for jet fuel.
4) Cracking also produces short alkenes like ethene, which are needed for making plastics (see p.42).



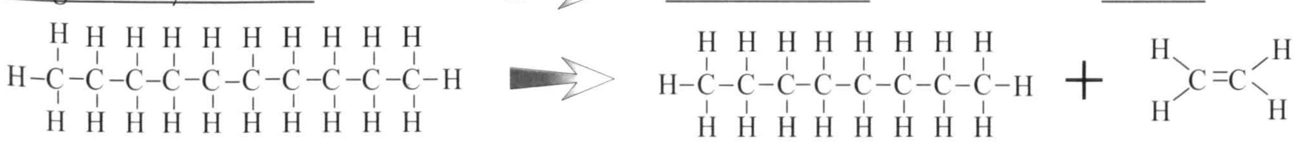
...by passing vapour over a hot catalyst

1) Cracking is a thermal decomposition reaction — breaking molecules down by heating them.
2) The first step is to heat the long-chain hydrocarbon to vaporise it (turn it into a gas).
3) Then the vapour is passed over a powdered catalyst at a temperature of about 400 °C – 700 °C.
4) Aluminium oxide is the catalyst used.
5) The long-chain molecules split apart or "crack" on the surface of the specks of catalyst.



6) Most of the products of cracking are alkanes and alkenes (see pages 37 and 38).

Long-chain hydrocarbon molecule → Shorter ALKANE molecule + ALKENE



E.g. Kerosene (ten C atoms) → Octane (eight C atoms) + ethene

(Too much of this in crude oil) (useful for petrol) (for making plastics)

Warm-Up and Exam Questions

You must be getting used to the routine by now — the warm-up questions run over the basic facts, then the exam questions show you the kind of thing you'll get on the day.

Warm-up Questions

1) What are hydrocarbons?
2) Name three fractions obtained from crude oil.
3) Why are alkenes described as unsaturated hydrocarbons?
4) What sort of hydrocarbon molecules are cracked, and why are they cracked?
5) Describe the conditions used for cracking hydrocarbons.
6) List three modern-day activities that depend on crude oil or its fractions.

Exam Questions

1 The bonds in alkanes are best described as

A carbon-carbon single bonds and carbon-hydrogen single bonds

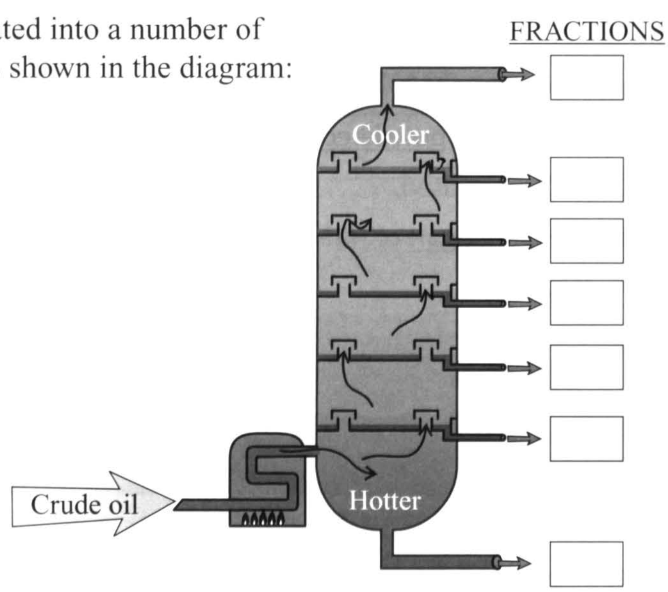
B carbon-carbon single bonds and carbon-hydrogen double bonds

C carbon-carbon double bonds and carbon-hydrogen single bonds

D carbon-carbon double bonds and carbon-hydrogen double bonds

(1 mark)

2 Crude oil can be separated into a number of different compounds as shown in the diagram:



(a) (i) Put an **M** in the box of the fraction with the largest hydrocarbon molecules.

(1 mark)

(ii) Put a **B** in the box of the fraction with the lowest boiling point.

(1 mark)

(b) Give **one** use for the kerosene fraction.

(1 mark)

(c) Briefly explain how the separation process works.

(3 marks)

Exam Questions

3 A test for an alkene is

A universal indicator is decolourised

B starch solution goes dark blue

C bromine water is decolourised

D bromine water remains brown

(1 mark)

4 Even though there are many environmental problems caused by using crude oil fractions, we continue to use them mainly because

A they are a renewable resource

B technology is always improving

C they are a readily available and concentrated energy source

D global warming is only a theory

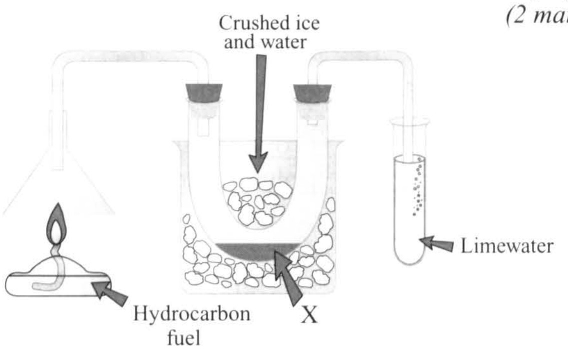
(1 mark)

5 (a) When a hydrocarbon fuel burns completely, it produces CO_2 and water.

(i) Write a balanced symbol equation for the complete combustion of ethane, C_2H_6.

(2 marks)

(ii) The apparatus shown can be used to identify the two products of complete combustion of a hydrocarbon, if a further test is carried out on the liquid X which collects in the U-tube.



Describe this test, giving the result you would expect, and explain how the other product is identified.

(2 marks)

(b) (i) What causes incomplete combustion?

(1 mark)

(ii) Name **two** products of incomplete combustion that are **not** produced in complete combustion.

(2 marks)

6 A renewable energy source for electricity production is

A sunlight

B LPG (liquefied petroleum gas)

C naphtha

D nuclear fuel

(1 mark)

Making Polymers

Plastics are made up of lots of molecules joined together. They're like long chains.

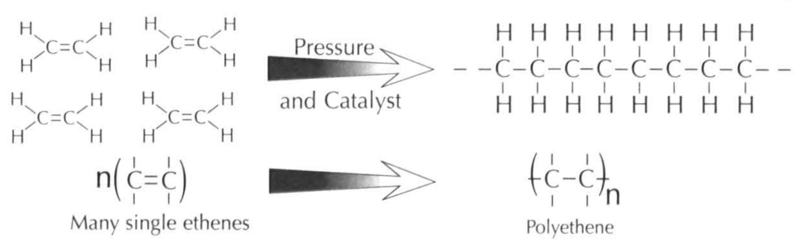
Plastics are long-chain molecules called polymers

1) Plastics are formed when lots of small molecules called monomers join together to give a polymer.
2) They're usually carbon based (and the monomers are very often alkenes — see page 38).

Addition polymers are made under high pressure

1) The monomers that make up addition polymers have a double covalent bond (i.e. they're unsaturated).
2) Under high pressure and with a catalyst (see pages 154-155) to help them along, many unsaturated small molecules open up those double bonds and "join hands" (polymerise) to form long saturated chains called polymers.

Ethene becoming polyethene or "polythene", is the easiest example:

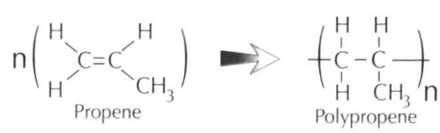


You'll need to be able to draw the formula of an addition polymer, given the formula of its monomer. Dead easy — the carbons just all join together in a row with no double bonds between them.

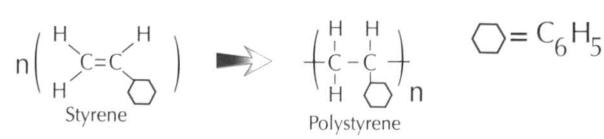
The 'n' just means there can be any number of monomers.

The name of the plastic comes from the type of monomer it's made from — you just stick the word "poly" in front of it:

Propene can form polypropene:



A molecule called styrene will polymerise into polystyrene:

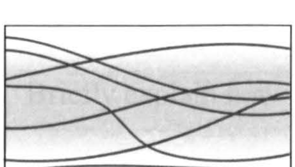


Forces between molecules determine the properties of plastics

Strong covalent bonds hold the atoms together in long chains. But it's the bonds between the different molecule chains that determine the properties of the plastic.

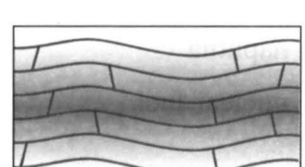
Weak Forces

Long chains held together by weak forces are free to slide over each other. This means the plastic can be stretched easily, and will have a low melting point.



Strong Forces

Plastics with stronger bonds between the polymer chains have higher melting points and can't be stretched, as the crosslinks hold the chains firmly together.



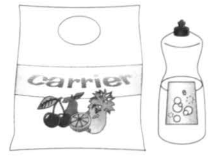
Uses of Polymers

Plastics are fantastically useful. You can make novelty football pencil sharpeners and all sorts.

Polymers' **properties** decide what they're **used for**

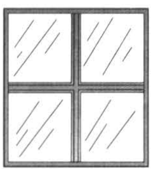
Different polymers have different physical properties — some are stronger, some are stretchier, some are more easily moulded, and so on. These different physical properties make them suited for different uses.

- Strong, rigid polymers such as high density polyethene are used to make plastic milk bottles.



- Light, stretchable polymers such as low density polyethene are used for plastic bags and squeezy bottles. Low density polyethene has a low melting point, so it's no good for anything that'll get very hot.

- PVC is strong and durable, and it can be made either rigid or stretchy. The rigid kind is used to make window frames and piping. The stretchy kind is used to make synthetic leather.



- Polystyrene foam is used in packaging to protect breakable things, and it's used to make disposable coffee cups (the trapped air in the foam makes it a brilliant thermal insulator).

- Heat-resistant polymers such as melamine resin and polypropene are used to make plastic kettles.

Non-biodegradable plastics cause **disposal problems**

1) Most polymers aren't "biodegradable" — they're not broken down by microorganisms, so they don't rot. This property is actually kind of useful until it's time to get rid of your plastic.

2) It's difficult to get rid of plastics — if you bury them in a landfill site, they'll still be there years later. Landfill sites fill up quickly, and they're a waste of land. And a waste of plastic.

3) When plastics are burnt, some of them release gases such as acidic sulfur dioxide and poisonous hydrogen chloride and hydrogen cyanide. So burning's out, really. Plus it's a waste of plastic.

4) The best thing is to reuse plastics as many times as possible and then recycle them if you can. Sorting out lots of different plastics for recycling is difficult and expensive, though.

5) Chemists are working on a variety of ideas to produce biodegradable polymers.

Uses of Polymers

Polymers are often used to *make clothes*

1) Nylon is a synthetic polymer often used to make clothes. Fabrics made from nylon are not waterproof on their own, but can be coated with polyurethane to make tough, hard-wearing and waterproof outdoor clothing.

2) One big problem is that the polyurethane coating doesn't let water vapour pass through it. So if you get a bit hot (or do a bit of exercise), sweat condenses on the inside. This makes skin and clothes get wet and uncomfortable — the material isn't breathable.

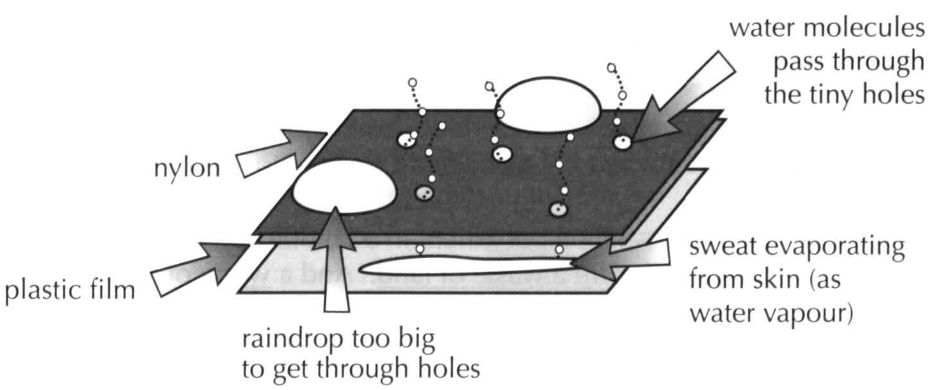
3) Breathable fabrics have all the useful properties of nylon/polyurethane ones, but they also let sweat out. If you sweat in a breathable material, water vapour can escape — so no condensation.

Polymers are used to make *breathable materials*

1) Some breathable fabrics are made by combining a thin film of a plastic with a layer of another fabric, such as nylon.

2) The plastic film has tiny holes which let water vapour pass through — so it's breathable. But it's waterproof, since the holes aren't big enough to let big water droplets through and the plastic repels liquid water.

3) This material is great for outdoorsy types — they can hike without getting rained on or soaked in sweat.



Polymers have a wide range of uses

If you're making a product, you need to pick your plastic carefully. It's no good trying to make a jacket out of a brittle, unbendy polymer — imagine trying to walk any distance in a jacket like that. The same goes for things like kettles — there's no point using a plastic that melts at 50 °C.

Warm-Up and Exam Questions

Warm-up Questions

1) Explain what monomers and polymers are.
2) How can a plastic's melting point tell you about the bonds between its polymer chains?
3) In a polymer molecule, what type of bonds hold the atoms together in a chain?
4) Melamine is a heat-resistant polymer. Suggest a use for melamine.

Exam Questions

1 Complete the passage using some of the following words.

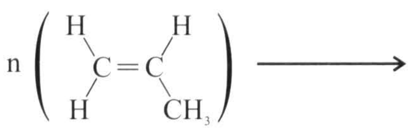
explosive **biodegradable** **hidden** **recycle**

expensive **toxic** **flammable** **burnt**

The majority of plastics are hard to dispose of because they are not .... They can be ...., but this often releases .... gases like hydrogen cyanide. Burying plastics also has its disadvantages. This is why it is best to .... plastics whenever possible, however this can be an .... process.

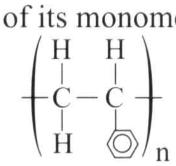
(5 marks)

2 (a) Complete this equation for the formation of polypropene.



(1 mark)

(b) The structural formula of polystyrene is shown below.
Draw the structural formula of its monomer.



(1 mark)

(c) PVC (polyvinyl chloride) is strong and durable and can be made in rigid or stretchy forms. Suggest a use for:

(i) the rigid form.

(1 mark)

(ii) the stretchy form.

(1 mark)

Revision Summary for Section Three

Now, my spies tell me that some naughty people skip these pages without so much as reading through the list of questions. Well, you shouldn't, because what's the point in reading that great big section if you're not going to check if you really know it or not? Look, just read the first 10 questions, and I guarantee there'll be an answer you'll have to look up. And when it comes up in the exam, you'll be so glad you did. Plus, if you don't do as you're told my spies will tell me, and then you won't get any toys. Or something.

1) Explain briefly the principle of fractional distillation.
2) Sketch the full diagram of the fractional distillation of crude oil, including the names of all seven main fractions.
3) Does fractional distillation involve a physical change, a chemical change or both?
4) Describe how four properties of hydrocarbons vary with the molecule size.
5) Give the general word equations for the complete and incomplete combustion of hydrocarbons.
6) Describe an experiment you could do to show which products are formed during complete combustion. Include all the equipment you would need.
7) Explain how incomplete combustion can be harmful to humans.
8)* Write down a balanced symbol equation for the incomplete combustion of ethane (C_2H_6).
9) Give three alternatives to using crude oil as a fuel. Give one advantage and one disadvantage of each of your alternatives, compared with crude oil.
10) Why do predictions about when crude oil will run out change over the years?
11) Describe how oil slicks can affect wildlife.
12) How else can using crude oil as a fuel have a negative effect on the environment?
13) What's the general formula for an alkane? What's the formula for a 5-carbon alkane?
14) What kind of carbon-carbon bond do all alkenes have?
    Give the general formula for an alkene containing one double bond.
15) Draw out the chemical structure of ethene.
16) Describe a test you can do to tell whether a hydrocarbon is an alkane or an alkene.
17) Give a typical example of a substance that might be cracked.
    Name the products you would get from cracking it.
18) What type of reaction is cracking?
19) Give the formula for the ethene monomer used to make polyethene, and the formula for the polyethene polymer.
20) Propene is an alkene containing three carbon atoms.
    What would a polymer made from propene be called?
21) Plastic bags stretch and melt easily. Are the forces between the polymer chains weak or strong?
22) Name one polymer that would be suitable for:
    a) plastic bags,
    b) window frames,
    c) kettles.
23) Explain how fabrics can be made both breathable and waterproof.
24) Give one disadvantage of burning plastics.
25) Explain why burying plastics is not a good way to dispose of them either.

*Answers on page 254.

Chemicals and Food

Cooking is just chemistry by another name — chemistry involving pies.

Some *foods* have to be *cooked*

There are loads of different ways to cook food — e.g. boiling, steaming, grilling, frying...

1) Many foods have a better taste and texture when cooked.
2) Some foods are easier to digest once they're cooked (e.g. potatoes, flour). See below for why.
3) The high temperatures involved in cooking also kill off those nasty little microbes that cause disease — this is very important with meat.
4) Some foods are poisonous when raw, and must be cooked to make them edible — e.g. red kidney beans contain a poison that's only destroyed by at least 10 minutes boiling (and 2 hours cooking in total).

Cooking causes *chemical changes*

Cooking food produces new substances. That means a chemical change has taken place. Once cooked, you can't change it back. The cooking process is irreversible.

e.g. Eggs and meat

1) Eggs and meat are good sources of protein.
2) Protein molecules change shape when you heat them. The energy from cooking breaks some of the chemical bonds in the protein, and this allows the molecule to take a different shape.
3) This gives the food a more edible texture.
4) The change is called denaturing — it's irreversible.

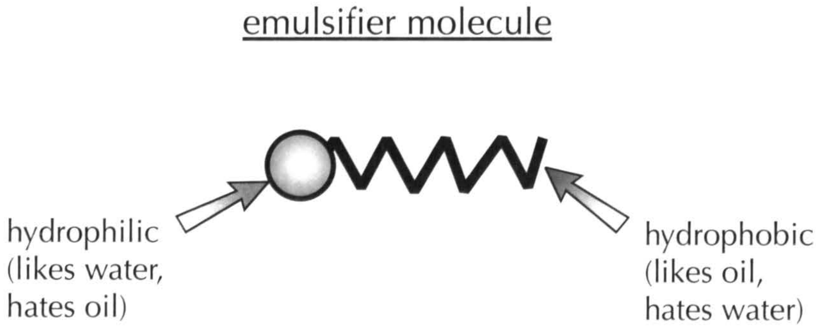
e.g. Potatoes

1) Potatoes are a good source of carbohydrates.
2) Potatoes are plants, so each potato cell is surrounded by a cellulose cell wall.
3) Humans can't digest cellulose, so this makes it difficult to get to the contents of the cells.
4) Cooking the potato breaks down the cell wall, making it a lot easier to digest.

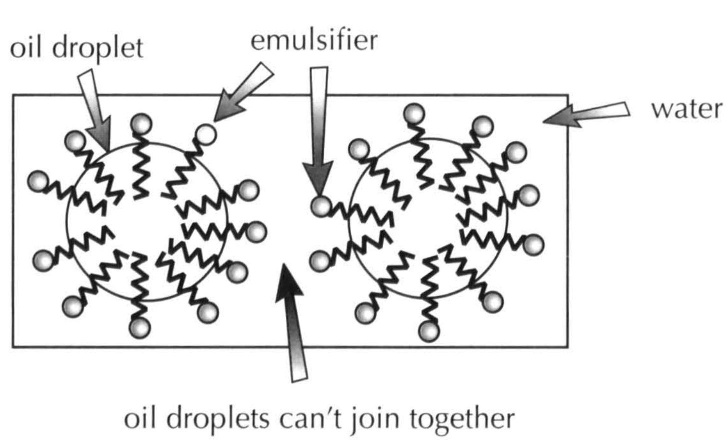
Chemicals and Food

Emulsifiers Help Oil and Water Mix

1) You can mix an oil with water to make an emulsion. Emulsions are made up of lots of droplets of one liquid suspended in another liquid.
2) Oil and water naturally separate into two layers with the oil floating on top of the water — they don't "want" to mix. Emulsifiers help to stop the two liquids in an emulsion from separating out.
3) Mayonnaise, low-fat spread and ice cream are foods which contain emulsifiers.
4) Emulsifiers are molecules with one part that's attracted to water and another part that's attracted to oil or fat. The bit that's attracted to water is called hydrophilic, and the bit that's attracted to oil is called hydrophobic.



5) The hydrophilic end of each emulsifier molecule latches onto water molecules.
6) The hydrophobic end of each emulsifier molecule cosies up to oil molecules.
7) When you shake oil and water together with a bit of emulsifier, the oil forms droplets, surrounded by a coating of emulsifier... with the hydrophilic bit facing outwards. Other oil droplets are repelled by the hydrophilic bit of the emulsifier, while water molecules latch on. So the emulsion won't separate out. Clever.



Cooking food makes it easier to digest and safer

When you cook something, you're bringing about chemical change. The changes are irreversible, as you'll know if you've ever tried to unscramble an egg.

Packaging and Smart Materials

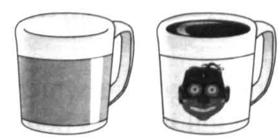
New materials are continually being developed. Their properties determine what they can be used for.

Smart materials can change their properties

Here are some examples of smart materials that behave differently depending on the conditions.

1) Nitinol is a "shape memory alloy". See page 26 for more information.

2) Some dyes contain thermochromic pigments that change colour or become transparent depending on the temperature. They're used on novelty mugs with designs that change when a hot drink is poured in. Similar materials can be used in food packaging.



3) There are also dyes that become more or less transparent depending on the light intensity. They're used in sunglasses that get darker in more intense sunlight.

4) Some materials expand or contract when you put an electric current through them. They also do the opposite — they produce electricity when they're squeezed. These materials are used in car airbag sensors. When squeezed by the forces of a car crash they produce electricity to activate the airbag.

5) There are liquids that turn solid when you put them in a magnetic field. They're used to control vibrations, e.g. in some car shock absorbers.

Active and intelligent packaging does more than wrap food

The food industry uses new technology to create "active packaging" to help food last for longer.

1) Active packaging and intelligent packaging don't only form a barrier between the food and the outside world. They can control, and even react to what's happening inside the package.
2) Active packaging is packaging which changes something inside the package.
3) Intelligent packaging can monitor the food and tell the customer whether the food is still okay or if it's gone off.

Smart materials react to outside influences

Smart materials are intelligent because they do something useful all on their own, whether it's expand, contract, change from liquid to solid, change colour, return to a previous shape or something else.

Packaging and Smart Materials

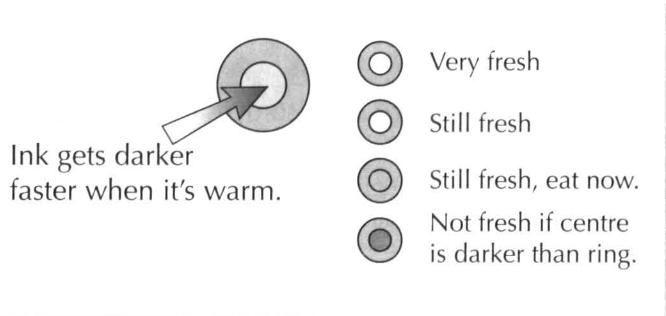
Here are lots of examples of active and intelligent packaging for you to learn.

Widgets release gas in beer cans

"Widgets" in beer cans release gas when the can's opened, to give the beer a foamy "head".

Thermochromic pigments are used to monitor food temperature

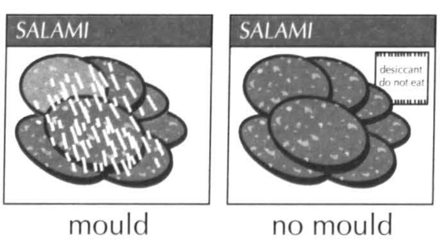
Thermochromic pigments which change colour faster the warmer they get can be used to tell if food's been warm for long enough for microbes to grow.



Cans can self-heat and self-cool

1) Self-heating cans use an exothermic reaction (see p.157) to produce heat and warm up the contents.
2) Self-cooling cans use evaporation to cool down their contents.
3) Some of these cans also use thermochromic pigments which change colour when the contents are at the correct temperature.

Silica gel removes water from inside a packet



1) Fresh food sometimes produces moisture inside the packaging. This allows bacteria or mould to grow.
2) Chemicals such as silica gel are desiccants and absorb water. Manufacturers put sachets of silica into some food packets to absorb excess water, which makes it harder for bacteria and mould to grow.

Intelligent packaging — make sure you understand what it does

Don't panic if you don't 'get' how stuff like the self-heating and self-cooling cans work — in the exam you'll be given the data you need, and you'll just have to interpret it.

Food Additives

Humans have been adding stuff to food for years. Before fridges were invented, we added salt to meat to stop it going off. Now we use additives not just to preserve food, but to make it look or taste different.

Processed foods often contain additives

1) Food manufacturers add various chemical compounds to food to improve its appearance, taste, texture and shelf life. These additives must be listed in the ingredients list on the back of the packet.
2) Most additives used in the UK have E-numbers — e.g. E127 is erythrosine (a red dye) and E201 is sodium sorbate (a preservative). Additives with E-numbers have passed safety tests and can be used in Europe.

- Preservatives help food stay fresh. Without them, more food would go off and need throwing away.

- Some foods 'go off' after reacting with oxygen — e.g. butter goes rancid. Antioxidants are added to foods that contain fat or oil to stop them reacting with oxygen.

- Colourings and flavourings make food look and taste better.

- Emulsifiers (see page 48) and stabilisers stop emulsions like mayonnaise from separating out.

- Sweeteners can replace sugar in some processed foods — helpful to diabetics and dieters.

There are natural and synthetic additives

1) Some food additives are of natural origin, e.g. lecithin from soya beans. Some synthetic additives are identical to natural substances. Others are completely new synthetic substances.
2) Some people think that some synthetic food colourings (e.g. sunset yellow) make children hyperactive. But many scientific studies haven't found any connection between additives and hyperactivity at all.
3) A small number of people are allergic to some additives, for example the food dye tartrazine.
4) Some additives aren't suitable for vegetarians. For example, the food colouring cochineal comes from crushed insects. And gelatin from animal bones is used to thicken and set some foods.

Food Additives

Artificial colours can be detected by chromatography

To identify different colourings in a food sample, you can use chromatography.

> Paper chromatography uses the fact that different dyes wash through wet filter paper at different rates.

Here's how you'd analyse food colourings...

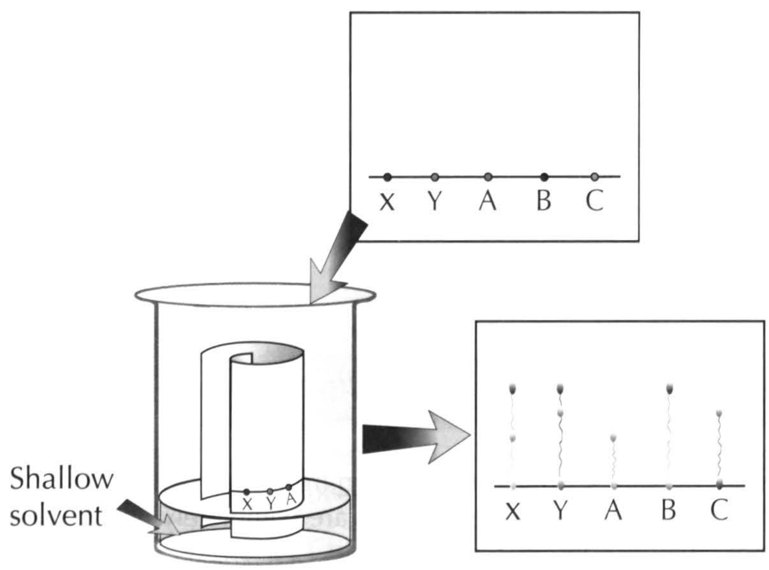
1) Extract the colour from each food sample by placing it in a small cup with a few drops of solvent (can be water, ethanol, salt water etc). Use a different cup for each different food sample.

2) Put spots of each coloured solution on a pencil baseline on filter paper. (Label them in pencil — don't use pen because it might dissolve in the solvent and confuse everything.)

3) Roll up the sheet and put it in a beaker with some solvent — but keep the baseline above the level of the solvent.

4) The solvent seeps up the paper, taking the food dyes with it. Different dyes form spots in different places.

5) Watch out though — a chromatogram with four spots means at least four dyes, not exactly four dyes. There could be five dyes, with two of them making a spot in the same place. (It can't be three dyes though, because one dye can't split into two spots.)



Learn all the different additives and why they're added to food

Chromatography can separate even complex mixtures if you choose the right equipment and conditions. On a different note, there's a lot about food additives in the media. Some statements are based on facts, others on rumour and prejudice. But without evidence to support a claim, it's not worth a bean.

Warm-Up and Exam Questions

I know that you'll be champing at the bit to get into the exam questions, but these warm-up questions are invaluable for getting the basic facts straight first.

Warm-Up Questions

1) How can cooking food make it safer for humans to eat?
2) What effect does cooking eggs have on the proteins in them?
3) Give an example of intelligent packaging.
4) What are antioxidants?

Exam Questions

1 Cooking potatoes breaks down

A cellulite

B protein

C cellulose

D celluloid

(1 mark)

2 Food colourings can be identified using chromatography. Outline a method you could use to identify the different colourings in a food sample.

(3 marks)

3 Mayonnaise contains emulsifiers — molecules that have a hydrophilic end and a hydrophobic end.

(a) On this diagram of an emulsifier molecule, label the hydrophilic and hydrophobic ends.

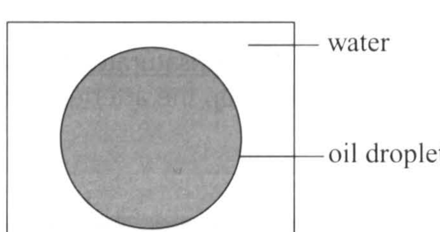
(1 mark)

(b) Explain the meanings of hydrophilic and hydrophobic.

(2 marks)

(c) The diagram below shows an oil molecule in water.
Show on the diagram how emulsifier molecules arrange themselves.

(1 mark)



(d) What effect do emulsifiers have on the mayonnaise?

(1 mark)

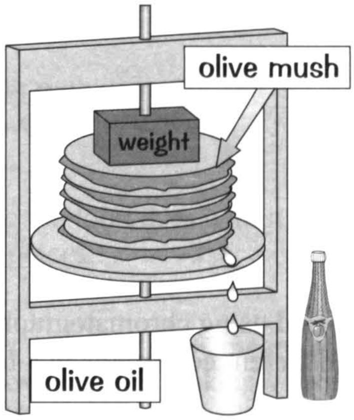
Plant Oils in Food

Plant oils come from plants. I know it's tricky, but just do your best to remember.

We can **extract oils** from **plants**

1) Some fruits and seeds contain a lot of oil. For example, avocados and olives are oily fruits. Brazil nuts, peanuts and sesame seeds are oily seeds (a nut is just a big seed really).

2) These oils can be extracted and used for food or for fuel.
3) To get the oil out, the plant material is crushed. The next step is to press the crushed plant material between metal plates and squash the oil out. This is the traditional method of producing olive oil.



4) Oil can be separated from crushed plant material by a centrifuge — rather like using a spin-dryer to get water out of wet clothes. Or solvents can be used to get the oil from the plant material.
5) Distillation is used to refine the oil, and it also removes water, solvents and impurities.

Vegetable oils are used in **food**

1) Vegetable oils provide a lot of energy.
2) There are other nutrients in vegetable oils too — for example, oils from seeds contain vitamin E.
3) Vegetable oils contain essential fatty acids, which the body needs for many metabolic processes.
4) Vegetable oils tend to be unsaturated, while animal fats tend to be saturated.
5) In general, saturated fats are less healthy than unsaturated fats (as saturated fats increase the amount of cholesterol in the blood, which can block up the arteries and increase the risk of heart disease).

Plant oils can be good for you

Before fancy stuff from abroad like olive oil, we fried our bacon and eggs in lard. Vegetable oils, like olive oil, are better for use than animal fats, like lard, because they're usually unsaturated. Too much lard in your diet could increase your risk of getting heart disease.

Plant Oils in Food

Oils and fats can be **saturated** or **unsaturated**

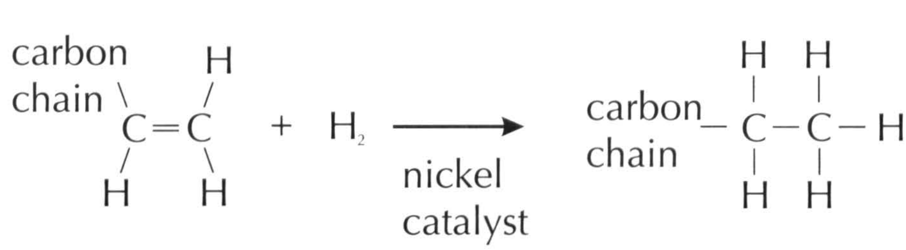
1) Oils and fats contain long-chain molecules with lots of carbon atoms.
2) They can be either saturated or unsaturated.

- Unsaturated oils contain double bonds between some of the carbon atoms in their carbon chains.
- Monounsaturated fats contain one C=C double bond somewhere in their carbon chains.
- Polyunsaturated fats contain more than one C=C double bond.

3) You can use the 'bromine water' test to check whether an oil or fat is saturated — see page 37.

Unsaturated oils can be **hydrogenated**

1) Unsaturated vegetable oils are liquid at room temperature.
2) They can be hardened by reacting them with hydrogen in the presence of a nickel catalyst at about 60 °C. This is called hydrogenation. The hydrogen reacts with the double-bonded carbons and opens out the double bonds.
3) Hydrogenated oils have higher melting points than unsaturated oils, so they're more solid at room temperature. This makes them useful as spreads and for baking cakes.
4) Margarine is usually made from partially hydrogenated vegetable oil — turning all the double bonds in vegetable oil to single bonds would make margarine too hard and difficult to spread. Hydrogenating most of them gives margarine a nice, buttery, spreadable consistency.
5) But partially hydrogenating vegetable oils means you end up with a lot of so-called trans fats. And there's evidence to suggest that trans fats are very bad for you.



You need to know the difference between saturated and unsaturated

This is tricky stuff. In a nutshell... there's saturated and unsaturated fats, which are generally bad and good for you (in that order) — easy enough. But... partially hydrogenated vegetable oil (which is unsaturated) is bad for you. Too much of the wrong types of fats can lead to heart disease. Got that...

Plant Oils as Fuel

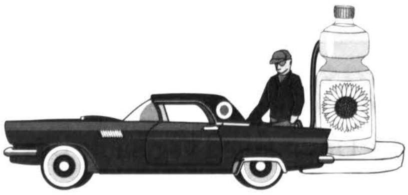
Fuel from vegetable oil is possible too — but as always, you have to weigh up the pros and cons.

Vegetable oils can be used to produce fuels

1) Vegetable oils such as rapeseed oil and soybean oil can be processed and turned into fuels.
2) Vegetable oil provides a lot of energy — that's why it's suitable for use as a fuel.
3) A particularly useful fuel made from vegetable oils is called biodiesel. Biodiesel has similar properties to ordinary diesel fuel — it burns in the same way, so you can use it to fuel a diesel engine.
4) Most diesel engines can burn 100% biodiesel, but usually biodiesel is mixed with ordinary diesel.
5) Engines burning biodiesel produce 90% as much power as engines burning ordinary diesel.

Biodiesel is a renewable fuel

1) Biodiesel comes from plant crops, which can be planted and harvested every year. You can always keep making biodiesel.
2) Compare this to ordinary diesel, which is made by distilling crude oil. Crude oil was formed millions of years ago and it'll take millions of years to make more — once it runs low that's it.



Biodiesel releases less pollution than ordinary diesel

1) Engines burning biodiesel produce much less sulfur dioxide pollution than engines burning diesel or petrol.
2) Burning biodiesel doesn't release as many "particulates" as burning diesel or petrol.

Particulates are little pieces of solid crud that you get in smoke and car exhausts.

3) Biodiesel is also biodegradable and it's less toxic than regular diesel.
4) Biodiesel engines do release the same amount of carbon dioxide (CO_2) as ordinary diesel engines. BUT biodiesel comes from recently grown plants. The plants took in carbon dioxide from the air when they were alive, and it's this same carbon which is released again when the biodiesel is burned. So net increase in carbon dioxide in the atmosphere: nil.
5) Regular diesel, on the other hand, comes from crude oil, which has been under the ground for millions of years. The carbon in crude oil was taken out of the atmosphere millions of years ago. Burning regular diesel does create a net increase in carbon dioxide in the atmosphere.

Biodiesel is expensive and it's difficult to make enough

1) We can't make enough biodiesel to replace regular diesel — there aren't enough veg oil crops. Biodiesel can be made from used vegetable oil, but there isn't enough of that either.
2) Because of this, biodiesel is expensive. Most people won't want to use it until it's cheaper.
3) Biodiesel has fewer drawbacks than some other "green" car fuels like biogas or electricity, though. Car engines need modification to run on gas — most diesel cars run on biodiesel without any tinkering. And biodiesel could use the same filling stations and pumps as diesel. (Compare this with electric cars, which would need a new network of recharging stations.)

Ethanol

There are different kinds of alcohol, but the one that's in beer, wine and so on is ethanol.

Ethanol can be made by fermentation

1) Fermentation is the process of using yeast to convert sugars into ethanol. Carbon dioxide is also produced. Here's the equation:

$$C_6H_{12}O_6 \xrightarrow{\text{yeast}} 2C_2H_5OH + 2CO_2$$

This is the formula for glucose — a common sugar.

The products are ethanol and carbon dioxide.

2) The yeast cells contain an enzyme called zymase. (Enzymes are naturally occurring catalysts.)

3) Fermentation happens fastest at a temperature of about 30 °C. At lower temperatures, the reaction slows down. If it's too hot the enzyme in the yeast is destroyed.

4) It's important to prevent oxygen getting to the fermentation process. If oxygen is present, a different reaction happens and you don't get ethanol.

5) When the concentration of alcohol reaches about 10 to 20%, the fermentation reaction stops, because the yeast gets killed off by the alcohol.

6) Different types of alcoholic drinks are made using sugars from different sources — usually from grains, fruits or vegetables, e.g. barley is used to make beer and grapes are used for making wine.

7) The fermented mixture can be distilled to produce more concentrated alcohol. Brandy is distilled from wine, whisky is distilled from fermented grain and vodka's distilled from fermented grain or potatoes.

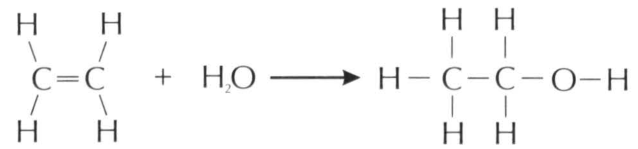
8) The ethanol produced this way can also be used as quite a cheap fuel in countries which don't have oil reserves for making petrol (see next page).

Ethanol

There's more than one way to make ethanol...

Ethene can be reacted with steam to produce ethanol

1) Ethene (C_2H_4) will react with steam (H_2O) to make ethanol.



2) The reaction needs a temperature of 300 °C and a pressure of 70 atmospheres.
3) Phosphoric acid is used as a catalyst.
4) At the moment this is a cheap process, because ethene's fairly cheap and not much of it is wasted.
5) The trouble is that ethene's produced from crude oil, which is a non-renewable resource and which will start running out fairly soon. This means using ethene to make ethanol will become very expensive.

Alcohol can be used as a fuel

1) Ethanol can be used as fuel. It burns to give just CO_2 and water.

2) Cars can be adapted to run on a mixture of about 10% ethanol and 90% petrol — 'gasohol'. Some countries (e.g. Brazil) make extensive use of gasohol. It's best used in areas where there's plenty of fertile land for growing the crops needed, and good crop-growing weather.

3) Using gasohol instead of pure petrol means that less crude oil is being used. Another advantage is the crops needed for ethanol production absorb CO_2 from the atmosphere in photosynthesis while growing. This goes some way towards balancing out the release of CO_2 when the gasohol is burnt.

4) But distilling the ethanol after fermentation needs a lot of energy, so it's not a perfect solution.

OH — it's all about ethanol

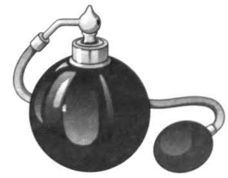
Make sure you learn the different ways that ethanol can be produced, that means you need to learn those equations as well, it'll also be useful if you can remember a few uses of alcohol too.

Perfumes

Some things smell nice, some don't... it's all down to the chemicals a substance contains.

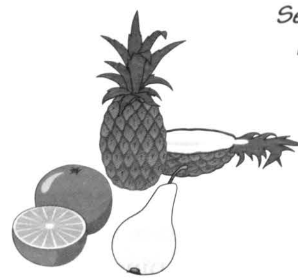
Perfumes can be natural or artificial

1) Chemicals that smell nice are used as perfumes and air fresheners. Esters are often used as perfumes as they usually smell quite pleasant.



2) Esters are pretty common in nature. Loads of common food smells (plus those in products like perfumes) contain natural esters.

3) Esters are also manufactured synthetically to enhance food flavours or aromas, e.g. there are esters (or combinations of esters) that smell of rum, apple, orange, pineapple, and so on. And esters are responsible for the distinctive smell of pear drops.



See pages 227-228 for more on esters and carboxylic acids.

Esters are made by esterification

1) Esters can be made by heating a carboxylic acid with an alcohol. (This is an example of esterification.)
2) An acid catalyst is usually used (e.g. concentrated sulfuric acid).

A carboxylic acid is an acid built around one or more carbon atoms.

$$\text{Acid} + \text{Alcohol} \rightarrow \text{Ester} + \text{Water}$$

Learn this equation.

Method:

1) Mix 10 cm³ of a carboxylic acid such as ethanoic acid with 10 cm³ of an alcohol such as ethanol.
2) Add 1 cm³ of concentrated sulfuric acid to this mixture and warm gently for about 5 minutes.
3) Tip the mixture into 150 cm³ of sodium carbonate solution (to neutralise the acids) and smell carefully (by wafting the smell towards your nose). The fruity-smelling product is the ester.

Perfumes

Not just any ester will do though...

Perfumes need **certain properties**

You can't use any old chemical with a smell as a perfume. You need a substance with certain properties:

1) Easily evaporates — or else the perfume particles won't reach your nose and you won't be able to smell it... bit useless really.

See page 7 for more about how we smell perfumes.

2) Non-toxic — it mustn't seep through your skin and poison you.
3) Doesn't react with water — or else it would react with the water in sweat.
4) Doesn't irritate the skin — or else you couldn't apply it directly to your neck or wrists. If you splash on any old substance you risk burning your skin.
5) Insoluble in water — if it was soluble in water it would wash off every time you got wet.

Don't forget that even if a substance has all these properties, it still might smell pretty bad and so be unsuitable for a perfume.

New perfumes and **cosmetics** have to be **tested**

Companies are always developing new cosmetic products to sell to us. Before they're released to the shops, they need to be tested thoroughly to make sure they're safe to use. They should be non-toxic and shouldn't irritate the eyes or skin. Pretty obvious, I'm sure you'll agree. But some tests are carried out using animals, which is a bit more controversial.

Advantages of testing new cosmetics on animals: We get an idea of whether they're likely to irritate the skin or be toxic before humans use them (though an animal test won't necessarily apply to humans).

Disadvantages of testing on animals: The tests could cause pain and suffering to the animals (especially if it turns out that the cosmetic is toxic). And animals can't choose whether or not to take part in the tests (so using human volunteers instead could be a possibility in certain circumstances).

Testing on animals is a controversial issue

Perfume needs to smell nice, but not everyone agrees on what smells nice. Perfume also needs to be safe, but not everyone agrees on the best way to test for this. That's life for you.

Warm-Up and Exam Questions

It's time to practise some more questions. If you struggle with the warm-up questions, do some more revision then try them again. Only do the exam questions when you think you know your stuff properly.

Warm-up Questions

1) Describe the process used to extract and refine oil from plants.
2) What conditions are used for the hydrogenation of unsaturated vegetable oils?
3) What useful fuel can be produced by fermentation?
4) Whisky, brandy and vodka all involve distillation in their production. What effect does the distillation process have?
5) What is gasohol? Name a country where it is used extensively.
6) Why do new perfumes and cosmetics have to be tested?

Exam Questions

1 Match the words **A**, **B**, **C** and **D** with numbers **1 – 4** in the sentences.

A oxygen

B sugar

C carbon dioxide

D glucose

The raw material for fermentation is ...**1**..., a type of ...**2**....

Fermentation produces the gas ...**3**... as well as alcohol.

However if ...**4**... is present a different reaction happens.

(4 marks)

2 Match words **A**, **B**, **C** and **D** with the numbers **1 - 4** in the sentences below.

A saturated

B unsaturated

C polyunsaturated

D monounsaturated

Animal fats tend to be ...**1**....

Vegetable oils are usually ...**2**....

...**3**... oils have only one C=C double bond in their carbon chains, whereas ...**4**... oils contain more than one C=C double bond.

(4 marks)

Exam Questions

3 Esters are often used as perfumes. It is possible to make an ester by reacting a carboxylic acid with an alcohol.

(a) This is an example of what type of reaction?

(1 mark)

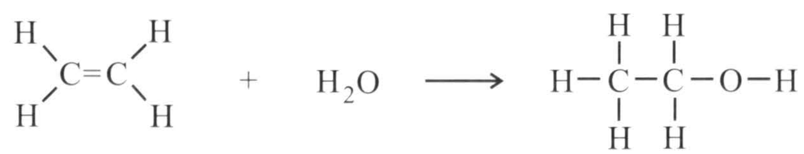
(b) Write a word equation for the reaction between a carboxylic acid and an alcohol.

(1 mark)

(c) List five properties that perfumes need to have.

(5 marks)

4 (a) Ethanol can be synthesised in the following reaction:



Reactant A is derived from crude oil. Explain why the cost of making ethanol by the synthesis method above is likely to increase.

(1 mark)

(b) Ethanol burns in oxygen to give carbon dioxide and water.

(i) Write a balanced symbol equation for this reaction.

(2 marks)

(ii) One method of ethanol production partially balances out the carbon dioxide produced when ethanol burns. Describe this method and explain why it helps balance out the carbon dioxide emissions from burning the ethanol.

(2 marks)

5 Biodiesel can be used in diesel engines.

(a) What is biodiesel made from?

(1 mark)

(b) (i) Explain why biodiesel doesn't add to the greenhouse effect.

(2 marks)

(ii) Give **two** other advantages of biodiesel.

(2 marks)

(c) Give **two** disadvantages of biodiesel.

(2 marks)

Revision Summary for Section Four

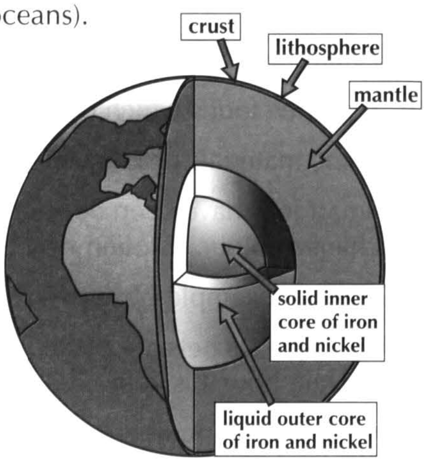
The only way that you can tell if you've learned this module is to test yourself. Try these questions and if there's something you don't know, it means you need to go back and learn it. Even if it is all that weird stuff about smart materials. And don't miss any questions out — you don't get a choice about what comes up on the exam so you need to be sure that you've learnt it all.

1) Give three reasons why many foods are cooked before eating.
2) Explain what is meant by an 'emulsion'.
3) List three foods which contain emulsifiers.
4) Some materials produce electricity when squeezed. How might these be used commercially?
5) Smart inks have been developed which get darker with time. The warmer they get, the faster they change colour. Explain why this is a useful property for a freshness indicator in food packaging.
6) Give an example of active packaging.
7) Silica gel sachets are placed into some chilled food packets. Explain how this keeps the food safe to eat for longer.
8) What is an E-number?
9) Explain why foods containing butter often have antioxidant chemicals added to them.
10) Give two advantages and two disadvantages of using food additives.
11) Describe how chromatography can be used to separate the different colours in a sweet.
12) Describe how olive oil is extracted from olives.
13) Describe the basic structure of a polyunsaturated fat.
14) How can unsaturated oils be hardened? Why is this done?
15) Explain why biodiesel has not replaced ordinary diesel.
16) Write the word equation for the fermentation reaction used to produce alcohol.
17) Explain why the temperature must be monitored carefully when brewing beer.
18) Describe a method of producing ethanol, other than by fermentation.
19) What is gasohol made from?
20) Give three properties that a substance must have in order to make a good perfume.
21) Many perfumes are esters. Describe one method of making an ester in the lab.
22) Write the general word equation for this esterification reaction.
23) Give one advantage and one disadvantage of testing new perfumes on animals.

The Earth's Structure

It's tricky to study the structure of the Earth — you can't just dig down to the Earth's centre. But after studying the evidence, this is what scientists think is down there...

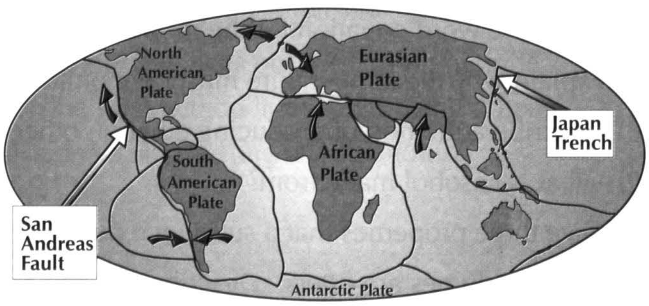
Crust, mantle, outer and inner core



1) The crust is Earth's thin outer layer of solid rock. There are two types of crust — continental crust (forming the land), and oceanic crust (under oceans).
2) The lithosphere includes the crust and upper part of the mantle below, and is made up of a jigsaw of 'plates'. The lithosphere is relatively cold and rigid.
3) The mantle extends from the crust almost halfway to the centre of the Earth. It's got all the properties of a solid but it can flow very slowly.
4) The core is just over half the Earth's radius. It's mostly iron and nickel, and is where the Earth's magnetic field originates.
5) The inner core is solid, while the outer core is liquid.
6) Radioactive decay creates a lot of the heat inside the Earth.
7) This heat causes convection currents, which cause the plates of the lithosphere to move (which is bad news for some people — see below).

The Earth's surface is made up of tectonic plates

1) The crust and the upper part of the mantle are cracked into a number of large pieces called tectonic plates. These plates are a bit like big rafts that 'float' on the mantle.
2) The plates don't stay in one place though. That's because the convection currents in the mantle cause the plates to drift. The map shows the edges of the plates as they are now, and the directions they're moving in (red arrows).
3) Most of the plates are moving at speeds of a few cm per year relative to each other.
4) Occasionally, the plates move very suddenly, causing an earthquake. Volcanoes often form at the boundaries between two tectonic plates too.



Scientists can't predict earthquakes and volcanic eruptions

1) Tectonic plates can stay more or less put for a while and then suddenly lurch forwards. It's impossible to predict exactly when they'll move.
2) Scientists are trying to find out if there are any clues that an earthquake might happen soon — things like strain in underground rocks. Even with these clues they'll only be able to say an earthquake's likely to happen, not exactly when it'll happen.
3) There are also clues that a volcanic eruption might happen soon. Before an eruption, molten rock rises up into chambers near the surface, causing the ground surface to bulge slightly. This causes mini-earthquakes near the volcano.
4) But sometimes molten rock cools down instead of erupting, so mini-earthquakes can be a false alarm.

Evidence for Plate Tectonics

A bloke called Alfred Wegener put forward his theory about the Earth's continents slowly drifting along in 1915, but not many people believed it. This was partly because he didn't have a good explanation for why it happened, partly because he wasn't a qualified geologist, and partly because the theory was so weird. But the truth will out, as they say — and the evidence now suggests the 'rocky raft' idea is correct.

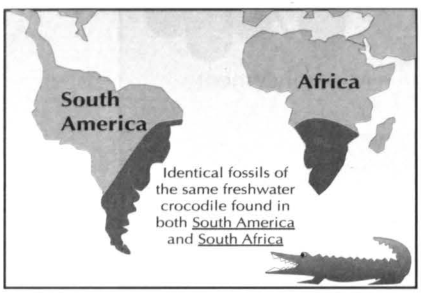
1) Jigsaw fit — the supercontinent 'Pangaea'

a) There's a very obvious jigsaw fit between Africa and South America.

b) The other continents can also be fitted in without too much trouble.

c) It's widely believed that they once all formed a single land mass, now called Pangaea.

2) Matching fossils in Africa and South America



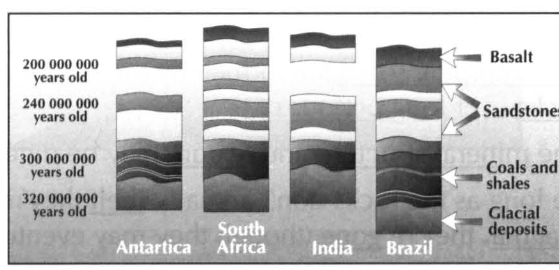
a) Identical plant fossils of the same age have been found in rocks in South Africa, Australia, Antarctica, India and South America, which strongly suggests they were all joined once upon a time.

b) Animal fossils support the theory too. There are identical fossils of a freshwater crocodile found in both Brazil and South Africa. It certainly didn't swim across.

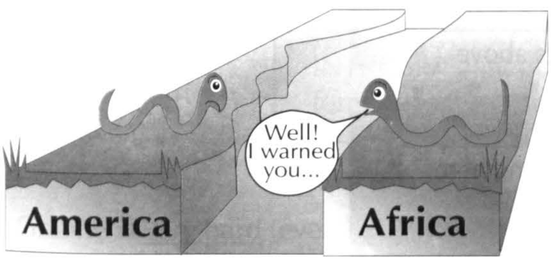
3) Identical rock sequences

a) Certain rock layers of similar ages in various countries show remarkable similarity.

b) This is strong evidence that these countries were joined together when the rocks formed.



4) Living creatures: The Earthworm



a) There are various living creatures found in both America and Africa.

b) One such beasty is a particular earthworm which is found living at the tip of South America and the tip of South Africa.

c) Most likely it travelled across ever so slowly on the big raft we now call America.

Learn about Wegener's Theory and all the evidence

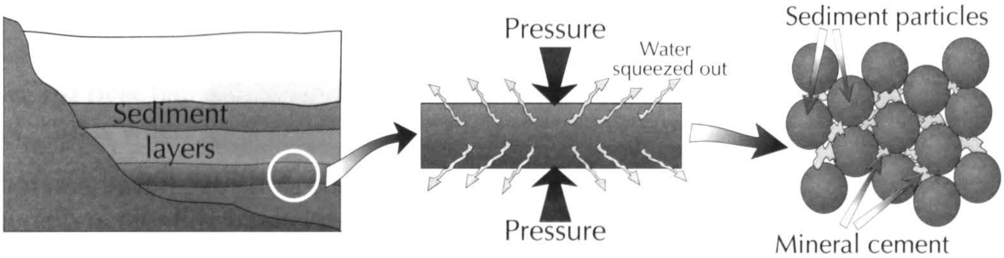
So there you go. Alfred Wegener's ideas were originally thought to be bonkers (the fact that he'd used some inaccurate data didn't help — one scientist claimed that the forces Wegener's theory needed would have stopped the Earth rotating). But as technology improved and more evidence was gathered (including from the bottom of the ocean), it turned out that Wegener's ideas were pretty convincing after all. But it took a while — it was only in the 1960s that scientists really accepted the theory.

The Three Different Types of Rock

You can use rocks (or stuff in them) to make all sorts of things, but not all rocks are the same. Scientists classify rocks according to how they're formed. The three different types are: sedimentary, metamorphic (on this page) and igneous (on the following page).

There are **three steps** in the formation of **sedimentary rock**

1) Sedimentary rocks are formed from layers of sediment laid down in lakes or seas.
2) Over millions of years the layers get buried under more layers and the weight pressing down squeezes out the water.
3) Fluids flowing through the pores deposit natural mineral cement.

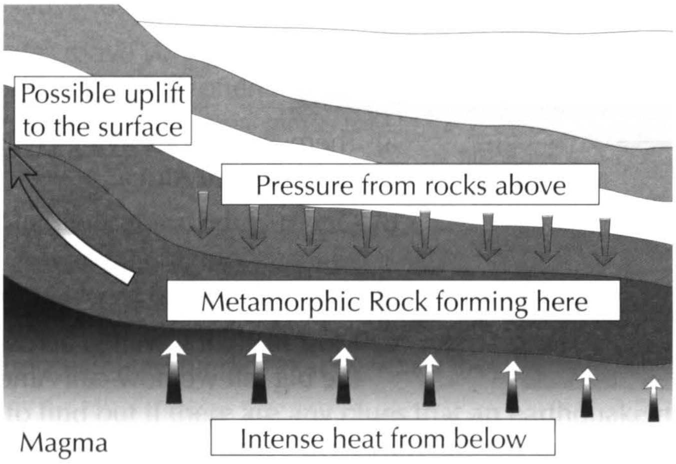


Limestone is a sedimentary rock formed from **seashells**

Limestone is that rather useful grey-white rock that you first encountered on page 15. Make sure you know all the important facts about it from that page.

Metamorphic rocks are formed from **other rocks**

1) Metamorphic rocks are formed by the action of heat and pressure on sedimentary (or even igneous) rocks over long periods of time.
2) The mineral structure and texture may be different, but the chemical composition is often the same.
3) So long as the rocks don't actually melt they're classed as metamorphic. If they melt and turn to magma, they're gone (though they may eventually resurface as igneous rocks — see next page).



Marble is a metamorphic rock formed from **limestone**

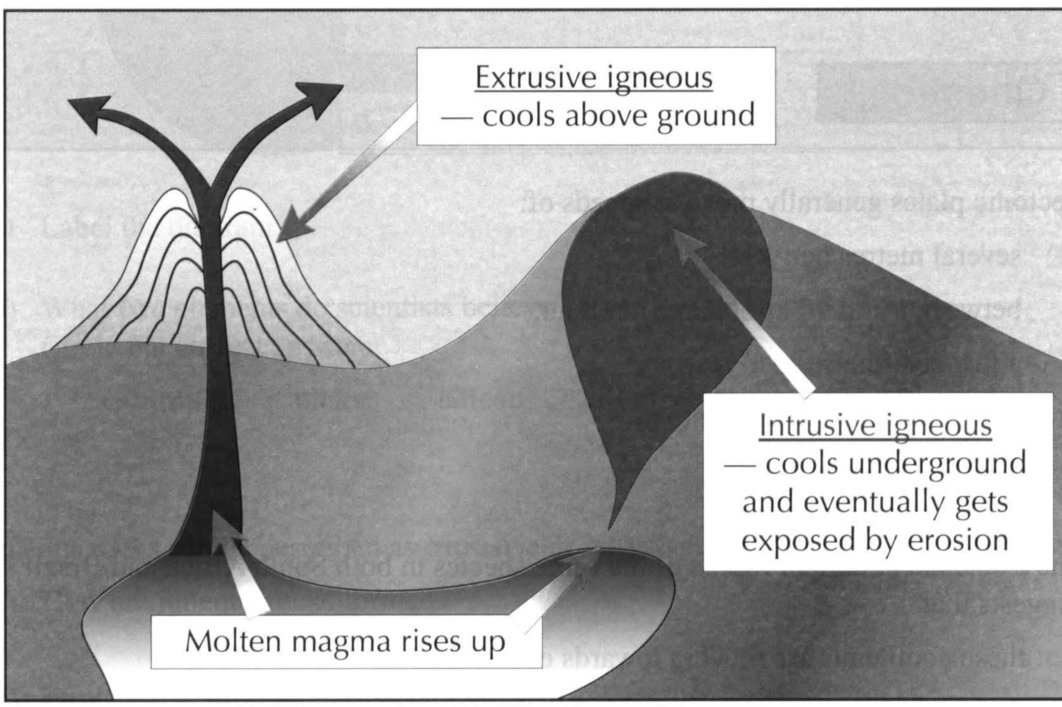
Marble is another form of calcium carbonate. Very high temperatures break down the limestone and it reforms as small crystals. This gives marble a more even texture and makes it much harder.

The Three Different Types of Rock

Sedimentary rocks are generally pretty soft, metamorphic rocks are harder, and igneous rocks are the hardest of the lot.

Igneous rocks *are formed from* ***fresh magma***

1) Igneous rocks form when molten magma pushes up into the crust (or right through it) before cooling and solidifying. They contain various different minerals in randomly arranged interlocking crystals.
2) There are two types of igneous rocks, extrusive and intrusive:



EXTRUSIVE igneous rocks cool QUICKLY ABOVE GROUND, forming SMALL crystals, e.g. basalt and rhyolite.

INTRUSIVE igneous rocks cool SLOWLY UNDERGROUND, forming BIG crystals, e.g. granite and gabbro.

Granite is very hard (even harder than marble). It's ideal for steps and buildings.

Igneous rocks are either intrusive or extrusive

The extrusive igneous rock basalt is made of the same minerals as the intrusive rock gabbro — both contain lots of iron. But gabbro is coarser because it cools more slowly, giving bigger crystals. Rhyolite and its coarser brother granite are made of the same stuff too — both contain lots of silicon.

Warm-Up and Exam Questions

So now it's time to see which facts stayed in your head, and which tumbled back out. Question time.

Warm-Up Questions

1) State one geological feature often seen at the boundary of two tectonic plates.
2) Give an example of a visible clue that a volcano might be about to erupt.
3) What was 'Pangaea'?
4) What are the three steps in the formation of sedimentary rock?
5) Name one common use of granite.

Exam Questions

1 Tectonic plates generally move at speeds of:

A several metres per year.

B between 5 and 10 kilometres per hour.

C a few centimetres per year.

D 1 or 2 millimetres per century.

(1 mark)

2 The existence of animal fossils of the same species in both South Africa and Brazil suggests that:

A these continents are moving towards each other.

B South America and Africa were once joined together.

C South America and Africa were never joined together.

D South America and Africa have changed places.

(1 mark)

3 Metamorphic rocks, such as marble, are formed from other types of rock.

(a) Describe how metamorphic rocks are formed from sedimentary rocks.

(2 marks)

(b) (i) What sedimentary rock is marble formed from?

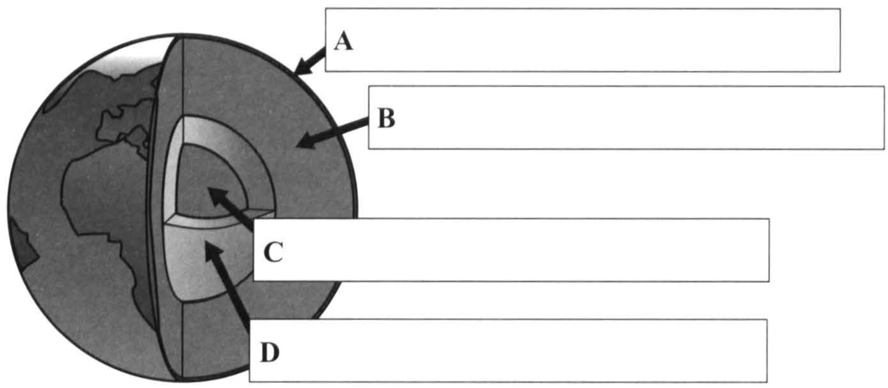
(1 mark)

(ii) Give two differences between the physical properties of marble and those of the rock it is formed from.

(2 marks)

Exam Questions

4 The following diagram shows the internal structure of the Earth.



(a) Label the diagram.

(4 marks)

(b) What **two** elements do scientists believe part **C** is largely made from? Circle the correct answers.

cadmium **nickel** **silicon** **aluminium** **iron**

(2 marks)

5 Igneous rocks can be described as extrusive or intrusive.

(a) How are igneous rocks formed?

(2 marks)

(b) Explain the difference between extrusive and intrusive igneous rocks, and give an example of each.

(3 marks)

6 Which of the following statements about limestone is **not** true?

A It is mostly formed from seashells.

B It is a metamorphic rock.

C It thermally decomposes when it is heated.

D It is virtually insoluble in water.

(1 mark)

7 Alfred Wegener came up with the theory of continental drift.

(a) Describe two pieces of evidence to support his theory.

(2 marks)

(b) Give one reason why it wasn't accepted at the time.

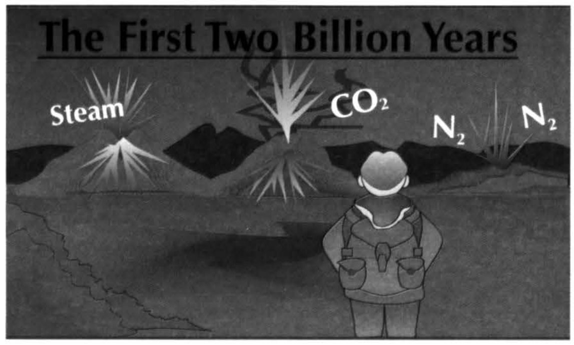
(1 mark)

Evolution of the Atmosphere

For 200 million years or so, the atmosphere has been about how it is now: 78% nitrogen, 21% oxygen, and small amounts of other gases, mainly CO_2 and noble gases. There can be a lot of water vapour too. But it wasn't always like this. Here's how the past 4.5 billion years may have gone:

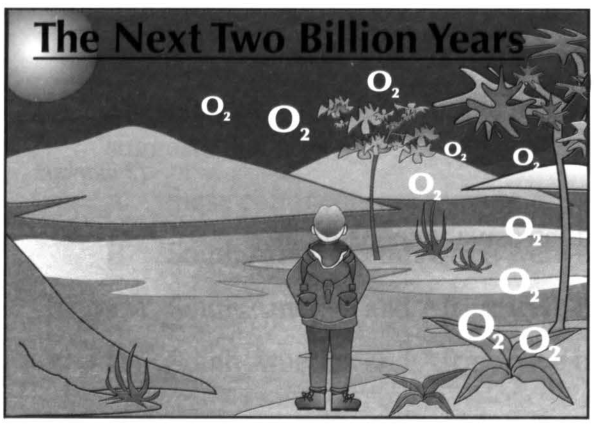
Phase 1 — Volcanoes gave out gases

1) The Earth's surface was originally molten for many millions of years. It was so hot that any atmosphere just 'boiled away' into space.
2) Eventually things cooled down a bit and a thin crust formed, but volcanoes kept erupting.
3) The volcanoes gave out lots of gas — including carbon dioxide, water vapour and nitrogen. We think this was how the oceans and atmosphere were formed.
4) According to this theory, the early atmosphere was probably mostly CO_2, with virtually no oxygen. This is quite like the atmospheres of Mars and Venus today.
5) The oceans formed when the water vapour condensed.



Holiday report: Not a nice place to be. Take strong walking boots and a good coat.

Phase 2 — Green plants evolved and produced oxygen



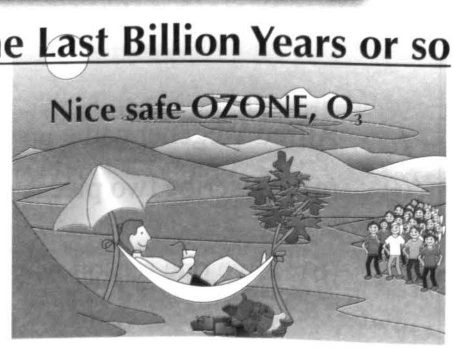
Holiday report: A bit slimy underfoot. Take wellies and a lot of suncream.

1) Green plants evolved over most of the Earth. They were quite happy in the CO_2 atmosphere.
2) A lot of the early CO_2 dissolved into the oceans. The green plants also removed CO_2 from the air and produced O_2 by photosynthesis.
3) When plants died and were buried under layers of sediment, the carbon they had removed from the air (as CO_2) became 'locked up' in sedimentary rocks as insoluble carbonates and fossil fuels.
4) When we burn fossil fuels today, this 'locked-up' carbon is released and the concentration of CO_2 in the atmosphere rises.

Phase 3 — Ozone layer allows evolution of complex animals

1) The build-up of oxygen in the atmosphere killed off some early organisms that couldn't tolerate it, but allowed other, more complex organisms to evolve and flourish.
2) The oxygen also created the ozone layer (O_3) which blocked harmful rays from the Sun and enabled even more complex organisms to evolve — us, eventually.
3) There is virtually no CO_2 left now.

Holiday report: A nice place to be. Visit before the crowds ruin it.



Atmospheric Change

Evidence for how the atmosphere evolved has been found in rocks and other sources. But no one was there to record the changes as they happened. So our ideas about the atmosphere are still theories...

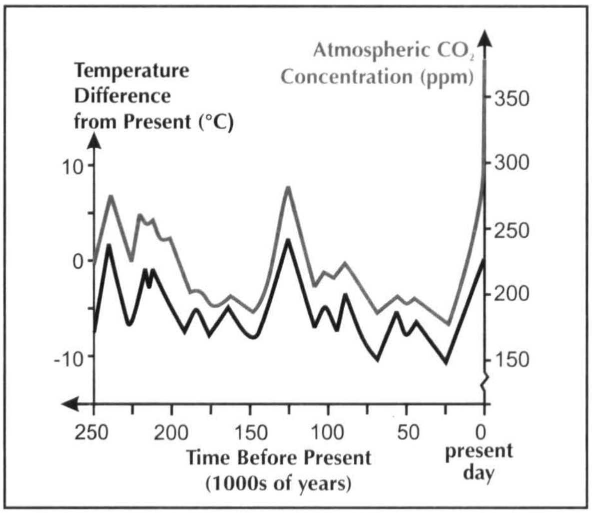
There are **competing theories** about **atmospheric change**

As well as the theory on the previous page, there are other theories about how the Earth's atmosphere changed millions of years ago. Ultimately, all the theories have to be judged on the evidence.

> For example, one theory says that the water on Earth came mainly from comets rather than volcanoes. When this theory was first suggested, it seemed far-fetched. But space science research soon suggested that lots of small icy comets really are hitting the Earth every day. So far so good. But studies of comets found that the water in comets isn't the same as the water on Earth (it's got more 'heavy water' in it). So current thinking is that most of Earth's water probably didn't come from comets.

The atmosphere **changes** all the time

1) This is a graph of CO_2 and global temperature data. It shows CO_2 levels rising rapidly over the last few thousand years, and a global temperature rise that's been more or less keeping up.



2) But the graph also shows that there have been huge changes in the climate before. (These changes are small compared with the changes described on page 70, when the entire composition of the atmosphere was changing... but they're still pretty big.)

3) For instance, there have been several ice ages over the last few million years. These happen for various reasons (e.g. things to do with the Earth's orbit, movement of continents, CO_2 in the atmosphere, and so on).

An ice age is a time when large areas of the Earth's surface are covered with ice.

4) So changes in the Earth's temperature aren't new — they happen all the time. However, we've recently become aware of the possibility of a faster warming of the planet, and what this could mean for us.

4 million years ago was a whole other world

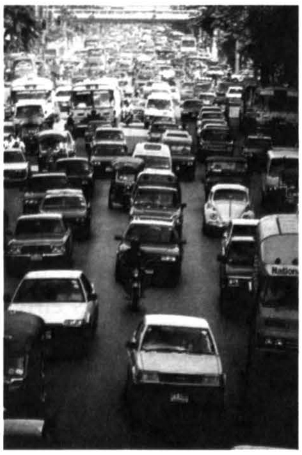
We've learned a lot about the past atmosphere from Antarctic ice cores. Each year, a layer of ice forms and bubbles of air get trapped inside it, then it's buried by the next layer. So the deeper the ice, the older the air — and if you examine the bubbles in different layers, you can see how the air has changed.

Atmospheric Change

There have been huge changes in the climate in the past. Now the climate is changing again, but there's still lots of debate about whether it is just a natural change or change brought about by human activity.

The atmosphere is **still changing**

Levels of CO_2 in the atmosphere are **increasing**



Levels of CO_2 in the atmosphere have increased by about 25% since 1750...

1) Burning fossil fuels releases CO_2 — and as the world has become more industrialised, more fossil fuels have been burnt in power stations and in car engines.
2) Carbon dioxide is a greenhouse gas — it traps heat from the Sun. You'd expect that more carbon dioxide would mean a hotter planet. (See page 75 for more info.)
3) However, a few scientists say that the concentration of CO_2 has increased and decreased a lot over the last 100s of millions of years, and argue that a little increase now might be just a blip.

The amount of **ozone** in the ozone layer has **decreased**

Over the last 50 years, the amount of ozone in the ozone layer has decreased...

1) Currently, holes in the Earth's ozone layer form over Antarctica and the Arctic each year.
2) Ozone is broken down by man-made gases called CFCs, widely used as aerosol propellants and fridge coolants between the 1930s and the 1980s. CFCs were phased out in the 1990s.
3) The ozone layer protects us from the harmful UV radiation which can cause skin cancer. It's difficult to test whether changes in the ozone layer are to blame for increases in skin cancer, though. Other factors affect skin cancer — people sunbathe more and have more beach holidays abroad, so they expose themselves to more UV radiation anyway.

See p.244-245 for more about CFCs.

The atmosphere is changing — it could be all our fault

Whether people believe scientific theories or not depends on the evidence that people produce to support them. Without evidence, a theory goes nowhere. Quite right too, I say.

Warm-Up and Exam Questions

Doing these warm-up questions will soon tell you if you've got the basic facts straight. If not, you'll really struggle, so take time to go back over the bits you don't know.

Warm-Up Questions

1) Name the two main gases that make up the Earth's atmosphere today.
2) Give one effect of the build-up of oxygen in the atmosphere over the last billion years.
3) Give one use of CFCs.

Exam Questions

1 Match the words for **A**, **B**, **C** and **D** with the numbers **1 - 4** in the sentences below.

A photosynthesis

B oxygen

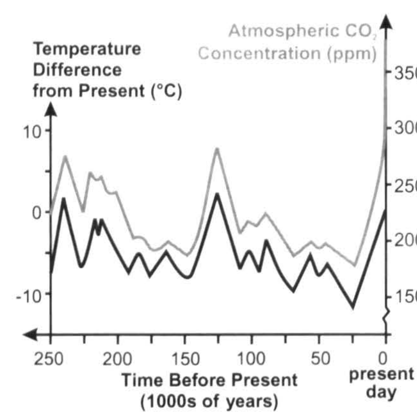
C carbon

D carbon dioxide

Once green plants had evolved, they thrived in an atmosphere rich in ...**1**....
These plants produced ...**2**... by the process of ...**3**....
...**4**... from dead plants eventually became 'locked up' in fossil fuels.

(4 marks)

2 The following graph shows how atmospheric CO_2 concentration and global temperature have varied over the last 250 000 years.



(a) Describe what the graph shows about temperature and CO_2 levels.

(2 marks)

(b) Mark with an X on the graph the time when the temperature was most different from its present value.

(1 mark)

3 (a) (i) Name two regions over which the Earth's ozone layer is thinner than normal.

(2 marks)

(ii) What effect do CFCs (chlorofluorocarbons) have on ozone?

(1 mark)

(b) Discuss the possible connection between ozone levels and a rise in the incidence of skin cancer in humans.

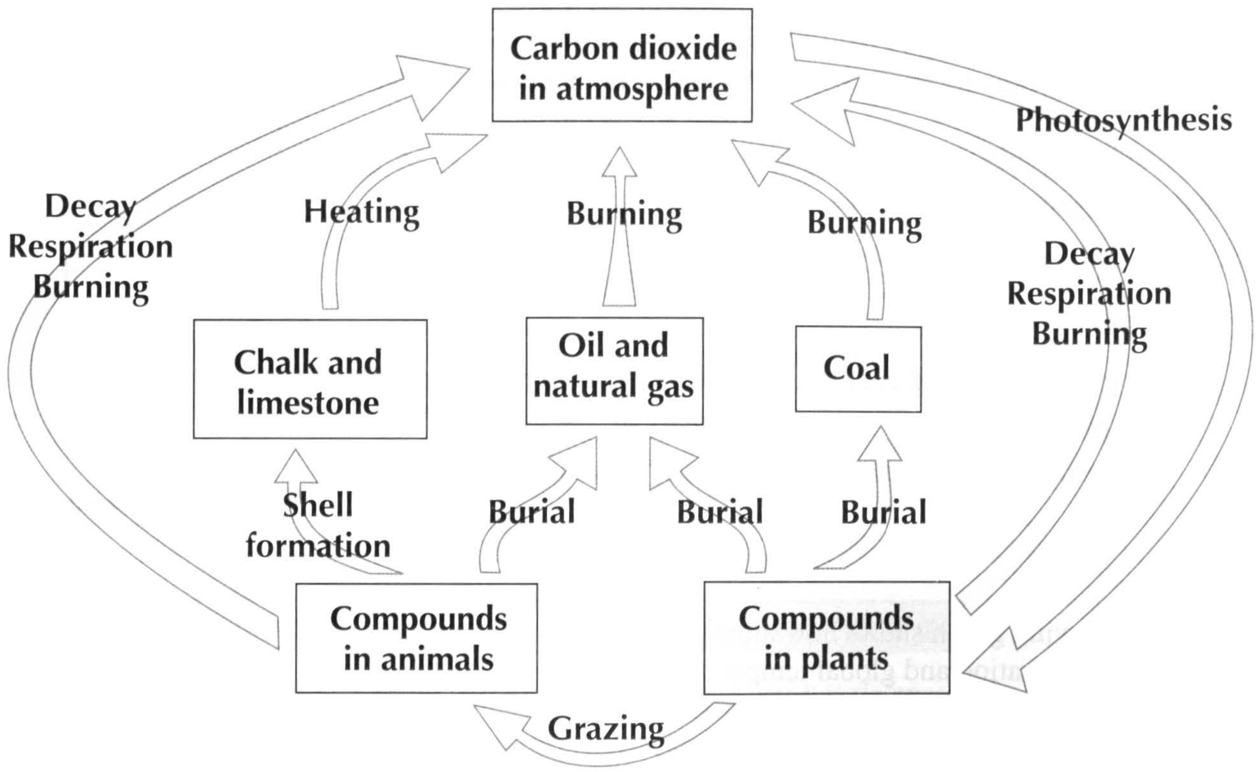
(3 marks)

The Carbon Cycle

Carbon flows through the Earth's ecosystem in the carbon cycle.
This cycle can help us see why the amount of carbon dioxide in the atmosphere has increased.

Carbon is constantly being recycled

Carbon is the key to the 'greenhouse effect' which keeps the Earth warm (see next page). It exists in the atmosphere as carbon dioxide gas (CO_2), and is also present in many other greenhouse gases such as methane (CH_4).



1) The carbon on Earth moves in a big cycle.
2) Respiration (in living cells), combustion (burning) and decay of plants and animals add carbon dioxide to the air and remove oxygen.
3) Photosynthesis does the opposite — it removes carbon dioxide and adds oxygen.
4) These processes should balance out. However, it looks like humans have upset the natural carbon cycle.

Eeeek — the carbon cycle's got a puncture

Releasing masses of CO_2 due to fossil fuel use is the worst problem here. But deforestation (cutting down large areas of forest) is also bad for the carbon cycle, in three ways. First off, carbon dioxide is released when trees are burnt to clear land. Secondly, microorganisms feeding on the dead wood release CO_2 due to respiration. And thirdly, fewer trees means there's less photosynthesis going on.

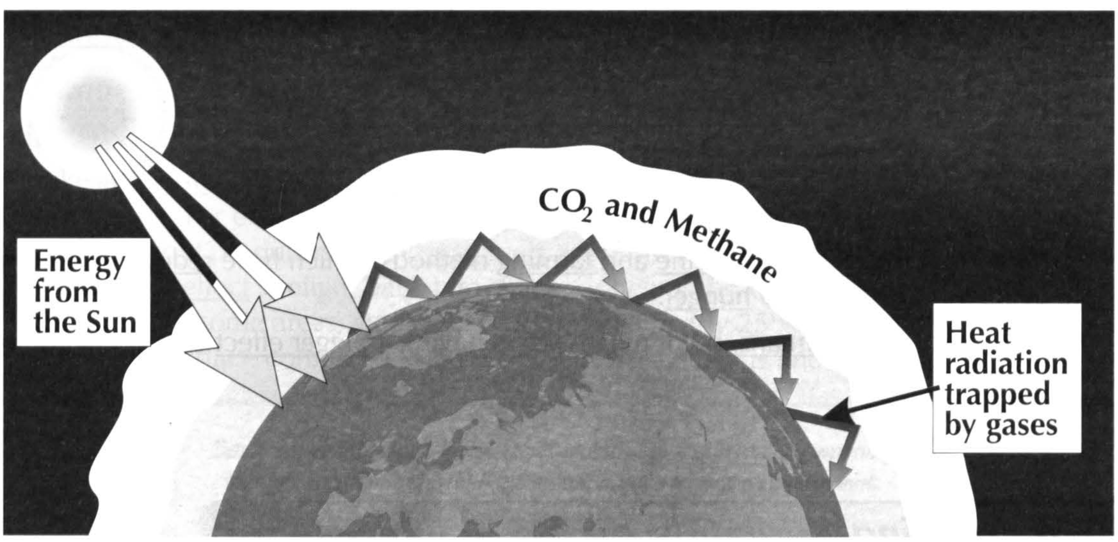
Global Warming

Most environmentalists and scientists now believe that human activities are changing the proportion of carbon dioxide in the atmosphere — and that that's going to have massive effects on life on Planet Earth. You need to understand the science behind the scary headlines — starting with the greenhouse effect.

Carbon dioxide and methane trap heat from the Sun

1) The temperature of the Earth is a balance between the heat it gets from the Sun and the heat it radiates back out into space.

2) Gases in the atmosphere absorb most of the heat that would normally be radiated into space, and re-radiate it in all directions (including towards Earth). If this didn't happen, then at night there'd be nothing to keep any heat in, and we'd quickly get very cold indeed. But recently we've started to worry that this effect is getting a bit out of hand...



3) There are several different gases in the atmosphere which help keep the heat in. They're called "greenhouse gases" (oddly enough) and the main ones whose levels we worry about are carbon dioxide and methane — because the levels of these two gases are rising quite sharply.

4) Humans release carbon dioxide into the atmosphere as part of our everyday lives — e.g. as we burn fossil fuels in power stations or cars.

5) This could be a big problem, but it's hard to be 100% sure (since Earth's climate is so complicated). For example, the Earth's temperature varies over the years anyway (see page 71) — so even if Earth is warming up, it might be nothing to do with humans and fossil fuels.

But nowadays, most scientists think that:

(i) Earth is gradually warming

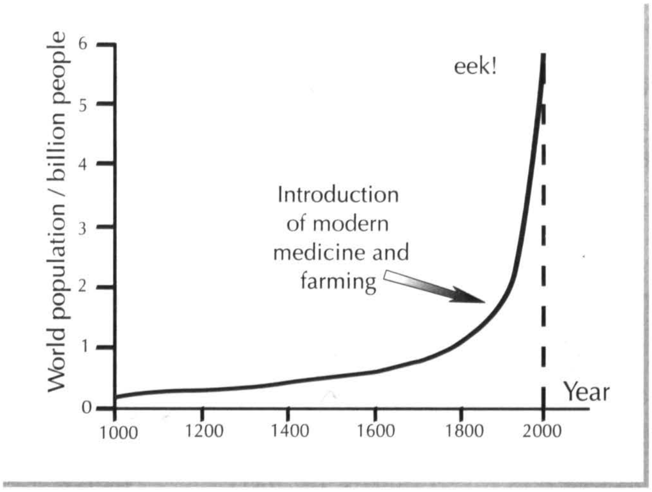
(ii) fossil fuel use has got something to do with it.

Human Impact on the Environment

We have an impact on the world around us — and the more humans there are, the bigger the impact.

There are **six billion people** *in the world...*

1) The population of the world is currently rising very quickly, and it's not slowing down — look at the graph...



2) This is mostly due to modern medicine and farming methods, which have reduced the number of people dying from disease and hunger.
3) This is great for all of us humans, but it means we're having a bigger effect on the environment we live in...

...with **increasing demands** *on the* **environment**

When the Earth's population was much smaller, the effects of human activity were usually small and local. Nowadays though, our actions can have a far more widespread effect.

1) Our rapidly increasing population puts pressure on the environment, as we take the resources we need to survive.
2) But people around the world are also demanding a higher standard of living (and so demand luxuries to make life more comfortable — cars, computers, etc.). So we use more raw materials (e.g. oil to make plastics), and we also use more energy for the manufacturing processes. This all means we're taking more and more resources from the environment more and more quickly.
3) Unfortunately, many raw materials are being used up quicker than they're being replaced. So if we carry on like we are, one day we're going to run out.

Human Impact on the Environment

We are damaging the environment in several different ways...

We're also producing **more waste**

As we make more and more things we produce more and more waste. And unless this waste is properly handled, more harmful pollution will be caused. This affects water, land and air.

Water

Sewage and toxic chemicals from industry can pollute lakes, rivers and oceans, affecting the plants and animals that rely on them for survival (including humans). And the chemicals used on land (e.g. fertilisers) can be washed into water.

Land

We use toxic chemicals for farming (e.g. pesticides and herbicides). We also bury nuclear waste underground, and we dump a lot of household waste in landfill sites.

Air

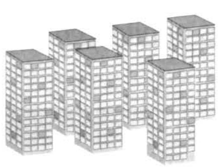
Smoke and gases released into the atmosphere can pollute the air (see pages 74, 75 and 78 for more). For example, sulfur dioxide can cause acid rain. Scientists also think that particles of soot produced when fossil fuels are burnt can stop sunlight reaching the Earth. The particles reflect sunlight back into space, or help to produce more clouds that do the same thing. In some areas, studies have found that nearly 25% less sunlight has been reaching the surface compared to 50 years ago. This effect is known as global dimming.

Some scientists don't believe that global dimming is really happening though, and blame the findings on inaccurate recording equipment.

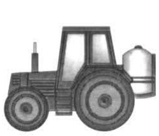
More people means **less land** for plants and other animals

Humans also reduce the amount of land and resources available to other animals and plants. The four main human activities that do this are:

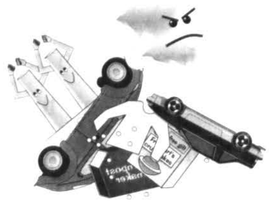
1) Building



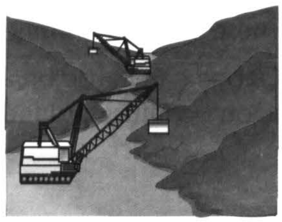
3) Farming



2) Dumping Waste



4) Quarrying



More people, more mess, less space, less resources

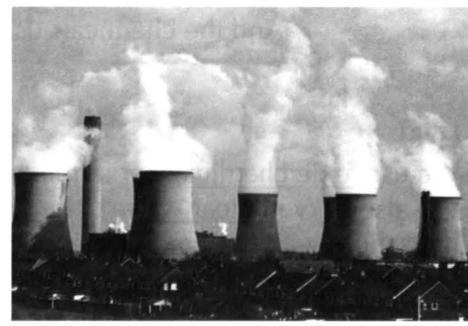
In the exam you might be given some data about environmental impact, so make sure you understand what's going on. Just keep your head and work out exactly what the data's saying. Job's a good 'un.

Air Pollution and Acid Rain

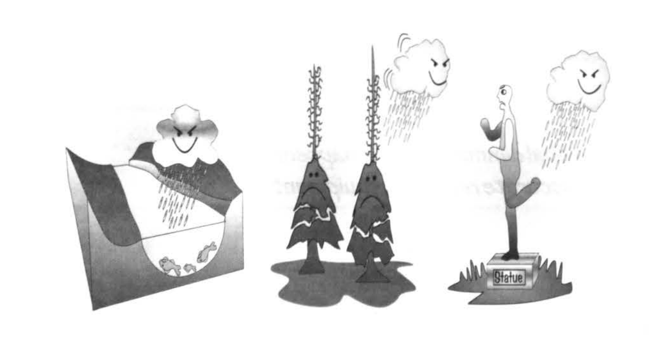
Carbon dioxide levels could be causing climate change. But CO_2 isn't the only gas released when fossil fuels burn — you also get other nasties like oxides of nitrogen and sulfur dioxide.

*Acid rain is caused by **sulfur dioxide** and **oxides of nitrogen***

1) When fossil fuels are burned they release mostly CO_2.
   But they also release other harmful gases — e.g. sulfur dioxide and various nitrogen oxides.
2) The sulfur dioxide (SO_2) comes from sulfur impurities in the fossil fuels.
3) However, the 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.)
4) When these gases mix with clouds they form dilute sulfuric acid and dilute nitric acid.
   This then falls as acid rain.
5) Power stations and internal combustion engines in cars are the main causes of acid rain.



Acid rain kills fish, trees and statues



1) Acid rain causes lakes to become acidic and many plants and animals die as a result.
2) Acid rain kills trees and damages limestone buildings and ruins stone statues. It also makes metal corrode. It's shocking.

Oxides of nitrogen also cause photochemical smog

1) Photochemical smog is a type of air pollution caused by sunlight acting on oxides of nitrogen.
2) These oxides combine with oxygen in the air to produce ozone (O_3).
3) Ozone can cause breathing difficulties, headaches and tiredness.
   (Don't confuse ground-level ozone with the useful ozone layer high up in the atmosphere.)

Acid rain — bad for the trees

But good for you, because acid rain can get you some nice easy marks in the exam. Make sure you remember the names of the two gases which cause acid rain, and where they come from.

Air Pollution and Acid Rain

Carbon monoxide is another pollutant that can be released from burning fossil fuels — it's pretty nasty. Pollution in the atmosphere such as CO can cause health problems, so it needs to be carefully controlled.

Carbon monoxide is a poisonous gas

1) Carbon monoxide (CO) can stop your blood doing its proper job of carrying oxygen around the body.
2) A lack of oxygen in the blood can lead to fainting, a coma or even death.
3) Carbon monoxide is formed when petrol or diesel in car engines is burnt without enough oxygen — this is incomplete combustion (see page 34 for more details).

It's important that atmospheric pollution is controlled

The build-up of all these pollutants can make life unhealthy and miserable for many humans, animals and plants. The number of cases of respiratory illnesses (e.g. asthma) has increased in recent years — especially among young people. Many people blame atmospheric pollution for this, so efforts are being made to improve things:

Catalytic converters

Catalytic converters on motor vehicles reduce the amount of carbon monoxide and nitrogen oxides getting into the atmosphere. The catalyst is normally a mixture of platinum and rhodium. It helps unpleasant exhaust gases from the car react to make things that are less immediately dangerous (though more CO_2 is still not exactly ideal).

carbon monoxide + nitrogen oxide → nitrogen + carbon dioxide

$$2CO + 2NO \rightarrow N_2 + 2CO_2$$

Flue Gas Desulfurisation (FGD)

Flue Gas Desulfurisation (FGD) technology in some fossil-fuel power stations removes sulfur dioxide from the exhaust gases.

Warm-Up and Exam Questions

You could just read through this page thinking 'yep, I can do that', but there'd not be much point. You need to write down an answer to every question, check them, and look up anything you didn't know.

Warm-up Questions

1) Carbon dioxide is a greenhouse gas. What does this mean?
2) Give one reason why the world's population is rising rapidly.
3) How is the rise in our standard of living affecting the environment?
4) Name two gases, apart from CO_2, that are often produced when fossil fuels burn.
5) Describe two effects of acid rain.

Exam Questions

1 The large human population is currently leading to

A a reduction in quarrying

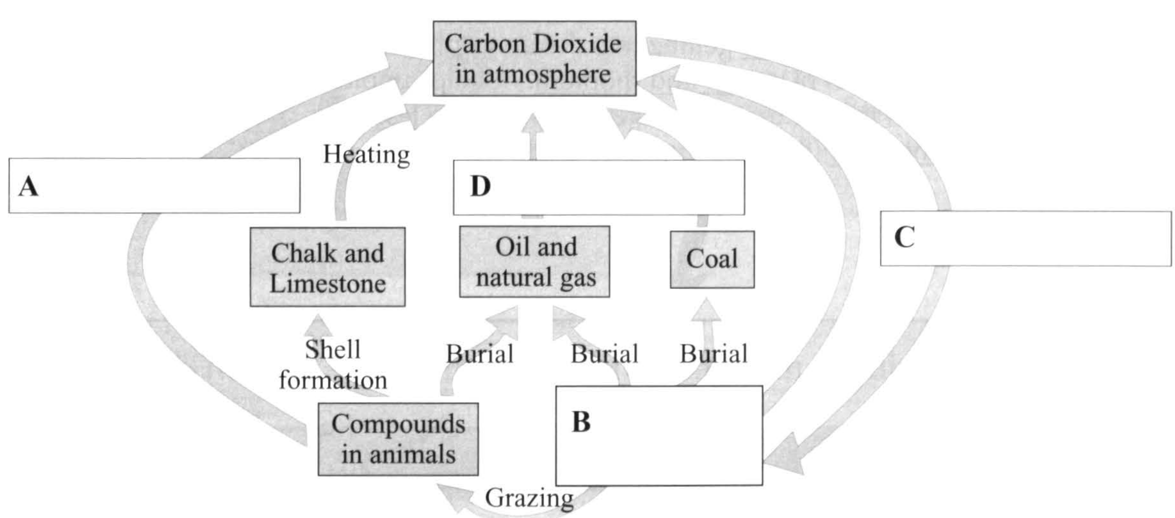
B less pollution

C rapid depletion of some resources

D more land being available for plants

(1 mark)

2 The diagram below shows the carbon cycle.



(a) Choose from the words and phrases below to fill in the blanks in the diagram.

respiration **compounds in plants** **burning** **photosynthesis** **burial**

(4 marks)

(b) Describe how large-scale deforestation affects CO_2 concentration in the atmosphere, and explain why.

(3 marks)

Exam Questions

3 Global warming happens when:

A more of the Sun's heat is reflected into space than is absorbed by the Earth.

B oceans absorb more and more CO_2.

C there is more CO_2 and methane in the atmosphere.

D we use a greater proportion of renewable energy sources.

(1 mark)

4 Match the words **A**, **B**, **C** and **D** with the numbers **1 - 4** in the sentences below.

A carbon monoxide

B nitrogen oxides

C oxygen

D smog

Acid rain is formed when ...**1**... and SO_2 mix with clouds.

Nitrogen oxides can cause photochemical ...**2**....

Breathing in another pollutant, ...**3**..., hinders the uptake of ...**4**... by the blood and is potentially fatal.

(4 marks)

5 Human activities can cause damage to water, land and air.

(a) (i) How might fertilisers used on land cause water pollution?

(1 mark)

(ii) Name **two** sources of water pollution.

(2 marks)

(b) How does household waste cause land pollution?

(1 mark)

(c) Name one source of land pollution other than household waste.

(1 mark)

6 Global dimming may be caused by:

A too much methane in the atmosphere.

B soot particles from burning fossil fuels, which reflect sunlight back into space.

C CO_2 particles from burning fossil fuels, which reflect sunlight back into space.

D the absorption of infrared radiation by CO_2.

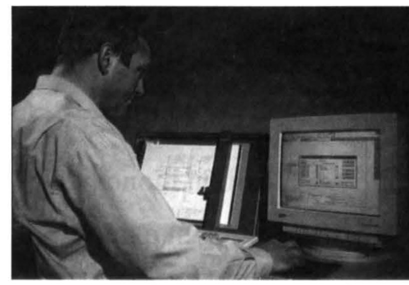
(1 mark)

Protecting the Atmosphere

What to do... what to do... How can we look after the atmosphere...?

Computer models are used to make predictions

1) Computer models are used to predict the temperature of the Earth's atmosphere in the future. They use data collected by thousands of monitoring stations all over the world.
2) The data is fed into the models, then millions of calculations are carried out.
3) However, computer models are only as good as the data you put into them, and the assumptions made when working out the calculations. If the assumptions are wrong, this could lead to false results. And one small error in an early calculation could be magnified if it's used to predict further into the future. (Having said that, early comparisons of computer predictions and observed events look pretty good.)



The precautionary principle — better safe than sorry

1) Various governments have agreed to apply the 'precautionary principle' to climate change.
2) The idea is that we should assume the worst and therefore reduce CO_2 emissions. If we turn out to be wrong, then the climate's safe anyway. But if we turn out to be right, we've taken early action.
3) There are two basic strategies for combating climate change...

(i) Burn less fossil fuels.
We should burn fossil fuels more efficiently, and also use other sources of energy (that don't emit greenhouse gases — e.g. nuclear, wind, etc.).

(ii) Burn fossil fuels but try to stop levels of greenhouse gases increasing so much.
We could capture some of the CO_2 before it's released into the atmosphere. And we could plant forests to absorb some of the CO_2 (as they photosynthesise — see page 74).

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.

Protecting the Atmosphere

As the demand on resources increases it is important to develop new, alternative fuels.

Alternative fuels are being developed

Some alternatives to fossil fuels already exist, and there are others in the pipeline (so to speak). They should reduce the amount of fossil fuels burnt.

BIOGAS is a mixture of methane and carbon dioxide. It's produced when microorganisms digest waste material. It can be produced on a large scale, or on a small scale where each family has its own generator.
Biogas is burned and the energy can be used for cooking, heating or lighting.

PROS: Waste material is readily available and cheap. It's 'carbon neutral'.

CONS: Biogas production is slow in cool weather.

And see page 59 about using ethanol as a fuel.

HYDROGEN GAS can also be used to power vehicles. You get the hydrogen from the electrolysis of water. There's plenty of water about but it takes electrical energy to split it up — however, this energy can come from a renewable source, e.g. solar.

PROS: Hydrogen combines with oxygen in the air to form just water — so it's very clean.

CONS: You need a special, expensive engine and hydrogen isn't widely available.
You still need to use energy from another source to make it.
Also, hydrogen's hard to store — it's very explosive.

There's lots to consider when choosing a fuel

1) Energy value (i.e. amount of energy) — funnily enough, this isn't always as important as it may seem.

2) Availability — there's not much point in choosing a fuel you can't get hold of easily.

3) Storage — some fuels take up a lot of space, and some produce flammable gases.

4) Cost — some fuels are expensive, but still good value in terms of energy content etc.

5) Toxicity — poisonous fumes are a problem.

6) Pollution — e.g. will you be adding to acid rain and the greenhouse effect? Or causing lots of smoke?

Example: You're at home and there's a power cut. You want a cup of tea. The only fuels you have in the house are candles or meths (in a spirit burner). Which one would you use to boil the water?

Fuel	Energy per gram	Rate of energy produced	Flame
Meths	28 kJ	15 kJ per minute	Clean
Candle	50 kJ	8 kJ per minute	Smoky

Even though a candle has more energy per gram, you'd probably choose meths because it's quicker and cleaner.

I produce a kind of biogas already

As it's mostly fossil fuel use that gives the atmosphere such a hard time, alternative fuels are a good thing to start looking for. And getting people to use less energy is also a pretty sensible idea too.

Recycling Materials

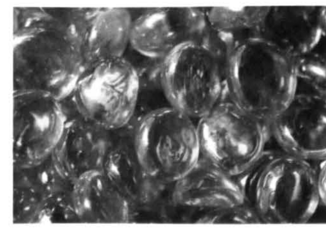
It's not all doom and gloom... if we do things sustainably e.g. by recycling, we'll be okay.

It's important to recycle

There are various reasons why...

1) Use less resources

There's a finite amount of materials (e.g. metals, oil for plastics) in the Earth. Recycling conserves these resources.



2) Use less energy

Mining, extracting and making materials (e.g. metals or glass) need lots of energy, which mostly comes from burning fossil fuels. Fossil fuels will run out one day, and they also cause pollution (see page 77 for more about global dimming, but also have a look at pages 71-79 for more general pollution info).
Recycling things like copper, aluminium and glass takes a fraction of the energy.

3) Use less money

Energy doesn't come cheap, so recycling saves money too.

4) Make less rubbish

Recycling also cuts down on the amount of rubbish that goes to landfill, which takes up space and pollutes the surroundings.

This is a reason why we should recycle paper.

*Hard work never killed anyone, but why take a chance**

You can calculate the financial benefits of recycling any material, but remember there are the 'resources', 'energy' and 'rubbish' benefits too. With paper, sustainable forests are good (where for every tree you cut down, you plant another one), but that doesn't reduce the amount of landfill.

* This is presented here in aid of Recycle an Elderly Joke Week

Recycling Materials

Recycling will benefit humans now and help to reduce our impact on future generations.

There may be **economic** and **environmental** benefits of recycling

1) Working out the cost benefits of recycling can get a bit tricky — there's lots to take into account.
2) For example, recycling isn't free. There are costs involved in collecting waste material, transporting it, sorting it, and then processing it.
3) But if you didn't recycle, say, aluminium, you'd have to mine more aluminium ore — 4 tonnes for every 1 tonne of aluminium you need. But mining makes a mess of the landscape (and these mines are often in rainforests). The ore then needs to be transported, and the aluminium extracted (which uses loads of electricity). And don't forget the cost of sending your used aluminium to landfill.

4) So it's a complex calculation, but for every 1 kg of aluminium cans you recycle, you save:

- 95% or so of the energy needed to mine and extract 'fresh' aluminium,
- 4 kg of aluminium ore,
- a lot of waste.

In fact, aluminium's about the most cost-effective metal to recycle.

5) But even if all these differences were very small, maybe it's still worth recycling — you're getting people involved in doing their bit for the environment. Can't be a bad thing.

Sustainable development needs careful **planning**

1) Human activities can damage the environment (e.g. pollution). And some of the damage we do can't easily be repaired (e.g. the destruction of the rainforests).
2) We're also placing greater pressure on our planet's limited resources (e.g. oil is a non-renewable resource so it will eventually run out).
3) This means that we need to plan carefully to make sure that our activities today don't mess things up for future generations — this is the idea behind sustainable development...

> SUSTAINABLE DEVELOPMENT meets the needs of today's population without harming the ability of future generations to meet their own needs.

4) This isn't easy — it needs detailed thought at every level to make it happen. For example, governments around the world will need to make careful plans. But so will the people in charge at a regional level.

Recycling is an example of sustainable development

Recycling materials is a great way in which we can start to reduce our impact on future generations. But there are lots of other ways too — e.g. managing forests sustainably and introducing fishing quotas.

Warm-Up and Exam Questions

These warm-up questions should ease you in gently then come the exam questions.
Unless you've learnt the facts you'll find the exams tougher than leather sandwiches.

Warm-up Questions

1) Why is hydrogen considered to be a 'clean fuel'.
2) Give four important criteria to consider when choosing a fuel.
3) Name two non-renewable resources.
4) In a cost-benefit analysis of recycling, name an important cost.
5) What is meant by 'sustainable development'?

Exam Questions

1 Recycling materials is one way of helping to ensure sustainable development.

(a) Explain how recycling glass and metals can help us use less energy.

(2 marks)

(b) Give two other reasons why it is important to recycle.

(2 marks)

(c) Name one process involved in recycling that costs money.

(1 mark)

2 Currently, fossil fuels provide about 60-70% of the world's electricity.

Since fossils fuels will eventually run out, it is important to find alternative energy sources for the future.

(a) Biogas is one fuel that may be more widely used in the future.

(i) What are the main constituents of biogas?

(1 mark)

(ii) How is biogas produced and what could it be used for?

(2 marks)

(iii) Give **one** advantage and **one** disadvantage of using biogas.

(2 marks)

(b) Another 'fuel for the future' could be hydrogen.

(i) Describe how hydrogen is produced on a large scale.

(1 mark)

(ii) Why is hydrogen difficult to store?

(1 mark)

(iii) Give **one** other disadvantage of using hydrogen as a fuel in a car compared with using petrol or diesel.

(1 mark)

Revision Summary for Section Five

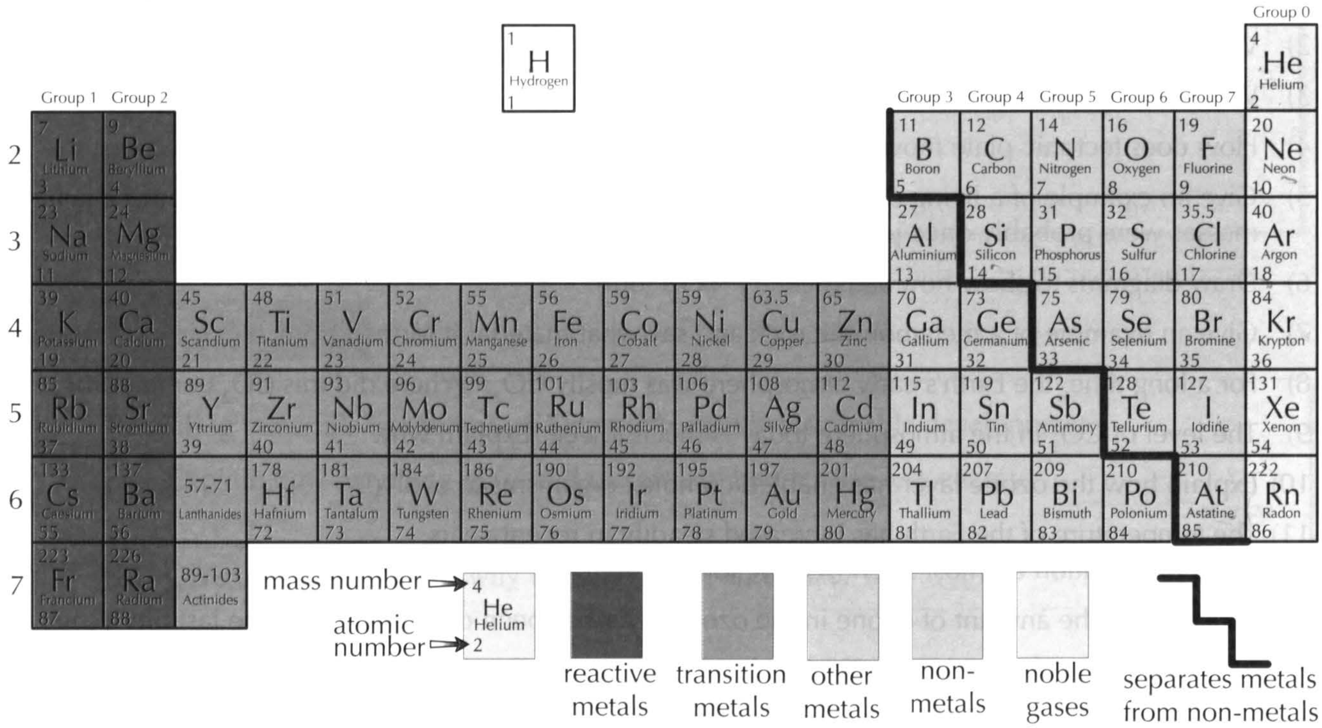
The structure and atmosphere of the Earth, rocks, pollution, alternative fuels — can they really belong in the same section, I almost hear you ask. Whether you find the topics easy or hard, interesting or dull, you need to learn it all before the exam. Try these questions... see how much you know.

1) What is the lithosphere?
2) What is the Earth's core made from?
3) What causes the Earth's tectonic plates to move?
4) How does tectonic plate movement cause: a) earthquakes? b) volcanoes?
5) Give an example of a living creature found in both America and Africa that shows that these land masses were probably once joined.
6) Draw diagrams to show how sedimentary rocks form.
7) Give an example of a metamorphic rock and say what material it formed from.
8) For a long time, the Earth's early atmosphere was mostly CO_2. Where did this CO_2 come from?
9) The level of CO_2 in the atmosphere today is much lower. Explain why.
10) Explain how the ozone layer has enabled complex organisms to evolve.
11) The temperature of the Earth has increased steadily in recent years. What explanation do most scientists give for this?
12) Explain why the amount of ozone in the ozone layer has been decreasing over the last fifty years.
13) Name two 'greenhouse gases'. Explain how these gases affect the temperature of the Earth.
14) Describe three ways in which an increasing population affects the environment.
15) Give three human activities that reduce the amount of land and resources for other organisms.
16) Which gases cause 'acid rain'? How do they get into the air?
17) Describe two ways of reducing acid rain.
18) Explain what is meant by 'photochemical smog'.
19) Name a poisonous gas that catalytic converters help to remove from car exhausts.
20) How can scientists predict what is likely to happen to the Earth's atmosphere in the future?
21) Explain what is meant by the precautionary principle, and how it can be applied to global warming.
22) Describe the benefits of using hydrogen to power vehicles.
23) Describe the benefits and difficulties of using biogas to power vehicles.

The Periodic Table and Electron Shells

Here's your old friend from Section One, back again — because you need to know it <u>really well</u>.

*The **periodic table** is a table of all the known **elements***



Hopefully you already know (from page 10) that:

1) the periodic table contains <u>all</u> of the 100 or so known <u>elements</u>...
2) ...in order of <u>ascending atomic number</u>...
3) ...and arranged into columns (<u>groups</u>) that share <u>similar properties</u>.

But that's really only half the story. Read on.

The atomic number is the number of protons, which conveniently also tells you the number of electrons. See page 6 for more.

*Elements in a **group** have the **same number** of **outer electrons***

1) The elements in each <u>group</u> all have the same number of <u>electrons</u> in their <u>outer shells</u>. Group 1 elements have one outer electron, Group 7 elements have seven outer electrons, and so on.
2) That's why they have <u>similar properties</u>. And that's why they're arranged in this way.
3) When only 50 or so elements were known, the periodic table was made by looking at the <u>properties</u> of the elements and arranging them in groups — the same groups that they are in today.
4) This next idea is <u>extremely important</u> to chemistry — so make sure you understand it:

> The properties of the elements are decided <u>entirely</u> by how many electrons they have. So <u>atomic number</u> is very significant, because it's equal to the number of electrons each atom has. But it's the number of electrons in the <u>outer shell</u> that's the most important thing.

The Periodic Table and Electron Shells

Electron shells are what chemistry is all about

The fact that electrons form shells around atoms is the basis for the whole of chemistry.

If they just whizzed round the nucleus any old how and didn't care about shells, there'd be no chemical reactions. No nothing in fact — because nothing would happen. The atoms would just sit there.

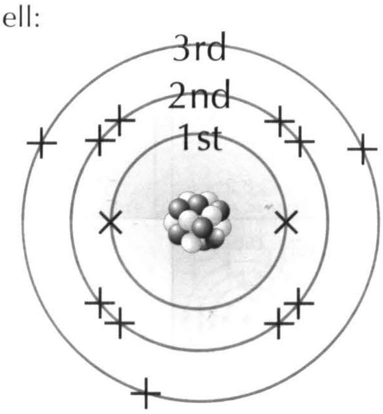
But amazingly, they do form shells (if they didn't, we wouldn't even be here to wonder about it), and the electron arrangement of each atom determines the whole of its chemical behaviour. Electron arrangements explain practically the whole Universe. Pretty amazing.

Electrons always follow the same pattern when filling shells

It's really important that you know these electron shell rules:

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 (see page 109).
5) In most atoms the outer shell is not full and this makes the atom want to react to fill it.



3rd shell still filling

Physicists can produce new elements in particle accelerators

All the new elements made this way are radioactive. Most only last a fraction of a second before they decay into other elements. They haven't even got round to giving most of them proper names yet, but then even "element 114" sounds pretty cool when you say it in Latin — ununquadium...

Electron Shells

You need to know the electron configurations for the first twenty elements
(things get a bit more complicated after that — luckily you don't have to worry about it).

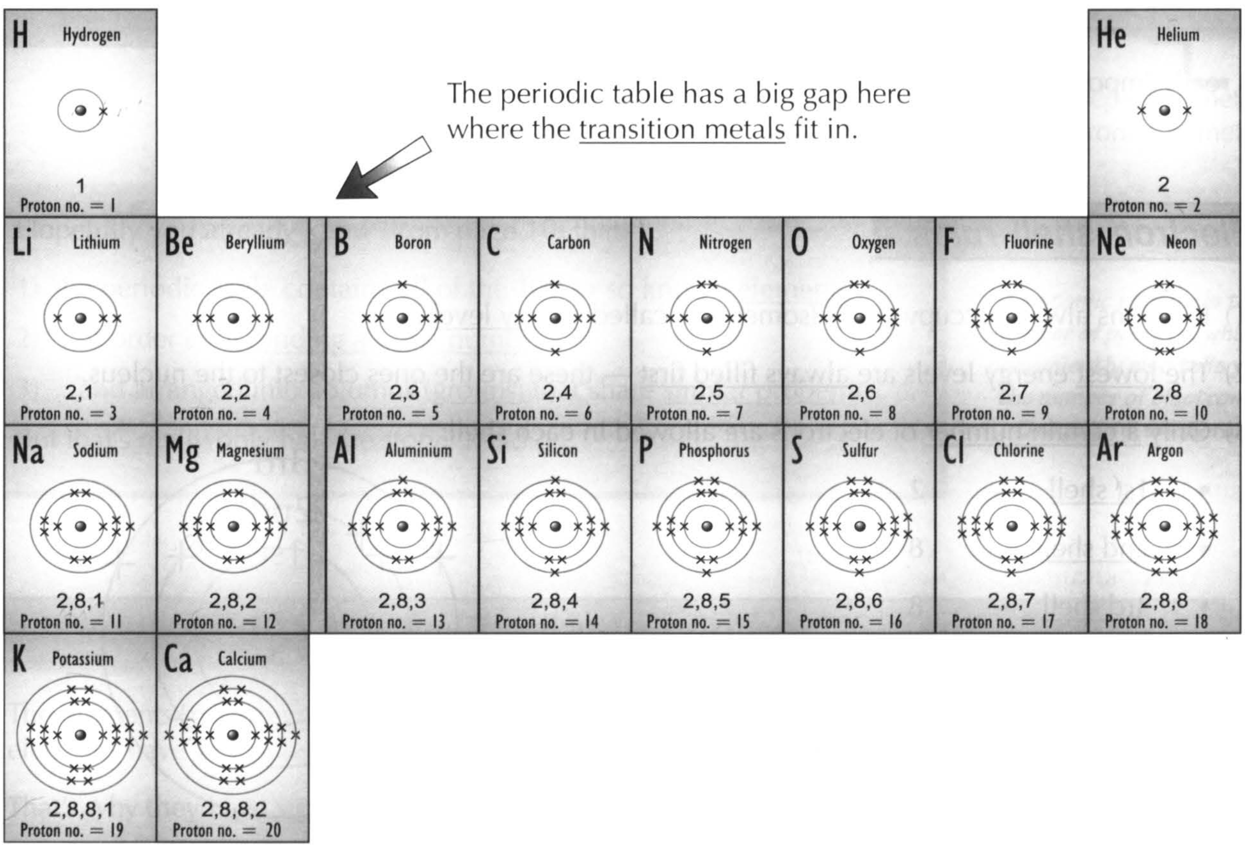
*Follow the rules to **work out** electron configurations*

Electron configurations are not hard to work out.

For a quick example, take nitrogen:

1) The periodic table tells you nitrogen has seven protons... so it must have seven electrons.
2) Follow the 'Electron Shell Rules' from the last page. The first shell can only take 2 electrons and the second shell can take a maximum of 8 electrons.
3) So the electron configuration for nitrogen must be 2, 5.

The best way to get better at working these out is to practise, so now you try it for argon.



Answer:

To calculate the electron configuration of argon, follow the rules. It's got 18 protons, so it must have 18 electrons. The first shell must have 2 electrons, the second shell must have 8, and so the third shell must have 8 as well. It's as easy as 2, 8, 8.

Ionic Bonding

In ionic bonding, atoms lose or gain electrons to form charged particles (called ions) which are then strongly attracted to one another (because of the attraction of opposite charges, + and –).

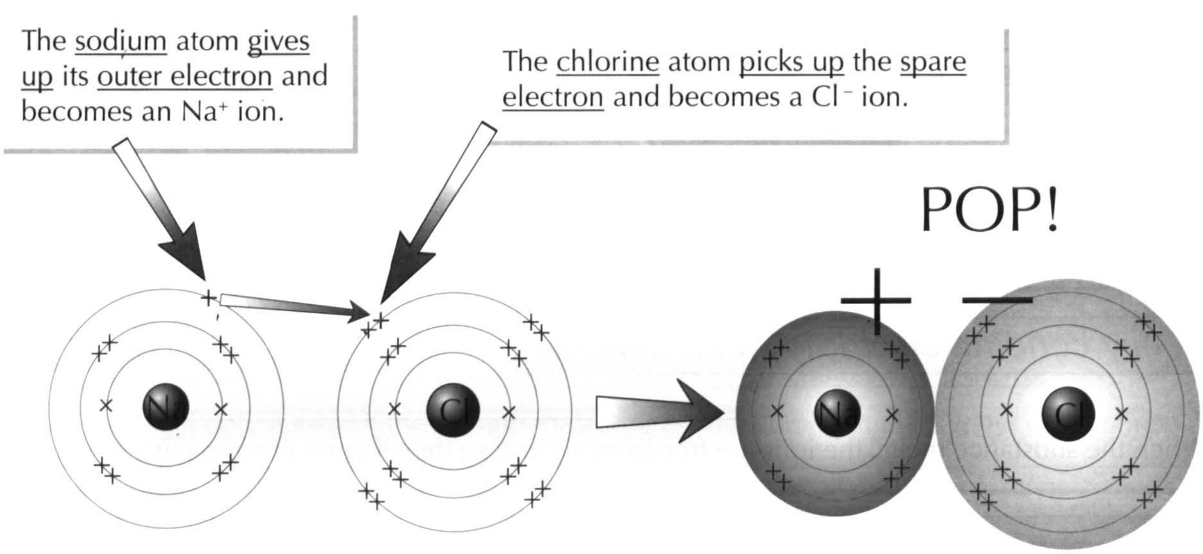
A shell with just **one** electron is **well keen to get rid...**

1) All the atoms over at the left-hand side of the periodic table, e.g. sodium, potassium, calcium, etc. have just one or two electrons in their outer shell.
2) They're pretty keen to get shot of these electrons, because then they'll only have full shells left, which is how they like it.
3) So given half a chance they do get rid, and that leaves the atom as an ion instead.
4) Now ions aren't the kind of things that sit around quietly watching the world go by. They tend to leap at the first passing ion with an opposite charge and stick to it like glue...

A **nearly full** shell is **well keen** to get that **extra electron...**

1) On the other side of the periodic table, the elements in Group 6 and Group 7, such as oxygen and chlorine, have outer shells which are nearly full.
2) They're obviously pretty keen to gain that extra one or two electrons to fill the shell up.
3) When they do of course they become ions and before you know it, pop, they've latched onto the atom (ion) that gave up the electron a moment earlier.

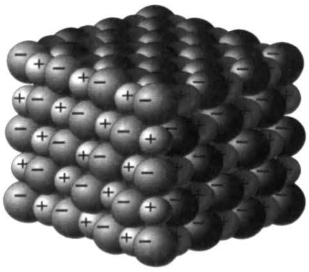
The reaction of sodium and chlorine is a classic case:



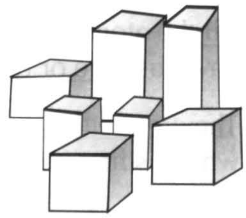
Ionic Bonding

Ionic bonds produce giant ionic structures.

Giant ionic structures don't melt easily — but when they do...



1) Ionic bonds always produce giant ionic structures.
2) The ions form a closely packed regular lattice arrangement.
3) There are very strong chemical bonds between all the ions.
4) A single crystal of salt is one giant ionic lattice, which is why salt crystals tend to be cuboid in shape.



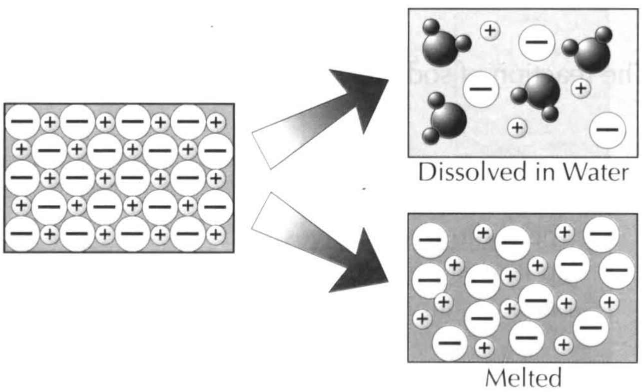
1) They have **high melting points** and **boiling points**

This is due to the very strong chemical bonds between all the ions in the giant structure.

2) They **dissolve** to form solutions that **conduct electricity**

When dissolved, the ions separate and are all free to move in the solution. These free-moving charged particles allow the solution to carry electric current.

Dissolved lithium salts are used to make rechargeable batteries.



3) They **conduct** electricity when **molten**

When the substance melts, the ions are free to move and so they'll carry electric current.

Batteries need to contain a conducting solution

Because they conduct electricity when they're dissolved in water, ionic compounds are used to make some types of battery. In the olden days, most batteries had actual liquid in, so they tended to leak all over the place. Now they've come up with a sort of paste that doesn't leak but still conducts. Clever.

Ions and Formulas

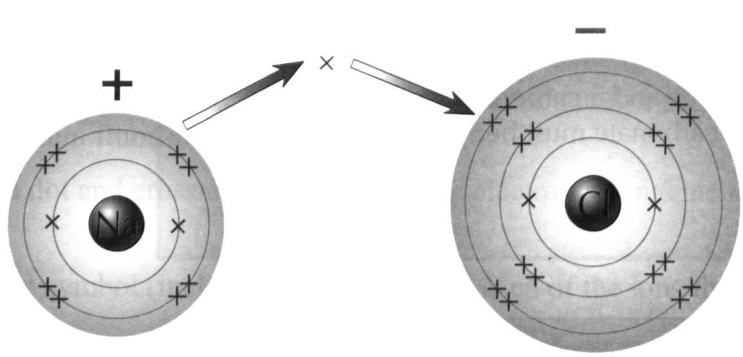
Atoms in Groups 1, 2, 6 and 7 always form ions with the same charges.
You need to know what these are.

Groups 1 & 2 and 6 & 7 are the most likely to form ions

1) Remember, atoms that have lost or gained an electron (or electrons) are ions.
2) The elements that most readily form ions are those in Groups 1, 2, 6 and 7.
3) Group 1 and 2 elements are metals and they lose electrons to form +ve ions or cations.
4) Group 6 and 7 elements are non-metals. They gain electrons to form –ve ions or anions.
5) Make sure you know these easy ones:

CATIONS		ANIONS	
Group 1	Group 2	Group 6	Group 7
Li^+ Na^+ K^+	Be^{2+} Mg^{2+} Ca^{2+}	O^{2-}	F^- Cl^- Br^-

6) When any of the above cations react with the anions, they form ionic bonds.
7) Only elements at opposite sides of the periodic table will form ionic bonds, e.g. Na and Cl, where one of them becomes a cation (+ve) and one becomes an anion (–ve).



Remember, the + and – charges, e.g. Na^+ for sodium, just tell you what type of ion the atom WILL FORM in a chemical reaction. In sodium metal there are only neutral sodium atoms, Na.

The Na^+ ions will only appear if the sodium metal reacts with something like water or chlorine.

A useful way of representing ions is to give the ion's name, then its electron configuration and the charge on the ion. For example, the electronic structure of the sodium ion Na^+ can be represented by Na $[2, 8]^+$. That's the electron configuration followed by the charge on the ion. Simple enough.

Ions and Formulas

You need to be able to write down the right chemical formulas for ionic compounds.

You need to know these chemical formulas

Knowing the chemical formulas for ionic compounds means you have to learn the stuff in the table below, and know how to use it.

The main thing to remember is that in compounds the total charge must always add up to zero.

POSITIVE IONS						NEGATIVE IONS	
Lithium	Li^{+}	Barium	Ba^{2+}	Zinc	Zn^{2+}	Chloride	Cl^{-}
Sodium	Na^{+}	Magnesium	Mg^{2+}	Manganese(II)	Mn^{2+}	Hydroxide	OH^{-}
Potassium	K^{+}	Iron(II)	Fe^{2+}	Aluminium	Al^{3+}	Oxide	O^{2-}
		Copper(II)	Cu^{2+}	Iron(III)	Fe^{3+}	Carbonate	CO_3^{2-}

1) Some metals (for example, copper, iron and manganese) can form different ions with different charges.
2) The number in brackets after the name tells you the size of the positive charge on the ion.
3) If you ever see them in compounds written without a number, assume 'manganese' is manganese(II) and 'copper' is copper(II).

EXAMPLE: Find the formula for zinc carbonate

A zinc ion has a +2 charge and a carbonate ion has a –2 charge. So the formula of zinc carbonate must be:

$ZnCO_3$

EXAMPLE: Find the formula for aluminium oxide

An aluminium ion is Al^{3+} and an oxide ion is O^{2-}. To balance the total charge you need two aluminium ions to every three oxide ions:

Al_2O_3

Don't forget, practice makes perfect

Learn which atoms will form 1^+, 1^-, 2^+ and 2^- ions, and why. Then have a go at these:

1) What ions will each of these elements form? Write out their electron configurations:
   a) potassium, b) aluminium, c) beryllium, d) sulfur, e) fluorine.
   (Use a periodic table. Answers on page 256.)

Warm-Up and Exam Questions

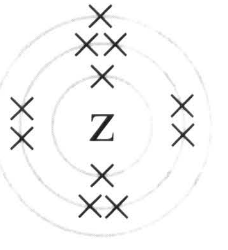
Without a good warm-up you're likely to strain a brain cell or two. So take the time to run through these simple questions before tackling the exam questions.

Warm-Up Questions

1) Magnesium is a metal found in Group 2 of the periodic table. How many electrons does it have in its outer shell?
2) How many electrons does it take to fill the first electron shell of an atom?
3) Sodium chloride has a giant ionic structure. Does it have a high or a low boiling point?
4) Do elements from Group 1 form cations or anions?
5) What is the formula of the compound containing Al^{3+} and OH^- ions only?

Exam Questions

1 The electron arrangement of an element, Z, is shown in the diagram:



(a) What is the atomic number of element Z?

(1 mark)

(b) Which group does Z belong to in the periodic table? Explain how you can tell.

(2 marks)

(c) How many electrons does an atom of Z need to lose so that it has a full outer shell?

(1 mark)

2 Magnesium (atomic number 12) and fluorine (atomic number 9) combine vigorously to form magnesium fluoride, an ionic compound.

(a) Draw a dot and cross diagram to show the electron arrangement of each atom.

(2 marks)

(b) Give the symbol (including the charge) for each of the ions formed.

(2 marks)

(c) Using your answer to (b), work out the formula of magnesium fluoride.

(1 mark)

(d) Once formed, explain why the ions remain together in a compound.

(1 mark)

(e) Magnesium fluoride has a giant ionic structure. Explain why:

(i) it doesn't melt easily.

(2 marks)

(ii) it conducts electricity when molten.

(1 mark)

Exam Questions

3 Iron forms two different compounds with oxygen, iron(II) oxide and iron(III) oxide.

(a) Give the charges on the two iron ions in these compounds.

(1 mark)

(b) What is the chemical formula of each compound?

(2 marks)

(c) Name two other elements that can form ions with different charges.

(1 mark)

4 When lithium reacts with oxygen it forms an ionic compound, Li_2O.

(a) Name the compound formed.

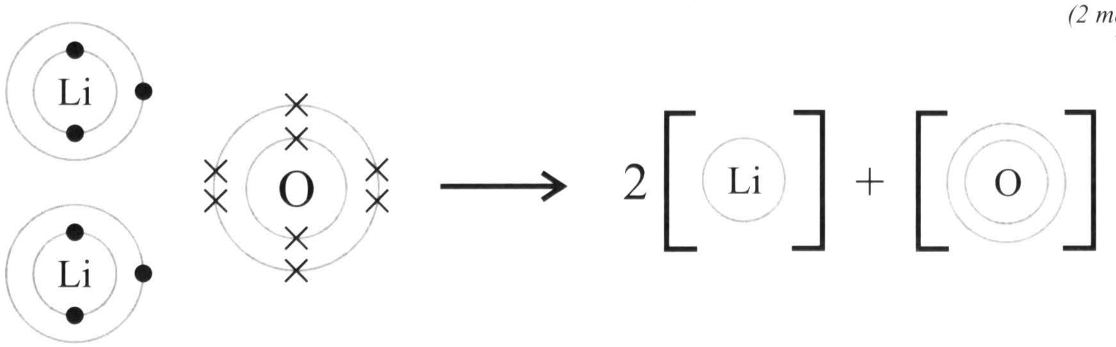
(1 mark)

(b) (i) Complete the diagram below using arrows to show how the electrons are transferred when Li_2O is formed.

(1 mark)

(ii) Show the electron arrangements and the charges on the ions formed.

(2 marks)



5 Potassium and chlorine react to form potassium chloride.

(a) Complete the following table.

	Potassium atom, K	Potassium ion, K^+	Chlorine atom, Cl	Chloride ion, Cl^-
Number of electrons	19			
Electron arrangement	2, 8, 8, 1			

(3 marks)

(b) Draw a dot and cross diagram to show the formation of potassium chloride.

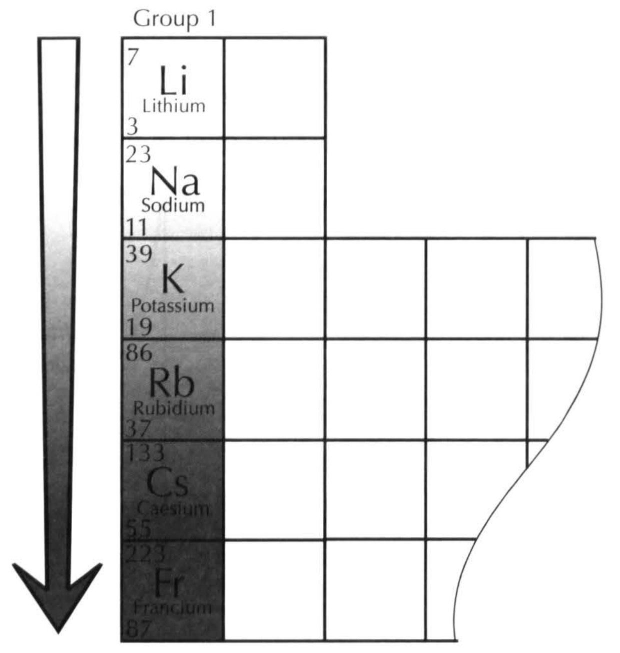
(2 marks)

Group 1 — Alkali Metals

Time to start getting to know a few of these periodic table groups a little better.
First up, alkali metals.

Group 1 metals are known as the 'alkali metals'

Group 1 metals include lithium, sodium and potassium... make sure you know those three names really well. They could also ask you about rubidium and caesium.



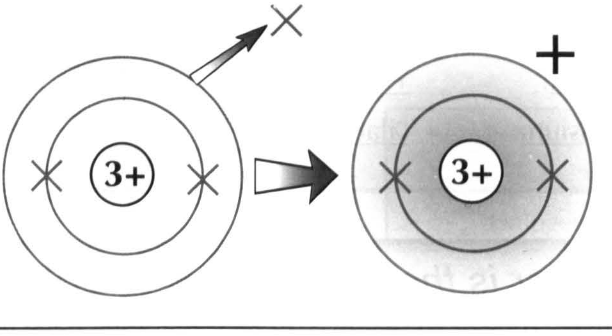
> As you go DOWN Group 1, the alkali metals become more reactive — the outer electron is more easily lost, because it's further from the nucleus and more shielded from it by the inner shells.

1) The alkali metals all have ONE outer electron. This makes them very reactive and gives them all similar properties.
2) They all have the following physical properties:
   - Low melting point and boiling point (compared with other metals),
   - Low density — lithium, sodium and potassium float on water,
   - Very soft — they can be cut with a knife.
3) The alkali metals always form ionic compounds. They are so keen to lose the outer electron that there's no way they'd consider sharing, so covalent bonding (see page 102) is out of the question.

Oxidation is the loss of electrons

1) Group 1 metals are keen to lose an electron to form a 1^+ ion with a stable electronic structure.
2) The more reactive the metal, the happier it is to lose an electron.
3) Loss of electrons is called OXIDATION.

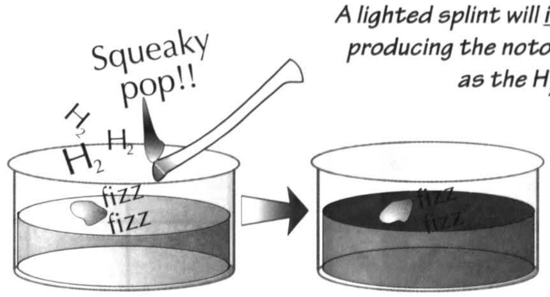
$$Li - e^- \rightarrow Li^+$$



Group 1 — Alkali Metals

Reaction with **cold water** produces **hydrogen gas**

1) When lithium, sodium or potassium are put in water, they react very vigorously.
2) They move around the surface, fizzing furiously.
3) They produce hydrogen.
4) The reactivity with water increases down the group — the reaction with potassium gets hot enough to ignite it.
5) Sodium and potassium melt in the heat of the reaction.
6) They form a hydroxide in solution, i.e. aqueous OH^- ions.



A lighted splint will indicate hydrogen by producing the notorious 'squeaky pop' as the H_2 ignites.

The solution becomes alkaline, which changes the colour of the pH indicator (see page 132) to purple.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Sodium + water → sodium hydroxide + hydrogen

Alkali metal compounds **burn** with **characteristic colours**

1) Dip a wire loop into some hydrochloric acid to clean it.
2) Put the loop into a powdered sample of the compound, then place the end in a blue Bunsen flame.
3) Alkali metal ions give pretty coloured flames — the colour of the flame shows which alkali metal is present:

Lithium:	Red flame
Sodium:	Yellow/orange flame
Potassium:	Lilac flame

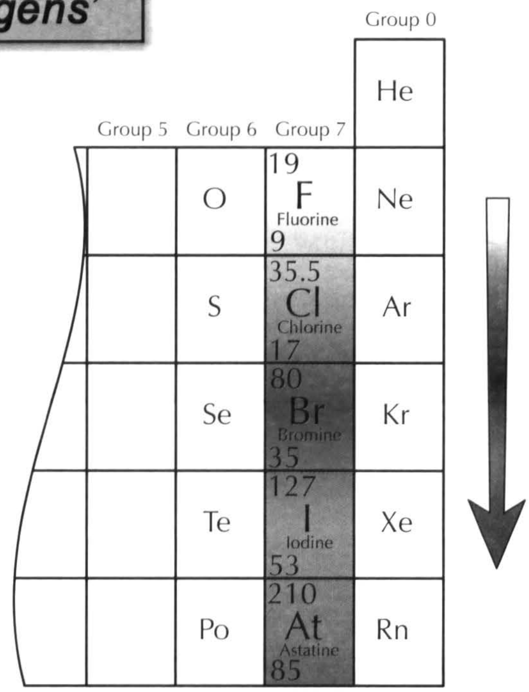
That reaction with water is the reason they're called alkali metals

Alkali metals are really reactive. They're so reactive, in fact, that they have to be stored in oil or they react with the air. Learn the trends and characteristics of alkali metals before carrying on.

Group 7 — Halogens

Next you'll be meeting the halogens. Besides appearing on this page, these little blighters also crop up on pages in this section about bonding. They can form covalent bonds (see pages 102 and 103) as well as ionic bonds (see pages 91 and 92).

Group 7 elements are known as the 'halogens'



Group 7 is made up of fluorine, chlorine, bromine, iodine and astatine.

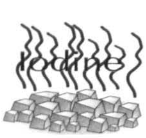
All Group 7 elements have seven electrons in their outer shells — so they've all got similar properties.

As you go DOWN Group 7, the halogens become less reactive — there's less to pull the extra electron in to fill the outer shell when it's further out and more shielded from the positive nucleus.

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

Bromine is a dense, poisonous, orange liquid.

Iodine is a dark grey crystalline solid.



Reduction is the gain of electrons

1) Halogens are keen to gain an electron to form a 1^- ion with a stable electronic structure.
2) The more reactive the halogen, the happier it is to gain an electron.
3) Gain of electrons is called REDUCTION.

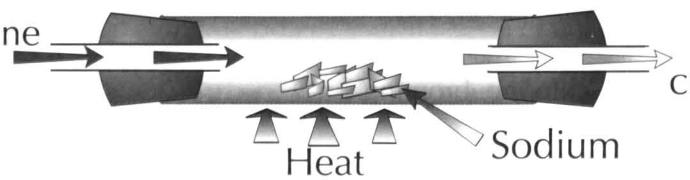
$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

Halogen molecule — Halide ion

Group 7 — Halogens

*The halogens react with **alkali metals** to form **salts***

They react vigorously with alkali metals to form salts called 'metal halides'.



$2Na + Cl_2 \rightarrow 2NaCl$

sodium + chlorine → sodium chloride

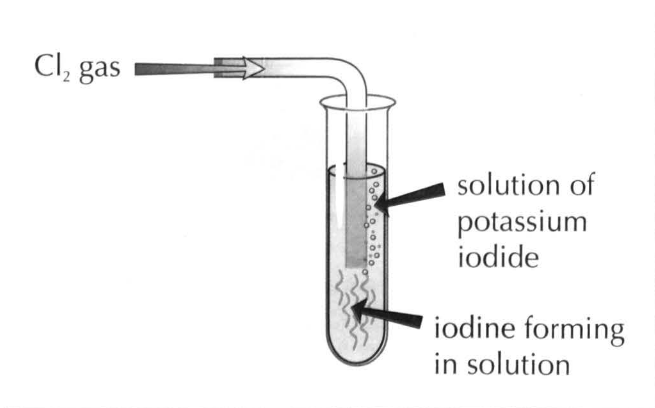
$2K + Br_2 \rightarrow 2KBr$

potassium + bromine → potassium bromide

*More reactive halogens will **displace** less reactive ones*

Chlorine can displace bromine and iodine from a solution of bromide or iodide.

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



$Cl_2 + 2KI \rightarrow I_2 + 2KCl$

chlorine + potassium iodide → iodine + potassium chloride

$Cl_2 + 2KBr \rightarrow Br_2 + 2KCl$

chlorine + potassium bromide → bromine + potassium chloride

Halogens — one electron short of a full shell

The halogens are another group from the periodic table, and just like the alkali metals (p.97-98) you've got to learn their trends and the equations on this page. Learn them, cover up the page, scribble, check.

Warm-Up and Exam Questions

These questions are all about the groups of the periodic table that you need to know about. Treat the exam questions like the real thing — don't look back through the book until you've finished.

Warm-Up Questions

1) In Group 1, as you go down the periodic table, does the reactivity increase or decrease?
2) In Group 7, what is the trend in physical state as you go down the group?
3) Which gas is produced when an alkali metal reacts with water?
4) Give an example of a salt produced when a Group 1 metal reacts with a Group 7 element.

Exam Questions

1 Match the words labelled **A**, **B**, **C** and **D**, with the numbers **1 - 4** in the sentences below.

A electron

B rate

C reaction

D hydrogen

All the alkali metals have a vigorous ...**1**... with water.
This is because they all have an outer ...**2**... which is easily lost.
Each metal reacts at a different ...**3**....
...**4**... is always produced in these reactions.

(4 marks)

2 The table shows some of the physical properties of four of the halogens.

Halogen	Properties			
	Atomic number	Colour	Physical state at room temperature	Boiling point
Fluorine	9	yellow		–188 °C
Chlorine	17	green		–34 °C
Bromine	35	red-brown		59 °C
Iodine	53	dark grey		185 °C

(a) Give the physical state at room temperature of all four halogens.

(4 marks)

(b) Draw an arrow next to the left hand side of the table to show the direction of increasing reactivity in the halogens.

(1 mark)

(c) This equation shows a reaction between chlorine and potassium iodide.

$$Cl_{2(g)} + 2KI_{(aq)} \rightarrow I_{2(aq)} + 2KCl_{(aq)}$$

(i) What type of reaction is this?

(1 mark)

(ii) Which is the less reactive halogen in this reaction?

(1 mark)

Covalent Bonding

Ionic bonding isn't the only way for an atom to get a nice full shell of electrons...

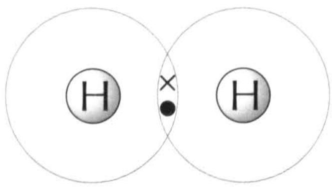
Covalent bonds — *sharing* electrons

1) Sometimes atoms prefer to make covalent bonds by sharing electrons with other atoms.
2) This way both atoms feel that they have a full outer shell, and that makes them happy.
3) Each covalent bond provides one extra shared electron for each atom.
4) Each atom involved has to make enough covalent bonds to fill up its outer shell.
5) Learn these seven important examples:

You only have to draw the outer shell of electrons.

1) Hydrogen, H_2

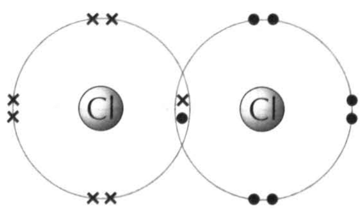
Hydrogen atoms have just one electron. They only need one more to complete their first shell, so they often form single covalent bonds to achieve this.



or **H—H**

2) Chlorine, Cl_2

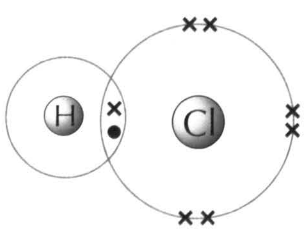
Chlorine atoms also need only one more electron to complete their outer shell, so they also form single covalent bonds.



or **Cl—Cl**

3) 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.



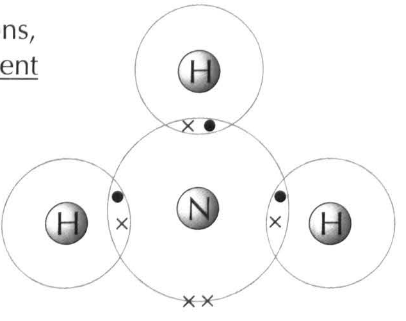
or **H—Cl**

Covalent Bonding

Four more covalent bonding examples to learn on this page.

4) Ammonia, NH_3

Nitrogen has five outer electrons, so it needs to form three covalent bonds to make up the extra three electrons needed.



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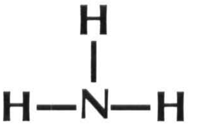
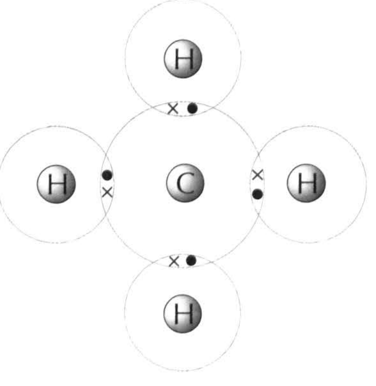


5) Methane, CH_4

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

To become a 4^+ or a 4^- ion is hard work, so it forms four covalent bonds to make up its outer shell.



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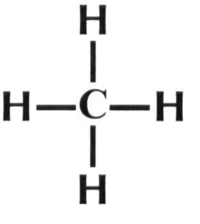
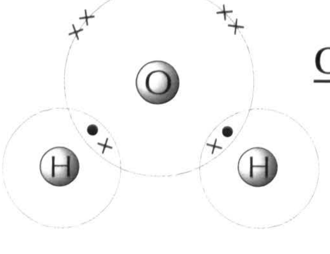


6) Water, H_2O

7) Oxygen, O_2

Oxygen atoms have six outer electrons. They sometimes form ionic bonds by taking two electrons to complete their outer shell.

However, they'll also cheerfully form covalent bonds and share two electrons instead. In water molecules, the oxygen shares electrons with the H atoms and in oxygen gas it shares with another oxygen atom.



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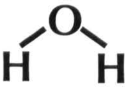
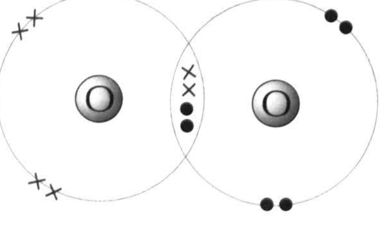




or

O=O

Covalent bonding involves sharing rather than giving electrons

Make sure you learn these seven really basic examples and why they work. Every atom wants a full outer shell, and they can get that either by becoming an ion (p.93–94) or by sharing electrons. Once you understand that, you should be able to apply it to any example they give you in the exam.

Giant Covalent Structures

Substances formed from covalent bonds can either be simple molecules (see p.106) or giant structures. The next two pages are all about the giant structures, and there are three examples to learn.

Giant covalent structures

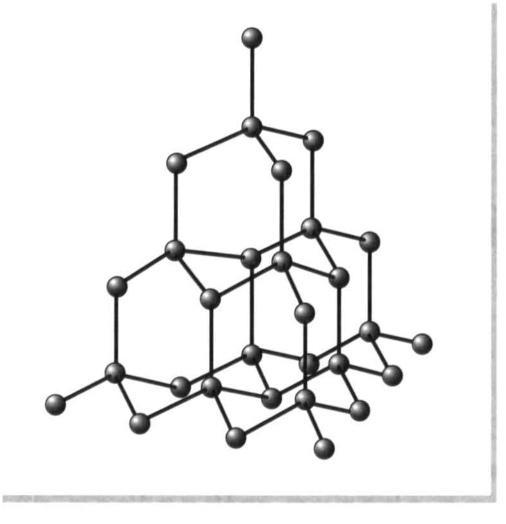
1) These are similar to giant ionic structures except that there are no charged ions.
2) All the atoms are bonded to each other by strong covalent bonds.
3) They have very high melting and boiling points.
4) They don't conduct electricity — not even when molten (except for graphite that is — see next page).
5) They're usually insoluble in water.
6) Important examples are diamond and graphite, which are both made only from carbon atoms.

Make sure you know these three examples

You can find the other two examples on the next page.

Diamond

1) Diamonds are sparkly, colourless and clear. Ideal for jewellery.
2) Each carbon atom forms four covalent bonds in a very rigid giant covalent structure, which makes diamond the hardest natural substance. This makes diamonds ideal as cutting tools.
3) All those strong covalent bonds give diamond a very high melting point.
4) It doesn't conduct electricity because it has no free electrons.

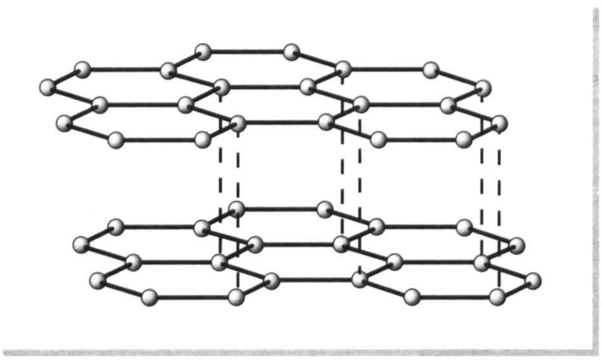


Giant Covalent Structures

Here are the other two examples of giant covalent structures that you need to learn.

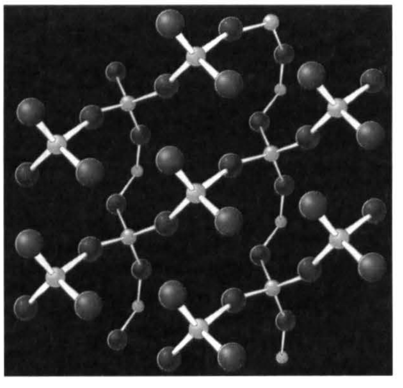
Graphite

1) Graphite is black and opaque, but still kind of shiny.
2) Each carbon atom only forms three covalent bonds, creating sheets of carbon atoms which are free to slide over each other. This makes graphite slippery, so it's useful as a lubricant.
3) The layers are held together so loosely that they can be rubbed off onto paper to leave a black mark — that's how pencils work.
4) Graphite has a high melting point — the covalent bonds need lots of energy before they break.
5) Only three out of each carbon's four outer electrons are used in bonds, so there are lots of spare electrons. This means graphite conducts electricity — it's used for electrodes. See p.217.



Silicon dioxide (silica)

1) Sometimes called silica, this is what sand is made of.
2) Each grain of sand is one giant structure of silicon and oxygen.
3) Silica can be melted down with sodium carbonate (Na_2CO_3) and limestone ($CaCO_3$) to make glass.



Graphite and diamond contain exactly the same atoms

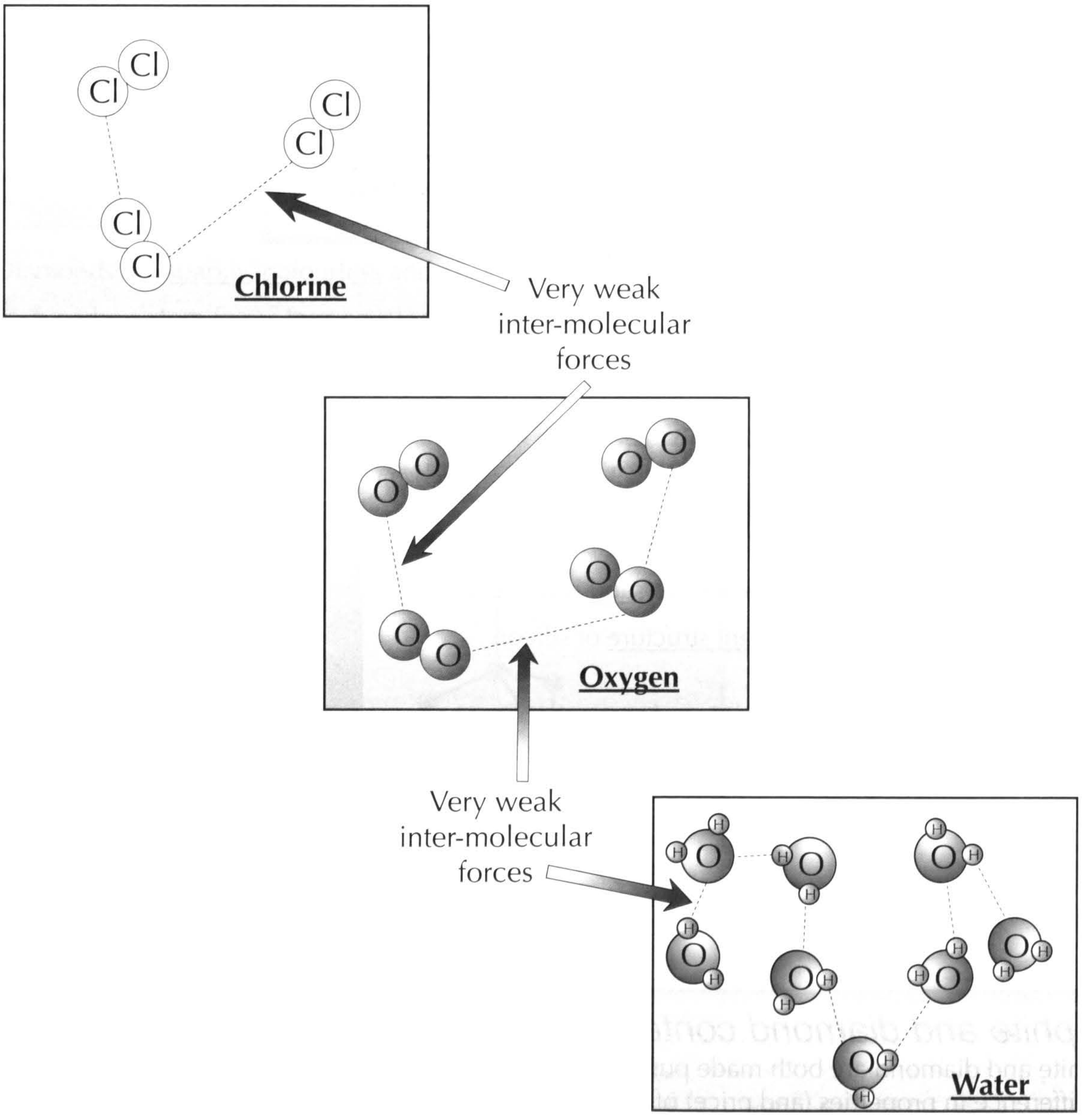
Graphite and diamond are both made purely from carbon — there's no difference at all in their atoms. The difference in properties (and price) of the two substances is all down to the way the atoms are held together. Different structural forms of the same element like this are called allotropes.

Simple Molecular Covalent Structures

Atoms that bond covalently don't all form giant structures.
Some form simple molecular substances.

Simple *molecular* substances

1) The atoms form very strong covalent bonds to form small molecules of two or more atoms.
2) By contrast, the forces of attraction between these molecules are very weak.
3) The result of these feeble inter-molecular forces is that 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) Molecular substances don't conduct electricity, simply because there are no ions.
6) You can usually tell a simple molecular substance just from its physical state, which is always kind of 'mushy' — i.e. liquid or gas or an easily-melted solid.



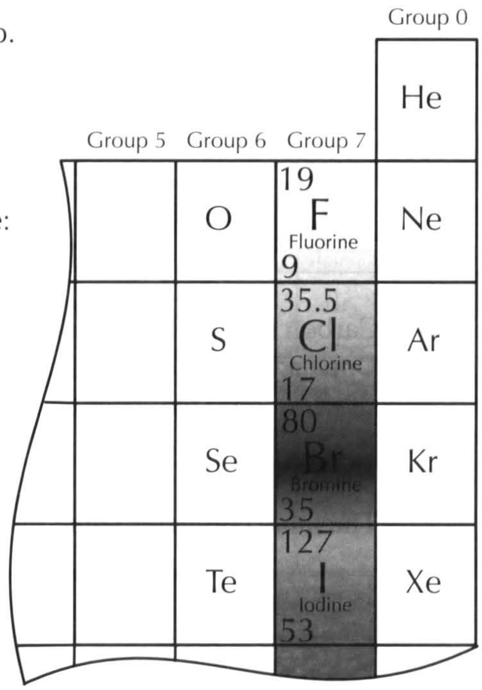
Simple Molecular Covalent Structures

Remember the halogens from pages 99-100? Well, they're all simple molecular covalent molecules. How convenient. Now you can just learn about their properties and the examiners will be happy.

*The **halogens** are **all** simple molecular substances*

1) The physical properties of the halogens change down the group. Particularly, melting point and boiling point both increase (see the table below).
2) This is because of the strength of the inter-molecular forces.
3) The pattern in properties in the table can be explained because:

- The halogens get bigger as you go down the group.
- The bigger the halogen molecule, the stronger the inter-molecular forces of attraction.
- The stronger the forces, the more energy it takes to separate the molecules, and so the higher their melting points and boiling points.



	Properties				
Group 7 Elements	Atomic number	Colour	Physical state at room temperature	Melting point	Boiling point
Fluorine	9	yellow	gas	–220 °C	–188 °C
Chlorine	17	green	gas	–102 °C	–34 °C
Bromine	35	red-brown	liquid	–7 °C	59 °C
Iodine	53	dark grey	solid	114 °C	184 °C

4) The little molecules of fluorine and chlorine have the weakest attraction to each other, so it takes very little energy to break them apart. That makes them gases at room temperature.
5) The bigger molecules of bromine have stronger attractions, so it's a liquid at room temperature. Iodine's molecules are larger still, its inter-molecular forces are the strongest of the four, and it's a solid.

Fluorine and chlorine are gases because of the weak forces

So, simple molecules are held together by weedy, pathetic inter-molecular forces. But these forces get gradually less weedy as the molecules get bigger. That's why the halogens get more solid down the group. It's also why the hydrocarbons (see page 32) get less runny as they get bigger.

Warm-Up and Exam Questions

Warm-Up Questions

1) How is covalent bonding different from ionic bonding?
2) Some molecules, e.g. oxygen (O_2), contain double bonds. Explain what this means.
3) Describe two differences in the physical properties of diamond and graphite.
4) Give another example of a substance that has a giant covalent structure.
5) Why does chlorine have a very low boiling point?

Exam Questions

1 Carbon dioxide is a covalently bonded molecule with the formula CO_2. Draw a dot and cross diagram for the carbon dioxide molecule, showing only the outer electrons.

(2 marks)

2 The table compares some physical properties of silicon dioxide, bromine and graphite.

Property	silicon dioxide	bromine	graphite
Melting point (°C)	1610	–7	3657
Electrical conductivity	poor	poor	good
Solubility in water	insoluble	slightly soluble	insoluble

(a) What is the structure and type of bonding in:

(i) silicon dioxide?

(1 mark)

(ii) graphite?

(1 mark)

(iii) bromine?

(1 mark)

(b) Why does silicon dioxide have poor electrical conductivity?

(1 mark)

(c) Why does graphite have good electrical conductivity?

(1 mark)

(d) Bromine is a liquid at room temperature (20 °C).

(i) Explain why bromine has such a low melting point compared with silicon dioxide and graphite.

(2 marks)

(ii) Explain why iodine, I_2, has a higher melting point than bromine.

(2 marks)

Group 0 — Noble Gases

The noble gases — stuffed full of every honourable virtue. They don't form covalent bonds or ionic bonds, 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) The noble gases were only discovered just over 100 years ago — it took so long because they have properties that make them hard to observe...
3) All elements in Group 0 are colourless gases at room temperature.
4) They're also more or less inert — they don't react with much at all. They don't bond with anything, not even with each other, and so they just wander about as single atoms. This is because they have a full outer shell, so they don't want to give up or gain or share electrons.
5) Luckily the noble gases all have a dead handy property that lets you see them — they each give out light if you pass an electric current through them. Each noble gas gives out a particular colour of light.

Group 6	Group 7	Group 0
		4 He Helium 2
O	F	20 Ne Neon 10
S	Cl	40 Ar Argon 18
Se	Br	84 Kr Krypton 36
Te	I	131 Xe Xenon 54
Po	At	222 Rn Radon 86

The noble gases have many *everyday uses*

Neon is used in *electrical discharge tubes*

Neon lights are used in tacky shop signs — the kind you'd expect to see if you visited Las Vegas. They don't use much current so they're cheap to run, and they give out a bright red light.

Noble gases are used in *lasers* too

There's the famous red helium-neon laser and the more powerful argon laser.

Helium is used in *airships* and *party balloons*

Helium has a lower density than air — so it makes balloons float. And it's a lot safer to use than hydrogen (the famous airship Hindenburg was filled with hydrogen and caught fire).

Argon is used in *filament lamps* (light bulbs)

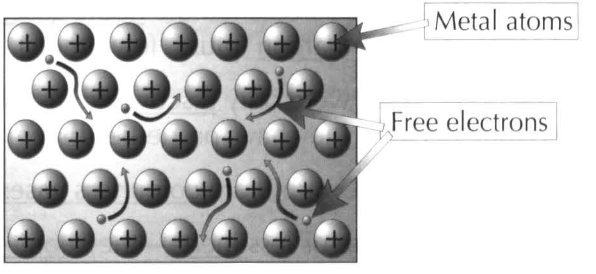
It provides an inert atmosphere which stops the very hot filament from burning away.

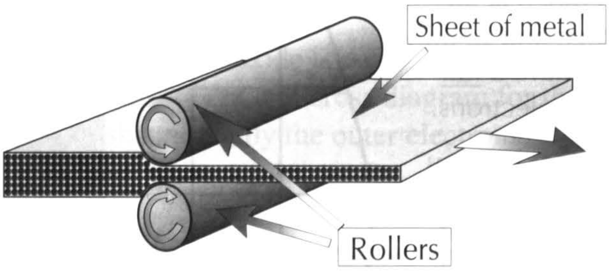
Metallic Structures

So you know about covalent and ionic bonds (plus the noble gases which don't bond at all). But there's another important type of bond in the periodic table — you may remember it from Section Two.

Metal properties are all due to the sea of free electrons

1) Metals also consist of a giant structure. It's held together with metallic bonds and involves the all-important 'sea' of free electrons (see page 24) which gives metals their properties — for example, conducting heat and electricity.





2) These electrons also hold the atoms together in a regular structure.

3) In addition, they allow the atoms to slide over each other, causing metals to be malleable.

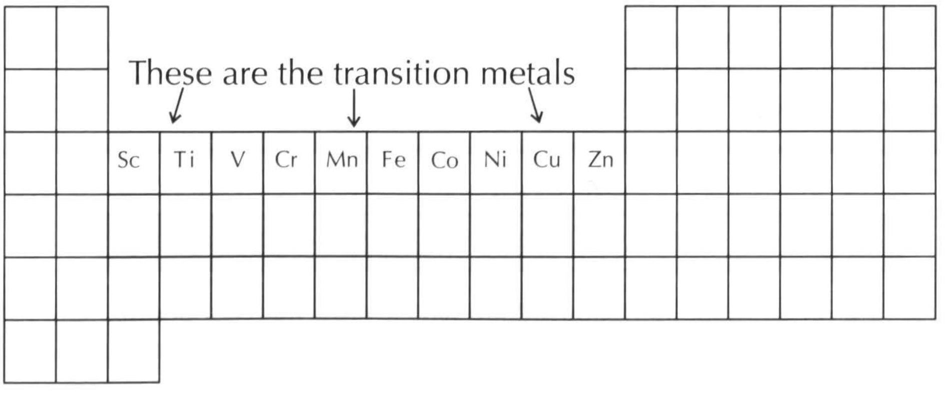
Metals in the middle of the periodic table are transition metals

A lot of everyday metals are transition metals (e.g. copper, iron, zinc, gold, silver, platinum) — but there are loads of others as well. Transition metals have typical 'metallic' properties.

Transition metals and their compounds make good catalysts:

1) Iron is the catalyst used in the Haber process for making ammonia (see page 167).

2) Nickel is useful for the hydrogenation of alkenes, e.g. to make margarine (see page 53).



Metallic Structures

If you make them cold enough, metals can start behaving in a pretty odd way.

At very low temperatures, some metals are superconductors

1) Normally, all metals have some electrical resistance, even really good conductors like copper.

2) That resistance means that whenever electricity flows through them, they heat up, and some of the electrical energy is wasted as heat.

3) If you make some metals cold enough, though, their resistance disappears completely. The metal becomes a superconductor.

4) Without any resistance, no energy is turned into heat, so none of it's wasted. That means you could start a current flowing through a superconducting circuit, take out the battery, and the current would carry on flowing forever.

Superconductors have the potential to be really useful

1) Using superconducting wires you can make:

   a) Power cables that transmit electricity without any loss of power.

   b) Really strong electromagnets that don't need a constant power source.

   c) Electronic circuits that work really fast, because there's no resistance to slow them down.

2) But here's the catch — metals only start superconducting at temperatures less than –265 °C. Getting things that cold is very hard, and very expensive.

3) Scientists are trying to develop room temperature superconductors. So far, they've managed to get some weird metal oxide things to superconduct at about –135 °C, but they've still got a long way to go.

Actually using superconductors isn't all that practical yet

Notice that although metals all share a similar structure and conduct well, they still behave in radically different ways. Take the alkali metals of Group 1 — they react vigorously with water, but transition metals don't do this. Lucky really, or the Statue of Liberty would disappear in the rain.

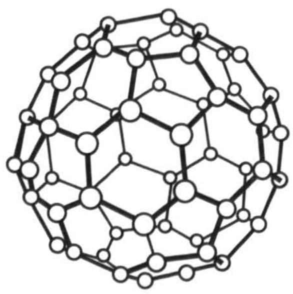
Nanomaterials

Hmm, nanomaterials. Very useful but, let's face it, pretty bizarre.

Nanomaterials are really really really really tiny

...in fact, they're smaller than that.

1) Really tiny particles, 1–100 nanometres across, are called 'nanoparticles' (1 nm = 0.000 000 001 m).
2) Nanoparticles include fullerenes. These are molecules of carbon, shaped like hollow balls or closed tubes. Each carbon atom forms three covalent bonds with its neighbours, leaving free electrons that can conduct electricity.
3) The smallest fullerene is buckminsterfullerene, which has 60 carbon atoms joined in a ball — its molecular formula is C_{60}.



buckminsterfullerene

Fullerenes can be joined together to form nanotubes — teeny tiny hollow carbon tubes, a few nanometres across:

a) All those covalent bonds make carbon nanotubes very strong. They can be used to reinforce graphite in tennis rackets and to make stronger, lighter building materials.
b) Nanotubes conduct electricity, so they can be used in tiny electric circuits for computer chips.

4) Nanoparticles like these have some other very useful properties:

a) They have a huge surface area, so they could help make great industrial catalysts (see pages 154-155) — individual catalyst molecules could be attached to carbon nanotubes.
b) With nanoparticles, you can build surfaces with very specific properties. That means you can use them to make sensors to detect one type of molecule and nothing else. These highly specific sensors are already being used to test water purity.

Nanomaterials

As with any new technology, there's a lot of muttering and worrying going on over nanomaterials. Which is probably a good thing. It's best to make sure new inventions are safe before you start unleashing them on the world. If only there had been a bit more concern about CFCs. Or DDT. Or...

Nanomaterials are becoming more and more widely used

Here are some other examples of useful nanomaterials:

- Nanoparticles of silver can stop viruses entering cells. Ordinary silver doesn't fight viruses.

- Nanoparticles of titanium dioxide and zinc oxide are used in some sunscreens. They reflect UV radiation from the Sun, but not visible light, so you can't see them. This reduces the need for larger particles which make the sunscreen thick and white, leaving white marks on the skin.

Some people are concerned about nanotechnology

1) Nanoparticles can be made by molecular engineering, but it's really hard. Molecular engineering is building a product molecule-by-molecule to a specific design — either by positioning each molecule exactly where you want it or by starting with a bigger structure and taking bits off it.
2) As this technology improves, it's possible that nanoparticles and nanotubes could one day be joined to make tiny nanomachines. These could make products molecule by molecule, very quickly and cheaply.
3) When this idea first appeared, there were scare stories about self-replicating nano-robots. People thought that self-replicating machines would keep making copies of themselves until they had used up all the world's resources. There were media articles about the world being covered with a 'grey goo' of nano-robots.
4) The scare stories about grey goo aren't really justified, though. Nanomachines and nanofactories that would be used to manufacture products aren't the same as self-replicating nano-robots. They'd only make what they were supposed to make — not copies of themselves.

News stories about nanotechnology (and other new technology too) often sensationalise it and make it seem further advanced and a lot scarier than perhaps it is in reality. They often take a fairly extreme view about what might happen in the future (either really optimistic or really pessimistic) — rather than looking at what's most likely to happen.

That doesn't mean that all nanotechnology is definitely safe though

There's lots of talk about using nanotechnology in foods — scientists are coming up with some bizarre Willy Wonka-style ideas. Which would be pretty cool, but some people are afraid of the particles acting strange and ending up where they shouldn't be (like in the brain) causing damage. Gulp.

Warm-Up and Exam Questions

There were some fairly strange bits to learn on the last few pages — superconductors and nanomaterials seem to belong on the Sci-Fi Channel rather than in a Chemistry book.
If you're feeling bemused, try these questions to make sure you won't be thrown in the exam.

Warm-Up Questions

1) Why did it take so long to discover the noble gases?
2) Give one use for helium and one use for neon.
3) What name is given to the block of metals in the middle of the periodic table?
4) Explain how the structure of metals allows them to conduct electricity.
5) What is the range of particle sizes in nanomaterials?

Exam Questions

1 Argon is a noble gas used to fill light bulbs. It has an atomic number of 18.

(a) What is the electron configuration of argon?

(1 mark)

(b) Argon atoms don't pair up to form molecules like many other gaseous elements do. Explain why not.

(2 marks)

(c) Explain how argon extends the life of a light bulb.

(2 marks)

2 The critical temperature (T_c) at which some materials become superconducting is shown in the table.

(a) What does 'superconducting' mean?

(2 marks)

(b) Which of the materials in the table is most likely to be useful in real-life applications of superconductors? Explain your answer.

(1 mark)

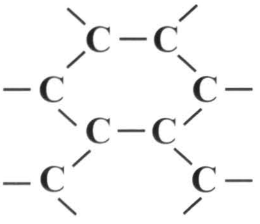
Type	T_c (°C)
zinc	–272
aluminum	–272
tin	–269
mercury	–269
metal oxide ceramic 1	–145
metal oxide ceramic 2	–123

(c) Give two potential applications of superconducting materials.

(2 marks)

3 Carbon nanotubes have a tube-like structure based on the covalent bonding between carbon atoms, as shown in the diagram. They can be used to make tiny electrical circuits.

Consider the number of bonds carbon normally forms, and use this to help explain why nanotubes conduct electricity.



(3 marks)

Revision Summary for Section Six

Hopefully you've committed everything from this section to memory now, and will have no trouble with it if it comes up in the exam. But there's only one way to check if you really know it, and that's to try some revision questions... like the ones on this page, for example.

1) What feature of atoms determines the order of the modern periodic table?
2) What feature of an atom determines the properties it will have?
3) How many electrons can occupy the third shell in an atom?
4) Give the electron configurations of potassium and sulfur.
(Use the periodic table at the front of the book.)
5) Draw a diagram to show the ionic bonding in sodium chloride.
6) List the main properties of ionic compounds.
7) Give the formula of an oxygen ion. Is this a cation or an anion?
8) Show the electronic structure of a magnesium ion using brackets.
9) Find the chemical formula for potassium hydroxide.
10) Do alkali metals form ionic or covalent compounds?
11) Write down the word and symbol equations for the reaction of lithium with water.
12) Which is more reactive, chlorine or iodine?
13) Write down the word and symbol equations for the reaction of chlorine with potassium.
14) Write down the word and symbol equations for the reaction of chlorine with potassium iodide.
15) What is covalent bonding?
16) Sketch dot and cross diagrams to show the bonding in molecules of:
17) a) hydrogen, b) hydrogen chloride, c) water, d)* carbon dioxide.
18) What type of structure does a molecule of silicon dioxide have?
19) Industrial diamonds are used in drill tips and precision cutting tools. What property of diamond makes it suitable for this use? Explain how the bonding in diamond gives it this physical property.
20) Graphite conducts electricity, which is unusual for a non-metal. Explain why it has this property.
21) Is a substance with a simple molecular structure more likely to be a solid or a gas? Why?
22) Explain why simple molecular structures don't conduct electricity.
23) Describe and explain the trend in boiling points as you go down Group 7.
24) Explain in terms of electrons why the noble gases are unreactive.
25) Give one example of a transition metal being used as a catalyst.
26)*Identify the type of bonding in each of the substances in the table:

Substance	Melting point (°C)	Electrical conductivity	Hardness on a scale of 0 – 10 (10 being diamond)
A	3410	Very high	7.5
B	2072	Zero	9
C	605	Zero in solid form High when molten	Low

27) How can you make some metals become superconductors?
28) What advantage is there in making power cables from superconductors?
29) Why don't they make the entire National Grid from superconducting metal wires?
30) What are nanoparticles? Describe two different applications of nanoparticles.
31) What is the smallest type of fullerene called? What element is it made from?

* Answers on page 257.

Relative Formula Mass

The biggest trouble with relative atomic mass and relative formula mass is that they sound so blood-curdling. Take a few deep breaths and just enjoy, as the mists slowly clear...

Relative atomic mass, A_r, is dead easy

1) This is just a way of saying how heavy different atoms are compared with the mass of an atom of carbon-12. So carbon-12 has an A_r of exactly 12.
2) It turns out that the relative atomic mass A_r is nothing more than the mass number of the element (to the nearest whole number).
3) In the periodic table, the elements all have two numbers. The smaller one is the atomic number (how many protons it has). But the bigger one is the mass number (how many protons and neutrons it has), which is also the relative atomic mass. See? Dead easy.

mass number, which is also relative atomic mass

$^{4}_{2}He$ $\quad$ $^{12}_{6}C$ $\quad$ $^{35.5}_{17}Cl$

Helium has $A_r = 4$. Carbon has $A_r = 12$. Chlorine has $A_r = 35.5$.

You may have noticed that the relative atomic mass of chlorine isn't a whole number like the others. That's because it has more than one stable isotope — it's all explained on p.117-118.

Relative formula mass, M_r, is also dead easy

If you have a compound like $MgCl_2$ then it has a relative formula mass, M_r, which is just all the relative atomic masses added together.

For $MgCl_2$ it would be:

$$MgCl_2: \quad 24 + (35.5 \times 2) = 95$$

So M_r for $MgCl_2$ is simply 95.

You can easily get A_r for any element from the periodic table (see inside front cover), but in a lot of questions they give you them anyway. Since it's nearly Christmas I'll run through another example:

Example:

Find the relative formula mass for calcium carbonate, $CaCO_3$, using the given data:
A_r for Ca = 40 $\quad$ A_r for C = 12 $\quad$ A_r for O = 16

ANSWER:

$$CaCO_3: \quad 40 + 12 + (16 \times 3) = 100$$

So the M_r for $CaCO_3$ is 100.

And that's it. Big fancy name like relative formula mass and all it means is "add up the mass numbers".

Isotopes

Some elements have more than one isotope.

Isotopes are the same except for an extra neutron or two

A favourite trick exam question: "Explain what is meant by the term isotope".

The trick is that it's impossible to explain what one isotope is. Nice of them that, isn't it. You have to outsmart them and always start your answer "Isotopes are..."

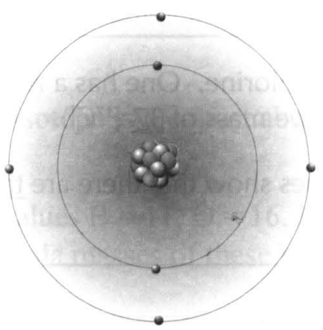
LEARN the definition:

> Isotopes are: different atomic forms of the same element, which have the SAME number of PROTONS but DIFFERENT numbers of NEUTRONS.

1) The upshot is: isotopes must have the same proton number but different mass numbers.
2) If they had different proton numbers, they'd be different elements altogether.
3) A very popular pair of isotopes are carbon-12 and carbon-14, used for carbon dating.

See page 6 for more about atomic structure.

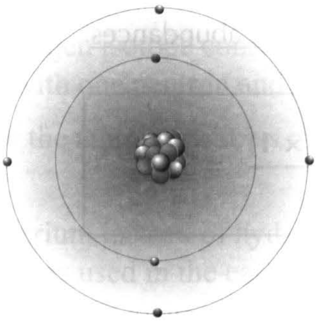
Carbon-12



$^{12}_{6}C$

6 PROTONS
6 ELECTRONS
6 NEUTRONS

Carbon-14



$^{14}_{6}C$

6 PROTONS
6 ELECTRONS
8 NEUTRONS

The number of electrons decides the chemistry of the element. If the atomic number (that is, the number of protons) is the same, then the number of electrons must be the same, so the chemistry is the same. The different number of neutrons in the nucleus doesn't affect the chemical behaviour at all.

Relative Atomic Mass

Remember the way chlorine had a relative atomic mass that wasn't a whole number? Well, this page explains why. It's unexpectedly straightforward, you'll be pleased to hear.

Relative atomic mass takes all stable isotopes into account

1) Relative atomic mass (A_r) uses the average mass of all the isotopes of an element.
2) It has to allow for the relative mass of each isotope and its relative abundance.
3) Relative abundance just means how much there is of each isotope compared to the total amount of the element in the world.
4) This can be a ratio, a fraction or a percentage and is easiest to see with an example:

element	relative mass of isotope	relative abundance
chlorine	35	3
	37	1

This means that there are two isotopes of chlorine. One has a relative mass of 35 (^{35}Cl) and the other has a relative mass of 37 (^{37}Cl).

The relative abundances of the two isotopes show that there are three atoms of ^{35}Cl for every one atom of ^{37}Cl.

1) First, multiply the mass of each isotope by its relative abundance.
2) Add those numbers together.
3) Divide by the sum of the relative abundances:

$$A_r = \frac{(35 \times 3) + (37 \times 1)}{3 + 1} = \underline{35.5}$$

Relative atomic masses don't usually come out as whole numbers or easy decimals, but they're often rounded to the nearest 0.5 in periodic tables (see page 116).

Don't forget, begin any definition of an isotope with "isotopes are..."

Some isotopes are unstable. That means they don't stay as they are forever, but change (decay) into other elements. When they do this, they release nuclear radiation.

Warm-Up and Exam Questions

Those pages were probably a lot more straightforward than you were expecting, but just to be on the safe side, try these questions to make sure you've really got the hang of A_r, M_r and isotopes.

Warm-Up Questions

What name is given to:

1) ... atoms of the same element with different mass numbers?
2) ... the sum of the number of protons and neutrons in an atom?
3) ... the average mass of the isotopes of an element?
4) ... the sum of the relative atomic masses of the atoms in a molecule?
5) ... the neutral particle in the nucleus?

Exam Questions

1 (a) Boron has two main isotopes, $^{11}_{5}B$ and $^{10}_{5}B$. Its A_r value is 10.8.

(i) What does A_r stand for?

(1 mark)

(ii) What is the difference between the two boron isotopes?

(1 mark)

(iii) Which isotope is the most abundant? Explain your reasoning.

(2 marks)

(b) Use the A_r values B = 11, O = 16, F = 19 and H = 1 to calculate the relative formula masses of these boron compounds:

(i) BF_3

(1 mark)

(ii) $B(OH)_3$.

(1 mark)

2 The nucleus of the most common hydrogen isotope consists of a single proton. Deuterium is an isotope of hydrogen with one neutron and one proton in its nucleus.

(a) Draw a labelled diagram to show the atomic particles in a deuterium atom.

(2 marks)

(b) Water molecules containing deuterium instead of hydrogen are written as D_2O. 'Heavy water', as D_2O is known, was used in the early development of nuclear weapons during World War II.

(i) What is the relative formula mass of deuterium oxide, D_2O? (A_r of oxygen = 16.)

(1 mark)

(ii) Why do you think D_2O is known as 'heavy water'?

(1 mark)

Percentage Mass

Although relative atomic mass and relative formula mass are easy enough, it can get just a tad trickier when you start getting into other calculations which use them. It depends on how good your maths is basically, because it's all to do with ratios and percentages.

Calculating % mass of an element in a compound

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

$$\text{Percentage mass of an element in a compound} = \frac{A_r \times \text{no. of atoms (of that element)}}{M_r \text{ (of whole compound)}} \times 100$$

If you don't learn the formula then you'd better be pretty smart — or you'll struggle.

Here's an example of how you'd use it:

Example:

Find the percentage mass of nitrogen in ammonium sulfate fertiliser, $(NH_4)_2SO_4$, using the following:

- A_r for H = 1
- A_r for O = 16
- A_r for N = 14
- A_r for S = 32

ANSWER: M_r of $(NH_4)_2SO_4$ $= 2 \times [14 + (1 \times 4)] + 32 + (16 \times 4)$

$= 132$

Now use the formula:

$$\textbf{Percentage mass} = \frac{A_r \times n}{M_r} \times 100$$

$$= \frac{14 \times 2}{132} \times 100 = \mathbf{21.2\%}$$

So there you have it. Nitrogen represents 21.2% of the mass of ammonium sulfate.

As usual with these calculations, practice makes perfect. You'll find some to do on the bottom of the next page. Don't skip it, you'll be glad you're perfect when it comes to exam day.

Empirical Formulas

*Finding the **empirical formula** (from masses or percentages)*

This also sounds a lot worse than it really is. Try this for a nice simple stepwise method:

1) List all the elements in the compound (there's usually only two or three).
2) Underneath them, write their experimental masses or percentages.
3) Divide each mass or percentage by the A_r for that particular element.
4) Turn the numbers you get into a nice simple ratio by multiplying and/or dividing them by well-chosen numbers.
5) Get the ratio in its simplest form, and that tells you the empirical formula of the compound.

Here's a nice example to give you a better idea of what I'm on about:

Example:

Find the empirical formula of the iron oxide produced when 44.8 g of iron react with 19.2 g of oxygen.
(A_r for iron = 56, A_r for oxygen = 16)

METHOD:

1) List the two elements:	**Fe**	**O**
2) Write in the experimental masses:	44.8	19.2
3) Divide by the A_r for each element:	$44.8/56 = 0.8$	$19.2/16 = 1.2$
4) Multiply by 10...	8	12
...then divide by 4:	2	3

5) So the simplest formula is 2 atoms of Fe to 3 atoms of O, i.e. $\mathbf{Fe_2O_3}$.

You need to realise (for the exam) that this empirical method (i.e. based on experiment) is the only way of finding out the formula of a compound. Rust is iron oxide, sure, but is it FeO, or Fe_2O_3? Only an experiment to determine the empirical formula will tell you for certain.

Don't learn that list of instructions — practise using it (it's much quicker)

Learn the formula on the last page and the five rules in the blue box. Then try these: (Answers on p.257.)
1) Find the percentage mass of oxygen in each of these: a) Fe_2O_3 b) H_2O c) $CaCO_3$ d) H_2SO_4.
2) Find the empirical formula of the compound formed from 2.4 g of carbon and 0.8 g of hydrogen.

Calculating Masses in Reactions

The three important steps — not to be missed...

1) Write out the balanced equation.
2) Work out M_r — just for the two bits you want.
3) Apply the rule: Divide to get one, then multiply to get all.
   (First for the substance they give info about, then for the other one!)

Don't worry — these steps should all make sense when you look at the example below.

Example:

What mass of magnesium oxide is produced when 60 g of magnesium is burned in air?

METHOD:

1) Write out the balanced equation: **$2Mg + O_2 \rightarrow 2MgO$**
2) Work out the relative formula masses: **$2 \times 24 \rightarrow 2 \times (24 + 16)$**
   (don't do the oxygen — you don't need it) **$48 \rightarrow 80$**
3) Apply the rule: Divide to get one, then multiply to get all:

The two numbers, 48 and 80, tell you 48 g of Mg react to give 80 g of MgO.
Here's the tricky bit. You've now got to be able to write this down:

48 g of Mg ...... reacts to give ...... 80g of MgO
1 g of Mg ...... reacts to give ......
60 g of Mg ...... reacts to give ......

The big clue is that they've said you want to burn "60 g of magnesium", i.e. they've told you how much Mg to have, and that's how you know to write down the left-hand side of it first, because:

You'll first need to ÷ by 48 to get 1 g of Mg, and then need to × by 60 to get 60 g of Mg.

Then you can work out the numbers on the other side (shown in blue below) by realising that you must divide both sides by 48 and then multiply both sides by 60.

÷ 48	48 g of Mg	80 g of MgO	**÷ 48**
	1 g of Mg	1.67 g of MgO	
× 60	60 g of Mg	100 g of MgO	**× 60**

The mass of product is called the yield of a reaction. You should realise that in practice you never get 100% of the yield, so the amount of product will be slightly less than calculated (see p.128).

This finally tells us that 60 g of magnesium will produce 100 g of magnesium oxide.

If the question had said, "Find how much magnesium gives 500 g of magnesium oxide", you'd fill in the MgO side first, because that's the one you'd have the information about.

You can't just read these pages — work through the examples too

The only way to get good at these is to practise. So have a go at these questions: (Answers on p.258.)
1) Find the mass of calcium which gives 30 g of calcium oxide (CaO) when burnt in air.
2) What mass of fluorine fully reacts with potassium to make 116 g of potassium fluoride (KF)?

The Mole

The mole can be really confusing. I think it's the word that puts people off. It's very difficult to see the relevance of the word "mole" to different-sized piles of brightly coloured powders.

"THE MOLE" is simply the name given to a certain number

Just like 'a million' is this many: 1 000 000,

or 'a billion' is this many: 1 000 000 000,

so 'a mole' is this many: 602 300 000 000 000 000 000 000 or 6.023×10^{23}.

1) And that's all it is. Just a number. The burning question, of course, is why is it such a silly long one like that, and with a six at the front?
2) The answer is that when you get precisely that number of atoms of carbon-12, it weighs exactly 12 g.
3) So, get that number of atoms or molecules, of any element or compound, and conveniently, they weigh exactly the same number of grams as the relative atomic mass, A_r (or M_r) of the element (or compound).
4) This is arranged on purpose of course, to make things easier.

Here's the definition of a mole written out nicely so you can learn it:

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.

And here are a few more examples to really drive the point home:

Iron has an A_r of 56.	So one mole of iron weighs exactly 56 g
Nitrogen gas, N_2, has an M_r of 28 (2 × 14).	So one mole of N_2 weighs exactly 28 g
Carbon dioxide, CO_2, has an M_r of 44.	So one mole of CO_2 weighs exactly 44 g

This means that 12 g of carbon, or 56 g of iron, or 28 g of N_2, or 44 g of CO_2, all contain the same number of particles, namely one mole or 6×10^{23} atoms or molecules.

The Mole

Of course, we can't leave you just knowing what a mole is. It's no fun without a few formulas...

Here's a formula for finding **the number of moles** in a **given mass**

$$\text{NUMBER OF MOLES} = \frac{\text{mass in g (of element or compound)}}{M_r \text{ (of element or compound)}}$$

Example:

How many moles are there in 42 g of carbon?

ANSWER:

No. of moles = mass (g) ÷ M_r = 42 ÷ 12 = 3.5 moles.

'Relative formula mass' is also 'molar mass'

1) You've been quite happy using the relative formula mass, M_r, all through the calculations so far.
2) In fact, that was already using the idea of moles because M_r is actually the mass of one mole in grams, or as it's sometimes called, the molar mass.

A '1M solution' contains 'one mole per litre'

The 'moles per litre' of a solution is sometimes called its 'molarity'.

This is pretty easy. So a 2 M solution of NaOH contains 2 moles of NaOH per litre of solution.

You need to know how many moles there'll be in a given volume:

NUMBER OF MOLES = volume in litres × moles per litre of solution

Example:

How many moles in 185 cm^3 of a 2 M solution?

ANSWER:

No. of moles = vol (l) × molarity = 0.185 × 2 = 0.37 moles

The M_r or A_r of a substance is the mass of one mole of it in grams

It's possible to do all the calculations on the previous pages without ever talking about moles.
You just concentrate on M_r and A_r instead — M_r and A_r represent the mass of one mole anyway.
Learn both the equations above. They'll make your life more complete (and they're useful in the exam).

Warm-Up and Exam Questions

Warm-Up Questions

1) Write down the definition of a mole.
2) What is the mass of one mole of oxygen gas, O_2?
3) How many moles are there in 4 g of sodium hydroxide? (A_r values: Na = 23, O = 16, H = 1.)
4) What is the mass of 2 moles of carbon dioxide, CO_2? (A_r values: C = 12, O = 16.)
5) What is the concentration of a solution containing 2 moles of a substance per litre?
6) How many moles are there in 500 cm^3 of a 0.5 M solution?

Exam Questions

1 An agricultural scientist needs to compare the amount of nitrogen in each of the three different fertilisers listed below.

Fertiliser	Formula
urea	$CO(NH_2)_2$
potassium nitrate	KNO_3
ammonium nitrate	NH_4NO_3

(A_r values: C = 12, O = 16, N = 14, H = 1, K = 39.)

(a) Work out the percentage mass of nitrogen in each of the three fertilisers.

(6 marks)

(b) Using your answers to part (a), explain which one of the three would you expect to make the best fertiliser.

(2 marks)

2 Analysis of an oxide of sulfur shows that it contains 60% oxygen by mass. (A_r values: S = 32, O = 16.)

(a) What is the percentage mass of sulfur in the oxide?

(1 mark)

(b) Work out the formula of the oxide.

(2 marks)

3 Heating a test tube containing 2 g of calcium carbonate produced 1.08 g of calcium oxide when it was reweighed. The equation for the reaction is:

$$\mathbf{CaCO_3(s) \rightarrow CaO(s) + CO_2(g)}$$

(M_r values: $CaCO_3$ = 100, CaO = 56.)

(a) Calculate the amount of calcium oxide you would expect to be formed from 2 g of calcium carbonate.

(1 mark)

(b) Compare the value to the amount obtained in the experiment. Suggest a possible reason for the difference.

(1 mark)

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 page 128) of the reaction.

'Atom economy' — % of reactants changed to useful products

1) A lot of reactions make more than one product.
2) Some of them will be useful, but others will just be waste, e.g. when you make quicklime from limestone, you also get CO_2 as a waste product.
3) The atom economy of a reaction tells you how much of the mass of the reactants ends up as useful products.

Learn the equation:

$$\textbf{atom economy} = \frac{\text{total } M_r \text{ of useful products}}{\text{total } M_r \text{ of reactants}} \times 100$$

Example:

Hydrogen gas is made on a large scale by reacting natural gas (methane) with steam.

$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$

Calculate the atom economy of this reaction.

METHOD:

1) Identify the useful product — **that's the hydrogen gas**.
2) Work out the M_r of the reactants and of the useful product:

CH_4:	$12 + (4 \times 1)$	$= 16$
H_2O:	$(2 \times 1) + 16$	$= 18$
$3H_2$:	$3 \times (2 \times 1)$	$= 6$

So M_r of useful products = 6, and M_r of reactants = 16 + 18 = 34

3) Use the formula to calculate the atom economy:

$$\textbf{atom economy} = \frac{6}{34} \times 100 = \underline{17.6\%}$$

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

In industry, the waste CO is reacted with more steam to make CO_2 (and a bit more H_2). That brings the overall atom economy down to only 15% — but the final waste product is much less nasty that way.

Atom Economy

As you may have guessed already, a high atom economy is good, and a low atom economy is not so good. Make sure you know all the reasons why.

*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.

3) At the same time, they produce loads of waste materials that have to be disposed of somehow.

4) That tends to make these reactions unsustainable — the raw materials will run out and the waste has to go somewhere.

5) For the same reasons, low atom economy reactions aren't usually profitable.

6) Raw materials are expensive to buy, and waste products can be expensive to remove and dispose of responsibly.

7) The best way around the problem is to find a use for the waste products rather than just throwing them away.

8) 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.

9) The reactions with the highest atom economy are the ones that only have one product — like the Haber process (see page 167).

10) Those reactions have an atom economy of 100%.

So why do they make hydrogen in that nasty, inefficient way you saw on the last page? Well, currently it's the best of a bad bunch. The other ways to make hydrogen on an industrial scale (like the electrolysis of brine, see page 145) use up huge amounts of energy and are too expensive to be worthwhile.

Atom economy — important, but not the whole story...

Atom economy isn't the only thing that affects profits — there are other costs besides buying raw materials and disposing of waste. There are energy and equipment costs, as well as the cost of paying people to work at the plant. You need to think about the percentage yield of the reaction too (p.128).

Percentage Yield

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

Percentage yield compares actual and predicted yields

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) The predicted yield of a reaction can be calculated from the balanced reaction equation (see page 122).

2) Percentage yield is given by the formula:

$$\text{percentage yield} = \frac{\text{actual yield (grams)}}{\text{predicted yield (grams)}} \times 100$$

3) Percentage yield is always somewhere between 0 and 100%.

4) A 100% percentage yield means that you got all the product you expected to get.

5) A 0% yield means that no reactants were converted into product, i.e. no product at all was made.

Yields are always less than 100%

1) In real life, you never get a 100% percentage yield.
2) Some product or reactant always gets lost along the way — and that goes for big industrial processes as well as school lab experiments.
3) How this happens depends on what sort of reaction it is and what apparatus is being used.

Lots of things can go wrong, but you can find the four you need to know about conveniently located on the next page.

Even with the best equipment, you can't get the maximum product

A high percentage yield means there's not much waste — which is good for preserving resources and keeping production costs down. If a reaction's going to be worth doing commercially, it generally has to have a high percentage yield or recyclable reactants, e.g. the Haber process.

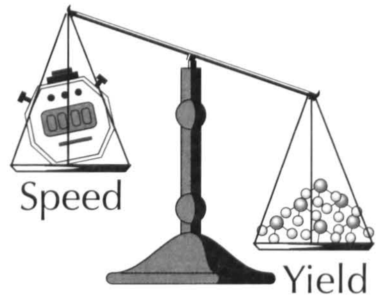
Percentage Yield

Learn these **four reasons** why yields **can't** be 100%

1) The reaction is **reversible**

In <u>reversible reactions</u> (like the Haber process, see page 167), not all the reactants change into product.

Instead, you get <u>reactants</u> and <u>products</u> in <u>equilibrium</u>. Increasing the temperature moves the <u>equilibrium position</u> (see pages 165-166), so heating the reaction to speed it up might mean a <u>lower yield</u>.



2) Filtration

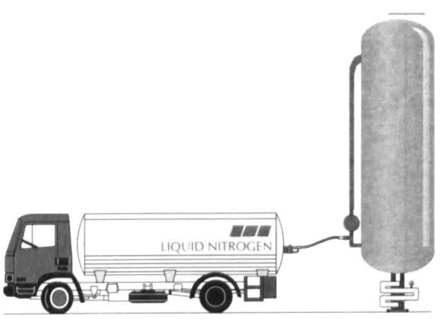
When you <u>filter a liquid</u> to remove <u>solid particles</u>, you nearly always lose a bit of liquid or a bit of solid.

1) If you want to <u>keep the liquid</u>, you lose the bit that remains with the solid and filter paper (as they always stay a bit wet).
2) If you want to <u>keep the solid</u>, some of it usually gets left behind when you scrape it off the filter paper — even if you're really careful.

3) **Transferring** liquids

You always lose a bit of liquid when you <u>transfer</u> it from one container to another — even if you manage not to spill it.

Some of it always gets left behind on the <u>inside surface</u> of the old container. Think about it — it's always wet when you finish.



4) **Unexpected reactions**

Things don't always go exactly to plan.

Sometimes you get unexpected reactions happening, so the yield of the <u>intended product</u> goes down. These can be caused by <u>impurities</u> in the reactants, but sometimes just changing the <u>reaction conditions</u> affects what products you make.

Warm-Up and Exam Questions

This is all about trying to judge how much useful product you're actually getting from a reaction. Try these questions to help make sure you won't get stuck in the exam. First a (fairly) gentle warm-up, and then some more exam-like questions to give you an idea of what you can expect.

Warm-Up Questions

1) What effect does a waste by-product have on the atom economy of a reaction?
2) What is the atom economy of the reaction shown? $\mathbf{2SO_2 + O_2 \rightarrow 2SO_3}$
3) Why might a reaction with a low atom economy be bad for the environment?
4) Why might a reaction with a low atom economy not be profitable?
5) What is the percentage yield of a reaction which produced 4 g of product if the predicted yield was 5 g?

Exam Questions

1 Ethanol produced by the fermentation of sugar can be converted into ethene, as shown below. The ethene can then be used to make polythene.

$$\mathbf{C_2H_6O\ (g) \rightarrow C_2H_4\ (g) + H_2O\ (g)}$$

Calculate the atom economy of this reaction.
(A_r values: C = 12, O = 16, H = 1.)

(3 marks)

2 A sample of copper was made by reducing 4 g of copper oxide with methane gas. When the black copper oxide turned orange-red, the sample was scraped out into a beaker. Sulfuric acid was added to dissolve any copper oxide that remained. The sample was then washed, filtered and dried. 2.8 g of copper was obtained.

(A_r values: Cu = 63.5, O = 16.)

The equation for this reaction is: $\mathbf{CH_4 + 4CuO \rightarrow 4Cu + 2H_2O + CO_2}$

(a) Use the equation to calculate the maximum mass of copper which could be obtained from the reaction (the predicted yield).

(3 marks)

(b) Calculate the percentage yield of the reaction.

(2 marks)

(c) Suggest three different reasons why the yield of the reaction was less than 100%.

(3 marks)

Revision Summary for Section Seven

Some more tricky questions to stress you out. The thing is though, why bother doing easy questions? These meaty monsters find out what you really know, and worse, what you really don't. Yeah, I know, some of them are a bit scary, but if you want to get anywhere in life you've got to face up to a bit of hardship. That's just the way it is. Take a few deep breaths and then try these.

1) Define the relative atomic mass of an element.
2)* Find the A_r or M_r for each of these (use the periodic table inside the front cover):
a) Ca b) Ag c) CO_2 d) $MgCO_3$
e) $Al(OH)_3$ f) ZnO g) Na_2CO_3 h) sodium chloride
3) Define the term isotope.
4) Explain how carbon-14 is different from carbon-12.
5)* The table below gives the masses and relative abundances of the isotopes of neon:

relative mass of isotope	relative abundance
20	91%
22	9%

Calculate the relative atomic mass of neon. Give your answer to 2 decimal places.
6)* Find the percentage mass of carbon in the following:
a) CH_4 b) K_2CO_3 c) CO_2
7)* Find the percentage mass of oxygen in the following:
a) CO b) K_2CO_3 c) CO_2
8)* Find the empirical formula of the compound formed when 21.9 g of magnesium, 29.3 g of sulfur and 58.4 g of oxygen react.
9)* Find the empirical formula of the compound formed when 227 g of calcium reacts with 216 g of fluorine.
10)*What mass of magnesium oxide is produced when 112.1 g of magnesium burns in air?
11)*What mass of sodium is needed to produce 108.2 g of sodium oxide?
12)*What mass of carbon will react with hydrogen to produce 24.6 g of propane (C_3H_8)?
13)*How many moles are there in 284 g of sodium sulfate, Na_2SO_4?
14)*What mass of chlorine is there in 2 moles of magnesium chloride, $MgCl_2$?
15) What is meant by a "3 M solution"?
16)*How many moles of barium chloride are there in 500 cm^3 of a 0.2 M solution of barium chloride?
17)*Iron is extracted from its ore using carbon monoxide in a process described by this equation:
$3CO + Fe_2O_3 \rightarrow 3CO_2 + 2Fe$
Using the periodic table, calculate the atom economy of this reaction.
18) Is it better to have a high atom economy or a low atom economy? Explain why.
19) What is the formula for percentage yield?
20)*The predicted yield of a reaction was found to be 12.5 g. However, when the reaction was carried out, the actual yield was only 8.25 g. Calculate the percentage yield of this reaction.
21) The percentage yield of a reaction is never 100%. Give four possible reasons for this.

* Answers on page 258.

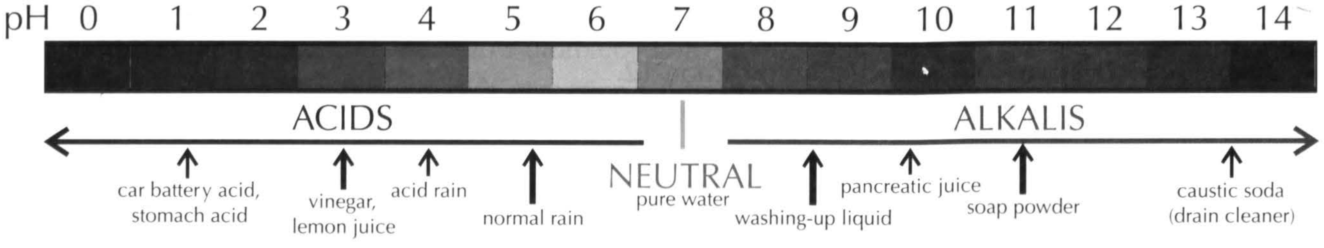
Acids and Bases

You'll find acids and bases at home, in industry and in the lab — they're an important set of chemicals.

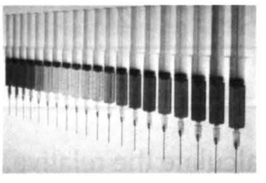
The pH scale and universal indicator

The dye in the indicator changes colour depending on whether it's above or below a certain pH. Universal indicator is a very useful combination of dyes which gives the colours shown below.

It's very handy for estimating the pH of a solution:



1) The pH scale goes from 0 to 14.
2) A very strong acid has pH 0. A very strong alkali has pH 14.
3) A neutral substance has pH 7 (e.g. pure water).



Acids and bases neutralise each other

An ACID is a substance with a pH of less than 7. Acids form H^+ ions in water.

A BASE is a substance with a pH greater than 7.

An ALKALI is a base that DISSOLVES IN WATER. Alkalis form OH^- ions in water.

The reaction between acids and bases is called neutralisation. Make sure you learn it:

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

Neutralisation can also be seen in terms of H^+ and OH^- ions like this, so learn this too:

$$H^+ + OH^- \rightarrow H_2O$$

When an acid neutralises a base (or vice versa), the products are neutral, i.e. they have a pH of 7.

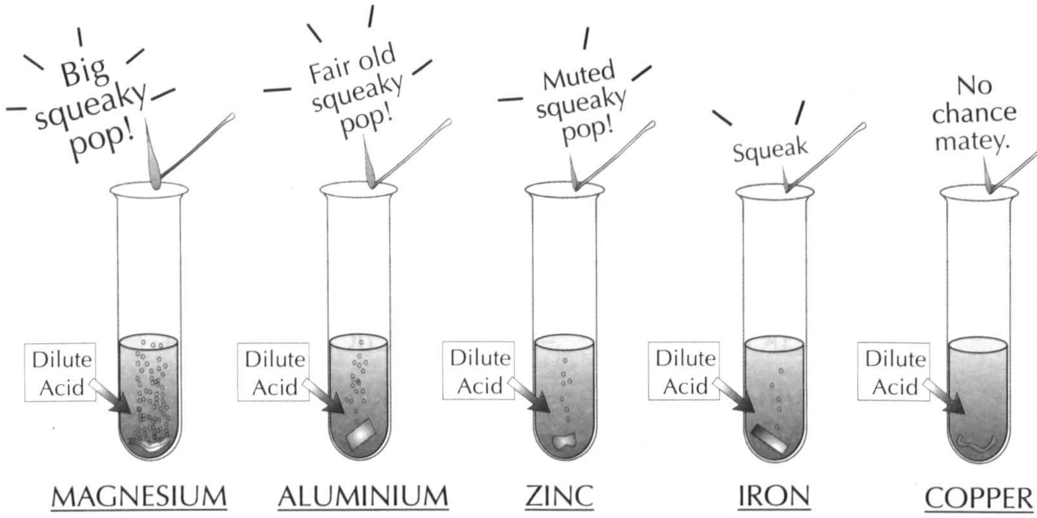
You need to know some of the uses of sulfuric acid

1) Sulfuric acid is used in car batteries, where it's concentrated enough to cause severe burns.
2) It's also used in many manufacturing processes, such as making fertilisers and detergents.
3) You can also use it to clean and prepare metal surfaces, e.g. before painting or welding. A metal surface is usually covered with a layer of insoluble metal oxide. Sulfuric acid reacts with this, forming soluble metal salts which wash away, nice and easily.

Acids Reacting with Metals

acid + metal → salt + hydrogen

That's written big because it's really worth remembering. Here's the typical experiment:



1) The more reactive the metal, the faster the reaction will go — very reactive metals (e.g. sodium) react explosively.
2) Copper does not react with dilute acids at all — because it's less reactive than hydrogen.
3) The speed of the reaction is indicated by the rate at which the bubbles of hydrogen are given off.
4) The hydrogen is confirmed by the burning splint test giving the notorious 'squeaky pop'.
5) The name of the salt produced depends on which metal is used, and which acid is used:

Hydrochloric acid will always produce chloride salts

$2HCl + Mg \rightarrow MgCl_2 + H_2$ (magnesium chloride)

$6HCl + 2Al \rightarrow 2AlCl_3 + 3H_2$ (aluminium chloride)

$2HCl + Zn \rightarrow ZnCl_2 + H_2$ (zinc chloride)

Chloride and sulfate salts are generally soluble in water (the main exceptions are lead chloride, lead sulfate and silver chloride, which are insoluble).

Sulfuric acid will always produce sulfate salts

$H_2SO_4 + Mg \rightarrow MgSO_4 + H_2$ (magnesium sulfate)

$3H_2SO_4 + 2Al \rightarrow Al_2(SO_4)_3 + 3H_2$ (aluminium sulfate)

$H_2SO_4 + Zn \rightarrow ZnSO_4 + H_2$ (zinc sulfate)

Nitric acid produces nitrate salts when NEUTRALISED, but...

Nitric acid reacts fine with alkalis to produce nitrates, but it can play silly devils with metals and produce nitrogen oxides instead, so we'll ignore it here.

Neutralisation Reactions

*Metal **oxides** and metal **hydroxides** are **bases***

1) Some metal oxides and metal hydroxides dissolve in water. These soluble compounds are alkalis.
2) Even bases that won't dissolve in water will still react 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

(These are neutralisation reactions, of course.)

hydrochloric acid	+	copper oxide	→	copper chloride	+	water
$2HCl$	+	CuO	→	$CuCl_2$	+	H_2O
sulfuric acid	+	potassium hydroxide	→	potassium sulfate	+	water
H_2SO_4	+	$2KOH$	→	K_2SO_4	+	$2H_2O$
nitric acid	+	sodium hydroxide	→	sodium nitrate	+	water
HNO_3	+	$NaOH$	→	$NaNO_3$	+	H_2O

*Acids and **carbonates** produce **carbon dioxide***

These are very like the ones above — they just produce carbon dioxide as well.

acid + carbonate → salt + water + carbon dioxide

hydrochloric acid	+	sodium carbonate	→	sodium chloride	+	water	+	carbon dioxide
$2HCl$	+	Na_2CO_3	→	$2NaCl$	+	H_2O	+	CO_2
hydrochloric acid	+	calcium carbonate	→	calcium chloride	+	water	+	carbon dioxide
$2HCl$	+	$CaCO_3$	→	$CaCl_2$	+	H_2O	+	CO_2

*Acids and **ammonia** produce **ammonium salts***

And lastly... **acid + ammonia → ammonium salt**

hydrochloric acid	+	ammonia	→	ammonium chloride
HCl	+	NH_3	→	NH_4Cl
sulfuric acid	+	ammonia	→	ammonium sulfate
H_2SO_4	+	$2NH_3$	→	$(NH_4)_2SO_4$
nitric acid	+	ammonia	→	ammonium nitrate
HNO_3	+	NH_3	→	NH_4NO_3

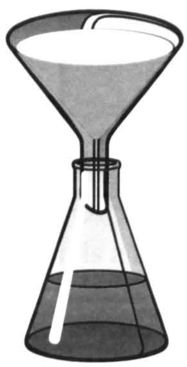
This last reaction with nitric acid produces ammonium nitrate fertiliser, much appreciated for its double dose of nitrogen (essential for healthy plant growth).

Making Salts

Most chlorides, sulfates and nitrates are soluble in water (the main exceptions are lead chloride, lead sulfate and silver chloride). Most oxides, hydroxides and carbonates are insoluble in water.

*Making **soluble** salts from **insoluble bases***

1) You need to pick the right acid, plus a metal carbonate or metal hydroxide, as long as it's insoluble.
2) You can't use sodium, potassium or ammonium carbonates or hydroxides, as they're soluble (so you can't tell whether the reaction has finished — see below).
3) You add the carbonate or hydroxide to the acid until all the acid is neutralised. (The excess carbonate or hydroxide will just sink to the bottom of the flask when all the acid has reacted.)
4) Then filter out the excess carbonate, and evaporate off the water — and you should be left with a pure, dry salt.



Filtering — to get rid of the excess carbonate or hydroxide.

For example, you can use copper carbonate and nitric acid to make copper nitrate:

$$CuCO_3(s) + 2HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

Ammonium nitrate** is a **soluble** salt from a **soluble base

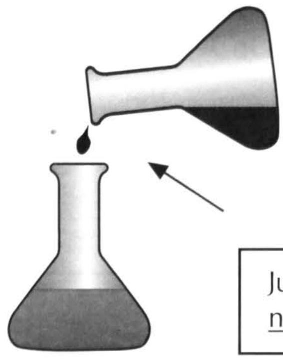
1) Ammonia itself is a base, but it's SOLUBLE, as are all other ammonium bases.
2) This means making soluble ammonium salts, such as ammonium nitrate, is a bit tricky.
3) You can't just add an excess of base and filter out what's left — you have to add exactly the right amount of base to just neutralise the acid.
4) You need to use an indicator so you can tell when the reaction has just finished.
5) Then you've got to repeat it without indicator, using exactly the same volumes of base and acid, so the salt isn't contaminated with indicator.
6) All this is obviously quite fiddly. But ammonium nitrate is a great fertiliser, so it's all worthwhile in the end (if you want nice big crops to grow, that is).

Making Salts

Making *insoluble* salts — *precipitation* reactions

1) If the salt you want to make is insoluble, you can use a precipitation reaction (see page 141).
2) You just need to pick the right acid and nitrate, then mix them together. For example, if you want to make lead chloride (which is insoluble), mix hydrochloric acid and lead nitrate.

$$Pb(NO_3)_2(aq) + 2HCl(aq) \longrightarrow PbCl_2(s) + 2HNO_3(aq)$$



Just mix an acid and a nitrate — simple as that.

3) Once the salt has precipitated out (and is lying at the bottom of your flask), all you have to do is filter it from the solution, wash it and then dry it on filter paper.
4) Precipitation reactions can be used to remove poisonous ions (e.g. lead) from drinking water. Calcium and magnesium ions can also be removed from water this way — they make water 'hard', which stops soap lathering properly.

Making salts by *displacement*

Have a look at page 20 for more on displacement.

1) If you put a more reactive metal like magnesium into a salt solution of a less reactive metal, like copper sulfate, then the magnesium will take the place of the copper — and make magnesium sulfate.
2) The "kicked-out" (or displaced) metal then coats itself onto the more reactive metal.
3) Once the magnesium has been completely coated with copper, the reaction stops, so this isn't a very practical way to make a salt.

Making soluble salts from soluble bases is the most fiddly

It's hard to find the precise neutral point using universal indicator. There's quite a wide range of greens between blue and yellow. There are more accurate indicators though — see p.178 for more on these.

Warm-Up and Exam Questions

Now try these questions — you're less likely to get a nasty surprise in the exam if you do.

Warm-Up Questions

1) What name is given to the type of reaction in which an acid reacts with a base?
2) Which two substances are formed when an acid reacts with a metal such as zinc?
3) Which two substances are formed when nitric acid reacts with copper oxide?
4) Explain what you would do to make a dry sample of a soluble salt from an insoluble base.
5) Why couldn't you make an iron salt from an aluminium salt by displacement?

Exam Questions

1 The table shows the results when five solutions, A–E, were tested with universal indicator.

Solution	Colour	pH
A		1
B	pale green	
C	orange	5
D	dark blue	
E		14

(a) Complete the blanks in the table.

(2 marks)

(b) Which solution is a weak acid?

(1 mark)

(c) Which solution is a strong alkali?

(1 mark)

(d) Which solution contains sodium chloride?

(1 mark)

(e) Which solution is battery acid?

(1 mark)

2 An experiment was carried out in which sodium hydroxide solution was added, 2 cm^3 at a time, to 10 cm^3 of sulfuric acid. The pH was estimated after each addition using universal indicator paper.

The results are shown in the table.

Volume of sodium hydroxide added (cm^3)	pH
0	1
2	1
4	2
6	4
8	12
10	13
12	13

(a) Plot the results on a graph, with pH on the vertical axis and volume of sodium hydroxide added on the horizontal axis.

Draw a best fit curve.

(2 marks)

(b) Estimate the volume of sodium hydroxide needed to neutralise the acid.

(1 mark)

(c) How do the results show that sulfuric acid is a strong acid?

(1 mark)

(d) Name the salt formed in the reaction.

(1 mark)

Exam Questions

3 An excess of different substances were added to a solution of hydrochloric acid containing universal indicator. The results are shown in the table below.

Substance	Formula	Observations: during the reaction	Observations: after excess added	Final pH
zinc oxide	ZnO	ZnO dissolves	ZnO settles out	7
potassium hydroxide	KOH	KOH dissolves	KOH stays in solution	14
ammonia	NH_3	NH_3 dissolves	NH_3 stays in solution	11
sodium chloride	NaCl	NaCl dissolves	NaCl stays in solution	1
magnesium metal	Mg	Mg dissolves, bubbling	Mg settles out	7

(a) Which substance didn't neutralise the acid?

(1 mark)

(b) Which two substances are alkalis?

(1 mark)

(c) Which substance is an insoluble base?

(1 mark)

(d) Complete the symbol equations for the following reactions. The ions involved are shown below to help you work out the formulas of the salts formed.

Ions: Mg^{2+}, Zn^{2+}, NH_4^+, K^+, Cl^-

(i) $Mg + 2HCl \rightarrow$ *(1 mark)*

(ii) $NH_3 + HCl \rightarrow$ *(1 mark)*

(iii) $ZnO + 2HCl \rightarrow$ *(1 mark)*

(iv) $KOH + HCl \rightarrow$ *(1 mark)*

4 Jenny wanted to make a dry sample of silver chloride, AgCl, by precipitation.

(a) What property must a salt have to be made by precipitation?

(1 mark)

(b) Jenny looked up the solubilities of some compounds she might use.

Write down one reaction using substances from the table that she could use to make silver chloride by precipitation.

(1 mark)

Compound	Formula	Solubility
silver oxide	Ag_2O	insoluble
silver nitrate	$AgNO_3$	soluble
silver carbonate	$AgCO_3$	insoluble
sulfuric acid	H_2SO_4	soluble
nitric acid	HNO_3	soluble
hydrochloric acid	HCl	soluble

(c) Outline the steps needed to give a pure dry sample of silver chloride after mixing the solutions.

(3 marks)

More Chemical Changes

There are loads of different kinds of chemical reactions — e.g. hydration, dehydration and thermal decomposition. There are many others too, but for now you just need to learn the ones on these pages.

Hydration reactions have water as a reactant

1) Water's dead useful stuff. It's used as a solvent because it dissolves so many substances.
2) Water's also used in hydration reactions.
3) Hydration reactions are ones where water reacts with another substance to form a new product.

Examples:

Calcium oxide (quicklime) reacts with water to produce calcium hydroxide (slaked lime).

calcium oxide + water → calcium hydroxide

$CaO + H_2O \rightarrow Ca(OH)_2$

Ethene (C_2H_4) will react with steam (H_2O) to make ethanol.

ethene + steam → ethanol

$C_2H_4 + H_2O \rightarrow C_2H_5OH$

Dehydration reactions have water as a product

Dehydration reactions are where water is removed from one or more substances, forming new products.

Example:

Dehydration reactions separate carbohydrates into carbon and water. Concentrated sulfuric acid is a strong dehydrating agent, and it'll grab water from other compounds. It acts as a catalyst to convert sucrose (sugar) into water and a spongy lump of black carbon.

sucrose $\xrightarrow{\text{catalyst of conc. sulfuric acid}}$ carbon + water

Your teacher might demonstrate this in the fume cupboard.

Reactions which join small molecules to make a bigger molecule sometimes also produce a molecule of water. These are usually called condensation reactions, but they're a kind of dehydration reaction too.

Example:

Alcohol reacts with organic acids to form an ester and water (see pages 227-228).

ethanol + ethanoic acid → ethyl ethanoate + water

More Chemical Changes

Thermal decomposition is breaking down with heat

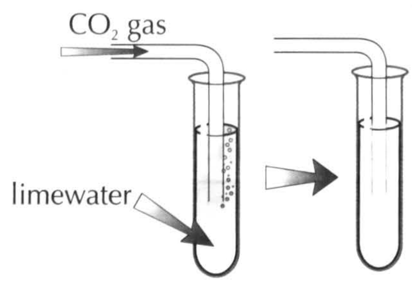
1) Carbonates and hydrogencarbonates release carbon dioxide gas (CO_2) when they're heated.
2) It's an example of thermal decomposition, which is when a substance breaks down into simpler substances when it is heated.
3) Learn the word equations for the thermal decomposition of carbonates and hydrogencarbonates:

   For example —

   calcium carbonate → calcium oxide + carbon dioxide

   sodium hydrogencarbonate → sodium carbonate + carbon dioxide + water

4) You can check it really is carbon dioxide that's released by testing it with limewater — CO_2 turns limewater cloudy when it's bubbled through.
5) Baking powder contains sodium hydrogencarbonate. Baking powder is added to cake mixtures — the carbon dioxide produced when it's heated in the oven makes the cake rise.



Rusting (corrosion) is an oxidation reaction

The word "rust" is only used for the corrosion of iron, not other metals.

Iron corrodes easily. In other words, it rusts.

1) When iron rusts, it's combining with oxygen (and also water). The iron gains oxygen to form iron(III) oxide. Water then becomes loosely bonded to the iron(III) oxide and the result is hydrated iron(III) oxide — which most people call rust.

   iron + oxygen + water → hydrated iron(III) oxide

2) The iron becomes oxidised in this reaction — it loses electrons in the reaction. The opposite process — the gain of electrons — is called reduction. There's more about reduction on page 217.

   $Fe \rightarrow Fe^{3+} + 3e^-$

3) Unfortunately, rust is a soft crumbly solid that soon flakes off to leave more iron available to rust. And if the water's salty or acidic, rusting will take place a lot quicker. Cars in coastal places rust a lot because they get covered in salty sea spray. Cars in dry deserty places hardly rust at all.

Only learn what you need to for the exam

By the way, there's actually more to the action of baking powder than just thermal decomposition*, but strangely the examiners don't care (so you don't have to either). The thing to remember about rusting is that it's an oxidation reaction — the iron loses electrons. When an atom gains electrons, that's a reduction reaction. Just remember OIL RIG — oxidation is loss, reduction is gain.

* Baking powder is sodium hydrogencarbonate and an acid (usually 'cream of tartar'), which react to make carbon dioxide. Baking soda is plain sodium hydrogencarbonate (aka bicarbonate of soda).

More Chemical Changes

Yes, that's right, yet more chemical changes — you also need to know about precipitation reactions. But you have to expect a lot of chemical changes in this book — this is chemistry, after all.

Precipitation — a solid forms in solution

1) A precipitation reaction is where two solutions react and an insoluble solid forms in the solution.
2) The solid is said to 'precipitate out' and, confusingly, the solid is also called 'a precipitate'.
3) Some soluble transition metal compounds react with sodium hydroxide to form an insoluble hydroxide, which then precipitates out.

Example:

Soluble copper sulfate reacts with sodium hydroxide to form insoluble copper hydroxide.

$$CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$$

copper sulfate + sodium hydroxide → copper hydroxide + sodium sulfate

Copper sulfate is soluble...

...but copper hydroxide is insoluble — so it precipitates out.

4) Since copper hydroxide is blue, you get a distinctive blue precipitate forming in the test tube.
5) You can also write the above equation in terms of ions:

$$Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2$$

The Cu^{2+} ions and the hydroxide ions combine to give you the insoluble copper hydroxide.

You can use precipitation to test for transition metal ions

1) Some insoluble transition metal hydroxides have distinctive colours.

   For example:

   - Copper hydroxide is a **blue** solid.
   - Iron(II) hydroxide is a **dark grey/green** solid.
   - Iron(III) hydroxide is an **orange** solid.

2) You can use this fact to test which transition metal ions a solution contains.
3) For example, if you add sodium hydroxide to an unknown soluble salt, and an orange precipitate forms, you know you had iron(III) ions in the solution.

Warm-Up and Exam Questions

By the time the big day comes you need to know all the facts in these questions like the back of your hand. It's not fun, but it's the only way to get good marks.

Warm-Up Questions

What name is given to a reaction in which:

1) ... water reacts with another substance to form a new product?
2) ... a substance is broken down by heating?
3) ... a substance loses electrons?
4) ... two solutions react to give an insoluble solid?
5) ... water is removed from a substance or substances to form a new product?

Exam Questions

1 If ethanol vapour is passed over hot aluminium oxide, it breaks down to give ethene and water. An equation for the reaction is given below:

$$C_2H_6O \rightarrow C_2H_4 + H_2O$$

Explain why this reaction could be considered to be:

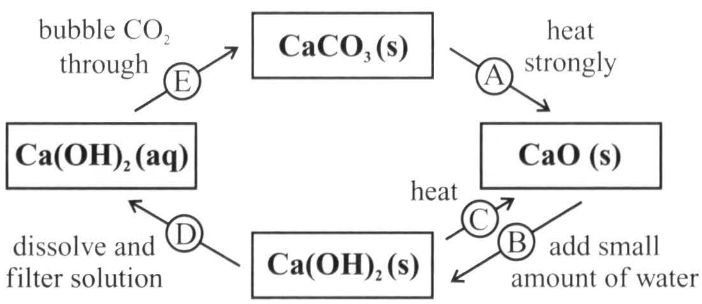
(a) thermal decomposition.

(1 mark)

(b) dehydration.

(1 mark)

2 The diagram below shows a series of reactions involving calcium.



Give the letter(s) for the reactions involving:

(a) thermal decomposition.

(2 marks)

(b) dehydration.

(2 marks)

(c) hydration.

(1 mark)

(d) the formation of a milky precipitate.

(1 mark)

Exam Questions

3 When iron rusts it reacts with oxygen and water to form hydrated iron(III) oxide.

(a) Write a word equation for the formation of rust.

(1 mark)

(b) Iron is converted into Fe^{3+} ions when it rusts.
Explain why this is an example of an oxidation reaction.

(1 mark)

(c) Suggest how hydrated iron(III) oxide might be converted into iron(III) oxide.

(1 mark)

(d) Which of the statements about the corrosion of iron below are correct?

A The layer of rust that forms on the surface of iron prevents further corrosion.

B Iron rusts more rapidly if exposed to salty water.

C Hydrated iron(III) oxide is a soft and crumbly solid.

D Rusting is an example of a thermal decomposition reaction.

(2 marks)

4 Baking powder contains sodium hydrogencarbonate.

(a) Write a word equation to show what happens to sodium hydrogencarbonate when it is heated in an oven.

(1 mark)

(b) What type of reaction is this?

(1 mark)

(c) Explain why baking powder is added to cake mixtures.

(2 marks)

5 In an experiment, iron filings were reacted with nitric acid and with sulfuric acid. The salts produced were tested by observing the precipitates formed when excess sodium hydroxide, NaOH, was added. The results are shown in the table below.

Reaction	Colour of precipitate when NaOH added	Name of precipitate
iron + sulfuric acid		iron(II) hydroxide
iron + nitric acid	orange-red	

Fill in the blanks in the table.

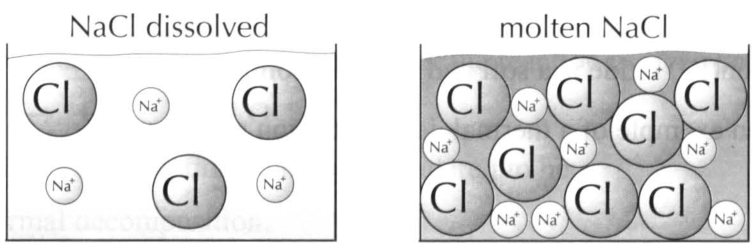
(2 marks)

Electrolysis and the Half-Equations

You need to know about the electrolysis of metals such as copper, and also about the electrolysis of salt solutions (covered on the next page).

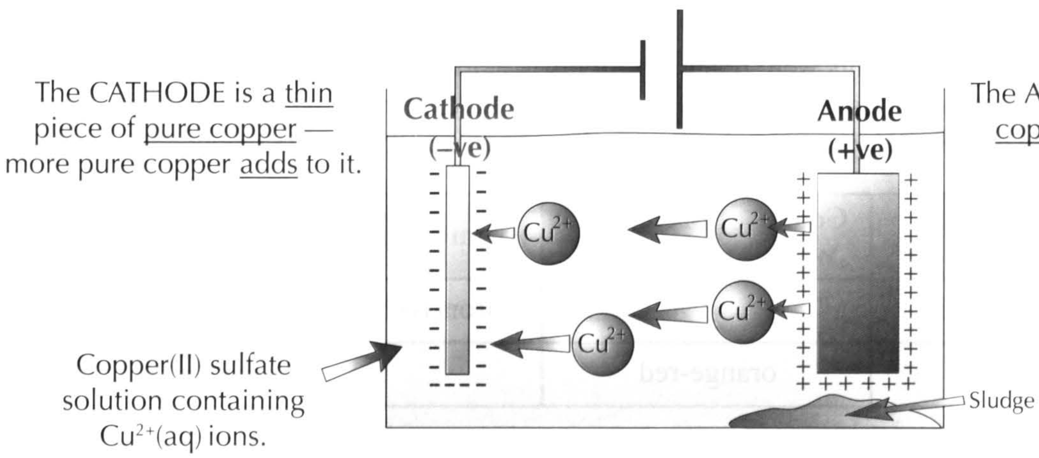
Electrolysis means 'splitting up with electricity'

1) Electrolysis is the breaking down of a substance using electricity.
2) It requires a liquid to conduct the electricity, called the electrolyte.
3) Electrolytes are usually free ions dissolved in water (for example, dissolved salts) or molten ionic substances.
4) In either case it's the free ions which conduct the electricity and allow the whole thing to work.
5) For an electrical circuit to be complete, there's got to be a flow of electrons. Electrons are taken away from ions at the positive anode and given to other ions at the negative cathode.
6) As ions gain or lose electrons they become atoms or molecules and are released.



Electrolysis is used to purify copper

1) The purer copper is, the better it conducts, so electrolysis is used to obtain very pure copper.
2) Electrons are pulled off copper atoms at the anode, causing them to go into solution as Cu^{2+} ions.
3) Cu^{2+} ions near the cathode gain electrons and turn back into copper atoms.
4) The impurities are dropped at the anode as a sludge, while pure copper atoms bond to the cathode.



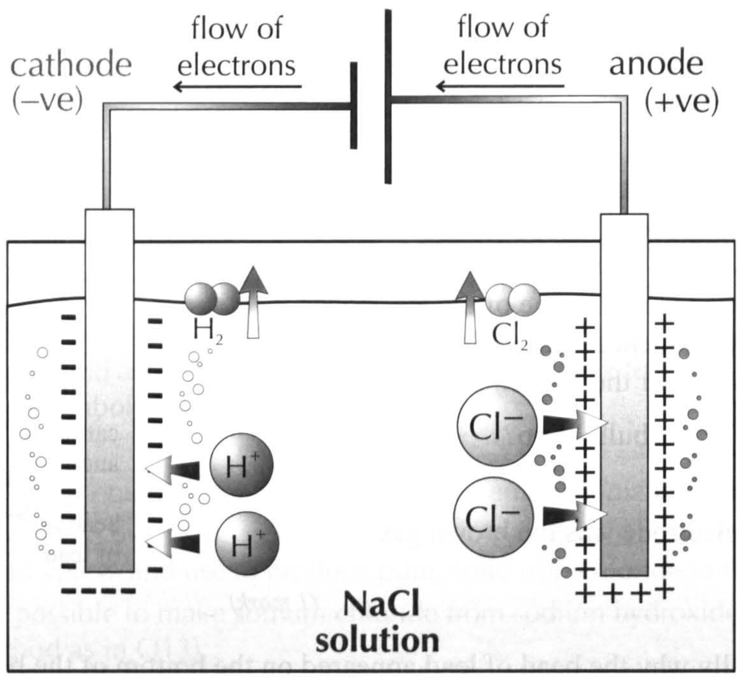
Electrolysis is used to purify other metals too

Electrolysis is also used for extracting aluminium from its ore (aluminium oxide). You have to melt the ore first so the ions can move. The positive Al^{3+} ions are attracted to the cathode where they pick up electrons and "zup", they turn into aluminium atoms. These then conveniently sink to the bottom.

Electrolysis and the Half-Equations

When common salt (sodium chloride) is electrolysed, it produces three very useful products. (See page 242 for more on these products).

*The diagram shows the **electrolysis** of a **salt solution***



Positive ions are called CATIONS because they're attracted to the negative cathode.

Hydrogen is produced at the cathode.

Negative ions are called ANIONS because they're attracted to the positive anode.

Chlorine is produced at the +ve anode.

1) At the cathode, two hydrogen ions accept two electrons to become one hydrogen molecule.

$$2H^+ + 2e^- \rightarrow H_2$$

2) At the anode, two chloride (Cl^-) ions lose their electrons and become one chlorine molecule.

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

3) NaOH is left in the solution.

Make sure the number of electrons is the same for both half-equations.

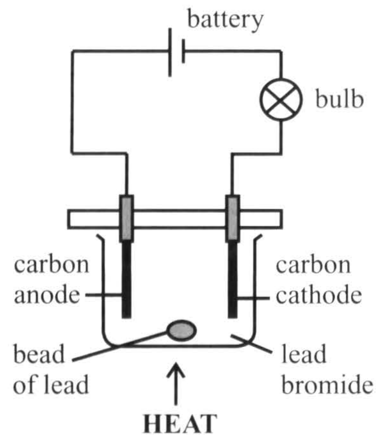
Warm-Up and Exam Questions

Warm-Up Questions

1) What state must an ionic compound be in if it's to be used as an electrolyte?
2) Why is it necessary to purify copper?
3) At which electrode are metals deposited during electrolysis?
4) Name another metal, besides copper, which is purified or extracted by electrolysis.
5) Name the three products obtained from the electrolysis of salt (NaCl) solution.

Exam Questions

1 Lead bromide was electrolysed as shown in the diagram. The lead bromide was heated strongly and the bulb lit up after about 5 minutes. A brown gas was formed at one of the electrodes and after the experiment a bead of lead was found at the bottom of the beaker.



(a) Explain why the bulb lit up?

(1 mark)

(b) At which electrode was the brown gas produced?

(1 mark)

(c) Explain fully why the bead of lead appeared on the bottom of the beaker.

(2 marks)

(d) Copy and complete the half-equations for the electrolysis of lead bromide:

(i) Pb^{2+} + _____ → _____ *(1 mark)*

(ii) ___ Br^- → _____ + ___ e^- *(1 mark)*

2 When sodium chloride solution is electrolysed a gas is produced at each electrode. Gas A produced a squeaky pop when tested with a lighted splint and gas B bleached a piece of damp litmus paper.

(a) (i) What is gas A?

(1 mark)

(ii) Which ion, present in water, is discharged when this gas is formed?

(1 mark)

(iii) Write a half-equation for the reaction at this electrode.

(1 mark)

(b) (i) What is gas B?

(1 mark)

(ii) Explain how this gas is formed at the electrode.

(2 marks)

Revision Summary for Section Eight

Have a go at these questions and see how much you can remember. If you're not sure about any of them, don't just skulk past them, check back to the relevant page, get it straight in your head and tackle the question again. It's the best way, I promise. Especially when you're trying to learn a section full of equations and reactions like that one.

1) Give the colour of universal indicator for a solution with a pH value of:
   a) 1, b) 5, c) 7, d) 10.
2) What type of ions are always present in: a) acids b) alkalis?
3) Write the general word equation for a neutralisation reaction.
4) Give two industrial uses of sulfuric acid.
5) Explain why you would not react sodium with hydrochloric acid in the lab.
6) Predict the outcome of adding copper to a dilute acid. Explain why this would be the result.
7) Describe a test for the gas produced when zinc reacts with sulfuric acid.
8)* Name the salts formed and write balanced symbol equations for the following reactions:
   a) hydrochloric acid with magnesium, b) sulfuric acid with aluminium.
9) What type of reaction is "acid + metal oxide", or "acid + metal hydroxide"?
10) What would the products of such a reaction be?
11)*Suggest a suitable acid and a suitable metal oxide/hydroxide to mix to form the following salts. Write out a balanced symbol equation for each reaction.
   a) copper(II) chloride b) calcium nitrate c) zinc sulfate
12) Write a balanced symbol equation for the reaction between ammonia and nitric acid.
13) Iron chloride can be made by mixing iron hydroxide (an insoluble base) with hydrochloric acid. Describe the method you would use to produce pure, solid iron chloride in the lab.
14) Explain why it isn't possible to make sodium chloride from sodium hydroxide and hydrochloric acid using the same method as in Q13).
15) Describe the method you would have to use to do this.
16) Describe one practical use of making insoluble salts by precipitation.
17) When concentrated sulfuric acid is added to sugar, the sugar turns into spongy black carbon. What kind of reaction is this?
18) What products are formed by the thermal decomposition of calcium carbonate?
19) Name the substance in baking powder that produces carbon dioxide to make cakes rise.
20) Describe a way to test solutions for transition metal ions.
21) What is the scientific term for rust?
22) Is rusting an example of an oxidation or a reduction reaction?
23) What is electrolysis? Explain why liquids have to be used in electrolysis.
24) Describe the process of purifying copper by electrolysis.
25) What is a cation?
26) Write balanced half-equations for the reactions at the anode and the cathode during the electrolysis of sodium chloride solution.

* Answers on page 259.

Rates of Reaction

This section is all about how fast chemical reactions go and the energy changes involved.

Reactions can go at all sorts of different rates

1) One of the slowest is the rusting of iron.
2) A moderate speed reaction is a metal (like magnesium) reacting with acid to produce a gentle stream of bubbles.
3) A really fast reaction is an explosion, where it's all over in a fraction of a second.

The rate of a reaction depends on four things

Make sure you learn them:

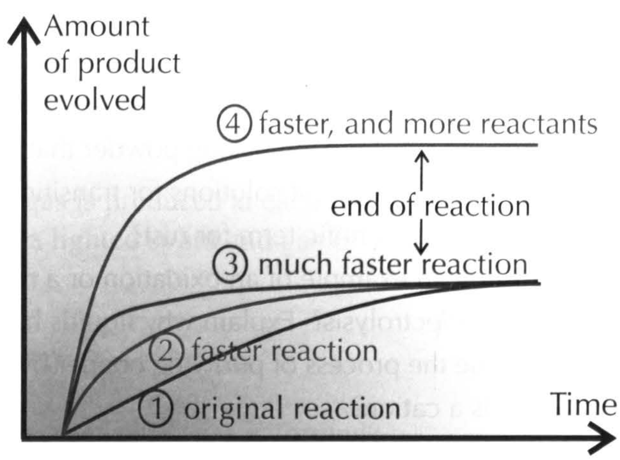
1) Temperature	2) Concentration	— (or pressure for gases)
3) Catalyst	4) Size of particles	— (or surface area)

Typical graphs for rate of reaction

The plot below shows how the speed of a particular reaction varies under different conditions. The quickest reaction is shown by the line that becomes flat in the least time. The line that flattens out first must have the steepest slope compared to the others, so you can spot the slowest and fastest reactions.

1) Graph 1 represents the original fairly slow reaction. The graph is not very steep.
2) Graphs 2 and 3 represent the reaction happening more quickly but with the same initial amounts of reactants. The slope of the graphs gets steeper.
3) The increased rate could be due to any of these:

> a) increase in temperature
> b) increase in concentration (or pressure)
> c) catalyst added
> d) solid reactant crushed into smaller bits



4) Graph 4 produces more product as well as going faster. This can only happen if more reactant(s) are added at the start. 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 get there.

It's really important that you understand the graph above

Industrial reactions generally use a catalyst and are done at high temperature and pressure. Time is money, so the faster an industrial reaction goes the better... but only up to a point. Chemical plants are quite expensive to rebuild if they get blown into lots and lots of teeny tiny pieces.

Measuring Rates of Reaction

You need to know some examples of how reaction rates are measured.

*There are **three ways** to measure the **speed** of a reaction*

The speed of a reaction can be observed either by how quickly the reactants are used up or by how quickly the products are formed. It's usually a lot easier to measure products forming.

You can calculate the rate of a reaction using the following equation:

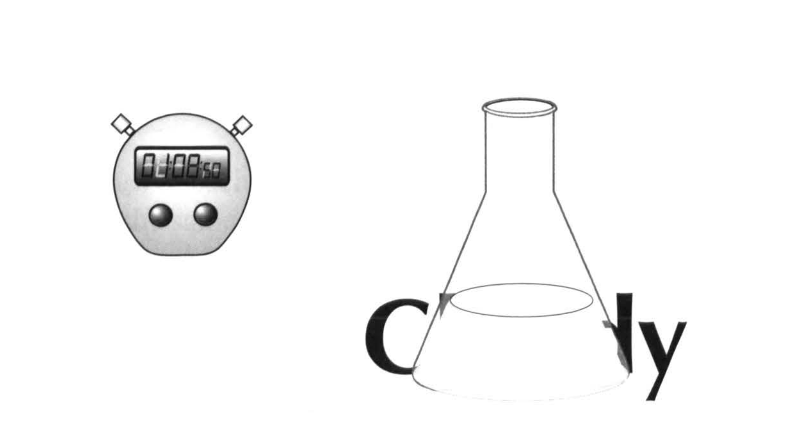
$$\textbf{Rate of reaction} = \frac{\textbf{amount of reactant used or amount of product formed}}{\textbf{time}}$$

There are different ways that the speed of a reaction can be measured.

Below and on the next page are three ways to measure the rate of a reaction. Learn them.

1) Precipitation

1) This is when the product of the reaction is a precipitate which clouds the solution.
2) Observe a marker through the solution and measure how long it takes for it to disappear.
3) The quicker the marker disappears, the quicker the reaction.
4) This only works for reactions where the initial solution is rather see-through.
5) The result is very subjective — different people might not agree over the exact point when the mark 'disappears'.



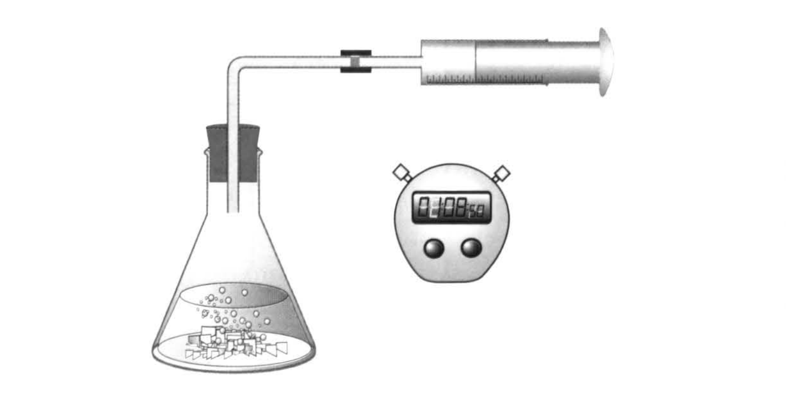
Measuring Rates of Reaction

2) *Change in mass (usually gas given off)*

1) Measuring the speed of a reaction that produces a gas can be done 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) Rate of reaction graphs are particularly easy to plot using the results from this method.
5) This is the most accurate of the three methods described 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 there is given off during a given time interval, the faster the reaction.
3) A graph of gas volume against time elapsed could be plotted to give a rate of reaction graph.
4) Gas syringes usually give volumes accurate to the nearest millilitre, so they're quite accurate. But if the reaction is too vigorous, you can blow the plunger out of the end of the syringe.



Each of these methods has pros and cons

The mass balance method is only accurate as long as the flask isn't too hot, otherwise you lose mass by evaporation as well as in the reaction. The first method isn't very accurate, but if you're not producing a gas you can't use either of the other two. Ah well.

Warm-Up and Exam Questions

So, let's see if you know the main factors that affect the rate of a reaction and how to measure them.

Warm-Up Questions

1) Give three ways of increasing the rate of a reaction between magnesium and sulfuric acid.
2) How would adding water to an acid alter the time taken for a piece of zinc to react with it?
3) Oxidation of lactose in milk makes it go 'sour'. How could this reaction be slowed down?
4) Give an example of a reaction that happens very slowly, and one that is very fast.
5) Describe one way of monitoring a reaction in which a gas is given off.

Exam Questions

1 Set volumes of sodium thiosulfate and hydrochloric acid were reacted at different temperatures. The time taken for a black cross to be obscured by the sulfur precipitated was measured at each temperature. The results are shown in the table.

Time (s)	Temperature (°C)
6	55
11	36
17	24
27	16
40	9
51	5

(a) Give two things that should be kept constant in this experiment.

(2 marks)

(b) Plot the results on a graph (with time on the x-axis) and draw a best-fit curve.

(2 marks)

(c) Explain the relationship illustrated by your graph.

(2 marks)

(d) How might the results change if the sodium thiosulfate concentration was reduced?

(2 marks)

2 The table shows the results of reactions between excess marble and 50 cm^3 of 1 M hydrochloric acid.

Time (min)	Mass of flask A (g)	Mass of flask B (g)
0	121.6	121.6
1	120.3	119.8
2	119.7	119.2
3	119.4	119.1
4	119.2	119
5	119.1	119
6	119	119
7	119	

(a) Why did the mass of the flasks and their contents decrease during the reaction?

(1 mark)

(b) Why did the mass of each flask and its contents fall by the same amount?

(1 mark)

(c) (i) Suggest what may have been different about flask B.

(1 mark)

(ii) Explain how this difference could lead to a change in the rate of the reaction.

(1 mark)

(d) In both reactions, the rate is fastest at the beginning. Explain why.

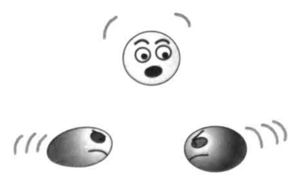
(1 mark)

Collision Theory

Reaction rates are explained perfectly by collision theory. It's really simple.

Collision theory shows why certain things increase reaction rates

1) Collision theory just says that the rate of a reaction depends on how often and how hard the reacting particles collide with each other.
2) The basic idea is that particles have to collide in order to react, and they have to collide hard enough (with enough energy).

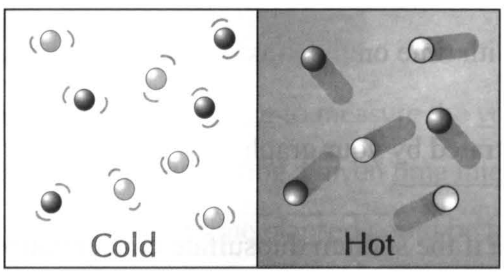


More collisions increases the rate of reaction

All four methods of increasing the rate of reactions (see also next page) can be explained in terms of increasing the number of successful collisions between the reacting particles:

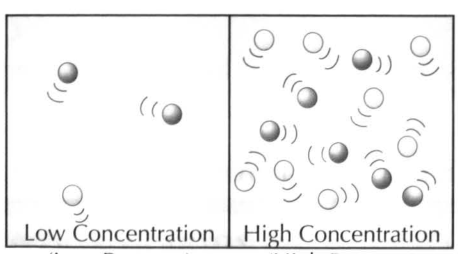
1) HIGHER TEMPERATURE increases collisions

When the temperature is increased the particles all move more quickly. If they're moving quicker, they're going to have more collisions.



2) HIGHER CONCENTRATION increases collisions

If a solution is made more concentrated it means there are more particles of reactant knocking about between the water molecules, which makes collisions between the important particles more likely. In a gas, increasing the pressure means the particles are more squashed up together so there are going to be more collisions.

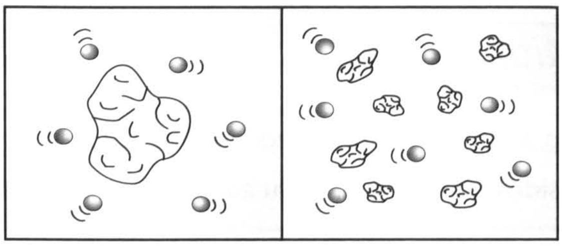


(Low Pressure) (High Pressure)

Collision Theory

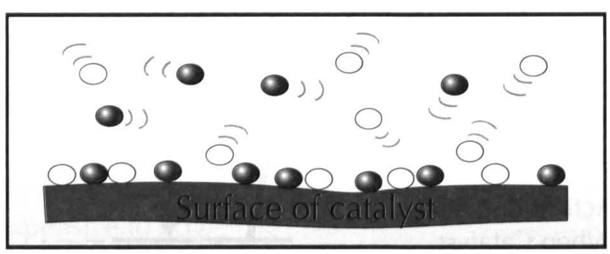
3) **LARGER SURFACE AREA** increases **collisions**

If one of the reactants is a solid then breaking it into smaller pieces will increase its surface area. This means the particles around it in the solution will have more area to work on, so there'll be more useful collisions.



4) **CATALYSTS** increase the **SUCCESSFUL collisions**

A solid catalyst works by giving the reacting particles a surface to stick to. They increase the number of SUCCESSFUL collisions by lowering the activation energy (see next page).

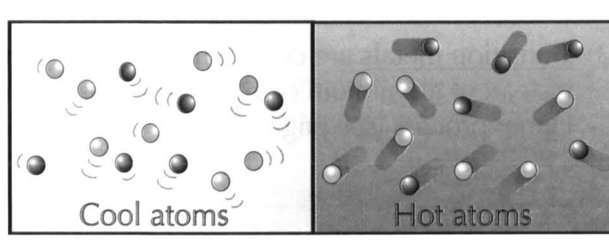


Faster collisions increase the rate of reaction

Faster collisions are ONLY caused by increasing the temperature

Reactions only happen if the particles collide with enough energy. At higher temperatures there are more particles colliding with enough energy to make the reaction happen.

This initial energy is known as the activation energy, and it's needed to break the original bonds.



It's easier to learn stuff when you know the reasons for it

Once you've learnt everything on these last two pages, the rates of reaction stuff should start making a lot more sense to you. The concept's fairly simple — the more often particles bump into each other, and the harder they hit when they do, the faster the reaction happens.

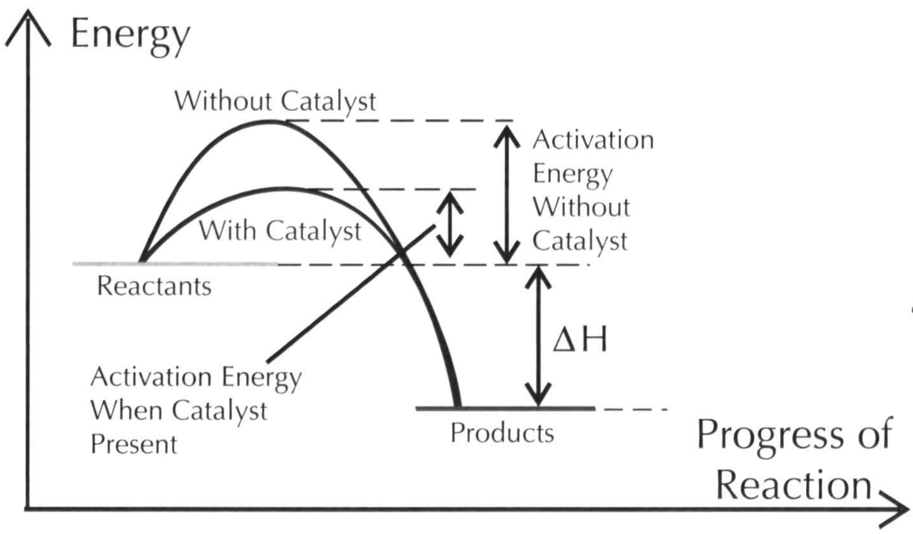
Catalysts

Many reactions can be speeded up by adding a catalyst.

> A catalyst is a substance which changes the speed of a reaction, without being changed or used up in the reaction.

Catalysts lower the activation energy

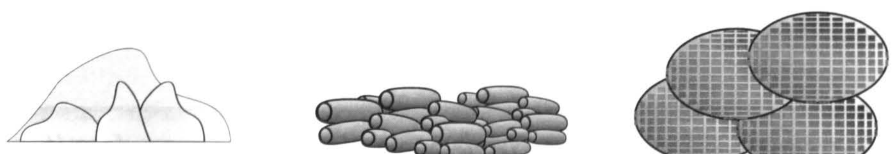
1) The activation energy is the minimum amount of energy needed for a reaction to happen.
2) It's a bit like having to climb up one side of a hill before you can ski / snowboard / sledge / fall down the other side.
3) Catalysts lower the activation energy of reactions, making it easier for them to happen.
4) This means a lower temperature can be used.



ΔH is the energy change.

Solid catalysts work best when they have a big surface area

1) Catalysts are usually used as a powder or pellets or a fine gauze.
2) This gives them a very large surface area to enable the reacting particles to meet up and react.
3) Transition metals are common catalysts in many industrial reactions, e.g. nickel can be used instead of aluminium oxide for cracking hydrocarbons (see page 39) and iron catalyses the Haber process (see page 167).



Catalyst Powder — Catalyst Pellets — Catalyst Gauzes

Catalysts

In industrial reactions, the main thing they're interested in is making a nice profit. Catalysts are helpful for this — they can reduce costs and increase the amount of product.

Catalysts help to **reduce costs** in industrial reactions

1) Catalysts are very important for commercial reasons — most industrial reactions use them.

2) Catalysts increase the rate of the reaction, which saves a lot of money simply because the plant doesn't need to operate for as long to produce the same amount of stuff.

3) Alternatively, a catalyst will allow the reaction to work at a much lower temperature. That reduces the energy used up in the reaction (the energy cost), which is good for sustainable development and can save a lot of money too.

Catalysts do sometimes have their **drawbacks** too

1) Nothing's perfect of course, and there are disadvantages to using catalysts. For starters, they can be very expensive to buy.

2) They often need to be removed from the product and cleaned each time the reaction is finished. They never get used up in the reaction though, so once you've got them you can use them over and over again.

3) Different reactions use different catalysts, so if you make more than one product at your plant you'll probably need to buy different catalysts for each of them.

4) Catalysts can be 'poisoned' by impurities, so they stop working — for example, sulfur can poison the iron catalyst used in the Haber process. That means you have to keep your reaction mixture very clean.

A big advantage of catalysts is that they can be used over and over

And they're not only used in industry... every useful chemical reaction in the human body is catalysed by a biological catalyst (an enzyme). If the reactions in the body were just left to their own devices, they'd take so long to happen, we couldn't exist. Quite handy then, these catalysts.

Warm-Up and Exam Questions

Try the questions below to check how much you've learnt about collision theory and catalysts.

Warm-Up Questions

1) According to the collision theory, what must happen in order for two particles to react?
2) What can you do to a reaction mixture to ensure more frequent collisions between particles?
3) Why does an increase in gas pressure increase the rate of a reaction between two gases?
4) Give a definition of a catalyst.
5) What is meant by the activation energy of a reaction?

Exam Questions

1 Hydrogen peroxide decomposes into water and oxygen. When lumps of a catalyst are dropped into a solution of hydrogen peroxide, bubbles of oxygen immediately start to form on the surface. Heating the solution increases the rate at which bubbles are formed.

(a) Use the collision theory to explain how heating increases the rate of the reaction.

(3 marks)

(b) Breaking the catalyst into smaller pieces also increases the rate. Explain why.

(2 marks)

(c) How else could the rate of the reaction be increased?

(1 mark)

2 Hydrocarbons can be cracked by passing their hot vapour over a nickel catalyst. The nickel used is in the form of small, hollow, cylindrical pellets. Using a catalyst reduces the operating temperature needed.

(a) Why do you think that the pellets used are hollow?

(1 mark)

(b) Why is it an advantage that a lower temperature can be used?

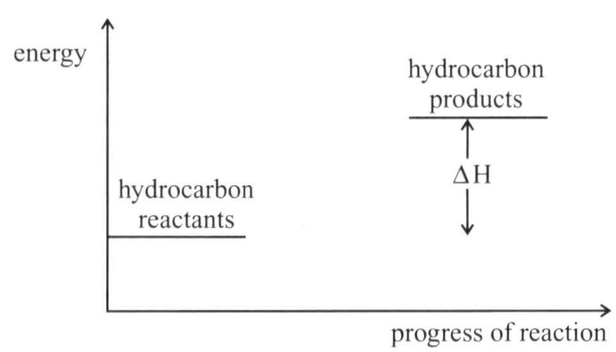
(2 marks)

(c) Suggest two possible drawbacks of using catalysts.

(2 marks)

(d) Copy and complete the diagram on the right and label it to show how the activation energy is different for catalysed and uncatalysed cracking of hydrocarbons.

(3 marks)



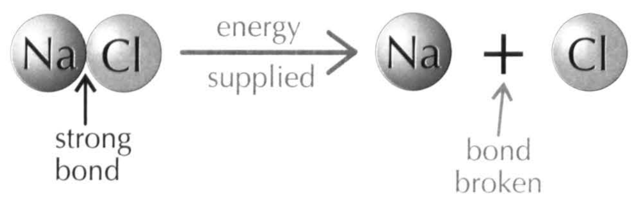
Energy Transfer in Reactions

In a chemical reaction, energy is usually transferred to or from the surroundings, and it's all about making and breaking bonds.

*Energy must always be **supplied** to **break bonds***

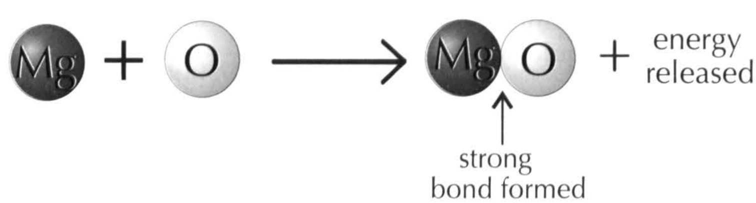
1) During a chemical reaction, old bonds are broken and new bonds are formed.
2) Energy must be supplied to break existing bonds — so bond breaking is an endothermic process.

BOND BREAKING – ENDOTHERMIC



3) Energy is released when new bonds are formed — so bond formation is an exothermic process.

BOND FORMING – EXOTHERMIC



*In an **exothermic** reaction, energy is **given out***

1) In an EXOTHERMIC reaction, the energy released in bond formation is greater than the energy used in breaking old bonds.

> An EXOTHERMIC reaction is one which GIVES OUT ENERGY to the surroundings, usually in the form of heat and usually shown by a RISE IN TEMPERATURE.

2) Burning fuels (COMBUSTION) gives out a lot of heat — it's very exothermic. That's because making new bonds in the products (water and carbon dioxide) gives out much more energy than it takes to break the bonds in the fuel.
3) Neutralisation reactions (acid + alkali) are also exothermic.

Energy Transfer in Reactions

In an endothermic reaction, energy is taken in from the surroundings. These are quite rare compared with exothermic reactions, but you still need to have a couple of examples memorised.

*In an **endothermic** reaction, energy is **taken in***

1) In an ENDOTHERMIC reaction, the energy required to break old bonds is greater than the energy released when new bonds are formed.

An ENDOTHERMIC reaction is one which TAKES IN ENERGY from the surroundings, usually in the form of heat and usually shown by a FALL IN TEMPERATURE.

2) Endothermic reactions are much less common than exothermic reactions.

3) Thermal decompositions are one good example:

Thermal decomposition of calcium carbonate

Heat must be supplied to break some of the bonds and make the compound decompose to form quicklime:

$$CaCO_3 \rightarrow CaO + CO_2$$

- A lot of heat energy is needed to make this happen.
- In fact, the calcium carbonate has to be heated in a kiln and kept at about 800 °C.
- It takes almost 18 000 kJ of heat to make 10 kg of calcium carbonate decompose.

That's pretty endothermic, I'd say.

4) Another example you could mention is photosynthesis. Think about it — the plant has to take in energy from the Sun to build sugars from carbon dioxide and water.

Exo = gives heat out, endo = takes heat in

This whole energy transfer thing is a fairly simple idea — don't be put off by the long words. Remember, "exo-" = exit, "-thermic" = heat, so an exothermic reaction is one that gives out heat. And "endo-" = erm... the other one. Okay, so there's no easy way to remember that one. Tough.

Bond Energies

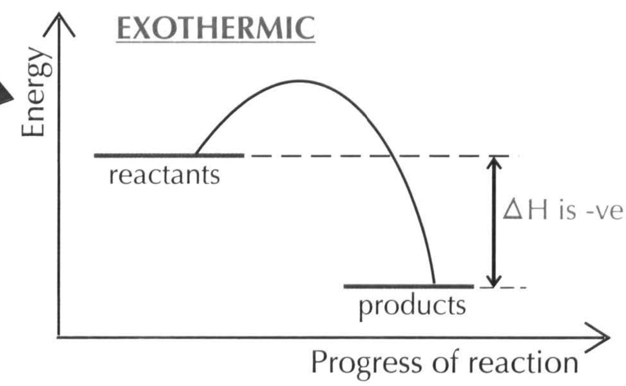
So you know what's meant by exothermic and endothermic — now things get a bit more technical.

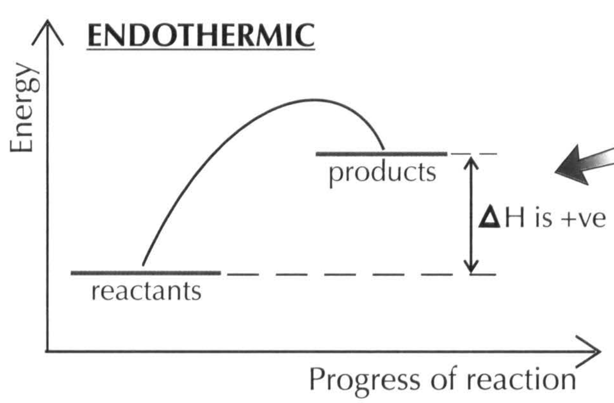
Energy level diagrams show if it's exo- or endothermic

In exothermic reactions, ΔH is –ve

1) This shows an exothermic reaction — the products are at a lower energy than the reactants. The difference in height represents the energy given out in the reaction (per mole). ΔH is negative here.
2) The initial rise in the line represents the energy needed to break the old bonds. This is the activation energy.

ΔH is the energy change.



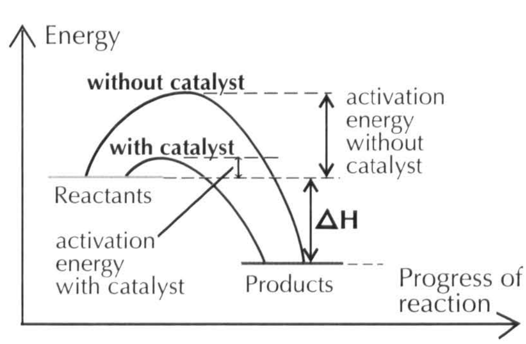


In endothermic reactions, ΔH is +ve

1) This shows an endothermic reaction because the products are at a higher energy than the reactants, so ΔH is positive.
2) The difference in height represents the energy taken in during the reaction.

Catalysts don't alter the overall energy change of a reaction

1) A catalyst reduces the activation energy needed for a reaction to happen (see p.154-155).
2) This is represented by the lower curve on the diagram showing a lower activation energy.
3) The overall energy change for the reaction, ΔH, remains the same though.



Bond Energies

You need to be able to work out ΔH for a particular reaction.

Bond energy calculations need to be practised

1) Every chemical bond has a particular bond energy associated with it.

2) This bond energy varies slightly depending on the compound the bond occurs in — but don't worry, you'll be given any you need to use in the exam.

3) You can use these known bond energies to calculate the overall energy change for a reaction.

You need to practise a few of these, but the basic idea is really very simple...

Example: The formation of HCl

Using known bond energies you can calculate the energy change for this reaction: $H_2 + Cl_2 \rightarrow 2HCl$

The bond energies you need are:

- H—H: +436 kJ/mol
- Cl—Cl: +242 kJ/mol
- H—Cl: +431 kJ/mol

1) Breaking one mole of H—H and one mole of Cl—Cl bonds requires:

   436 + 242 = 678 kJ

2) Forming two moles of H—Cl bonds releases 2 × 431 = 862 kJ.

3) Overall more energy is released than is used to form the products:

   862 – 678 = 184 kJ/mol released.

4) Since this is energy released, if you wanted to show ΔH you'd need to put a negative sign in front of it to indicate that it's an exothermic reaction, like this:

$$\Delta H = -184 \text{ kJ/mol}$$

You're given the bond energies, but you must know how to use them

I admit — it's a bit like maths, this. But think how many times you've heard energy efficiency mentioned over the last few years. Well, this kind of calculation is used in working out whether we're using resources efficiently or not. So even if it's not exciting, it's useful at least.

Measuring the Energy Content of Fuels

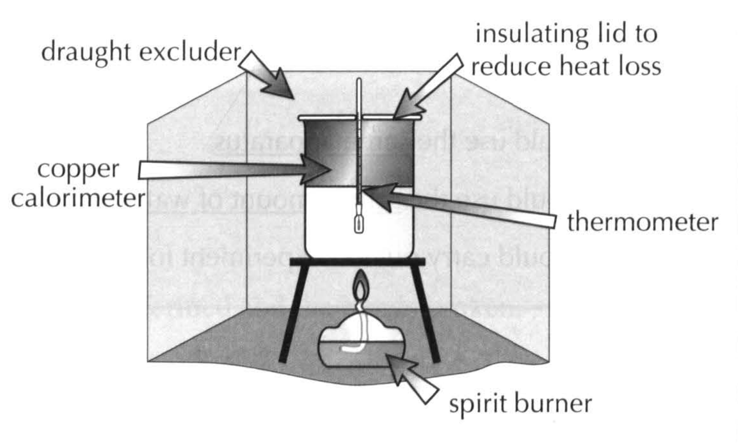
Different fuels give out different amounts of energy when they burn. One way to measure the energy content of a fuel is by using a none-too-fancy copper cup (or a "calorimeter", to give it its proper name).

Use *specific heat capacity* to calculate *energy transferred*

1) This "calorimetric" experiment involves heating water by burning a liquid fuel.
2) If you measure:
   (i) how much fuel you've burned, and
   (ii) the temperature change of the water,
   you can work out how much energy is supplied by each gram of fuel.
3) You also need to know water's specific heat capacity — this is the amount of energy needed to raise the temperature of 1 gram of water by 1 °C. The specific heat capacity of water is 4.2 J/g/°C — so it takes 4.2 joules of energy to raise the temperature of 1 g of water by 1 °C.
4) If you do the same experiment with different fuels, you can compare their energies transferred per gram. If a fuel has a higher energy content per gram, you need less fuel to cause the same temperature rise.

Calorimetric method — reduce heat loss as much as possible

1) It's dead important to make as much heat as possible go into heating up the water. Reducing draughts is the key here — use a screen to act as a draught excluder (and don't do it next to an open window).
2) Put some fuel into a spirit burner (or use a bottled gas burner if the fuel is a gas) and weigh the burner full of fuel.
3) Measure out, say, 200 cm^3 of water into a copper calorimeter.
4) Take the initial temperature of the water — then put the burner under the calorimeter and light the wick.
5) While the water's heating up, stir it every now and then to distribute the heat evenly.
6) When the heat from the burner has made the water temperature rise by 20–30 °C, blow out the spirit burner and make a note of the highest temperature the water reaches.
7) Re-weigh the burner and fuel.
8) If you're comparing two fuels, repeat the procedure with the second fuel.



Measuring the Energy Content of Fuels

Once you've done your calorimeter experiment and made a note of the mass of water used, the temperature change and the mass of fuel used, you can use some handy calculations to turn all that information into a nice value for energy per gram of fuel. Like this...

Three calculations to find the energy output per gram of fuel

1) You find the mass of fuel burned by subtracting the final mass of fuel and burner from the initial mass of fuel and burner. Simple.

2) The amount of energy transferred to the water is given by:

$$\text{Energy transferred (in J)} = \text{mass of water (in g)} \times \text{specific heat capacity of water } (= 4.2) \times \text{temperature change (in °C)}$$

3) Then the energy given out per gram of fuel is given by:

$$\text{Energy given out per gram (in J/g)} = \frac{\text{energy transferred (in J)}}{\text{mass of fuel burned (in g)}}$$

This is assuming that all the energy given out by the burning fuel is absorbed by the water — which in reality is actually quite unlikely. So you've got to do your best to make sure it's as true as possible, as described on the last page.

Make it a fair comparison by keeping conditions the same

1) To compare the energy content of different fuels you need to do the same experiment several times, but using a different fuel in the burner each time.
2) For the comparison to be fair, everything (except the fuel used) should be the same.
3) This means that:

   (i) you should use the same apparatus,

   (ii) you should use the same amount of water each time,

   (iii) you should carry out the experiment for the same length of time.

Make sure you know those two formulas

In the exam they might give you data from simple calorimetric experiments involving the combustion of fuel to compare, and you'll have to use it to say which fuel releases the most energy. Pretty easy.

Warm-Up and Exam Questions

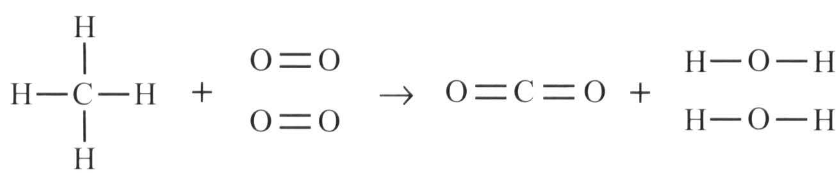
Bond energies can seem quite a strange idea at first. Hopefully these questions will get you used to it.

Warm-Up Questions

1) What word is used to describe a reaction which gives out heat?
2) An endothermic reaction happens when ammonium nitrate is dissolved in water. Predict how the temperature of the solution will change during the reaction.
3) Describe the type of energy change that happens when new chemical bonds form.
4) Which symbol is used to represent the energy change per mole in a reaction?
5) The specific heat capacity of water is 4.2 J/g/°C. What does this mean?

Exam Questions

1 When methane burns in air it produces carbon dioxide and water, as shown in the diagram:



The bond energies for each bond in the above molecules are given below.

Bond energies (kJ/mol): C–H +414 O=O +494 C=O +800 O–H +459

(a) How could you tell just by observing the reaction that it is exothermic?

(1 mark)

(b) (i) Which two types of bond are broken during the reaction?

(1 mark)

(ii) Which of these bonds needs the most energy before it will break? Suggest why its bond energy is higher.

(1 mark)

(c) Calculate an energy value (in kJ/mol) for:

(i) the total bonds broken.

(1 mark)

(ii) the total bonds formed.

(1 mark)

(iii) the difference between the bonds formed and the bonds broken.

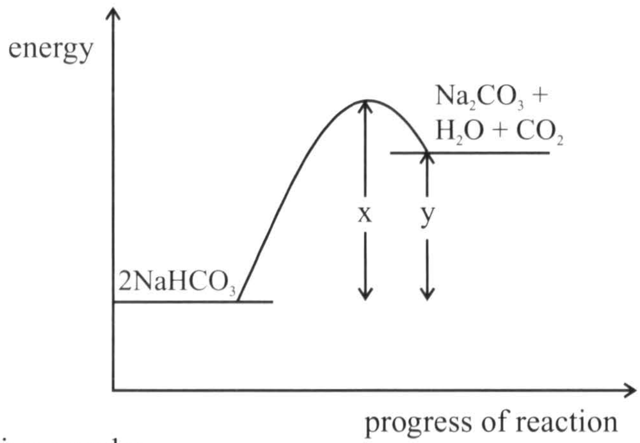
(1 mark)

(d) Use the values from part (c) to explain why the reaction is exothermic.

(1 mark)

Exam Questions

2 An energy level diagram for the decomposition of sodium hydrogencarbonate is shown.



(a) Complete the diagram by naming x and y.

(2 marks)

(b) Is this an exothermic or an endothermic reaction?
Give a reason for your choice.

(2 marks)

(c) Why doesn't the compound decompose at room temperature?

(1 mark)

3 The amount of energy produced by two different fuels was compared.
1 g of each fuel was burnt and the heat produced was used to increase the temperature of 100 cm^3 of water. The temperature rise for fuel A was 21 °C and for fuel B it was 32 °C. (The specific heat capacity of water is 4.2 J/g/K.)

(a) Why must the same volume of water be used each time?

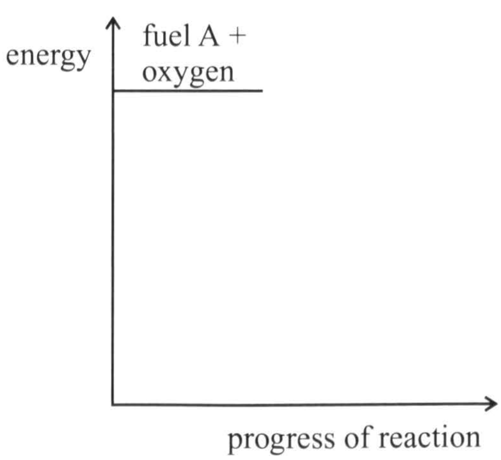
(1 mark)

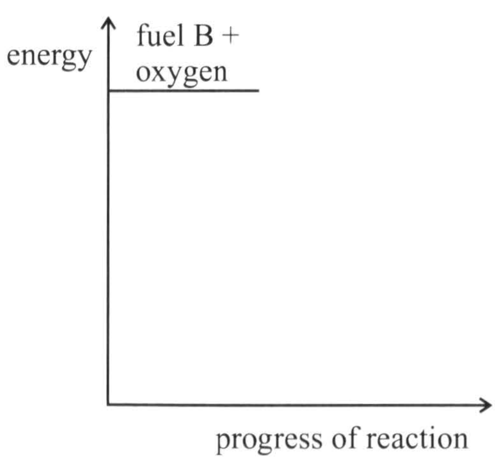
(b) Calculate the heat energy transferred to the water from Fuel A, if the water weighs 100 g.

(2 marks)

(c) Complete the diagrams to compare the energy changes caused by the two fuels.

(2 marks)





Reversible Reactions

In most of the reactions covered so far in this book, you mix some reactants and after a while you get the products. Makes sense. But of course, real life isn't always so simple.

Reversible reactions go in both directions

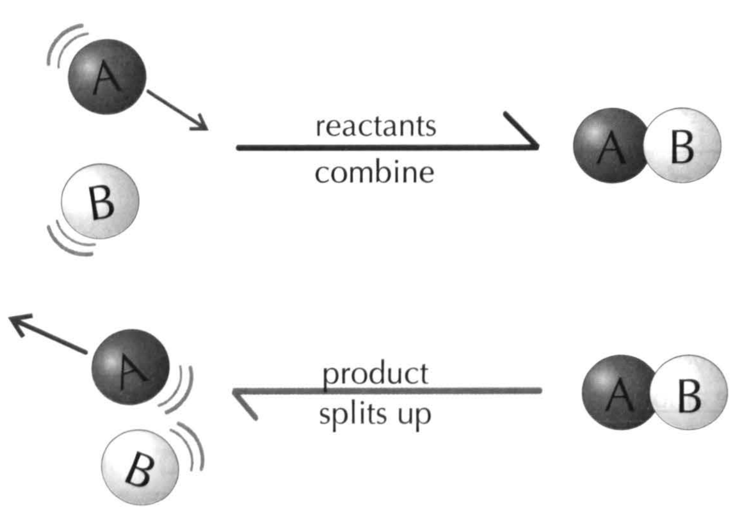
A reversible reaction is one where the products of the reaction can react with each other and convert back into the original reactants. In other words, it can go both ways.

A reversible reaction is one where the products of the reaction can themselves react to produce the original reactants:

$$A + B \rightleftharpoons C + D$$

Reversible reactions will reach dynamic equilibrium

1) If a reversible reaction happens in a closed system then a state of equilibrium will always be reached.
2) Equilibrium means that the relative (%) quantities of reactants and products will reach a certain balance and then stay there. (A 'closed system' just means that none of the reactants or products are able to escape.)
3) It is in fact a DYNAMIC EQUILIBRIUM, which means that the reactions are still taking place in both directions, but the overall effect is nil because the forward and reverse reactions cancel each other out. The reactions are taking place at exactly the same rate in both directions.



Reversible Reactions

Reversible reactions always reach equilibrium eventually, but by changing the conditions you can change the position of the equilibrium — i.e. shift it over so you end up with more products.

You can change **temperature** and **pressure** to get **more product**

1) In a reversible reaction the 'position of equilibrium' (the relative amounts of reactants and products) depends very strongly on the temperature and pressure surrounding the reaction.
2) If you deliberately alter the temperature and pressure you can move the 'position of equilibrium' to give more products and fewer reactants.

Temperature

All reactions are exothermic in one direction and endothermic in the other.

- If you raise the temperature, the endothermic reaction will increase to use up the extra heat.
- If you reduce the temperature, the exothermic reaction will increase to give out more heat.

Pressure

Many reactions have a greater volume on one side, either of products or reactants (greater volume means there are more molecules and less volume means there are fewer molecules).

- If you raise the pressure it will encourage the reaction which produces less volume.
- If you lower the pressure it will encourage the reaction which produces more volume.

Adding a **catalyst doesn't change** the equilibrium position

1) Catalysts speed up both the forward and backward reactions by the same amount.
2) So, adding a catalyst means the reaction reaches equilibrium quicker, but you end up with the same amount of product as you would without the catalyst.

Remember — catalysts DON'T affect the equilibrium position...

Changing the temperature always changes the equilibrium position, but that's not true of pressure. If your reaction has the same number of molecules on each side of the equation, changing the pressure won't make any difference at all to the equilibrium position (it still affects the rate of reaction though).

The Haber Process

This is an important industrial process. It produces ammonia (NH_3), which is used to make fertilisers.

Nitrogen and hydrogen are needed to make ammonia

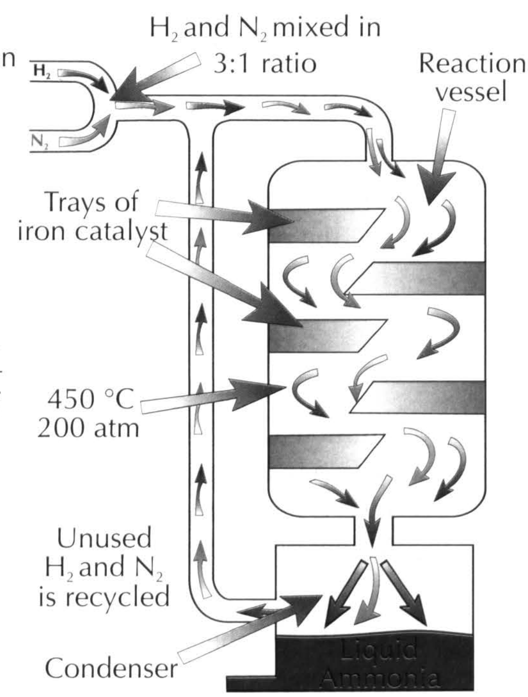
$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad (+\text{ heat})$$

1) The nitrogen is obtained easily from the air, which is 78% nitrogen (and 21% oxygen).
2) The hydrogen comes from natural gas or from other sources like crude oil.
3) Because the reaction is reversible — it occurs in both directions — not all of the nitrogen and hydrogen will convert to ammonia. The reaction reaches a dynamic equilibrium.

Industrial conditions: pressure = 200 atmospheres; temperature = 450 °C; catalyst: iron.

The reaction is reversible, so there's a compromise to be made

1) Higher pressures favour the forward reaction (since there are four moles of gas on the left-hand side for every two moles on the right).
2) So the pressure is set as high as possible to give the best percentage yield, without making the plant too expensive to build (it'd be too expensive to build a plant that'd stand pressures of over 1000 atmospheres, for example). Hence the 200 atmospheres operating pressure.
3) The forward reaction is exothermic, which means increasing the temperature will move the equilibrium the wrong way — away from ammonia and towards N_2 and H_2. So the yield of ammonia would be greater at lower temperatures.
4) The trouble is, lower temperatures mean a slower rate of reaction. So what they do is increase the temperature anyway, to get a much faster rate of reaction.
5) The 450 °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) The ammonia is formed as a gas but as it cools in the condenser it liquefies and is removed.
7) The unused hydrogen, H_2, and nitrogen, N_2, are recycled so nothing is wasted.

The iron catalyst speeds up the reaction and keeps costs down

1) The iron catalyst makes the reaction go faster, which gets it to the equilibrium proportions more quickly. But remember, the catalyst doesn't affect the position of equilibrium (i.e. the % yield).
2) Without the catalyst the temperature would have to be raised even further to get a quick enough reaction, and that would reduce the % yield even further. So the catalyst is very important.

Warm-Up and Exam Questions

It's easy to predict what will happen to the position of an equilibrium if you just remember that it always shifts to oppose any change in the conditions. Bear that in mind as you work through these questions.

Warm-Up Questions

1) What could you do to speed up a reaction without changing the position of equilibrium?
2) What can you say about the forward and backward reaction rates at dynamic equilibrium?
3) How does increasing the pressure alter the equilibrium position of a reaction which produces fewer moles of gas molecules in the forward direction?
4) Name the catalyst used in the Haber process.
5) What happens to leftover reactants that are not converted to product in the Haber process?

Exam Questions

1 When calcium carbonate is heated to a high temperature in a closed system, an equilibrium is reached:

$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ (ΔH is +ve)

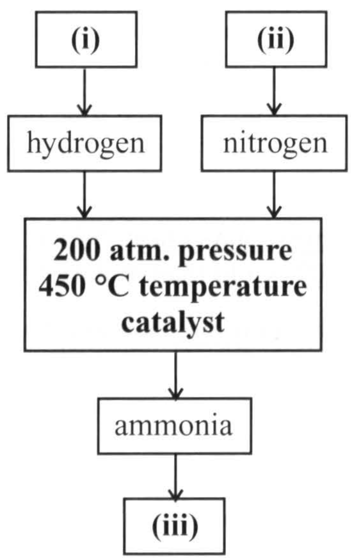
(a) Why is a closed system needed for this reaction to reach equilibrium?

(1 mark)

(b) Give two ways in which the equilibrium could be changed to increase the proportion of products present.

(2 marks)

2 Ammonia, NH_3, is made by combining nitrogen and hydrogen at a pressure of 200 atm, a temperature of 450 °C and in the presence of a catalyst.
A flow diagram is shown for the reaction:



(a) Write labels for boxes (i) and (ii) to show where the nitrogen and hydrogen come from.

(2 marks)

(b) Write an equation with state symbols for the reaction between nitrogen and hydrogen.

(3 marks)

(c) Write a label for box (iii) to show what ammonia is used to make.

(1 mark)

(d) The reaction is exothermic. Explain why a high temperature is still used.

(2 marks)

Revision Summary for Section Nine

Well, I don't think that was too bad. Four things affect the rate of reactions, there are loads of ways to measure reaction rates and it's all explained by collision theory. Reactions can be endothermic or exothermic, and quite a few of them are reversible. Easy. Ahem.

Anyway, here are some more of those nice questions that you enjoy so much. If there are any you can't answer, go back to the appropriate page, do a bit more learning, then try again.

1) What are the four factors that affect the rate of a reaction?
2) Describe three different ways of measuring the rate of a reaction.
   Give one advantage and one disadvantage of each method.
3)* A student carries out an experiment to measure the effect of surface area on the reaction between marble and hydrochloric acid. He measures the mass of gas given off at regular intervals.
   He uses four samples for his experiment:
   - Sample A – 10 g of powdered marble
   - Sample B – 10 g of small marble chips
   - Sample C – 10 g of large marble chips
   - Sample D – 5 g of powdered marble

   Sketch a typical set of graphs for this experiment, with time on the horizontal axis and 'mass of gas evolved' on the vertical axis.
4) Explain how the frequency of collisions between particles is increased by:
   a) increasing the temperature,
   b) using more concentrated solutions,
   c) increasing the surface area of the reactants.
5) Catalysts are often used in the form of a powder or fine gauze. Explain why.
6) Give one advantage and one disadvantage of using a catalyst in an industrial process.
7) What is an exothermic reaction? Give two examples.
8) Give an example of an endothermic reaction.
9) a) Draw energy level diagrams for exothermic and endothermic reactions.
   b) Explain how bond breaking and bond forming relate to these diagrams.
10) Mark the activation energy on your exothermic energy level diagram above.
11) How does a catalyst affect: a) activation energy, b) overall energy change for a reaction?
12)*a) Calculate the energy change for the following reaction: $2H_2 + O_2 \rightarrow 2H_2O$
   You need these bond energies: H–H: +432 kJ/mol, O=O: +494 kJ/mol, O–H: +459 kJ/mol
   Hint: There are 2 O–H bonds in each molecule of water.
   b) Is this an exothermic or an endothermic reaction?
13)*A calorimeter was used to measure how much energy was released when pentane was burnt.
   It takes 4.2 joules of energy to heat 1 g of water by 1 °C.
   Use the following data to calculate the amount of energy per gram of pentane.

Mass of empty copper can	64 g
Mass of copper can + water	116 g

Initial temp. of water	17 °C
Final temp. of water	47 °C

Mass of sprit burner + pentane before burning	97.72 g
Mass of sprit burner + pentane after burning	97.37 g

14) Describe one precaution you should take when using the calorimetric method.
15) What is a reversible reaction? Explain what is meant by a dynamic equilibrium.
16) How does changing the temperature and pressure of reversible reactions alter the position of the equilibrium?
17) How does this influence the choice of pressure for the Haber process?

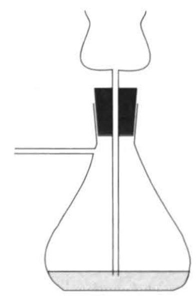
* Answers on page 261.

Gas Tests

There are lots of clever ways of testing for different gases. Sometimes the hardest part is collecting the gas in the first place...

Gases can be collected in several ways

A side-arm flask is the standard apparatus to use when you're trying to collect gases. But what you connect the side arm to depends on what it is you're trying to collect.

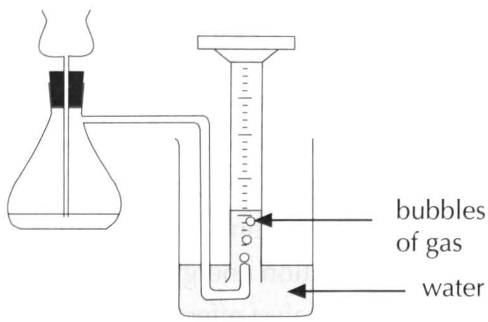


1) Gas syringe

You can use a gas syringe to collect pretty much any gas.

2) Collection over water

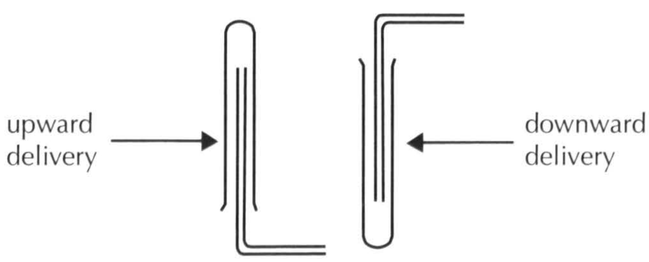
1) You can use a delivery tube to bubble the gas into an upside-down measuring cylinder or gas jar filled with water.
2) This method only works for insoluble gases — it's no good for collecting things like hydrogen chloride or ammonia because they just dissolve in the water.



3) Upward or downward delivery

This all depends on the density of the gas relative to the density of air.

1) Use upward delivery to collect 'lighter than air' gases (e.g. H_2).
2) Use downward delivery to collect 'heavier than air' gases (e.g. CO_2, Cl_2).



Gas Tests

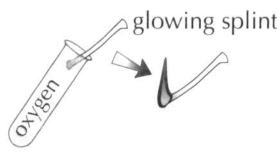
It's no good carefully collecting a gas if you've then got no idea what you've collected...

There are tests for five 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.)

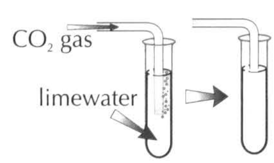
2) Oxygen



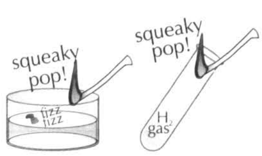
Oxygen relights a glowing splint.

3) Carbon dioxide

Carbon dioxide turns limewater cloudy — just bubble the gas through a test tube of limewater and watch what happens.



4) Hydrogen



Hydrogen makes a "squeaky pop" with a lighted splint. (The noise comes from the hydrogen burning with the oxygen in the air to form H_2O.)

5) Ammonia

Ammonia turns damp red litmus paper blue. (It also has a very strong smell.)

It's probably a good idea to decide how you'd collect these five gases

The rules for collection may seem a bit complicated, but they're not too bad. Think "is it soluble?" and "is it lighter or heavier than air?". That'll do you. The only way to handle the gas tests is to get on and learn them. Cover up the page, and scribble down the tests. See what you know.

Tests for Positive Ions

Say you've got a compound, but you don't know what it is. Well, you'd want to identify it... that's only natural. And that's what the next pages are all about. Tests for positive ions first...

*Add **sodium hydroxide** and look for a **coloured precipitate***

This is the first test for positive ions you need to know about. It's a bit complicated, so concentrate...

1) Many metal hydroxides are insoluble and precipitate out of solution when formed. Some of these hydroxides have a characteristic colour.
2) 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.
3) If you get a coloured insoluble hydroxide you can then tell which metal was in the compound.

Usually, anyway... the result for NH_4^+ is a bit different.

CALCIUM (Ca^{2+}) gives a **WHITE** precipitate. The ionic reaction is:

$$Ca^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Ca(OH)_{2(s)}$$

COPPER(II) (Cu^{2+}) gives a **BLUE** precipitate. The ionic reaction is:

$$Cu^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Cu(OH)_{2(s)}$$

IRON(II) (Fe^{2+}) gives a **SLUDGY GREEN** precipitate. The ionic reaction is:

$$Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2(s)}$$

IRON(III) (Fe^{3+}) gives a **RED-BROWN** precipitate. The ionic reaction is:

$$Fe^{3+}_{(aq)} + 3OH^-_{(aq)} \rightarrow Fe(OH)_{3(s)}$$

ALUMINIUM (Al^{3+}) first gives a **WHITE** precipitate, then redissolves in excess NaOH to give a colourless solution. The ionic reactions are:

$$Al^{3+}_{(aq)} + 3OH^-_{(aq)} \rightarrow Al(OH)_{3(s)}$$
then $$Al(OH)_{3(s)} + OH^-_{(aq)} \rightarrow Al(OH)^-_{4(s)}$$

AMMONIUM (NH_4^+) doesn't give any precipitate, but on heating it gives off ammonia, which has a distinctive smell (see also the test for ammonia on page 171).

Tests for Positive Ions

First it's time to focus a bit more on those ionic equations you met on the last page.
Then there's another type of test for positive ions — luckily this one's a bit simpler.

Ionic equations show just the useful bits of reactions

The reactions in the blue boxes on the last page are ionic equations. Ionic equations are 'half' a full equation, if you like. They just show the bit of the equation you're interested in — nothing else.

Example:

$$Ca^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Ca(OH)_{2(s)}$$

1) This shows the formation of (solid) calcium hydroxide from the calcium ions and the hydroxide ions in solution. And it's the formation of this that helps you to identify the compound.
2) The full equation in the above reaction would be (if you started off with, for example, calcium chloride):

   $$CaCl_{2(aq)} + 2NaOH_{(aq)} \longrightarrow Ca(OH)_{2(s)} + 2NaCl_{(aq)}$$

3) But the formation of sodium chloride is of no great interest here — it's not helping to identify the compound, after all.
4) So the ionic equation just concentrates on the good bits.

Flame tests — spot the colour

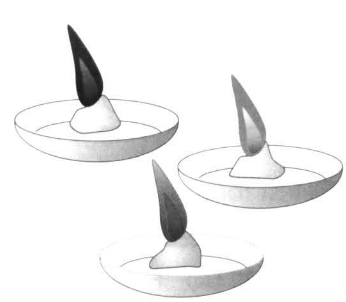
Compounds of some metals give a characteristic colour when heated.
This is the idea behind flame tests — see also page 98.

(i) **Sodium**, Na^{+}, gives an orange/yellow flame.

(ii) **Potassium**, K^{+}, gives a lilac flame.

(iii) **Calcium**, Ca^{2+}, gives a brick-red flame.

(iv) **Copper**, Cu^{2+}, gives a blue-green flame.



Ionic equations might seem odd at first, but they do save time

The flame test gubbins ought to be familiar to you already. Pay full attention to the new bits — you have to learn the ionic equations for each test with NaOH as well as all the test results. Just think of an ionic equation as a bit like Match of the Day — an edited highlights package.

Tests for Negative Ions

It's not just positive ions you can test for. Yep, you can also test for negative ions. So the fun goes on...

Hydrochloric acid can help detect carbonates and sulfites

The gases given off when carbonates (CO_3^{2-}) and sulfites (SO_3^{2-}) react with HCl can be used to identify these substances.

SO_3^{2-} = sulfite

SO_4^{2-} = sulfate

Carbonates give off CO_2 with HCl

With dilute hydrochloric acid, carbonates (CO_3^{2-}) give off carbon dioxide.

$$CO_3^{2-}{}_{(s)} + 2H^+{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

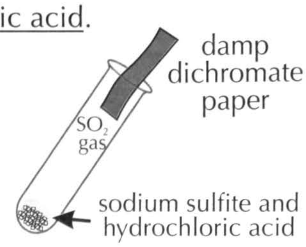
You can test for carbon dioxide using limewater — see page 171.

Sulfites give off SO_2 with HCl

Sulfites (SO_3^{2-}) give off sulfur dioxide when mixed with dilute hydrochloric acid.

$$SO_3^{2-}{}_{(s)} + 2H^+{}_{(aq)} \rightarrow SO_{2(g)} + H_2O_{(l)}$$

You can test for sulfur dioxide using damp potassium dichromate(VI) paper. The paper turns from orange to green.



Test for sulfates with HCl and barium chloride

Sulfate ions (SO_4^{2-}) produce a white precipitate

To test for a sulfate ion (SO_4^{2-}), add dilute HCl, followed by barium chloride solution, $BaCl_2(aq)$.

$$Ba^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$$

A white precipitate of barium sulfate means the original compound was a sulfate.

(The hydrochloric acid is added to get rid of any traces of carbonate or sulfite ions before you do the test. Both of these would also produce a precipitate, so they'd confuse the results.)

Tests for Negative Ions

You can also test for halides and nitrates. Make sure you know how.

Test for halides (Cl^-, Br^-, I^-) with nitric acid and silver nitrate

To test for halide ions, add dilute nitric acid (HNO_3), followed by silver nitrate solution, $AgNO_3$(aq).

A chloride gives a white precipitate of silver chloride.

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$$

A bromide gives a cream precipitate of silver bromide.

$$Ag^+_{(aq)} + Br^-_{(aq)} \rightarrow AgBr_{(s)}$$

An iodide gives a yellow precipitate of silver iodide.

$$Ag^+_{(aq)} + I^-_{(aq)} \rightarrow AgI_{(s)}$$

(Again, the acid is added to get rid of carbonate or sulfite ions before the test. You use nitric acid in this test, though, not HCl.)

The test for nitrates (NO_3^-) produces ammonia

1) Mix some of your mystery compound with a little aluminium powder.
2) Then add a few drops of sodium hydroxide solution and heat the mixture.
3) If you started off with a nitrate, it'll be reduced to ammonia.
4) As always, test for ammonia using your nose or, better, damp red litmus paper (which will turn blue).

You've just got to get your head down and learn this, I'm afraid

Don't just stare at these pages on tests for negative ions till your eyes swim and you don't want to see the word "precipitate" ever again. They're handily divided into four subsections, so learn them that way.

Warm-Up and Exam Questions

Lots to remember on those six pages. Try these questions to see how good your memory really is.

Warm-Up Questions

1) Why can't you collect a sample of ammonia by collection over water?
2) Suggest a gas that could be collected using downward delivery.
3) How would you test for ammonia gas?
4) Name the gas that will relight a glowing splint.
5) What colour is the precipitate formed when sodium hydroxide is added to a solution of copper(II) ions?

Exam Questions

1 Brian reacted hydrochloric acid with four different chemicals.
In each case a gas was produced.

(a) With substance A, the gas formed turned limewater cloudy. What was the gas?

(1 mark)

(b) With substance B, the gas formed burned with a squeaky pop. What was the gas?

(1 mark)

(c) With substance C, the gas turned a piece of damp litmus paper white. What was it?

(1 mark)

(d) Next, Brian reacted hydrochloric acid with sodium sulfite. This reaction produced the gas sulfur dioxide. Copy and complete the ionic equation for this reaction:

$$\mathbf{SO_3^{2-}(s) + \_\_\_\_\_\_\_\_(aq) \rightarrow \_\_\_\_\_\_\_\_(g) + H_2O(l)}$$

(2 marks)

2 Kelly carried out flame tests on compounds of four different metal ions.
Copy and complete the table below showing her results.

Flame colour	Metal ion
blue-green	
	K^+
orange-yellow	
	Ca^{2+}

(4 marks)

3 You are provided with a solution of a halide salt. Describe how you would test this solution to identify whether the solution is of a chloride, bromide or iodide.

(4 marks)

Exam Questions

4 William conducted a series of tests on several solutions of ionic compounds to identify the positive ions. Complete the table by writing the correct symbol for the positive ion that William has identified in each case.
The first one has been done for you.

TEST	OBSERVATION	ION
sodium hydroxide solution	reddish-brown precipitate	Fe^{3+}
sodium hydroxide solution	white precipitate that redissolves with excess sodium hydroxide	
sodium hydroxide solution	no precipitate, but when heated a strong-smelling gas is released	
sodium hydroxide solution	blue precipitate	
sodium hydroxide solution and then flame test	white precipitate with sodium hydroxide and brick-red flame	

(4 marks)

5 A bottle of a chemical solution is labelled 'iron(II) sulfate'.

(a) Describe a chemical test to confirm that the solution contains iron(II) ions.

(2 marks)

(b) Describe a chemical test to confirm that the solution contains sulfate ions.

(2 marks)

6 The table below shows the results of a series of chemical tests conducted on two unknown compounds, X and Y.

TEST	OBSERVATION	
	COMPOUND X	COMPOUND Y
sodium hydroxide solution	blue precipitate	no precipitate
hydrochloric acid & barium chloride solution	no precipitate	no precipitate
flame test	blue-green flame	lilac flame
aluminium powder & NaOH solution (any gas released on heating is tested with damp red litmus)	litmus paper stays red	litmus paper turns blue
nitric acid & silver nitrate solution	white precipitate	no precipitate

(a) What is the chemical name of compound X?

(2 marks)

(b) What is the chemical name of compound Y?

(2 marks)

Tests for Acids and Alkalis

The other ions you need to be able to test for are H^+ and OH^- ions — that is, acids and alkalis.

An *indicator* is just a *dye* that changes *colour*

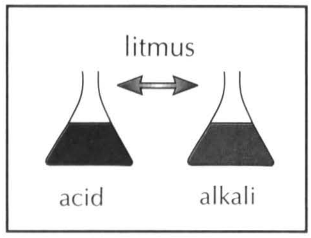
The dye in an acid/base indicator changes colour depending on whether it's above or below a certain pH.

Common indicators are:

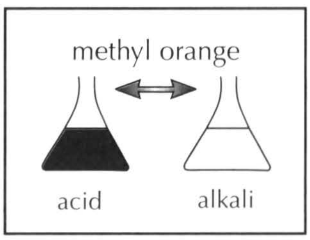
1) Litmus

Testing for $H^+(aq)$ and $OH^-(aq)$ ions can be done using red or blue litmus indicator.

- Blue litmus turns red if lots of $H^+(aq)$ ions are present — i.e. if the solution is an acid.
- Red litmus turns blue if lots of $OH^-(aq)$ ions are present — i.e. if the solution is an alkali.



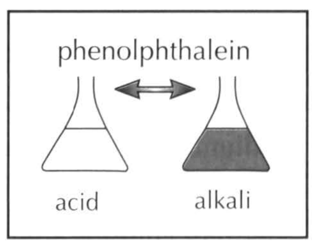
2) Methyl orange



Methyl orange is yellow in alkalis but red in acids.

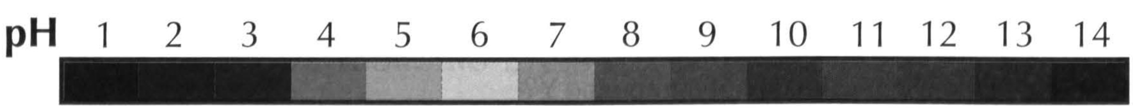
3) Phenolphthalein

Phenolphthalein is pink in alkalis but colourless in acids.



Universal indicator is useful for estimating pH

Universal indicator is a very useful combination of dyes which gives the colours shown below. It's very good for estimating the pH of a solution. (See page 132 for more.)



Tests for Acids and Alkalis

Indicators aren't the only way to test for acids and alkalis.

Add the substance to a **metal** to test for **acids (H^+ ions)...**

When an acid reacts with a metal, it gives off hydrogen gas.

This is a typical acid reaction — a hallmark of an acid, if you like.

The metal gives electrons to the H^+ ions, which then form hydrogen gas.

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$$

Not all metals will react with acids — it's only the ones that are more reactive than hydrogen. It's good to use a fairly reactive metal, like magnesium — so that you get plenty of hydrogen even for a very dilute or weak acid.

You can use the 'squeaky pop' test to detect the hydrogen (see page 171).

...or add it to a **carbonate**, also to test for **H^+ ions**

Alternatively, you could react the substance with a carbonate. When an acid reacts with a carbonate, it gives off carbon dioxide gas:

$$CO_3^{2-}{}_{(s)} + 2H^+_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

You can test for the carbon dioxide using limewater (see page 171).

Heat it with an **ammonium salt** to test for **alkalis (OH^- ions)**

Hydroxides give off ammonia gas when heated gently with an ammonium compound.

$$NH_4^+{}_{(aq)} + OH^-_{(aq)} \rightarrow NH_{3(g)} + H_2O_{(l)}$$

You should know how to test for ammonia gas by now — but if you've forgotten, look back at page 171.

Remember that acids mean H^+ ions and alkalis mean OH^- ions

The stuff about litmus, phenolphthalein, methyl orange and universal indicator is just easy facts. Learn it. And don't think you can get away without learning the tests for H^+ ions and the ammonium test for OH^- ions — including the equations. They could mean some nice easy marks come exam time.

Tests for Organic Compounds

The previous pages were about testing inorganic compounds (things not built around a chain of carbon atoms). But your mystery substance might just as easily be organic. In that case, here's what you do...

*Organic compounds **burn** when **heated***

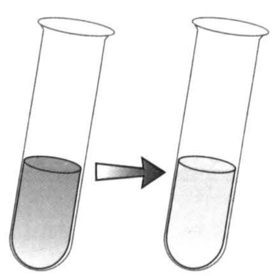
1) Organic compounds burn in air, with a yellowy-orange and/or blue flame.
2) The greater the proportion of carbon in the compound, the more yellow and smoky the flame is.
3) When there's plenty of air available, burning a hydrocarbon produces carbon dioxide and water.
4) If the amount of air is reduced, then carbon monoxide (a poisonous gas), and carbon (soot) can also be produced.
5) Solid organic compounds will char — in other words, their surface will get scorched with black marks of carbon.

A hydrocarbon is an organic compound containing only carbon and hydrogen, remember.

*Compounds with **C=C** bonds **decolourise bromine water***

The test for C=C double bonds is a piece of cake (though not literally).

1) If your organic compound is unsaturated (i.e. it has double or triple bonds between carbon atoms), it'll decolourise bromine water.
2) If your organic compound is saturated (i.e. there are no double or triple bonds), the bromine water will stay brown.



...so this one's unsaturated.

3) You can do this test on margarine, which has C=C bonds. Shake 1 cm^3 of bromine water with a small amount of melted margarine, and the bromine water decolourises.

Tests for Organic Compounds

You can do more than just tell if an organic compound is saturated or unsaturated — you can even work out its empirical formula with a few simple(ish) calculations.

Find the **empirical formula** of an organic compound by burning it

An empirical formula shows the ratios of all the elements in a substance (see also page 121).

It's possible to work out the empirical formula of an organic compound by burning a known mass of it completely in oxygen, and measuring the masses of all the products.

With a hydrocarbon, all the carbon ends up in CO_2 and all the hydrogen ends up in water. So...

Step 1 Find the mass of each element in the compound.

- To find the mass of carbon in the compound, multiply the mass of CO_2 produced by the proportion of C in CO_2.
- To find the mass of hydrogen in the compound, multiply the mass of H_2O produced by the proportion of H in H_2O.

Using relative atomic masses, the proportion of C in CO_2 is $12 \div 44 = 0.2727...$

And the proportion of H in H_2O is $2 \div 18 = 0.1111...$

Step 2 Divide these masses of C and H by the atomic masses of C and H (to find the no. of moles).

Step 3 Divide both answers by the smallest one to get the simplest ratio of atoms of each element.

Example:

0.4 g of an organic hydrocarbon is burnt completely in oxygen. 1.1 g of carbon dioxide and 0.9 g of water are formed. What is the compound's empirical formula?

- Step 1 — Find the mass of carbon in the compound: $1.1 \times (12 \div 44) = 0.3$ g
  Do the same for hydrogen: $0.9 \times (2 \div 18) = 0.1$ g
- Step 2 — The relative atomic mass of carbon is 12, so: $0.3 \div 12 = 0.025$ mol
  The relative atomic mass of hydrogen is 1, so: $0.1 \div 1 = 0.1$ mol
- Step 3 — Divide the biggest answer by the smallest one to get the ratio of carbon to hydrogen:
  The simplest whole number ratio of atoms of each element is $0.1 \div 0.025 = 4$
  (meaning there is 1 carbon to 4 hydrogens).

  This gives an empirical formula for this compound of CH_4.

A bit of practice and you'll soon be breezing through these

Finding an empirical formula involves an awful lot of sums. Sure, they're simple sums taken one by one, but it'd be all too easy to get confused, do them in the wrong order, and end up with completely the wrong answer. Learn the three steps and follow them — mass, then moles, then ratio.

Instrumental Methods

Nowadays you can turn to machines to do the donkey work of identifying substances, if need be.

Machines can also analyse unknown substances

1) Machines are useful for medical purposes, police forensic work, environmental analysis, drugs testing, analysis of products in industry, and so on.
2) Rapid advances in electronics and computing have made more advanced analysis possible.

Advantages of using machines

'Lab methods' means doing tests like the ones earlier in the section.

- Can be operated by technicians. Lab methods need trained chemists to do everything.
- More accurate than lab methods, and can detect even the tiniest amounts of substances.
- Much faster than lab methods, and tests can be automated.

Disadvantages of using machines

- It's very expensive to buy, run and maintain the machines.

Atomic absorption spectroscopy identifies metals

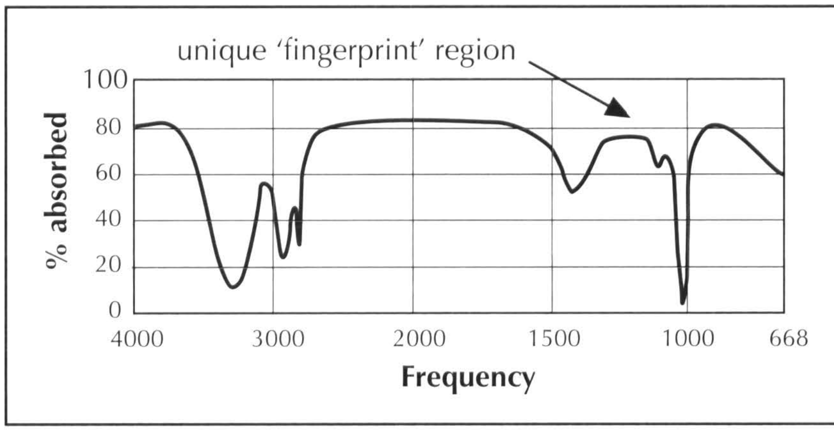
1) Atomic absorption spectroscopy is a bit like a flame test machine, and it's used to identify metals.
2) The patterns of light absorbed by the metals in the sample are analysed.
3) Each metal present in the sample produces a different pattern.
4) It's much faster and much more reliable than anything that can be done with the human eye.
5) The steel industry uses atomic absorption spectroscopy to check the composition of the steels. (Each kind of steel has to have the right composition to make it suitable for its particular use.)
6) This only takes minutes, compared to days with the lab method.

Instrumental Methods

Other techniques identify **elements** or **compounds**

1) Infrared (IR) Spectroscopy

This identifies which frequencies of infrared radiation are absorbed — the pattern of absorbance is unique for every compound. This 'fingerprint' allows identification of individual compounds.



2) Ultraviolet (UV) Spectroscopy

This is similar to infrared spectroscopy, but with ultraviolet light instead of infrared.

3) Nuclear Magnetic Resonance (NMR) Spectroscopy

This method is used for organic compounds. It shows what atoms the hydrogen atoms are connected to. This helps find the structure of the molecule, by telling you if there are –OH groups, $-NH_2$ groups, etc.

4) Gas-Liquid Chromatography

This uses a similar principle to paper chromatography. It's used to identify gases and liquids.

5) Mass Spectrometry

This method can be used for both elements and compounds. It tells you the mass of each molecule or particle. For elements, this tells you exactly what element you've got, and for larger molecules the mass is a good clue.

Warm-Up and Exam Questions

Time to test your knowledge again — first a warm-up, and then some very realistic exam questions.

Warm-Up Questions

1) What ion is present in all acids?
2) What colour does litmus indicator turn in alkaline solutions?
3) What gas is given off when an acid reacts with a metal?
4) An organic solid is heated strongly. How would the appearance of its surface change?
5) Suggest an instrumental method that could be used to identify a solid non-metal element.

Exam Questions

1 One way to test whether a solution is acidic is to react it with a carbonate.

(a) What gas is released when an acid reacts with a carbonate?

(1 mark)

(b) Describe a test to identify the gas released in this reaction.

(2 marks)

(c) Copy and complete the ionic equation for the reaction.

$$\mathbf{CO_3^{2-}(s) + \_\_\_\_\_\_(aq) \rightarrow \_\_\_\_\_\_(g) + H_2O(l)}$$

(2 marks)

2 Alkaline solutions contain hydroxide ions. Hydroxide ions in a solution can be detected by adding an ammonium compound to the solution and then gently heating the mixture.

(a) If hydroxide ions were present in a solution, what gas would be released if the above procedure were followed?

(1 mark)

(b) Complete the ionic equation for the reaction.

$$\mathbf{NH_4^{+}(aq) + \_\_\_\_\_\_(aq) \rightarrow \_\_\_\_\_\_(g) + H_2O(l)}$$

(2 marks)

(c) Describe a test to identify the gas released in this reaction.

(1 mark)

3 A sample of an organic hydrocarbon was burnt completely in air. 4.4 g of carbon dioxide and 1.8 g of water were formed.

(a) Calculate the number of moles of carbon in the hydrocarbon. (Relative atomic masses: H = 1, C = 12, O = 16.)

(2 marks)

(b) Calculate the number of moles of hydrogen in the hydrocarbon.

(2 marks)

(c) Use your answers to work out the empirical formula of the hydrocarbon.

(1 mark)

Exam Questions

4 An organic hydrocarbon was burnt completely in air. 1.1 g of carbon dioxide and 0.675 g of water were formed.

(a) Calculate the empirical formula of the hydrocarbon. Show all your working out. (Relative atomic masses: H = 1, C = 12, O = 16)

(3 marks)

(b) The hydrocarbon is an alkane. Describe a chemical test that could be used to demonstrate that it is saturated.

(2 marks)

5 Meg is a materials scientist and she works with metals in order to try to develop new catalysts. She frequently uses instrumental methods to test the quality and composition of metal samples.

(a) One of the techniques that Meg uses to test her samples is atomic absorption spectroscopy. Which of the following best describes how this technique works?

A Ultraviolet light is used to excite electrons in the metallic 'sea' of free electrons. The amount of energy absorbed by these electrons is different for each metal.

B The metal sample is vaporised and a type of gas chromatography is used to identify the different metals in the sample.

C The patterns of light absorbed by the metals in a sample are analysed — each metal produces its own characteristic pattern.

(1 mark)

(b) Give two advantages of using modern machine-based instrumental methods compared to older, lab-based methods.

(2 marks)

6 Ethanoic acid, $C_2H_4O_2$, is an organic acid. It is both an organic compound and an acid.

(a) Name an indicator that could be used to show that ethanoic acid is an acid. What colour change would you expect to see?

(1 mark)

(b) When ethanoic acid reacts with a metal such as magnesium, a gas is produced.

(i) What is the name of this gas?

(1 mark)

(ii) Write an ionic equation to show how the metal's electrons are used in the formation of the gas.

(1 mark)

(c) Propanoic acid, $C_3H_6O_2$, is another organic acid. Suggest two instrumental methods that could be used to distinguish between ethanoic acid and propanoic acid.

(2 marks)

Revision Summary for Section Ten

And that's it... the end of another section. Which means it's time for some more questions. There's no point in trying to duck out of these — they're the best way of testing that you've learned everything in this topic. For these questions you'll need to know some common chemical tests. If you find you don't know them, look back in the book. If you can't do all this now, you won't be able to in the exam either.

1) Which of these gases can be collected by downward delivery — carbon dioxide, chlorine, hydrogen?
2) What's the test for each of the following:
   a) chlorine, b) hydrogen, c) oxygen, d) carbon dioxide, e) ammonia?
3) Crystals of an ionic compound are heated in a Bunsen flame and produce a brick-red flame. What positive ion is present?
4) A compound gives a blue-green flame in a flame test. Predict the result if a few drops of sodium hydroxide are added to a solution of this compound.
5) A student makes a solution of a mystery compound, and adds a few drops of sodium hydroxide. He gets a white precipitate. He adds more sodium hydroxide and the precipitate dissolves.
   a) What positive ion is present?
   b) Write down an ionic equation for the formation of the white precipitate.
6) Iron(II) chloride forms a sludgy green precipitate with sodium hydroxide. Write down an ionic equation for this reaction.
7) Which acid is usually used to test for carbonates and sulfites?
8) Sulfites give off sulfur dioxide when mixed with dilute acid.
   a) What is the test for sulfur dioxide?
   b) Write an ionic equation for the reaction between sulfites and dilute acid.
9) What's the test for sulfates?
10) a) What's the test for a halide?
   b) What colour precipitate does a solution containing iodide ions produce in the halide test?
   c) What colour precipitate does a solution containing bromide ions produce?
11) Describe a method you could use to test for nitrates.
12) Describe the litmus test for H^+ and OH^- ions.
13) What colour is phenolphthalein in an acid? And in an alkali?
14) a) A few drops of methyl orange are added to 50 cm^3 of HCl. What colour does the solution go?
   b) A few drops of methyl orange are added to 50 cm^3 of KOH. What colour does the solution go?
15) What gas is given off when a metal reacts with acid? What's the test for this gas?
16) What gas is given off when a solution of OH^- ions is heated with an ammonium compound?
17) Explain how you could distinguish between butane and butene.
18) What's an empirical formula?
19)*An organic hydrocarbon is burnt completely in oxygen. 4.4 g of carbon dioxide and 1.8 g of water are formed. What is the compound's empirical formula?
20) Give three advantages of instrumental analysis over traditional lab methods.
21) What type of substance can be identified using atomic absorption spectroscopy?
22) What kind of information does NMR spectroscopy give you about an unknown substance?
23) Would mass spectrometry be used for elements, compounds or for both of these?

* Answer on page 261.

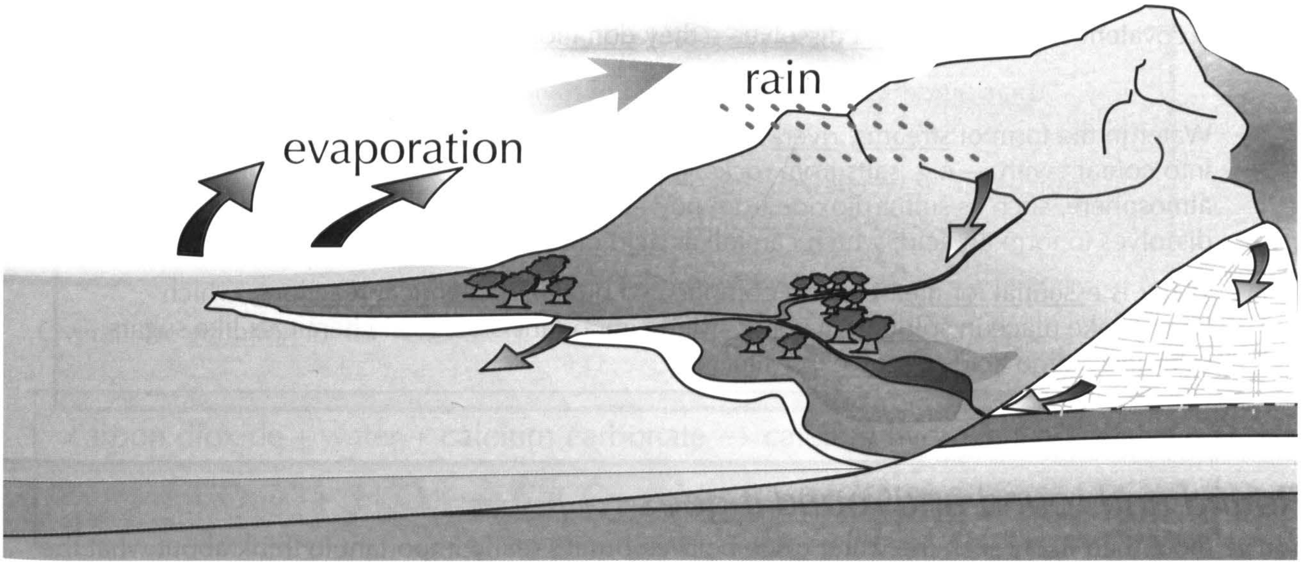
Water

Without water, there'd be no swimming, no cups of tea, no power showers — a nightmare. Plus you wouldn't even exist — you need water to live.

*The **water cycle** means water is endlessly **recycled***

This stuff about the water cycle probably won't come as a complete surprise...

1) The Sun causes evaporation of water from the sea.
2) The water vapour is then carried upwards as the warm air rises.
3) As the water vapour rises it cools — due to the general cooling of the lower part of the atmosphere at higher altitudes.
4) This fall in temperature means the water condenses to form clouds.
5) When the condensed droplets get too big they fall as rain.
6) Then the water runs back to the sea.
7) As it does this, at some stage it's going to come into contact with the rocks on (or underneath) the ground — meaning that water in different places will dissolve different minerals.
8) Then the cycle starts over again.



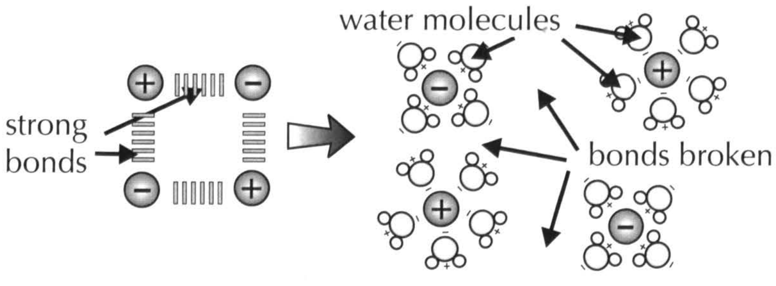
Water

Water doesn't just dissolve minerals from the ground — in fact, it's known as the universal solvent.

Water's a solvent — it dissolves many other chemicals

So many substances dissolve in water that sometimes it's called the universal solvent.

1) Water dissolves most ionic compounds. Water molecules start to surround the ions and disrupt the ionic bonding, so the solid structure of the ionic compound gradually falls apart.
2) Water molecules are polar — they've got a positive hydrogen side and a negative oxygen side.
3) The slightly negative side attracts the positive ions and the slightly positive side attracts the negative ions.



4) The following ionic compounds dissolve in water — LEARN them:

a) Salts of SODIUM (Na), POTASSIUM (K) or AMMONIUM (NH_4). ALL of these dissolve.
b) NITRATES (NO_3). ALL of these dissolve.
c) CHLORIDES (Cl), except for silver and lead.
d) SULFATES (SO_4), except for barium and lead. Calcium sulfate is only slightly soluble.

Plus there are one or two others, but you don't need to know about them.

5) Some substances that exist as small molecules are soluble in water, e.g. CO_2, SO_2 and Cl_2. Many covalent compounds don't dissolve — they don't form ions and their molecules are too big.

Water in the form of streams, rivers and rain dissolves a lot of substances that it comes into contact with — e.g. salts from rocks, fertilisers from fields, and gases in the atmosphere, such as sulfur dioxide from power stations and car exhausts. Sulfur dioxide dissolves to form an acid, which can fall as acid rain.

Water is essential for life. Life is a complicated bunch of chemical reactions, which largely take place in solution in water. Many important biological chemicals like sugars, salts and amino acids dissolve in water.

Round and round and round it goes

You've more than likely seen the water cycle before. But it's really important to think about what the water comes into contact with as it goes round — what it dissolves will affect the properties of the water. You'll have to learn the rules for which ionic compounds dissolve, I'm afraid. No easy shortcuts.

Hard Water

The water in the area where you live might be hard or it might be soft.
It depends on the type of rocks your water meets on its way to you.

Hard water *makes* ***scum*** *and* ***scale***

1) Hard water won't easily form a lather with soap. It makes a nasty scum instead. So to get a decent lather you need to use more soap.

2) Hard water also forms limescale (calcium carbonate) on the insides of pipes, boilers and kettles.

3) Limescale is a thermal insulator. This means that a kettle with limescale on the heating element takes longer to boil than a clean non-scaled-up kettle. Scale can even eventually block pipes.

4) Worst of all, hard water also causes a horrible scum to form on the surface of tea.

Hardness *is caused by* ***Ca^{2+}*** *and* ***Mg^{2+}*** *ions*

Hard water contains calcium ions (Ca^{2+}), magnesium ions (Mg^{2+}), or both. As water flows over rocks and through soils containing calcium and magnesium compounds, these ions dissolve in it.

1) Magnesium sulfate, $MgSO_4$, dissolves in water — and so does calcium sulfate, $CaSO_4$ (though only a little bit).

2) Calcium carbonate commonly exists as chalk, limestone or marble.
   - It doesn't dissolve in water, but it will react with acids.
   - And since CO_2 from the air dissolves in rainwater (forming carbonic acid, $CO_2 + H_2O \rightarrow H_2CO_3$), rainwater is slightly acidic.
   - This means that calcium carbonate can react with rainwater to form calcium hydrogencarbonate ($H_2CO_3 + CaCO_3 \rightarrow Ca(HCO_3)_2$), which is soluble.

Overall the equation for the reaction is:

carbon dioxide + water + calcium carbonate → calcium hydrogencarbonate

$$CO_2(g) + H_2O(l) + CaCO_3(s) \rightarrow Ca(HCO_3)_2(aq)$$

Hard Water

So hard water doesn't do you any harm, but it can be a bit annoying.
Luckily it's possible to remove the hardness from water.

Temporary hardness can be removed by *boiling*

There are two kinds of hardness — temporary and permanent.

Temporary hardness is caused by the hydrogencarbonate ion, HCO_3^-, in $Ca(HCO_3)_2$.
Hardness caused by dissolved calcium sulfate (among other things) is permanent hardness.

1) Temporary hardness is removed by boiling.
2) The calcium hydrogencarbonate decomposes to form $CaCO_3$.
   (This calcium carbonate precipitate is the 'limescale' on your kettle — it's insoluble.)
3) This won't work for permanent hardness, though. Heating a sulfate ion does nowt.

calcium hydrogencarbonate → calcium carbonate + water + carbon dioxide

$$Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$$

Both types of hardness can be removed using **washing soda**

1) Both types of hardness are removed by adding washing soda — sodium carbonate, Na_2CO_3.
2) The carbonate ions join onto the calcium ions and make an insoluble precipitate of calcium carbonate.
3) This works whether the hardness is due to calcium sulfate or calcium hydrogencarbonate.

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

1) Both types of hardness can also be removed by 'ion exchange columns'.
2) These clever bits of chemistry have lots of sodium ions (or hydrogen ions) and 'exchange' them for calcium or magnesium ions.

Make sure you know all those equations

One thing that I've never understood is that they sell water softeners in areas that already have soft water. Hmm... For the exam, you're supposed to know how the salts that cause hard water get into the water in the first place, and how they can be removed. So make sure you know it.

Warm-Up and Exam Questions

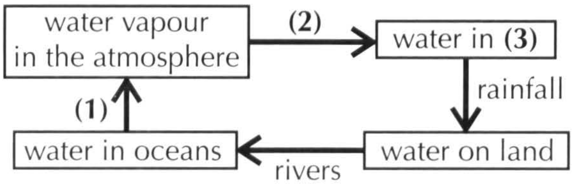
Warm-Up Questions

1) By what process does water in the oceans become water vapour in the atmosphere?
2) What property of water molecules makes water a good solvent for ionic compounds?
3) Hard water can lead to the formation of limescale. What is the chemical name of the main compound found in limescale?
4) Name one compound that causes permanent hardness in water.
5) What two methods can be used to remove both temporary and permanent hardness?

Exam Questions

1 (a) The diagram shows a simple form of the water cycle. Write down three words to fill in the blanks (1), (2) and (3) in the diagram.

(3 marks)



(b) Rainwater is relatively pure, but seawater contains large amounts of dissolved substances. Explain how these dissolved substances enter the water.

(2 marks)

2 Kate is investigating the solubility of potassium nitrate in water.

(a) She finds that potassium nitrate dissolves readily in water. What does this suggest about the type of bonding found in potassium nitrate?

(1 mark)

(b) Kate then investigated how the solubility of potassium nitrate in water is affected by temperature. She recorded her results in a table:

Temperature (°C)	20	30	40	50	60
Solubility (g/100g water)	30	45	65	85	110

What is the relationship between the temperature and the solubility?

(1 mark)

3 (a) Hard water in many areas is caused by dissolved Ca^{2+} ions. Give two disadvantages of living in a hard water area.

(2 marks)

(b) Temporary hard water can be caused by dissolved calcium hydrogencarbonate. Copy and complete the equation below to show the effect of heat on calcium hydrogencarbonate:

$Ca(HCO_3)_2(aq) \rightarrow$ ________ (s) + ________ (l) + ________ (g)

(3 marks)

Water Quality

It's easy to take water for granted... turn on the tap, and there it is — nice, clean water.
The water you drink has been round the block a few times though — so there's some fancy chemistry needed to make it drinkable.

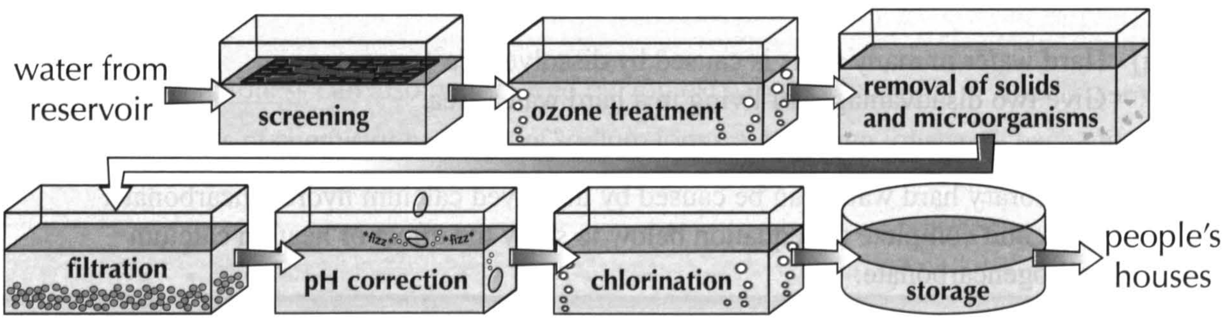
Drinking water needs to be good quality

Water is essential for life, but it must be free of poisonous salts (for example, phosphates and nitrates) and harmful microorganisms. Microorganisms in drinking water can cause diseases such as cholera and dysentery.

Most of our drinking water comes from reservoirs. Water flows into reservoirs from rivers and ground-water — water companies choose to build reservoirs where there's a good supply of clean water. Government agencies keep a close eye on pollution in reservoirs, rivers and ground-water.

Water from reservoirs is treated at the water treatment works

1) The water passes though a mesh screen to remove big bits like twigs.
2) Next, it's treated with ozone or chlorine to kill microorganisms.
3) Chemicals are added to make solids and microorganisms stick together and fall to the bottom. Sometimes iron is added to remove dissolved phosphates. Bacteria are used to remove nitrates.
4) The water is filtered through gravel beds to remove all the solids. Nasty tastes and odours can also be removed by passing the water through "activated carbon" filters or with "carbon slurry".
5) The pH is corrected if the water is too acidic or too alkaline.
6) Water is chlorinated to kill off any harmful microorganisms left.



Water Quality

The only way to get totally pure water is by distillation. So how come they don't do that...?

Water quality is constantly monitored

To monitor water quality, water companies take samples of water — from the water entering the treatment works right though to the taps in consumers' houses.

Some people still aren't satisfied. They buy filters that contain carbon or silver to remove substances from their tap water. Carbon in the filters removes the chlorine taste and silver is supposed to kill bugs.

Some people in hard water areas buy water softeners which contain ion exchange resins (see page 190 for more on this).

Totally pure water with nothing dissolved in it can be produced by distillation — boiling water to make steam and condensing the steam. But this process is too expensive to use for producing tap water — bags of energy would be needed to boil all the water we use (which wouldn't be great for the environment either). Distilled water is used in chemistry labs though.

You'd use pure water to make a solution of (say) KBr, because you wouldn't want any other ions mucking it up.

Clean water is essential for life

1) Not everyone has clean water. The World Health Organisation (WHO) and the United Nations estimated in 1995 that a billion people in the world don't have access to clean drinking water.
2) In many developing countries it's very expensive to get clean water. Some people in developing countries live in isolated rural areas, and have to walk miles to get any water at all.
3) It's a fact that the biggest increases in life expectancy in most countries' histories (including the UK's) are linked with the ability to supply clean water — not with medical advances or anything like that. Clean water is that vital.
4) In November 2004 the WHO said that improving drinking water quality could reduce diarrheal disease by up to 40%. Currently, approximately 1.8 million people around the world die each year of diarrheal diseases (such as cholera).
5) Some water purifying processes can damage the environment, which is worth bearing in mind. Clean water is important, but if possible it's best to obtain it in a 'green' way.

The water you drink has been through seven people already

Well, it's possible. It's also possible that the water you're drinking used to be part of the Atlantic Ocean. Or it could have been drunk by Alexander the Great. Or been part of an Alpine glacier. It gets about a bit, does water. And remember... tap water isn't pure — but it's drinkable, and that's the main thing.

Solubility

Something is soluble if it dissolves — like sugar when you put it in tea (hurrah).
Something is insoluble if it doesn't dissolve — like sand when you put it in tea (boo, hiss).

*All **gases** are **soluble** — to some extent, anyway*

1) "Chlorine water" is (unsurprisingly) a solution of chlorine gas in water. It's used as bleach in the paper and textile industries, and also to sterilise water supplies (it kills bacteria).
2) The amount of gas that dissolves depends on the pressure of the gas above it — the higher the pressure, the more gas there is dissolving.

> Fizzy drinks initially contain a lot of carbon dioxide dissolved in water (carbonated water). But when you take the cap off, the pressure's released and a lot of the carbon dioxide fizzes out of solution.

3) But... gases become less soluble as the temperature of the solvent increases, which is exactly the opposite of solids (see the next page).

> Aquatic life needs dissolved oxygen, but oxygen levels in rivers can be lowered by pollution and a rise in temperature (caused by warm water discharged from towns and industry).

Solubility** — learn the proper **definitions

> The solubility of a substance in a given solvent is the number of grams of the solute (usually a solid) that dissolve in 100 g of the solvent (the liquid) at a particular temperature.

E.g. at room temperature (20 °C), about 36 g of sodium chloride (NaCl) will dissolve in 100 g of water.

> The solubility of (solid) solutes usually increases with temperature.

E.g. at 60 °C, about 37 g of sodium chloride (NaCl) will dissolve in 100 g of water.

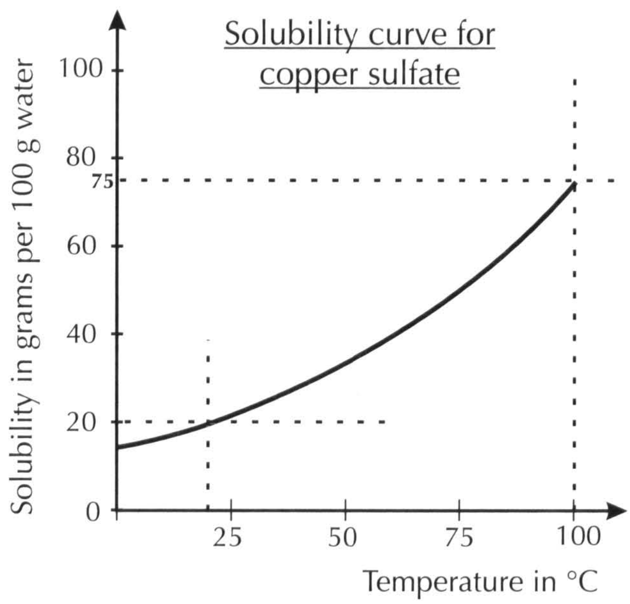
> A saturated solution is one that cannot hold any more solid at that temperature — and you have to be able to see solid on the bottom to be certain that it's saturated.

Solubility

Solubility curves are useful for investigating how much solid can dissolve in a particular solution.

Solubility curves *show when a solution is* ***saturated***

1) A solubility curve plots the mass of solute dissolved in a saturated solution at various temperatures.
2) The solubility of most solids increases as the temperature increases.
3) This means that cooling a saturated solution will usually cause some solid to crystallise out — that means it separates from the solution.
4) The mass of crystals formed by cooling a solution a certain amount can be calculated from a solubility curve...



Draw lines perpendicular to both axes through the temperatures in the question, then subtract the smaller mass from the larger — that difference will precipitate out on cooling.

This graph is for 100 g of water — so if you had 1000 g of water instead, you'd just multiply your answer by 10. Simple.

Example:

What mass of solid copper sulfate will crystallise out when a saturated solution containing 100 g of water is cooled from 100 °C to 20 °C?

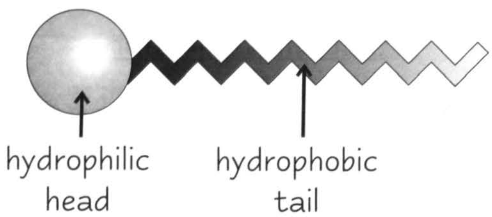
Answer: 75 g – 20 g = 55 g

Detergents and Dry-Cleaning

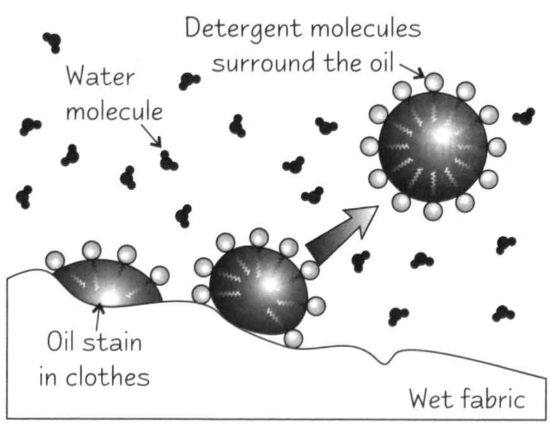
Cleanliness is next to godliness... or so they say.

Detergents work by sticking to both water and grease

1) Some dirt will dissolve in water without the help of a detergent, but most won't (see below).
2) Detergents help water and oil to mix.
3) Detergents contain molecules that have a hydrophilic (water-loving) head, and a hydrophobic (water-hating) tail.
4) The hydrophilic heads form intermolecular bonds with water.
5) And the hydrophobic tails bond to the fat molecules in greasy dirt.



6) When you (or a washing machine) swish the fabric around, the detergent molecules find their way in between the grease and the wet fabric.
7) The detergent molecules eventually surround the grease completely, and bond to it, with their hydrophilic heads around the outside like a coat.
8) This hydrophilic coat stops the grease droplets re-attaching themselves to the fabric, and they're pulled away into the wash water.
9) Then, when you rinse the fabric, the grease and dirt are rinsed away along with the water.



Modern synthetic detergents are mostly made using big organic molecules from crude oil. An acid group is added to one end of the molecule to make an organic acid. This is then neutralised with a strong alkali, usually sodium or potassium hydroxide, to form a salt.

Detergents and Dry-Cleaning

Sometimes a simple detergent won't be able to do the trick, and you need something extra.

Washing at **low temperatures** saves **energy** (and your clothes)

1) Some natural fabrics (e.g. wool) shrink, and some artificial fabrics (e.g. nylon) quickly lose their shape if they're washed at too high a temperature.
2) Also, some dyes will run in high-temperature washes — brightly coloured clothes can quickly fade and stop looking new.
3) Nowadays you can get biological detergents with enzymes in them. The enzymes digest stains without the need for high temperatures, which protects your clothes.
4) Low-temperature washes also save energy, which is better for the environment and for your energy bill too.

Different solvents dissolve different stains

1) You can't dissolve every solid in every liquid. Different solids need different solvents.
2) To remove a stain, you have to use the right solvent to dissolve it off the fabric. And then...
3) ... the solvent molecules form strong intermolecular bonds with the 'stain' molecules (the solute), pulling apart the solute-solute bonds.
4) So the stain breaks up and the solvent molecules surround the solute molecules completely — making a solution.
5) A lot of stains aren't soluble in water — especially greasy stains, paints and varnishes.
6) Sometimes using a detergent can remove the stain, but sometimes you need to use a different, dry-cleaning solvent.
7) Paints, varnishes and other organic chemicals will often dissolve in an organic solvent.

Water is called the universal solvent, but it can't dissolve everything

Yes, I'm afraid you are going to have to do some washing at some point — you will run out of clothes eventually. And when you do, you'll have the pleasure of knowing how it all works. Isn't that nice.

Warm-Up and Exam Questions

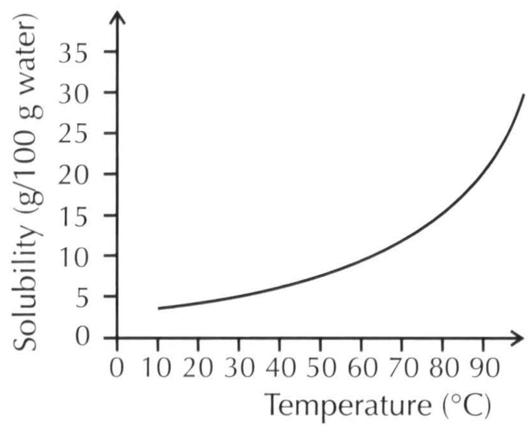
Well, those last few pages were a bit of a mixed bag — first you're cleaning water, then you're dissolving stuff in it and finally you're washing your jeans. These questions will help you get it all straightened out.

Warm-Up Questions

1) Why is drinking water treated by filtration?
2) Why is drinking water treated with chlorine?
3) How can totally pure water be produced?
4) What is the relationship between water temperature and the solubility of gases in the water?
5) What are contained in biological detergents that are not found in other detergents?

Exam Questions

1 Charlie works for a food company and investigated the solubility of a sweetener, compound X, in water. The solubility curve for compound X is shown on the graph.



(a) Why does Charlie not record any data for temperatures above 100 °C?

(1 mark)

(b) What is the maximum mass of X that can dissolve in 100 g of water at 25 °C?

(1 mark)

(c) Charlie has a beaker containing 250 g of water at 80 °C. What is the maximum mass of compound X that Charlie will be able to dissolve in this water?

(2 marks)

(d) What mass of solid compound X will crystallise out when a saturated solution of compound X in 100 g of water is cooled from 90 °C to 10 °C?

(2 marks)

2 Detergent molecules are able to remove dirt, oil and grease from fabrics during washing.

(a) Explain how the chemical structure of the molecules helps them to remove oil.

(3 marks)

(b) Detergents like soaps can be made by reacting fatty acids to produce their sodium salts.

(i) Name one sodium compound that can be used to react with the fatty acids in such a reaction.

(1 mark)

(ii) What type of reaction is this?

(1 mark)

(c) Some stains have to be removed using dry cleaning. Explain what dry cleaning involves and how it removes difficult stains.

(2 marks)

Changing Equilibrium

If you cast your mind back to page 165, you'll remember that some reactions are reversible, and that these reactions end up reaching a dynamic equilibrium. The position of the equilibrium might lie to the right (meaning lots of products and not many reactants) or to the left (lots of reactants and not many products). But you can change the position of the equilibrium by changing the conditions.

The **equilibrium** tries to **minimise** any changes you make

TEMPERATURE

All reactions are exothermic in one direction and endothermic in the other (see p.157).

1) If you decrease the temperature, the equilibrium will move to try and increase it — it moves in the exothermic direction to produce more heat.
2) If you raise the temperature, the equilibrium will move to try and decrease it — the equilibrium moves in the endothermic direction.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The forward reaction is exothermic — a drop in temperature moves the equilibrium to the right (more products).

PRESSURE

Changing this only affects an equilibrium involving gases.

1) If you increase the pressure, the equilibrium tries to reduce it — the equilibrium moves in the direction where there are fewer moles of gas.
2) If you decrease the pressure, the equilibrium tries to increase it — it moves in the direction where there are more moles of gas.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

There are 4 moles on the left, but only 2 on the right. So, if you increase the pressure, the equilibrium shifts to the right.

CONCENTRATION

Same reaction again...

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

1) If you increase the concentration of N_2 or H_2, the equilibrium tries to decrease it by shifting to the right (making more NH_3).
2) If you increase the concentration of NH_3, the equilibrium tries to reduce it again by shifting to the left (making more N_2 and H_2).

If you decrease the concentration of N_2, H_2 or NH_3, the equilibrium moves to try and increase the concentration again.

Changing Equilibrium

It's important that you can interpret any data you might get in an exam about equilibria.

Make sure you can read **equilibrium tables** and **graphs**

You might be asked to interpret data about equilibrium, so you'd better know what you're doing. The Haber process (see page 167) is a great example of all this...

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

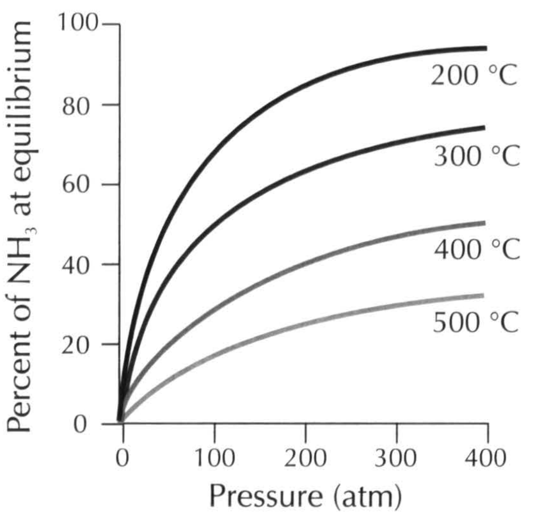
The forward reaction is exothermic.

First off, a table...

Pressure (atmospheres)	100	200	300	400	500
% of ammonia at 450 °C	14	26	34	39	42

As the pressure increases, the proportion of ammonia increases (exactly what you'd expect — since increasing the pressure shifts the equilibrium to the side with fewer moles of gas — here, the right).

And now a graph...



1) This time, each different line represents a different temperature.
2) As the temperature increases, the proportion of ammonia decreases (the backward reaction is endothermic, so this speeds up to try and reduce the temperature again).
3) The conditions that will give you most ammonia are high pressure and low temperature.

An equilibrium is like a particularly stubborn mule

You do one thing, the reaction does the other. Sounds pretty annoying, but actually it's what gives you control over what happens. And in industry, control is what makes the whole shebang profitable.

Warm-Up and Exam Questions

Some of this stuff can be a bit tricky — have a go at these to make sure you've got it.

Warm-Up Questions

1) What is the effect on an equilibrium of increasing the temperature if the forward reaction is endothermic?
2) What is the effect on an equilibrium of decreasing the temperature if the forward reaction is exothermic?
3) For the reaction: $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$, what would be the effect on the equilibrium of changing the gas pressure?
4) For the reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, what is the effect on the reaction of removing the NH_3 produced from the reaction vessel?

Exam Questions

1 The reaction between sulfur dioxide and oxygen to form sulfur trioxide is reversible:

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ The forward reaction is exothermic.

(a) What is the effect on the yield of sulfur trioxide of:

(i) increasing the temperature?

(1 mark)

(ii) decreasing the pressure? Explain your answer.

(3 marks)

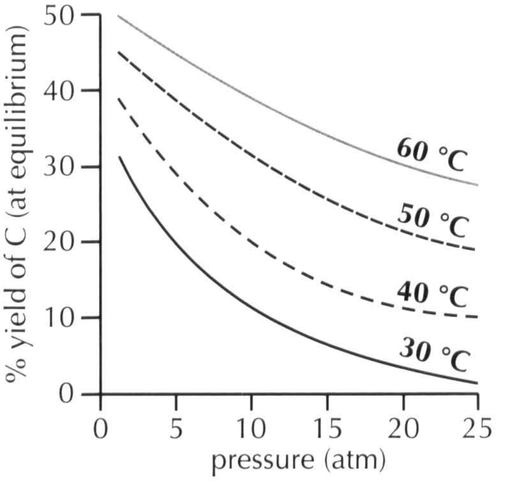
(b) For an industrial plant producing sulfur trioxide, a design is proposed where the SO_3 is removed from the reaction vessel as soon as it forms. Explain why this will help to increase the production of sulfur trioxide.

(1 mark)

2 Charlotte investigates the effect of pressure and temperature on the chemical equilibrium:

$A(g) + B(g) \rightleftharpoons 3C(g)$

where A, B and C are gas molecules.
She presents her findings in a graph.



(a) At 10 atm pressure, what is the percentage yield of C at 40 °C?

(1 mark)

(b) What is the effect on the equilibrium of increasing the pressure? Explain why this happens.

(3 marks)

(c) (i) Explain the effect on the equilibrium of increasing the temperature.

(2 marks)

(ii) Is the forward reaction endothermic or exothermic?

(1 mark)

Acid-Base Theories

Theories about what makes an acid an acid, and a base a base, have evolved a bit over the years.

*Arrhenius said acids **release hydrogen ions** in water*

1) A guy called Arrhenius studied acids and bases in water. His theory was that when mixed with water, all acids release hydrogen ions, H^+ (a H^+ ion is a proton).

Example:

$$HCl\,(g) + water \rightarrow H^+(aq) + Cl^-(aq)$$
$$H_2SO_4\,(l) + water \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

But HCl doesn't release hydrogen ions until it meets water — so hydrogen chloride gas isn't an acid.

2) He also said that alkalis form OH^- ions (hydroxide ions) when in water.

Example:

$$\text{ammonia: } NH_3(g) + H_2O\,(l) \rightarrow NH_4^+(aq) + OH^-(aq)$$

Not all bases dissolve in water, but those that do are called alkalis.

*The idea **only** worked for acids and bases that **dissolved in water***

Arrhenius' idea worked pretty well, but it only worked for acids and bases that dissolved in water.

However, ammonia gas can react as a base even when it isn't dissolved in water, which was one reason why these ideas weren't immediately accepted.

(A hypothesis is always less likely to be accepted when there are lots of exceptions which the hypothesis can't explain.)

Also, back in the 1880s when Arrhenius first suggested that molecules ionise in water, many scientists didn't think it was possible. Charged subatomic particles hadn't been discovered yet, so the idea of charged ions seemed strange. Scientists couldn't imagine how Cl^- could be different from Cl_2 gas.

Acid-Base Theories

Arrhenius had made a good start, but his theory couldn't explain all of the known facts. However, other scientists who followed him were able to use his ideas and improve on them.

Lowry and Brønsted said acids are proton donors

1) Lowry and Brønsted (working separately) made things a bit more general. They came up with definitions that work for both soluble and insoluble bases:

> Acids release H^+ ions — i.e. they're proton donors.
> Bases accept H^+ ions — i.e. they're proton acceptors.

In fact, it'd be more accurate to say that acids have their proton taken away from them.

2) These ideas were readily accepted because they explained the behaviour of acids and bases in solvents other than water.

3) Also, they were an adaptation of an idea which already kind of worked. When Arrhenius came up with his idea it was totally new, so people took more convincing.

Protons are hydrated in water

Anyway... for a substance to act as an acid or as a base, you usually need water. This is what happens...

In acidic solutions:

The acid molecules dissociate, releasing lots of H^+ ions. These H^+ ions (protons) become hydrated (surrounded by water molecules). The protons are now given the fancy name 'hydrated protons' and can be represented by '$H^+(aq)$'. And it's these hydrated protons that make acids acidic, if you like.

In basic solutions:

Water molecules can dissociate into H^+ and OH^- ions, although they almost never do in pure water. But some base molecules, like ammonia (NH_3), can take hydrogen ions from water, causing more molecules to dissociate, and leaving an excess of OH^- ions behind. Other bases, like potassium hydroxide (KOH), release hydroxide ions straight into the solution.

This is an example of how theories gradually develop into facts

It's another example of how scientific knowledge progresses — lots of people contributing ideas to fit the available evidence. Some ideas are better than others, but usually the rubbish ones are quickly forgotten and you don't have to learn about those (e.g. acidic behaviour being due to magic pixies).

Strong and Weak Acids

Right then. Acids. Brace yourself...

Acids can be **strong** or **weak**

1) Strong acids (e.g. sulfuric, hydrochloric and nitric) ionise almost completely in water. This means almost every hydrogen atom is released — so there are loads of H^+ ions.

2) Weak acids (e.g. ethanoic, citric, carbonic) ionise only very slightly. Only some of the hydrogen atoms in the compound are released — so only small numbers of H^+ ions are formed.

Examples:

Strong acid: $HCl \rightarrow H^+ + Cl^-$

Weak acid: $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$

Use a 'reversible reaction' arrow for a weak acid.

3) The ionisation of a weak acid is a reversible reaction. Since only a few H^+ ions are released, the equilibrium lies well to the left.

4) The pH of an acid or alkali is a measure of the concentration of H^+ ions in the solution. Strong acids typically have a pH of about 1 or 2, while the pH of a weak acid might be 4, 5 or 6.

5) The pH of an acid or alkali can be measured with a pH meter or with universal indicator paper (or can be estimated by seeing how fast a sample reacts with, say, magnesium).

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 many moles of acid molecules there are in a litre (1 dm^3) of water. Concentration is basically how watered down your acid is.

3) Note that concentration describes the total number of acid molecules — not the number of molecules that release hydrogen ions.

4) The more moles of acid molecules per dm^3, the more concentrated the acid is.

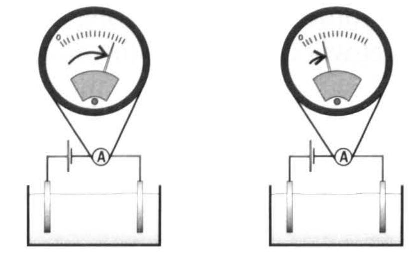
5) So you can have a dilute but strong acid, or a concentrated but weak acid.

Strong and Weak Acids

Strong acids and weak acids sometimes behave differently.

Strong acids are ***better electrical conductors*** *than weak acids*

1) Hydrochloric acid has a much higher electrical conductivity than the same concentration of ethanoic acid. It's all to do with the concentration of the ions.
2) It's the ions that carry the charge through the acid solutions. So the greater concentration of ions in the strong acid means more charge can be carried. Simple.



hydrochloric acid ($1\ mol/dm^3$) ethanoic acid ($1\ mol/dm^3$)

Strong acids *react* ***faster*** *than* ***weak acids***

Strong and weak acids react with reactive metals and with carbonates in the same way.

1) Both hydrochloric acid (strong) and ethanoic acid (weak) react with magnesium to give hydrogen. And both hydrochloric acid and ethanoic acid react with calcium carbonate to give carbon dioxide.

$$2HCl + Mg \rightarrow MgCl_2 + H_2$$
$$2CH_3COOH + Mg \rightarrow Mg(CH_3COO)_2 + H_2$$

$$2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$$
$$2CH_3COOH + CaCO_3 \rightarrow Ca(CH_3COO)_2 + H_2O + CO_2$$

2) The difference between the reactions of the two acids will be the rate of reaction. Ethanoic acid will react more slowly than hydrochloric acid of the same concentration.
3) It's all to do with the equilibrium in the weak acid reaction ($CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$)...
4) When you put a weak acid into water, it releases a few H^+ ions (but not that many compared to what you'd get with a strong acid). When you add magnesium (or calcium carbonate), these H^+ ions react.
5) This means the concentration of H^+ ions decreases, so the equilibrium shifts to compensate — meaning more H^+ ions are released. These ions then react, so the equilibrium shifts... and so on. As more ions are removed, more are supplied — kind of a drip-feed arrangement.
6) This is completely different from what you get with a strong acid, where all the acid molecules will react pretty darn quickly — since all the H^+ ions are just sitting there ready and waiting to go.

Strong and weak acids both react to form the same kind of products

Hydrochloric acid is really nasty stuff. It's corrosive and irritating, difficult to store, damages almost everything it touches and will severely burn skin. Yet we have it in our stomachs. Surprising. But luckily we have a thick layer of mucus to protect our stomach walls. Neat, eh.

Strong and Weak Acids

Weak acids react more slowly than strong acids — but this can actually make them more useful.

The amount of gas produced depends on the amount of acid

1) Hydrochloric acid (strong) will react faster than ethanoic acid (weak), but the amount of product you get will be the same (if you start with the same amount and they're the same concentration, etc.).
2) This is because if the concentrations are the same, the number of molecules in a litre (say) of water will be the same.
3) And each of these molecules can let go of one H^+ ion:

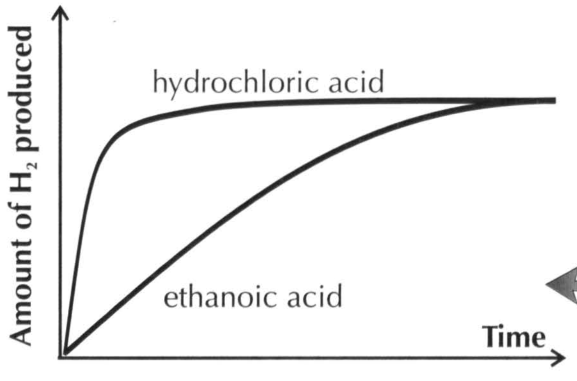
$$HCl \longrightarrow H^+ + Cl^-$$

and

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

It's just that hydrochloric acid lets go of them all at once, whereas ethanoic acid lets go gradually.

4) But since the total number of H^+ ions available is the same, the amount of product will be the same (it's the H^+ ions that are the important bits in acid reactions).



HCl reacts faster (steeper curve), but both reactions give the same amount of hydrogen.

Weak acids can be really useful

Weak acids can be more useful than dilute strong acids.

1) A strong acid reacts very fast — all its H^+ ions are released straight away and are ready for action.
2) But this might not always be as useful as it sounds — for example, if you want a slightly more controllable reaction, like with kettle descalers.
3) A strong acid would react very quickly with the scale — but before you knew it, it might also react with the metal of the kettle.
4) A weak acid still removes the scale, but the lower concentration of H^+ ions means the reaction will be slower and easier to control. And you can just tip the solution away when the scale's gone.

Warm-Up and Exam Questions

Warm-Up Questions

1) When sulfuric acid, H_2SO_4, dissolves in water, what ions are found in the solution?
2) According to the ideas of Lowry and Brønsted, what is the definition of an acid?
3) Citric acid is an example of a weak acid. What is a weak acid?
4) Write the equation for the ionisation of ethanoic acid, CH_3COOH.
5) What is the difference between a strong acid and a concentrated acid?

Exam Questions

1 Ammonia dissolves in water to give a weak alkali.

(a) In the 1880s, Arrhenius proposed that when molecules are dissolved in water they can form ions. Copy and complete the ionic equation below to show how ammonia can dissolve in water to form a hydroxide ion.

$$NH_3(g) + \_\_\_\_\_\_\_ (l) \rightarrow \_\_\_\_\_\_\_ (aq) + OH^-(aq)$$

(2 marks)

(b) Arrhenius's ideas were not widely accepted at first. One reason was that ammonia can act as a base even when it is not dissolved in water. Explain how the ideas of **Lowry and Brønsted** in the 1920s showed why ammonia gas acts as a base.

(2 marks)

(c) Ammonia solution is a weak alkali. How are weak alkalis different from strong alkalis?

(2 marks)

2 Asha compared the properties of a strong acid, hydrochloric acid, and a weak acid, ethanoic acid. She made 1 mol/dm^3 solutions of each acid and she tested them by measuring their electrical conductivity and by reacting them with calcium carbonate.

(a) Why did Asha prepare solutions of equal concentration?

(1 mark)

(b) Which acid solution would have the highest electrical conductivity? Explain your choice.

(2 marks)

(c) Asha reacted equal volumes of each solution with equal masses of powdered calcium carbonate. She found that each acid produced the same final volume of carbon dioxide gas, but that one acid produced the gas much more quickly than the other.

(i) Which acid would you expect to produce the gas faster?

(1 mark)

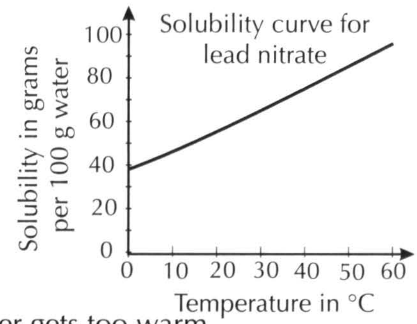
(ii) Ethanoic acid is a weak acid. Explain why it produced the same final volume of carbon dioxide as the hydrochloric acid did. Use ideas about chemical equilibria in your answer.

(3 marks)

Revision Summary for Section Eleven

Bit of a mixed bag — one minute you're pondering equilibria, the next you're worrying about how many millions of people are without clean drinking water. The one thing that's constant and unchanging is the need to learn it all for the exam you've got coming up. So test yourself on these little beauties.

1) Explain why water vapour condenses and falls as rain.
2) Sea water contains dissolved minerals. Where do these come from?
3)* Which of the following will dissolve in water?
   a) lead nitrate b) $PbCl_2$ c) ammonium chloride
   d) potassium sulfate e) Ag_2SO_4 f) silver chloride
   g) $CuSO_4$ h) barium sulfate i) $Ba(NO_3)_2$
4) What are the main ions that cause water hardness?
5) Give two disadvantages of living in an area with hard water.
6) How are microorganisms removed from drinking water?
7) How are poisonous phosphates removed from drinking water?
8) Tap water is not pure water. Why don't we make sure that all our drinking water is pure water?
9) Why is having a clean water supply so important?
10) What is a saturated solution?
11)*The graph shows the solubility of lead nitrate in 100 g of water.
   a) How much lead nitrate will dissolve in 100 g of water at 40 °C?
   b) At what temperature will 70 g of lead nitrate dissolve in 100 g of water?
   c) What mass of solid lead nitrate will crystallise when a saturated solution containing 100 g of water is cooled from 60 °C to 40 °C?



12) Why does a bottle of lemonade fizz up when you open it?
13) Explain why aquatic animals might not be able to survive if their water gets too warm.
14) Describe the advantages of washing clothes at low temperatures.
15) Draw a labelled diagram to show the typical structure of a molecule of detergent.
16) Describe how a detergent removes greasy dirt from clothes.
17) Explain why clothing stained with nail varnish might need to be dry-cleaned.
18) A reaction is exothermic in the forward direction. What should you do to the temperature of the reaction mixture in order to get more product?
19) Explain how altering the pressure can affect the position of an equilibrium involving gases.
20) Give one other factor that can be changed in order to shift the position of an equilibrium.
21) Write a brief description of Arrhenius's theory about acids and bases.
22) Why weren't scientists willing to accept his ideas at first?
23) What's the difference between a weak acid and a strong acid? Give one example of each.
24) Are strong acids or weak acids better conductors of electricity?
25) Explain why weak acids react more slowly than strong acids do.
26)*A weak acid and a strong acid of the same concentration react with the same mass of magnesium. The reaction with the weak acid produces 7 cm^3 of hydrogen gas. How much hydrogen gas would you expect the reaction with the strong acid to produce?

* Answers on page 262.

Concentration

A rather dull and boring page to start the section with. But at least there are some calculations on it.

Concentration is a measure of how crowded things are

The concentration of a solution can be measured in moles per dm^3 (i.e. moles per litre).
So 1 mole of stuff in 1 dm^3 of solution has a concentration of 1 mole per dm^3 (or 1 mol/dm^3).

> The more solute you dissolve in a given volume, the more crowded the solute molecules are and the more concentrated the solution.

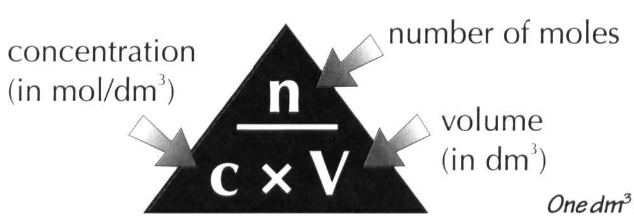
1 litre
= 1000 cm^3
= 1 dm^3

Concentration can also be measured in grams per dm^3. So 56 grams of stuff dissolved in 1 dm^3 of solution has a concentration of 56 grams per dm^3.

There's a calculation you can do to convert moles per dm^3 to grams per dm^3 (see next page). In the exam, look out for which one the question's asking for.

Concentration = no. of moles ÷ volume

Here's a nice formula triangle for you to learn:



One dm^3 is a litre.

> concentration
> = no. of moles ÷ volume

Example 1:

> What's the concentration of a solution with 2 moles of salt in 500 cm^3?

Answer: Easy — you've got the number of moles and the volume, so just use the formula...

$$\text{Concentration} = \frac{2}{0.5} = \underline{4 \text{ moles per dm}^3}$$

Convert the volume to litres (i.e. dm^3) first by dividing by 1000.

Example 2:

3 molar is sometimes written '3 M'.

> How many moles of sodium chloride are in 250 cm^3 of a 3 molar solution of sodium chloride?

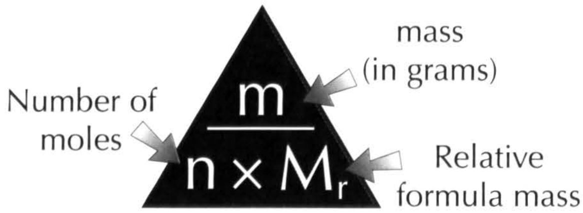
Answer: Well, 3 molar just means it's got 3 moles per dm^3. So using the formula...

$$\text{Number of moles} = \text{concentration} \times \text{volume} = 3 \times 0.25 = \underline{0.75 \text{ moles}}$$

Concentration

Converting moles per dm³ to grams per dm³ (and vice versa)

They might ask you to find out a concentration in grams per dm³. If they do, don't panic — you just need another formula triangle:



number of moles = mass ÷ relative formula mass

Example 1:

You have a solution of sulfuric acid of 0.04 mol/dm³.
What is the concentration of this solution in GRAMS per dm³?

Step 1: Work out the relative formula mass for the solute (you should be given the relative atomic masses, e.g. H = 1, S = 32, O = 16):

So, $H_2SO_4 = (1 \times 2) + 32 + (16 \times 4) = 98$

Step 2: Convert the concentration in moles into concentration in grams.

So, in 1 dm³: mass in grams = moles × relative formula mass = 0.04 × 98 = 3.92 g

So the concentration in g/dm³ = 3.92 g/dm³

Example 2:

The concentration of a solution of sulfuric acid is 19.6 grams/dm³. What is it in MOLES per dm³?

Step 1: The relative formula mass of H_2SO_4 = 98

Step 2: moles = mass in grams ÷ relative formula mass = 19.6 ÷ 98 = 0.2 moles

So the concentration in mol/dm³ = 0.2 mol/dm³

If you remember the formula triangles, the rest will follow

High concentration's like a whole rugby team in a mini. Or everyone in Britain on the Isle of Wight. Low concentration's like a guy stranded on a desert island, or a small fish in a big lake. Poetic, no?

Calculating Volumes

Run for it now, while you've still got the chance — it's yet more equations and stuff.

Avogadro's law — one mole of any gas occupies 24 dm^3

Learn this fact — you're going to need it:

24 dm^3 is the molar volume at RTP.

> One mole of any gas always occupies 24 dm^3 (= 24 000 cm^3) at room temperature and pressure (RTP — 25 °C and 1 atmosphere).

Remember, dm^3 is just a fancy way of writing 'litre', so 1 dm^3 = 1000 cm^3

This means you can use the following formula to convert a mass of any gas into a volume:

$$\text{VOLUME OF GAS (in dm}^3\text{)} = \frac{\text{MASS OF GAS}}{M_r \text{ OF GAS}} \times 24$$

So you can make this handy formula triangle using:

$\frac{\text{mass}}{M_r}$ = no. of moles (from the last page):

$$\frac{\text{volume}}{\text{moles} \times 24}$$

Example 1:

What's the volume of 4.5 moles of chlorine at RTP?

Answer: 1 mole = 24 dm^3, so 4.5 moles = 4.5 × 24 dm^3 = 108 dm^3

Example 2:

How many moles are there in 8280 cm^3 of hydrogen gas at RTP?

Answer: no. of moles = $\frac{\text{volume of gas}}{\text{volume of 1 mole}} = \frac{8.28}{24}$ = 0.345 moles

Don't forget to convert from cm^3 to dm^3.

This business about 1 mole of any gas taking up the same volume is true for other temperatures and pressures too. For example, at 0 °C and 1 atmosphere, 1 mole of any gas will occupy 22.4 dm^3 (lower temperature, so a smaller volume).

If you need any other numbers, you'll be told them in the exam though — so just learn the above.

Calculating Volumes

Just to make things really complicated, sometimes they'll ask you a question where you need to combine using this new formula with finding a reacting mass (see page 122).

You can calculate **volumes in reactions** if you know the **masses**

For this type of question there are two stages:

1) Find the reacting mass (just like you did on page 122).
2) Then convert the mass into a volume using the formula on the last page.

Example:

Find the volume of carbon dioxide produced (at RTP) when 2.7 g of carbon is completely burned in oxygen.
(A_r of carbon = 12, A_r of oxygen = 16)

ANSWER:

Step 1: Write out the balanced equation: $C + O_2 \rightarrow CO_2$

Step 2: Write down the M_r for each: **12 32 44**

Step 3: Divide for one, times for all:

÷12	**1**	**3.666...**	÷12
×2.7	**2.7**	**9.9**	×2.7

Step 4: So 2.7 g of C gives 9.9 g of CO_2. Now using the formula from the last page:

$$\text{volume} = \frac{\text{MASS}}{M_r} \times 24$$

volume = (MASS/M_r) × 24
= (9.9/44) × 24
= 5.40 dm^3

That's ANY gas — oxygen, methane, carbon dioxide, ANY gas

All this stuff ties in with p.122 — if you're not comfortable working out the reacting masses (write down a balanced equation, write down the M_r for each reactant and product, divide for one and times for all), look there first. The only new thing here is the molar volume business: 1 mole of gas = 24 dm^3.

Titrations

Titrations have a bad reputation — but they're not as difficult as they're sometimes made out to be.

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). Here's how you do a titration...

1) Using a pipette and pipette filler, add some alkali (usually about 25 cm^3) to a conical flask, along with two or three drops of indicator.

2) Fill a burette with the acid. Make sure you do this BELOW EYE LEVEL — you don't want to be looking up if some acid spills over.

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

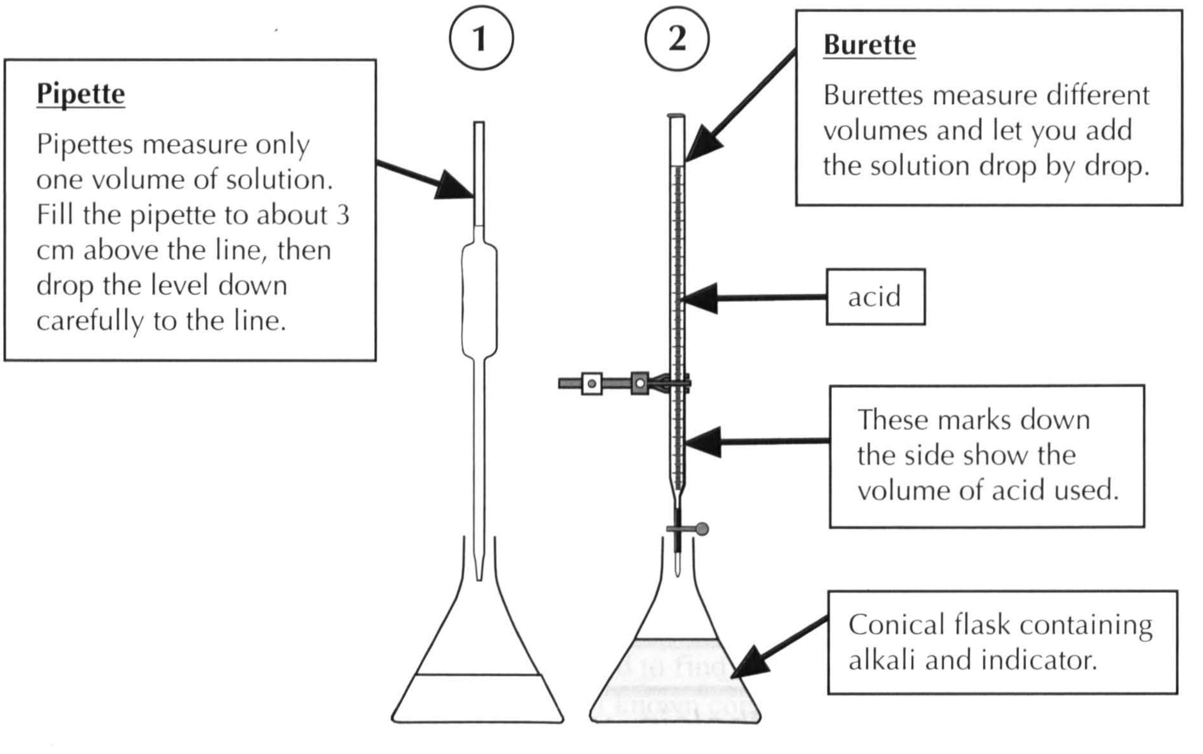
3) Using the burette, add the acid to the alkali a bit at a time — giving the conical flask a regular swirl.

4) Go especially slowly when you think the end-point (colour change) is about to be reached.

5) The indicator changes colour when all the alkali has been neutralised, e.g. phenolphthalein is pink in alkalis but colourless in acids.

6) Record the volume of acid used to neutralise the alkali.

7) It's best to repeat this process a few times, making sure you get (pretty much) the same answer each time — this makes for more reliable results.



Titrations

Titrations are often done in order to find out the concentration of a mystery solution of an acid or alkali. If you can find out the volume needed to neutralise a known solution, then all you need to do is slot all the numbers into a handy formula. The example below should make everything clear.

The calculation — work out the numbers of moles

Here goes... basically, you're trying to find the number of moles of each substance.

A formula triangle is pretty handy here, I reckon.
(And it's the same one as on page 209, conveniently.)

$$\frac{n}{c \times V}$$

Example:

> Suppose you start off with 25 cm^3 of sodium hydroxide in your flask, and you know that its concentration is 0.1 moles per dm^3.
>
> You then find from your titration that it takes 30 cm^3 of sulfuric acid (of an unknown concentration) to neutralise the sodium hydroxide.
>
> Find the concentration of the acid.

Step 1: Work out how many moles of the 'known' substance you have:

number of moles = concentration × volume = 0.1 × (25 / 1000) = 0.0025 moles

Step 2: Write down the equation for the reaction:

$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$$

...and work out how many moles of the 'unknown' stuff you must have had:

Using the equation, you can see that for every two moles of sodium hydroxide you had, there was just one mole of sulfuric acid.

So if you had 0.0025 moles of sodium hydroxide...

...you must have had 0.0025 ÷ 2 = 0.00125 moles of sulfuric acid.

Step 3: Work out the concentration of the 'unknown' stuff.

Concentration = number of moles ÷ volume

= 0.00125 ÷ (30 / 1000)

= 0.0417 moles per dm^3

If you need the concentration in g/dm^3, convert your answer using the method on page 209.

If you can spell phenolphthalein, you deserve a GCSE

The indicator's job is to tell you when the reaction is finished. Phenolphthalein is good for acids and alkalis, but other indicators are possible too. However, don't use universal indicator — it's too hard to tell accurately when the reaction is over. You want an indicator that gives a sudden colour change.

Warm-Up and Exam Questions

Come on now, don't look at me like that — it's been ages since you had two pages of questions to do at once. Anyway, I'll let you in on a secret — this is the last double page of questions in the book. Seriously.

Warm-Up Questions

1) A solution has a volume of 0.15 litres. Convert this volume into cm^3.
2) How many moles of hydrochloric acid are there in 25 cm^3 of a 0.1 mol/dm^3 solution?
3) A solution of sodium carbonate, Na_2CO_3, has a concentration of 2.65 g/dm^3. What is the concentration of this solution in mol/dm^3?
4) A sample of nitrogen gas occupies a volume of 280 cm^3 at room temperature and pressure. How many moles of nitrogen are there in the sample?
5) In an acid-base titration, what is the end-point?

Exam Questions

1 Paul works for a soft drinks manufacturer and is investigating the use of phosphoric acid, H_3PO_4, as a flavouring agent.

(a) Paul prepares a solution of phosphoric acid by dissolving 4.9 g of the acid in 250 cm^3 of water.

(i) What is the concentration of the acid solution in g/dm^3?

(1 mark)

(ii) What is the concentration of the acid solution in mol/dm^3?
Relative atomic masses: H = 1, P = 31, O = 16.

(2 marks)

(b) Paul wants to prepare a phosphoric acid solution with a concentration of 0.1 mol/dm^3. What mass of phosphoric acid should he dissolve in 1 dm^3 of water?

(1 mark)

2 Amrita heats 37.2 g of copper carbonate powder in a thermal decomposition reaction.

$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

(a) If all the copper carbonate reacts, what mass of carbon dioxide will be produced?
Relative atomic masses: Cu = 64, C = 12, O = 16.

(2 marks)

(b) At room temperature and pressure, what volume will the carbon dioxide occupy?
(1 mole of gas occupies a volume of 24 000 cm^3 at RTP.)

(2 marks)

3 You are asked to find the concentration of a bottle of hydrochloric acid solution. Describe how you could use a titration method to find the concentration of the acid if you had a solution of sodium hydroxide of known concentration. Include names of special equipment needed for this procedure.

(5 marks)

Exam Questions

4 In a titration, 27.5 cm^3 of a solution of 0.1 mol/dm^3 sodium hydroxide was required to neutralise 25 cm^3 of a solution of hydrochloric acid.

(a) Calculate the number of moles of sodium hydroxide used in the titration.

(2 marks)

(b) Use the equation for the reaction shown below to work out the number of moles of hydrochloric acid used in the titration:

$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$

(1 mark)

(c) Calculate the concentration of the hydrochloric acid solution.

(2 marks)

5 In a titration, 30.3 cm^3 of a solution of 1.0 mol/dm^3 sodium hydroxide was required to neutralise 25 cm^3 of a solution of sulfuric acid.

(a) Calculate the number of moles of sodium hydroxide used in the titration.

(2 marks)

(b) Work out the number of moles of sulfuric acid used in the titration.

The equation is: $2NaOH\ (aq) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + 2H_2O\ (l)$

(1 mark)

(c) Calculate the concentration of the sulfuric acid solution.

(2 marks)

6 Jonah is concerned about the amount of acid in soft drinks. He decides to use a titration method to find the acid content of his favourite lemonade. He uses a solution of 0.1 mol/dm^3 sodium hydroxide in titrations with 25 cm^3 samples of the lemonade. His results are shown in the table.

	Initial burette reading (cm^3)	Final burette reading (cm^3)	Vol. of NaOH needed (cm^3)
1	0.0	9.4	9.4
2	9.4	18.4	9.0
3	18.4	27.4	9.0

Jonah calculates that the average volume of 0.1 mol/dm^3 NaOH needed is 9.0 cm^3.

(a) The first titration value was not included in calculating the average. Why not?

(1 mark)

(b) The equation for the reaction in the titration can be written:

$HA + NaOH \rightarrow NaA + H_2O$, where HA is the acid present in the lemonade.

Calculate the concentration of acid HA present in the lemonade.

(3 marks)

(c) The calculated acid concentration in (b) is relatively high. However, this isn't a cause for too much concern. Use ideas about the properties of different acids in solution to explain why the high concentration isn't dangerous.

(2 marks)

Electrolysis

Think back — you've seen electrolysis before... electrical current, ions moving to electrodes, positive ions to the cathode, negative ions to the anode... If this is all a bit fuzzy, go back and read p.144-145 first.

Electrolysis reactions involve **oxidation** and **reduction**

1) Oxidation is a loss of electrons. Reduction is a gain of electrons. So 'oxidation' doesn't have to involve oxygen.

Oxidation Is Loss **Reduction Is Gain**

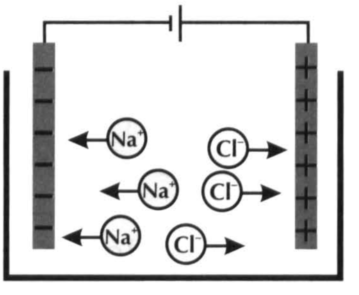
Remember it as OIL RIG.

2) Electrolysis ALWAYS involves an oxidation and a reduction.

For example, if electrodes are placed in molten sodium chloride and a current is passed through...

(i) At the (positive) anode, chloride ions (Cl^-) lose electrons and form chlorine molecules — this is oxidation.

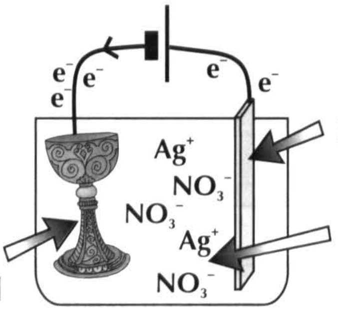
(ii) At the (negative) cathode, sodium ions (Na^+) gain electrons and become sodium atoms — this is reduction.



3) So, in electrolysis, ions are discharged. They lose their negative or positive charge and become atoms.

Graphite electrodes or **metal electrodes** are used

1) Electrodes are made of different materials depending on what you're trying to do.
2) When splitting up an ionic salt by electrolysis, graphite electrodes are most often used (as it's cheap).
3) But when using electrolysis to split water into hydrogen and oxygen, platinum electrodes are usually used (since it doesn't react with the oxygen produced).
4) Electrolysis is used for electroplating. Here, the cathode is the metal object you want to plate, and the anode is the pure metal you want it to be plated with. You also need the electrolyte to contain ions of the plating metal. (The ions that plate the metal object come from the solution, while the anode keeps the solution 'topped up'.)



Example: To electroplate silver onto a brass cup, you'd make the brass cup the cathode (–ve), a lump of pure silver the anode (+ve) and dip them in a solution of silver ions, e.g. silver nitrate.

Electrolysis

If you electrolyse certain solutions, you get different products from what you might expect...

Sometimes H^+ and OH^- ions from water are discharged instead

1) In aqueous solutions there are hydrogen ions (H^+) and hydroxide ions (OH^-) from the water as well as the ions from the solute.
2) Sometimes it's easier to discharge the ions from the water instead of the ones from the solute.
3) So hydrogen could be produced at the cathode, and oxygen at the anode.

Example:

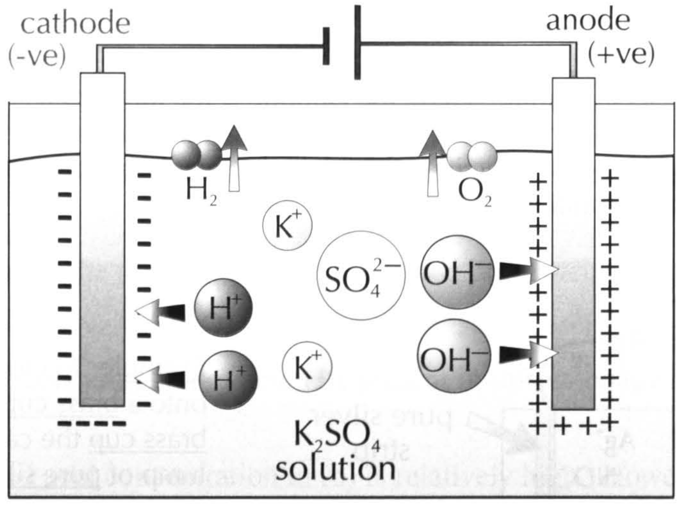
A solution of aqueous potassium sulfate (K_2SO_4) contains four different ions:
K^+, SO_4^{2-}, H^+ and OH^-.

- H^+ ions (from water) accept electrons more easily than K^+ ions.
- So hydrogen's given off at the cathode.

$$2H^+ + 2e^- \rightarrow H_2$$

- OH^- ions (from water) can lose electrons more easily than SO_4^{2-} ions.
- So oxygen's given off at the anode.

$$4OH^- - 4e^- \rightarrow O_2 + 2H_2O$$



The same thing happens with potassium nitrate solutions

So with aqueous solutions, there are ions from the dissolved substance and from the water. Remember the K_2SO_4 example, and the same thing happens with KNO_3 too — you get hydrogen and oxygen.

Electrolysis — Calculating Masses

Will these calculations never end? (Well, yes, it's nearly the end of the section, in fact.)

No. of electrons transferred increases with time and current

1) The amount of product made by electrolysis depends on the number of electrons that are transferred.

2) If you increase the number of electrons, you increase the amount of substance produced.

3) This can be achieved by:

- electrolysing for a longer time,
- increasing the current.

Coulombs and faradays are amounts of electricity

1) One amp flowing for one second means a charge of one coulomb has moved.
2) Generally, the amount of charge (Q, measured in coulombs) flowing through a circuit is equal to the current (I) multiplied by the time in seconds (t):

$$Q = It$$

3) 96 000 coulombs (amps × seconds) is called one faraday.
4) One faraday (F) contains one mole of electrons.

1 A for 1 s = 1 C

Q = I × t (seconds)

96 000 C = 1 faraday

1 faraday = 1 mole of electrons

The more time you spend on this page, the more you'll learn

This stuff isn't easy. So take your time over it. Read it through once. If you don't get it, read it through again. If you still don't get it, have a cup of tea before reading it again. That should help.

Electrolysis — Calculating Masses

You can use the relationships on the last page to predict the mass of a substance that you'll get if you carry out electrolysis with a known current for a known amount of time. It's quite clever really.

One mole of product needs n moles of electrons

1) A sodium ion needs one electron to make a sodium atom.
2) So one mole of sodium ions is going to need one mole of electrons (one faraday) to make one mole of sodium atoms.
3) But an ion with a 2^+ charge needs two moles of electrons to make one mole of atoms, and, guess what, three for a 3^+ charge...

$Na^+ + e^- \rightarrow Na$	1 mole of sodium ions + 1 mole of electrons → 1 mole of sodium atoms
$Zn^{2+} + 2e^- \rightarrow Zn$	1 mole of zinc ions + 2 moles of electrons → 1 mole of zinc atoms
$Al^{3+} + 3e^- \rightarrow Al$	1 mole of aluminium ions + 3 moles of electrons → 1 mole of aluminium atoms

Use these steps in calculations

Example:

Find the mass of lead liberated if 5 amps flows for 20 minutes during the electrolysis of lead(II) chloride ($PbCl_2$).

Step 1: Write out the BALANCED HALF-EQUATION for each electrode:

$Pb^{2+} + 2e^- \rightarrow Pb$, and $2Cl^- \rightarrow Cl_2 + 2e^-$

Writing the half-equations is easier if you remember that the full equation is: $PbCl_2 \rightarrow Pb + Cl_2$

Step 2: Calculate the NUMBER OF FARADAYS:

First calculate amps × seconds = 5 × 20 × 60 = 6000 coulombs.

Number of faradays = 6000 / 96 000 = 0.0625 F

Step 3: Calculate the NUMBER OF MOLES OF PRODUCT — divide the number of faradays by the number of electrons in the half-equation:

0.0625 ÷ 2 = 0.03125 moles of lead atoms.

Step 4: WRITE IN THE M_r VALUES from the periodic table to work out the mass of solid products:

Mass of lead = M_r × no. of moles = 207 × 0.03125 = 6.5 g

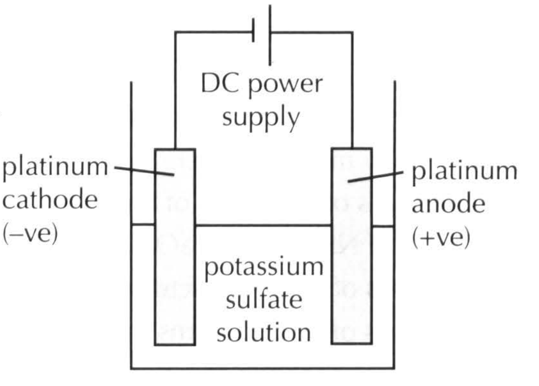
Warm-Up and Exam Questions

Warm-Up Questions

1) In electrolysis, what is meant by the terms oxidation and reduction?
2) At which electrode (anode or cathode) should an object requiring electroplating be placed?
3) During the electrolysis of water, what gas is formed at the anode?
4) What is a faraday?
5) For how long must a current of 2.5 A run if a charge of 200 000 C is to pass through a cell?

Exam Questions

1 The diagram shows the electrolysis cell for an aqueous solution of potassium sulfate, K_2SO_4. During electrolysis a gas forms at each electrode.



(a) Four ions are present in the electrolyte. Write the formula of each ion.
(2 marks)

(b) (i) Write the half-equation showing the production of the gas at the cathode.
(1 mark)

(ii) Write the half-equation showing the production of the gas at the anode.
(1 mark)

(c) Why is platinum used as a material for the electrodes?
(1 mark)

2 Electrolysis can be used to purify copper. Impure copper is used as one electrode in the cell. During electrolysis, this electrode dissolves as copper ions enter the solution.

(a) Is this dissolving an example of oxidation or reduction? Explain your answer.
(1 mark)

(b) At the other electrode pure copper forms.
The half-equation for the reaction at this electrode is: $Cu^{2+} + 2e^- \rightarrow Cu$
A direct current of 4 A runs for 2 hours through the cell.

(i) What is the total charge that passes through the cell in this time? (1 hr = 3600 s.)
(2 marks)

(ii) How many moles of electrons pass through the cell? (1 F = 96 000 C.)
(1 mark)

(iii) How many moles of pure copper are formed at the electrode?
(1 mark)

(iv) What mass of pure copper is formed? (Relative atomic mass: Cu = 64.)
(1 mark)

Revision Summary for Section Twelve

Ah, the revision summary... my favourite part of the section. And yours no doubt, since they're always at the end. There's a lot of calculations in this section, but that's good (honest), because you can expect a fair few in the exam as well. And as a wise man once said... it's best to practise before the exam, because once you're in there, it's a bit late really. So get your calculator fired up, and away you go...

1)* What is the concentration of a solution with 0.5 moles of sodium chloride in 200 cm^3?

2)* How many moles of barium chloride are in 500 cm^3 of a 0.2 molar solution of barium chloride?

3)* Calculate the concentration of the solution (in g/dm^3) formed when 7.5 g of calcium hydroxide, $Ca(OH)_2$, is dissolved in:
a) 1 dm^3 of water, b) 2 dm^3 of water.

4)* A 5 cm^3 sample of solution is heated until all the water has evaporated.
The mass of the evaporating basin and remaining solid is 52.7 g.
The mass of the evaporating basin is 52.1 g. What was the concentration of the solution in g/dm^3?

5)* What is the concentration in g/dm^3 of a 0.1 molar solution of NaOH?

6) What volume is occupied by 1 mole of nitrogen at room temperature and pressure?

7)* a) What's the volume of 6 moles of methane gas at RTP?
b) How many moles are there in 3600 cm^3 of chlorine gas at RTP?

8)* Give the mass of one mole of the following compounds:
a) CO_2, b) NH_3, c) NaOH, d) NaCl, e) H_2SO_4.

9)* a) What mass of NaOH reacts with HCl to give 23.4 g of NaCl?
b) What mass of carbon burns in oxygen to give 11 g of CO_2?
c) What mass of lithium reacts with water to give 48 g of LiOH?

10)*What volume of H_2 is formed when 5 g of calcium reacts with hydrochloric acid at RTP?

11) Why do you need to get several consistent readings in titrations?

12) a) Briefly describe how you would carry out a titration between 25 cm^3 of 0.2 mol/dm^3 KOH and an unknown concentration of HCl.
b)*You carry out the titration and find that it takes 49.8 cm^3 of HCl to neutralise 25 cm^3 of 0.2 mol/dm^3 KOH. What is the concentration of HCl used?

13) Why is an indicator like phenolphthalein used in titrations, rather than universal indicator?

14)*In a titration, 22.5 cm^3 of nitric acid was required to neutralise 25 cm^3 of potassium hydroxide with a concentration of 0.15 moles per dm^3. Calculate the concentration of the nitric acid in:
a) mol/dm^3, b) g/dm^3.

15) What kind of electrodes would be used for the electrolysis of sodium chloride solution?

16) Why is hydrogen released during the electrolysis of $K_2SO_4(aq)$?

17) When a sulfate solution is electrolysed, what gas comes off at the positive electrode?

18) Describe two ways in which you could increase the amount of product made during electrolysis.

19)*If 2 amps of current flows for 3 seconds, how much charge is that in coulombs?

20) What's the name for the amount of charge equal to 1 mole of electrons?
How many coulombs is this?

21) Write the half-equations at each electrode for the electrolysis of $PbI_2(l)$.

22)*If 3 amps flows for 30 minutes in the electrolysis of copper(II) chloride solution, find:
a) the mass of copper formed, b) the volume of chlorine formed (at RTP).

* Answers on page 263.

Chemical Production

Fast reactions and high percentage yields are nice in industry, but the most important thing is keeping costs down.

The type of manufacturing process depends on the product

There are two main types that you need to know about.

Continuous production

This is used in large-scale industrial manufacture of chemicals, for example the Haber process (see page 167).

1) Production never stops, so you don't waste time emptying and setting up the reactor.
2) It runs automatically — you only need to interfere if something goes wrong.
3) The quality of the product is very consistent.
4) The plant normally only makes one product, so there's a low risk of contamination.
5) But start-up costs to build the plant are huge, and it isn't cost-effective to run it at less than full capacity.

Batch production

This is used to make small quantities of specialist chemicals, for example pharmaceutical drugs. Products are often made on demand (i.e. only when there's an order for them, not all the time).

1) It's flexible — several different products can be made using the same equipment.
2) Start-up costs are relatively low — small-scale, multi-purpose equipment can be bought off the shelf.
3) It's very labour-intensive — the equipment needs to be set up and manually controlled for each batch.
4) 'Downtime' between batches means sometimes you're producing nothing.
5) It's harder to keep the quality consistent. And there's more chance of contamination as the same equipment is used to make different things. But any problem can easily be traced to a specific batch.

Pharmaceutical drugs are complicated to make and there's relatively low demand for them. So, batch production is often the most cost-effective way to produce small quantities of different drugs to order.

Chemical Production

Production cost depends on five main factors

It's these factors that companies have to consider when deciding if, and then how, to make a chemical.

1) Price of energy

Companies need to keep energy bills as low as possible. If a reaction needs a high temperature, the running costs will be higher.

2) Cost of raw materials

This is kept to a minimum by recycling any materials that haven't reacted, like in the Haber process (see page 167).

3) Labour costs (wages)

Everyone who works for a company has to be paid, so labour-intensive processes (those that involve many people), can be very expensive. Automation cuts running costs by reducing the number of people involved, but companies always have to weigh any savings they make on their wage bill against the initial cost and running costs of the machinery.

4) Plant costs (equipment)

The cost of equipment depends on the conditions it has to cope with. For example, it costs far more to make something to withstand very high pressures than something which only needs to work at atmospheric pressure.

5) Rate of production

Usually the faster the reaction goes, the better it is in terms of reducing the time and cost of production, so rates of reaction are often increased by using catalysts. But the increase in production rate has to balance the cost of buying the catalyst and replacing any that gets lost.

Optimum conditions are chosen to give the lowest cost

1) Optimum conditions are those that give the lowest production cost per kg of product — even if this means compromising on the speed of reaction or percentage yield.
2) But the speed and percentage yield must both be high enough to make enough product each day.
3) Don't forget, a low percentage yield is okay, as long as the starting materials can be recycled.

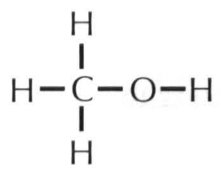
Alcohols

You need to learn the structure, physical properties, chemical properties and uses of alcohols.

Alcohols have an '-OH' functional group and end in '-ol'

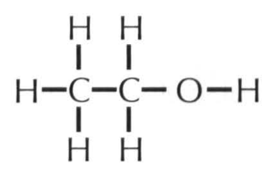
1) The general formula of an alcohol is $C_nH_{2n+1}OH$.
2) So an alcohol with two carbons has the formula C_2H_5OH.
3) The basic naming system is the same as for alkanes — but replace the final '-e' with '-ol'.

methanol



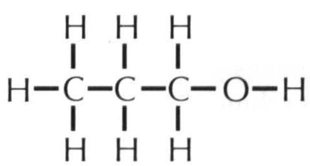
CH_3OH

ethanol



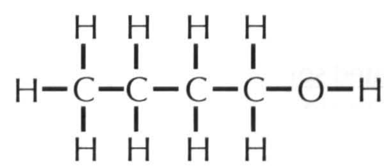
C_2H_5OH

propanol



C_3H_7OH

butanol



C_4H_9OH

4) Don't write CH_4O instead of CH_3OH, or C_2H_6O instead of C_2H_5OH — it doesn't show the functional –OH group.
5) The functional group is what makes all members of a family of compounds (e.g. alcohols, alkenes) react in a similar way.

The *first five alcohols* are *clear colourless liquids* (at room temp.)

1) Alcohols are flammable. They burn to produce carbon dioxide and water.
2) Methanol and ethanol evaporate easily and give off fumes (i.e. they're volatile). This means they should be stored in closed containers away from heat sources (e.g. naked flames). You wouldn't want to accidentally set fire to a cloud of alcohol vapour. Very unpleasant.
3) The first three alcohols all mix completely with water.
4) Alcohols react with oxygen to produce carboxylic acids (see page 227 for the equation). And alcohols react with carboxylic acids to produce esters (page 228).
5) All alcohols are toxic to some degree. Methanol is much more toxic than ethanol, and causes blindness if it's drunk. Ethanol (the alcohol in alcoholic drinks) damages the liver and brain.

Alcohols

Just a few uses and a dehydration reaction to commit to memory before you move on to a new topic.

Alcohols *are used as* ***solvents***

1) Alcohols such as methanol and ethanol can dissolve most compounds that water dissolves, but they can also dissolve substances that water can't dissolve — e.g. hydrocarbons, oils and fats.

   The 'carbon chain' end of the alcohol molecule can mix with oils and other carbon-chain compounds. And the –OH group can mix with water and ionic compounds.

   This makes ethanol, methanol and propanol very useful solvents in industry.

2) Ethanol is the solvent for perfumes and aftershave lotions. It can mix with both the oils (which give the fragrance) and the water (that makes up the bulk of the perfume or lotion).

3) 'Methylated spirit' (or 'meths') is ethanol with other chemicals (such as methanol) added to it. It's used to clean paint brushes and as a fuel (among other things). It's poisonous to drink, so a purply-blue dye is also added (to try and prevent people drinking it by mistake).

Ethanol can be ***dehydrated*** *back to ethene*

1) The plastics and polymers industry uses lots of ethene.
2) Countries which have no oil but plenty of land for growing crops for fermentation can make ethene through the dehydration of ethanol.
3) Ethanol vapour is passed over a hot aluminium oxide catalyst.

ethanol → ethene + water

$$C_2H_5OH \rightarrow C_2H_4 + H_2O$$

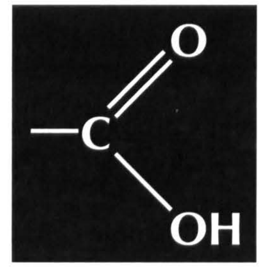
You get off lightly here — just one reaction to learn

Alcohols don't have many chemical reactions that you need to know about for GCSE. The examiners will be happy if you know the formulas, the physical properties and the uses (including why they're so good at dissolving oily substances). They're not usually so easy to please, so enjoy it while you can.

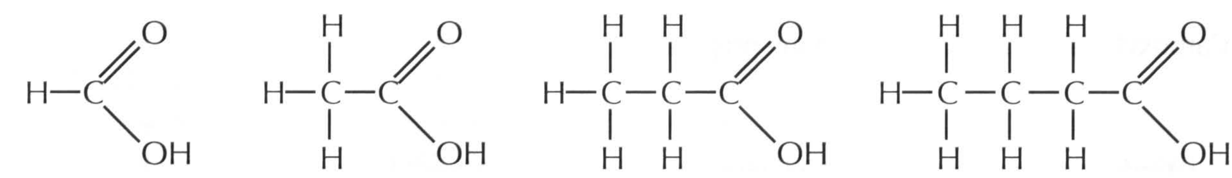
Carboxylic Acids and Esters

So what if carboxylic is a funny name — these are easy.

Carboxylic acids have the functional group –COOH



1) Carboxylic acids have '–COOH' as a functional group.
2) They're often called organic acids (since they're built around carbon atoms).
3) Their names end in '-anoic acid' (and start with the normal 'meth/eth/prop/but').



methanoic acid **ethanoic acid** **propanoic acid** **butanoic acid**

4) Organic acids mix with water, and with solvents like alcohols, alkanes, etc.
5) However, the longer the hydrocarbon chain, the less soluble in water an organic acid is.

Some carboxylic acids are fairly common

1) Ethanoic acid is the acid in vinegar, which is used for flavouring and preserving foods.
2) If wine or beer is left open to the air, the ethanol is oxidised to ethanoic acid. This is why wine that's been open for a few days tastes like vinegar — basically it is vinegar.

$$\text{ethanol} + \text{oxygen} \rightarrow \text{vinegar} + \text{water}$$
$$CH_3CH_2OH + O_2 \rightarrow CH_3COOH + H_2O$$

You can also write C_2H_5OH as CH_3CH_2OH — it shows the structure more clearly.

3) Citric acid (another organic acid) is present in oranges and lemons, and is manufactured in large quantities to make fizzy drinks. It's also used as a descaler (in kettles, etc.).
4) Aspirin is a man-made organic acid. It's useful in medicine, e.g. as a painkiller — see page 233.

Carboxylic Acids and Esters

Carboxylic acids have lots of uses, from making soaps and fabrics to preparing esters.

Organic acids are used in industry to make soaps and esters

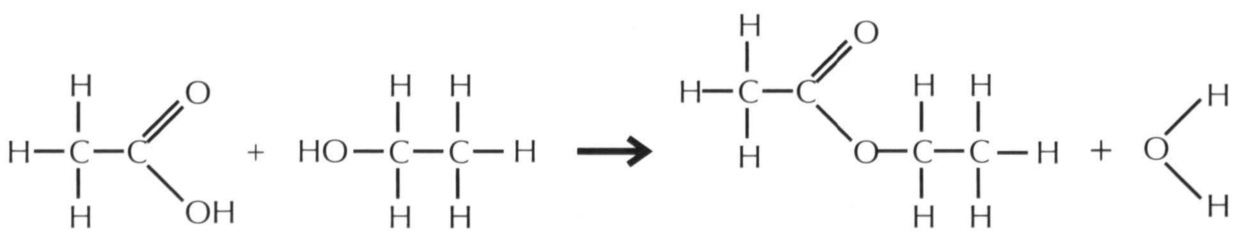
1) Carboxylic acids were first obtained from fats. The carboxylic acids with longer chains of carbon atoms are still often called fatty acids. These fatty acids are used in soaps and detergents.

2) Ethanoic acid is used in the manufacture of a form of rayon, a synthetic fibre.

3) Ethanoic acid is a very good solvent for many organic molecules. But ethanoic acid isn't usually chosen as a solvent because it would make the solution acidic.

4) Carboxylic acids are also used in the preparation of esters:

Carboxylic acids can be mixed with alcohols to make esters by esterification (see page 59). An example is shown below.



CH_3COOH	C_2H_5OH	$CH_3COOC_2H_5$	H_2O
ethanoic acid	ethanol	ethyl ethanoate	water

The alcohol forms the first part of the ester's name, and the acid forms the second part. For example:

- methanol + ethanoic acid → methyl ethanoate + water
- butanol + propanoic acid → butyl propanoate + water

Ethanoic acid — it's not just for putting on your chips

The esterification reaction's the one to learn. Elsewhere, carboxylic acids react just like any old weak acid (and have a pH less than 7). Examiners like you to be able to give some real-life uses too.

Warm-Up and Exam Questions

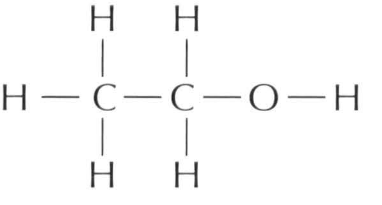
Just like day follows night, exam questions follow warm-up questions.

Warm-Up Questions

1) Give two advantages of using batch production rather than continuous production.
2) Optimum conditions for production are often described as a 'compromise'. Explain why.
3) What is the general formula of an alcohol?
4) Give two uses of methylated spirit.
5) Name three carboxylic acids and give one use of each.

Exam Questions

1 Alcohols are an important group of organic chemicals. The most widely used alcohol is ethanol. Its displayed formula is shown in the diagram.



(a) Write down the functional group of ethanol.

(1 mark)

(b) Ethanol can be dehydrated to form ethene, which is used to produce polymers.

(i) Copy and complete the equation showing the dehydration of ethanol.

$C_2H_5OH \rightarrow$ __________ + __________

(2 marks)

(ii) Industrially, ethene can also be produced by cracking hydrocarbon fractions from crude oil. When deciding which method to choose, the cost of raw materials (i.e. ethanol or crude oil) is important. Name two other chemical production factors that could influence the choice of which method to use.

(2 marks)

(iii) Both methods would favour a continuous production process. Explain why.

(2 marks)

2 Carboxylic acids are a widely used family of organic chemicals.

(a) Ethanoic acid is better known by the name of its dilute solution — vinegar. Draw its displayed (full structural) formula.

(2 marks)

(b) One use of carboxylic acids is in the production of esters.

(i) Name the ester formed when ethanoic acid is reacted with ethanol.

(1 mark)

(ii) Name the ester formed when propanoic acid reacts with methanol.

(1 mark)

(iii) When carboxylic acids and alcohols react to form esters, one other product is formed. Name this product.

(1 mark)

Drug Development

Lots of chemists end up working in the pharmaceutical industry — it's a huge business.
If a company can come up with the next wonder drug it stands to make a fortune (and save the odd life into the bargain.) So loads of time, effort and money is spent by drug companies on researching and developing new drugs.

Drugs are developed using staged synthesis

Imagine A, B, C and D represent chemicals that can react with each other to make a new compound ABCD. You can make the new compound in stages, like this:

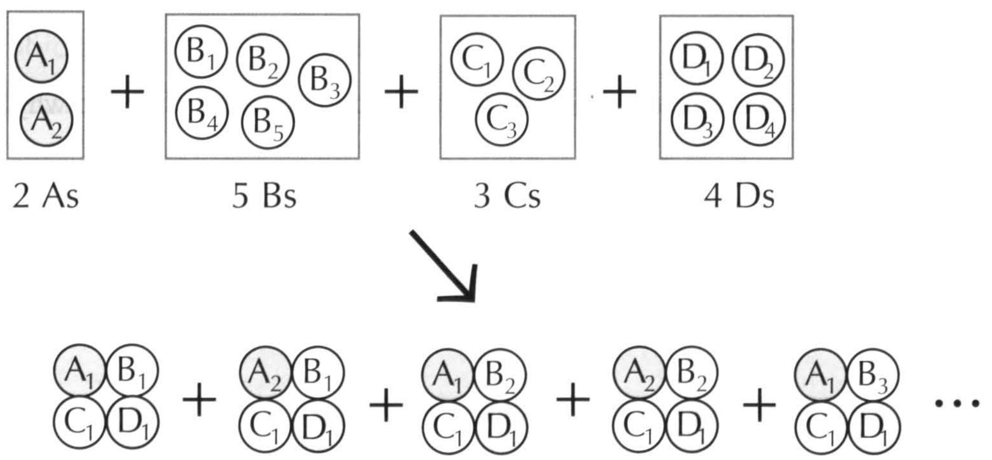
A + B → AB, then **AB + C → ABC**, then **ABC + D → ABCD**

1) When a company is trying to find a new drug, they like to have lots of possible substances to test. It's more cost-effective that way, and they don't risk missing out on a great new drug that was just a little bit different to a substance already being tested.

2) Staged synthesis is a way of making lots of similar substances, very quickly.

3) To make a 'family' of compounds similar to ABCD, a drug company would react a family of substances similar to A (call them A_1, A_2, etc.) with a family similar to B (B_1, B_2, etc.) and so on...

4) They're reacted in such a way as to give every possible combination — for example:



5) You can calculate the total number of products from a reaction like this by multiplying together the number of reactants in each family, i.e. the numbers of As, Bs, Cs and Ds involved in the reaction.

So in this example, you would end up with 2 × 5 × 3 × 4 = 120 different products.

Drug Development

Several factors affect the cost of pharmaceutical drugs

It takes on average 12 years and £900 million to develop a new drug and get it onto the market.

Market research

Market research tries to identify whether there's a market for a new drug. It might happen alongside research and development. It aims to answer questions like: Is there any competition already out there? Is there enough demand for it to be worthwhile developing the new drug?

Research and development

Research and development involves finding a suitable compound, testing it, modifying it, and then testing it again until it's ready. This involves the work of lots of highly paid scientists.

Trialling

No drug can be sold until it's gone through loads of time-consuming tests including animal trials and human trials to prove that it works and that it's safe.

Marketing

Marketing could involve advertising in medical magazines and buttering up doctors. The company might also fund clinical trials to help prove that their drug works better than a competitor.

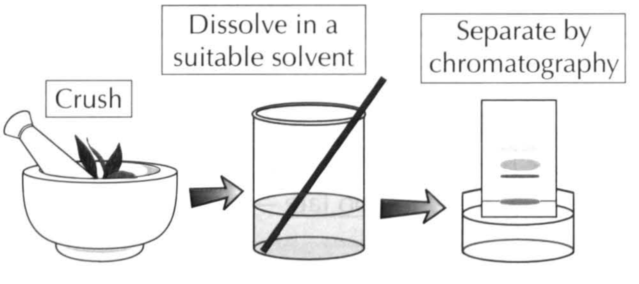
Manufacture

Multi-step batch production is labour-intensive and can't be automated. Other costs include energy and raw materials. The raw materials for pharmaceuticals are often rare and sometimes need to be extracted from plants (an expensive process — see below).

The actual price per dose of a new drug depends on the demand and how long the company is willing to wait to get back its initial investment. A company only holds a drug patent for 20 years — after that anyone can make it. Some drugs can cost thousands for just one dose.

Active ingredients can be extracted from plants

To extract a substance from a plant, it has to be crushed and dissolved in a suitable solvent. Then you can extract the substance you want by chromatography.



Once the active ingredient has been isolated, it can be analysed and its chemical structure worked out.

It's often possible to make a synthetic version of the chemical.

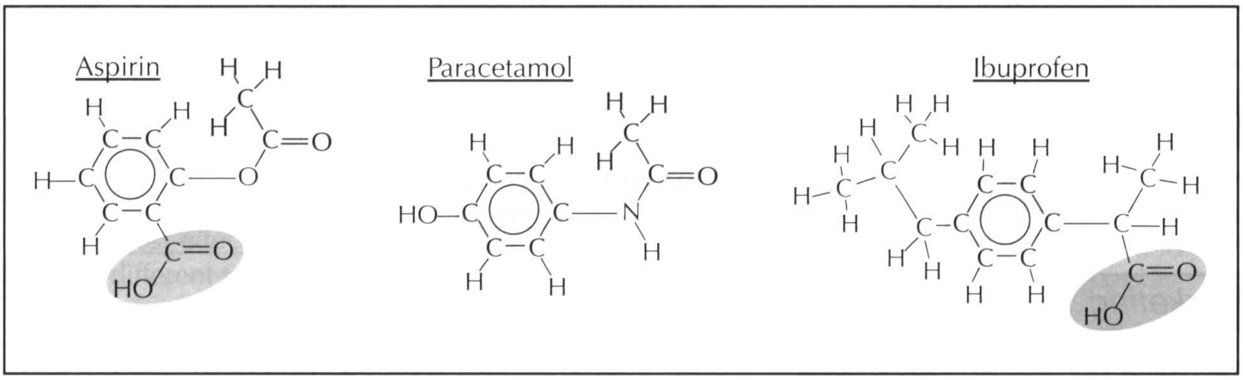
Painkillers

Drugs include anything externally administered that alters the chemical reactions in your body.

Analgesics are painkillers

1) Analgesics are drugs used to reduce pain, e.g. aspirin, paracetamol and ibuprofen.
2) The chemicals used in making the analgesics (or any drug) must be very pure. Any impurity could produce chemicals which cause unwanted or dangerous side effects.

Here are the displayed formulas of aspirin, paracetamol and ibuprofen:



Don't panic — you'll be given these formulas if you're asked questions about them.

1) There are similarities between them, e.g. they all have a benzene ring (in blue).
2) Aspirin and paracetamol both have a $-COCH_3$ group (highlighted in green).
3) Aspirin and ibuprofen both have a $-COOH$ group (highlighted in pink).
4) Paracetamol's different as it contains an N atom and an OH group's attached directly to the benzene ring.

Painkillers can be dangerous

An overdose of ASPIRIN can lower blood pressure, raise heart rate and cause breathing problems. It can also irritate the stomach, causing nausea and vomiting and even internal bleeding. An overdose can be fatal, but if it's caught in time there's a fairly good chance of recovery.

PARACETAMOL overdose causes horrendous liver damage. As little as 10–15 g (20–30 tablets) taken in one go can be fatal. It's particularly dangerous because the damage sometimes isn't apparent for 4–6 days after the drug's been taken. By that time, it's too late — the patient dies from liver failure.

Painkillers

Time to focus a bit more on aspirin — that's the painkiller you need to know most about.
It's the most interesting really, because it's made from trees and there's even a soluble version.
And if that hasn't sold it to you, well, I don't know what will...

Aspirin is made from salicylic acid

1) Salicylic acid occurs naturally in willow trees.
2) It was first isolated for use as an analgesic way back in 1828.
3) Unfortunately, it also caused mouth ulcers and reacted with the stomach lining.
4) Aspirin (acetylsalicylic acid) is manufactured from salicylic acid, but doesn't cause these side effects.
5) This doesn't mean that aspirin is completely problem-free — it can still aggravate existing stomach ulcers, and other possible side effects include headaches, dizziness and ringing in the ears.

Industrial method used to manufacture aspirin

- Salicylic acid is mixed with a substance called ethanoic anhydride and an organic solvent.
- The mixture is heated to 90 °C for 24 hours.
- It is then cooled for 3–4 days.

Soluble aspirin works faster than normal aspirin

1) Prostaglandins are chemicals which cause swelling and are involved in the feeling of pain.
2) Aspirin works at the site of a painful area by stopping prostaglandins being made.
3) Aspirin molecules are not very soluble, so they dissolve only slightly in the blood and so get to the painful area quite slowly.
4) Soluble aspirin allows quick absorption into the blood and faster relief of symptoms.
5) It is produced by reacting ordinary aspirin with sodium hydroxide or sodium carbonate.
6) Acetylsalicylic acid reacts with the alkali to give a salt and water.
7) The salt is the soluble form of aspirin.
8) Ordinary aspirin is completely covalent — there are no ions available for water to latch on to. But soluble aspirin is a negatively charged ion, so water can latch on and dissolve it.

Aspirin's covalent structure means it's not soluble

Remember, charged ionic structures are easier to surround with polar water molecules and so dissolve.

Making Sulfuric Acid

And here's another example where getting the conditions right makes for a better yield.

The Contact process is used to make sulfuric acid

1) The first stage is to make sulfur dioxide (SO_2) — usually by burning sulfur in air or by roasting sulfide ores.

$$\text{sulfur} + \text{oxygen} \rightarrow \text{sulfur dioxide}$$
$$S + O_2 \rightarrow SO_2$$

If you use, say, zinc sulfide, the equation is: $2ZnS + 3O_2 \rightarrow 2SO_2 + 2ZnO$

2) The sulfur dioxide is then oxidised (with the help of a vanadium pentoxide catalyst) to make sulfur trioxide (SO_3).

$$\text{sulfur dioxide} + \text{oxygen} \rightleftharpoons \text{sulfur trioxide}$$
$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

3) Next, the sulfur trioxide is dissolved in concentrated sulfuric acid to form fuming sulfuric acid, or oleum.

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$

You can't just dissolve SO_3 in water — because the reaction gives out enough heat to evaporate the sulfuric acid.

4) Finally, oleum is diluted with measured amounts of water to form concentrated sulfuric acid.

$$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$$

Making Sulfuric Acid

The second stage of making sulfuric acid is reversible so it needs just the right conditions.

The **conditions** used to make SO_3 are **carefully chosen**

The reaction in step 2 is reversible. So, the conditions used can be controlled to get more product.

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

See p.166 for more on changing the position of equilibrium.

Temperature

1) Oxidising sulfur dioxide to form sulfur trioxide is exothermic (it gives out heat).
2) So to get more product you'd think the temperature should be reduced (so the equilibrium will shift to the right to replace the heat).
3) Unfortunately, reducing the temperature slows the reaction right down — not much good.
4) So an optimum temperature of 450 °C is used, as a compromise.

Pressure

1) There are two moles of product, compared to three moles of reactants.
2) So to get more product, you'd think the pressure should be increased (so that the equilibrium will shift to the right to reduce the pressure).
3) But increasing the pressure is expensive, and as the equilibrium is already on the right, it's not really necessary. (And increasing the pressure liquefies the SO_2, so it's no use anyway.)
4) In fact, atmospheric pressure is used.

Catalyst

1) To increase the rate of reaction a vanadium pentoxide catalyst (V_2O_5) is used.
2) This DOESN'T change the position of the equilibrium.

With a fairly high temperature, a low pressure and a vanadium pentoxide catalyst, the reaction goes pretty quickly and you get a good yield of SO_3 (about 99%).

Conditions for the Contact Process

- **Temperature: 450 °C.**
- **Pressure: 1-2 atmospheres.**
- **Catalyst: vanadium pentoxide.**

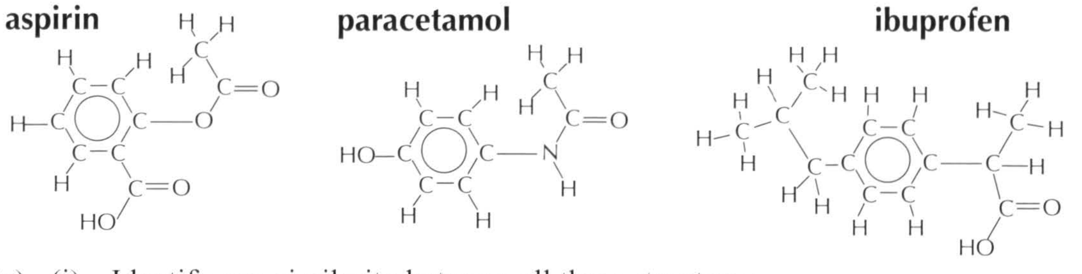
Warm-Up and Exam Questions

Warm-Up Questions

1) What is the advantage of using staged synthesis in drug development?
2) Give five costs for a company involved in developing and bringing to market a new drug.
3) What is an analgesic?
4) How is soluble aspirin made from 'ordinary' aspirin?
5) What catalyst is used in the manufacture of sulfuric acid?

Exam Questions

1 The diagram shows the displayed formulas of three common painkillers.



(a) (i) Identify one similarity between all three structures.

(1 mark)

(ii) Identify one similarity between the structures of aspirin and paracetamol that is not found in ibuprofen.

(1 mark)

(iii) What is the molecular formula of paracetamol?

(1 mark)

(b) A pharmaceuticals company wants to develop a new painkiller. As part of its research and development it needs to extract and isolate a compound from a plant. Describe an experimental method for how this could be done.

(3 marks)

2 In industry, sulfuric acid is made from sulfur in a multi-step process.

(a) What is the name of this process?

(1 mark)

(b) One step in the process involves a reversible reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

(i) The forward reaction is exothermic and so low temperatures would favour a high yield, but the reaction vessel is kept at about 450 °C. Explain why.

(1 mark)

(ii) What pressure is the reaction vessel kept at? Explain why this is chosen.

(3 marks)

(c) The sulfur trioxide produced reacts with sulfuric acid to produce a compound called oleum. Write a balanced symbol equation to show this reaction.

(2 marks)

Fuel Cells

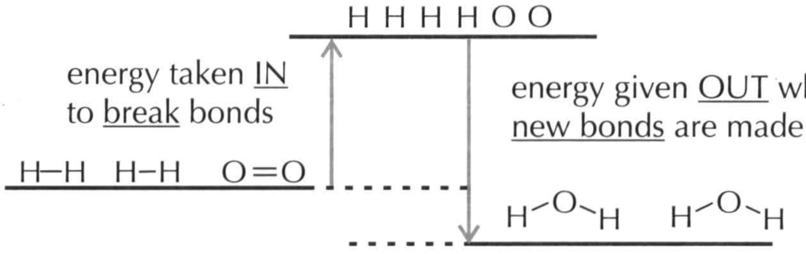
Fuel cells are great — they use hydrogen and oxygen to make electricity.

Hydrogen and oxygen give out energy when they react

Remember the lab tests for hydrogen and oxygen. Hydrogen plus a lighted splint gives a squeaky pop. Oxygen relights a glowing splint.

1) Hydrogen and oxygen react to produce water — which isn't a pollutant.
2) The reaction between hydrogen and oxygen is exothermic — it releases energy.
3) Put these two facts together, and you get something useful: you can get energy by reacting hydrogen and oxygen — and it doesn't produce any nasty pollutants, only nice clean water...

Energy Level Diagram



Fuel cells use fuel and oxygen to produce electrical energy

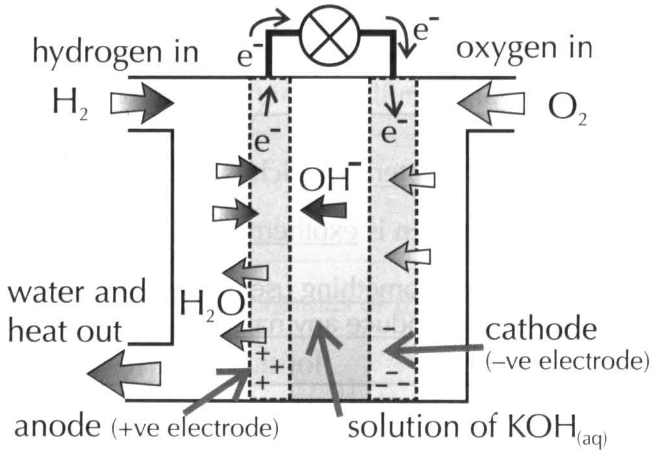
A fuel cell is an electrical cell that's supplied with a fuel and oxygen and uses energy from the reaction between them to generate an electrical voltage.

I'd learn that if I were you.

1) Fuel cells were developed in the 1960s as part of the space programme, to provide electrical power on spacecraft — they were more practical than solar cells and safer than nuclear power. (They're still used on the Space Shuttle missions.)
2) Unlike a battery, a fuel cell doesn't run down or need recharging from the mains. It'll produce energy in the form of electricity and heat as long as fuel is supplied.
3) There are a few different types of fuel cells, using different fuels and different electrolytes. The one they want you to know about is the hydrogen-oxygen fuel cell.

Fuel Cells

Hydrogen-oxygen fuel cells involve a redox reaction



1) The electrolyte is often a solution of potassium hydroxide, and the electrodes are often porous carbon with a catalyst.
2) Hydrogen goes into the anode compartment and oxygen goes into the cathode compartment.
3) At the negative cathode, oxygen gains electrons (from the cathode) and reacts with water (from the electrolyte) to make OH^- ions.

$$O_2 + 4e^- + 2H_2O \rightarrow 4OH^-$$

4) The oxygen gas is gaining electrons — this is reduction.
5) OH^- ions in the electrolyte move to the positive anode.
6) At the positive anode, hydrogen combines with the hydroxide ions to produce water and electrons.

$$2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$$

7) The hydrogen gas loses electrons — this is oxidation.
8) The electrons flow through an external circuit from the anode to the cathode — this is the current.
9) The overall reaction is hydrogen plus oxygen, which gives water.

$$2H_2 + O_2 \rightarrow 2H_2O$$

They can ask you for this equation in the exam — luckily it's a nice simple one.

There's reduction at the cathode and oxidation at the anode, so the whole thing is a REDOX reaction.

Fuel cells don't produce nasty carbon emissions or poisonous gases

This page is tough. If you want the marks, you'll have to learn it — even the nasty fuel cell diagram and equations. They can ask what happens at each electrode, including oxidation and reduction.

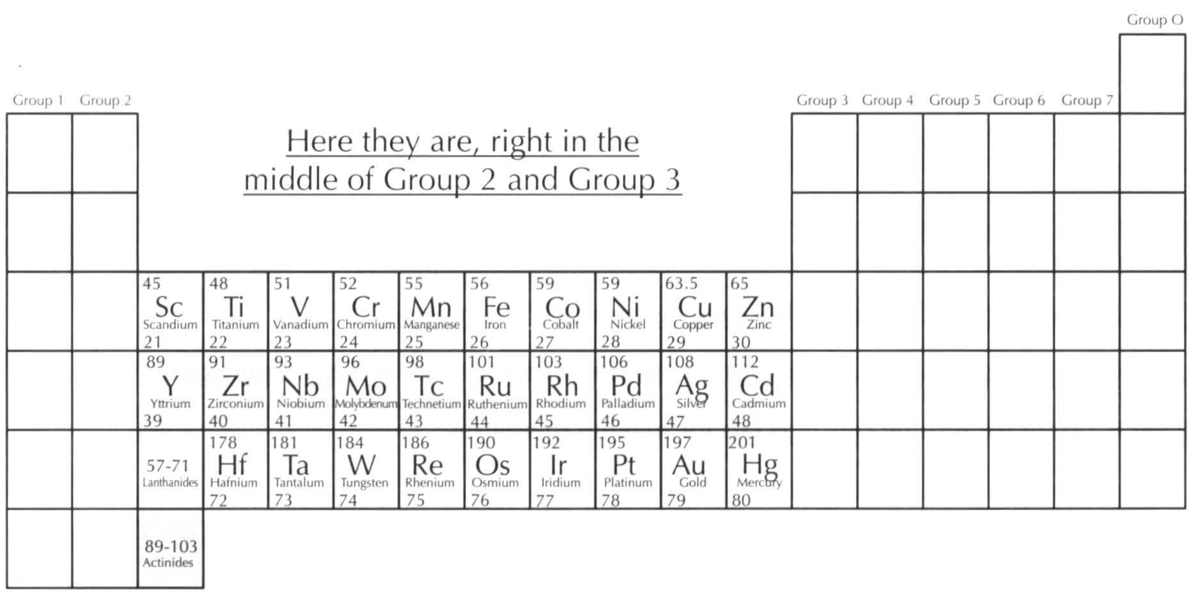
Transition Metals

Transition elements make up the big clump of metals in the middle of the periodic table.

Transition elements tend to have the *properties* of a *typical metal*

Transition metals are typical metals and generally have the properties you'd expect of a 'proper' metal:

1) They're good conductors of heat and electricity.
2) They're very dense, strong and shiny.
3) Transition metals are much less reactive than Group 1 metals — they don't react very much with water or oxygen, for example.
4) They're also much denser, stronger and harder than the Group 1 metals, and have much higher melting points (except for mercury, which is a liquid at room temperature). For example, iron melts at 1500 °C, copper at 1100 °C and zinc at 400 °C.



Transition metals often have *more than one ion*, e.g. Fe^{2+}, Fe^{3+}

1) Two other examples are:
   - copper: Cu^{+} and Cu^{2+}
   - chromium: Cr^{2+}, Cr^{3+} and Cr^{6+}

2) The different ions usually form different-coloured compounds too:
   - Fe^{2+} ions usually give green compounds.
   - Fe^{3+} ions usually form red/brown compounds (e.g. rust).

Transition Metals

The compounds are very colourful

1) The compounds are colourful due to the transition metal ion they contain — for example:
   - Potassium chromate(VI) is yellow.
   - Potassium manganate(VII) is purple.
   - Copper(II) sulfate is blue.

2) Transition metals are responsible for the colours in:
   - People's hair.
   - Gemstones, like blue sapphires and green emeralds.
   - Some pottery glazes.

   And weathered (oxidised) copper is a lovely colourful green.

Transition metals and their compounds make good catalysts

1) Iron is the catalyst used in the Haber process for making ammonia.
2) Manganese(IV) oxide is a good catalyst for the decomposition of hydrogen peroxide.
3) Nickel is useful for turning oils into fats for making margarine.

Their properties are due to the way their electron shells fill

1) In an atom, as you get further from the nucleus, energy levels get closer together until they start to overlap. This first happens between energy levels 3 and 4. It affects the way the electron shells fill.
2) Potassium has 19 electrons — but the 19th electron goes into the 4th energy level, not the 3rd. The electron arrangement's 2, 8, 8, 1. Same thing with the next element, calcium — which is 2, 8, 8, 2.
3) The next ten elements (the transition metals) put their electrons into the overlapping third energy level until it's full.

You don't need to know how this causes their various properties, just that it does.

Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
2,8,9,2	2,8,10,2	2,8,11,2	2,8,13,1 ↑	2,8,13,2	2,8,14,2	2,8,15,2	2,8,16,2	2,8,18,1 ↑	2,8,18,2

(Chromium (Cr) and copper (Cu) fill up a bit differently. The reason's complicated (A2-level), so for now just learn the numbers.)

Shiny metals, pretty colours, electrons — these elements have it all...

Most common everyday metals are transition elements — for example, iron, nickel, copper, silver, gold, and so on. There are a lot of facts to learn here about colour and melting points. Learn the weird fact about electron shells and impress the examiners.

Warm-Up and Exam Questions

You know the drill, do the warm-up questions, then have a crack at the exam questions.

Warm-Up Questions

1) What is a fuel cell?
2) In the hydrogen-oxygen fuel cell, what solution is often used as the electrolyte?
3) In an atom, what is the maximum number of electrons that the third energy level can hold?
4) Give three typical properties of transition metals.

Exam Questions

1 (a) Fuel cells involve redox reactions.

(i) One of the half-equations for the hydrogen-oxygen fuel cell is:

$O_2 + 4e^- + 2H_2O \rightarrow 4OH^-$

Is this an oxidation or a reduction? Explain your choice.

(1 mark)

(ii) What is the overall chemical equation for the hydrogen-oxygen fuel cell?

(2 marks)

(b) Fuel cells were developed to use on spacecraft. Give one advantage of using fuel cells to provide energy on long space journeys.

(1 mark)

(c) Why might the car industry be investing heavily in researching fuel cells as a future source of power for vehicles?

(1 mark)

2 (a) These are the electronic configurations of five elements in the periodic table:

Element A: 2.8.18.5 Element B: 2.8.8.2 Element C: 2.8.9.2

Element D: 2.8.8.1 Element E: 2.8.16.2

Which element(s) are transition metals?

(2 marks)

(b) Iron is a transition metal. Sodium is a Group I metal (alkali metal).

(i) Give two differences between the physical properties of iron and sodium.

(2 marks)

(ii) Give one difference in the chemical properties of iron and sodium.

(1 mark)

(c) Copper can exist as two ions, Cu^+ and Cu^{2+}.

(i) Which copper ion is present in copper sulfate solution, $CuSO_4(aq)$?

(1 mark)

(ii) What feature of the appearance of copper sulfate solution, $CuSO_4(aq)$, is characteristic of transition metal compounds?

(1 mark)

Industrial Salt

In hot countries they get salt by pouring sea water into big flat open tanks and letting the Sun evaporate the water, leaving the salt behind. This is no good in Britain though — there isn't enough sunshine.

Salt is mined from underneath Cheshire

1) In Britain, the salt comes from underground deposits left behind millions of years ago when ancient seas evaporated. There are massive deposits of this rock salt under Cheshire and Teeside.
2) Rock salt is a mixture of salt and impurities. It's drilled, blasted, dug out and brought to the surface.
3) It can also be mined by pumping hot water underground. The salt dissolves and the salt solution is forced to the surface by the pressure of the water.

Salt for roads is just blasted and dug out and used like that. Salt for food and chemicals has to be extracted with hot water.

4) Rock salt can be used in its raw state on roads to stop ice forming, or the salt can be filtered out and used to enhance the flavour in food or for making chemicals. If salt's going to be used to make chemicals, it is electrolysed (see page 145). This gives some very useful products.

The electrolysis of brine gives three very useful products

On page 145 you learnt that electrolysis of salt gives three useful products. And here they are...

1) Chlorine

(Don't forget the simple lab test for chlorine — it bleaches damp litmus paper.)

Used in: a) disinfectants b) killing bacteria (e.g. in swimming pools) c) household bleach d) plastics (e.g. PVC) e) HCl f) insecticides

2) Hydrogen

Used in: a) the Haber process to make ammonia
b) changing oils into fats for making margarine

3) Sodium hydroxide

This is a very strong alkali used widely in the chemical industry — e.g. it's used to make:

a) soap b) ceramics c) organic chemicals
d) paper pulp e) oven cleaner f) household bleach (see below)

ABOUT BLEACH — Household bleach is made by reacting chlorine with sodium hydroxide.

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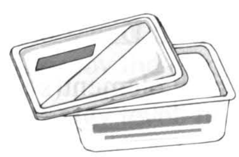

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Industrial Salt

You saw on page 218 that the electrolysis of some solutions, like potassium sulfate and potassium nitrate, leads to ions from the water being discharged rather than ions from the solute. Well, if your solution of sodium chloride is dilute enough, the same thing happens.

Dilute brine produces oxygen — not chlorine

1) A solution of sodium chloride contains sodium ions and chloride ions, but it also contains a few hydrogen ions (H^+) and hydroxide ions (OH^-) from the water.

2) As it happens, the OH^- is the most easily discharged anion. In other words, it'll come out of solution first, before other negative ions (see page 218 for a bit more info).

3) So in a dilute solution of brine, the OH^- ions are discharged before the chloride ions, and so oxygen is produced at the anode.

$$4OH^- - 4e^- \rightarrow O_2 + 2H_2O$$

4) In a concentrated solution of brine, there are loads of chloride ions and very few OH^- ions. The chloride ions win by sheer numbers — chlorine is formed at the anode.

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

Molten brine produces sodium — not hydrogen

Remember that if you electrolyse molten sodium chloride, you get sodium at the cathode and chlorine at the anode. There's no hydrogen produced because there's no water to provide the H^+ ions.

You need to learn the equations for the reactions at the electrodes:

$$Na^+ + e^- \rightarrow Na$$

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

Salt — it's a tasty little resource for industrial chemists...

It's not much trouble to learn a few products made from chlorine, hydrogen and sodium hydroxide. And make sure you know how bleach is made. One last thing... get the differences between electrolysis of dilute NaCl solution and electrolysis of concentrated NaCl solution sorted (the OH^- ions stuff).

CFCs and the Ozone Layer

Chlorofluorocarbons (CFCs) are organic molecules containing carbon, chlorine and fluorine, e.g. CCl_2F_2. At first they were thought to be a great invention and were used for all kinds of stuff, like coolants in fridges and propellants in aerosols. But it turned out that they were actually damaging the ozone layer, a layer of O_3 (a form of oxygen) which protects us by absorbing UV light from the Sun. Now that it's getting thinner, there's a greater risk of sunburn, skin cancer, premature ageing of the skin, cataracts...

Free radicals are made by breaking covalent bonds

1) A covalent bond, remember, is one where two atoms share electrons between them, like in H_2.
2) A covalent bond can break unevenly to form two ions, e.g. **H–H → H^+ + H^-**.
   The H^- has both of the shared electrons, and the poor old H^+ has neither of them.
3) But a covalent bond can also break evenly — and then each atom gets one of the shared electrons, e.g. **H–H → H· + H·** — the H· is called a free radical. (The unpaired electron is shown by a dot.)
4) The unpaired electron makes the free radical very, very reactive.

Chlorine free radicals from CFCs damage the ozone layer

1) Ultraviolet light makes CFCs break up to form free radicals:

$$CCl_2F_2 \rightarrow CClF_2\cdot + Cl\cdot$$

Free radicals...

2) This happens high up in the atmosphere (in the stratosphere), where the ultraviolet light from the Sun is stronger.

3) Chlorine free radicals from this reaction react with ozone (O_3), turning it into ordinary oxygen molecules (O_2):

$$O_3 + Cl\cdot \rightarrow ClO\cdot + O_2$$

4) The chlorine oxide molecule ClO is very reactive, and reacts with ozone to make two oxygen molecules and another Cl· free radical:

$$ClO\cdot + O_3 \rightarrow 2O_2 + Cl\cdot$$

5) This Cl· free radical now reacts with another ozone molecule. It's a chain reaction, so just one Cl· free radical from one CFC molecule can break up a lot of ozone molecules.

CFCs and the Ozone Layer

CFCs have already been banned in many countries, but unfortunately they're still destroying ozone.

CFCs stay in the stratosphere for ages

1) CFCs are not very reactive and they will only react with one or two of the chemicals that are present in the atmosphere.

2) They only break up to form chlorine atoms in the stratosphere, where there's plenty of high-energy ultraviolet light around. They won't do it in the lower atmosphere.

3) This means that the CFCs in the atmosphere now will take a long time to be removed.

4) Remember, each CFC molecule produces one chlorine atom which can react with an awful lot of ozone molecules. Thousands of them, in fact.

5) So the millions of CFC molecules that are present in the stratosphere will continue to destroy ozone for a long time — even after all CFCs have been banned (they already have been in many countries).

6) Each molecule will stay around for a long time, and each molecule will destroy a lot of ozone molecules.

Alkanes and HFCs are safe alternatives to CFCs

1) Alkanes don't react with ozone, so they can provide a safe alternative to CFCs.
2) Hydrofluorocarbons (HFCs) are compounds very similar to CFCs — but they contain no chlorine. It's the chlorine in CFCs that attacks ozone, remember.
3) Scientists have investigated the compounds that could be produced by breakdown of HFCs in the upper atmosphere, and none of them seem to be able to attack ozone. The evidence suggests that HFCs are safe to use.

The Montreal Protocol was an agreement to stop using CFCs

After discovering a hole in the ozone layer, many countries got together and decided to reduce CFC production and eventually ban CFCs completely — the agreement was called the Montreal Protocol).

Warm-Up and Exam Questions

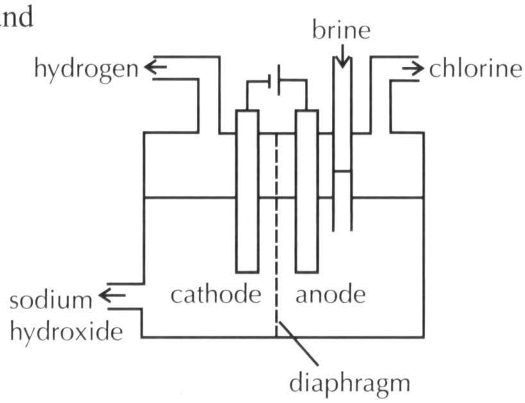
Give this last set of questions your very best shot, and then you can go into the exam full of confidence.

Warm-Up Questions

1) Name one region in Britain where salt mining takes place.
2) What three products are obtained from the electrolysis of brine (concentrated NaCl solution)?
3) If dilute sodium chloride solution is electrolysed, what two gases are produced?
4) What is a free radical?
5) Why will CFCs continue to deplete the ozone layer for many years after their use is banned?

Exam Questions

1 The diagram shows a cell used for the electrolysis of brine.



(a) Hydrogen is produced at the cathode. Copy and complete the half-equation for this reaction:

_______ + _______ $\rightarrow H_2$

(2 marks)

(b) Chlorine is produced at the anode. Copy and complete the half-equation:

_______ $- 2e^- \rightarrow$ _______

(2 marks)

(c) Give one important industrial use of:

(i) hydrogen (ii) chlorine

(2 marks)

(d) If the electrolysis is done with molten sodium chloride, what will the products be?

(2 marks)

2 CFCs (chlorofluorocarbons) were widely used as refrigerants for many decades, but it is now known that CFCs have a destructive effect on the stratospheric ozone layer.

(a) Many CFC molecules only break up when they reach the stratosphere — why?

(1 mark)

(b) One product of the break-up of CFCs is the chlorine free radical, Cl•. Copy and complete the following two symbol equations showing its effect on ozone:

(i) O_3 + Cl• $\rightarrow$ _______ + _______

(ii) _______ + $O_3 \rightarrow 2O_2$ + _______

(4 marks)

(c) A small amount of Cl• can break up a lot of ozone molecules. Explain why.

(2 marks)

(d) Scientists have researched alternatives to CFCs to use in refrigeration. Name two groups of compounds that are now used instead.

(2 marks)

Revision Summary for Section Thirteen

Time to test yourself. If you can't answer these now, you won't magically know them in the exam. But this is almost the end of the book. So when you've done these you've only got one last section and then you can just sit back and wait for your exam. Well, you could... but that would be silly. You need to keep your brain in the chemistry mood all the way through to the exam — don't let any of that hard-earned knowledge just dribble away. Go back and try the questions on the earlier modules again. Just to check you've still got what it takes.

1) What are the advantages and disadvantages of continuous production and batch production?
2) The production cost of a chemical depends on five main factors. What are they?
3) Explain what is meant by the 'optimum conditions' for the industrial production of a chemical.
4) Write down three physical properties of a typical alcohol such as ethanol.
5) Why can alcohols mix with both water and oils? Give one use of this property.
6) What functional group do all carboxylic acids have in common?
7) Name the ester that would be made by mixing together methanoic acid and propanol.
8)* A family of substances is being synthesised using staged synthesis.

Family of reactants	X	Y	Z
Number of reactants in family	7	6	12

The process involves two stages — X reacts with Y to form XY, then XY reacts with Z to form XYZ. Given the data in the table, how many possible products are there from this reaction?

9) How would you extract the active ingredient needed to make a new drug from a plant?
10) It can take 12 years and £900 million to bring a new drug onto the market. Explain why.
11) Give three examples of analgesics.
12) Describe the industrial method for the preparation of aspirin.
13) What's the difference in structure between aspirin and soluble aspirin?
14) Write the symbol equations for the four reactions in the Contact process.
15) What temperature is used for the Contact process? Why is this said to be a compromise?
16) Give two advantages of hydrogen fuel cells over conventional ways of generating electricity.
17) Write down the overall reaction in a hydrogen-oxygen fuel cell.
18) Give an industrial use for transition metals.
19) Write down the electron configuration of: a) titanium, b) cobalt, c) zinc.
20) Briefly describe two methods of salt mining.
21) What is bleach made from?
22) How are free radicals formed?
23) Write an equation for the reaction between an ozone molecule and a chlorine free radical.
24) One CFC molecule can destroy thousands of ozone molecules. Why is this?

* Answer on page 264.

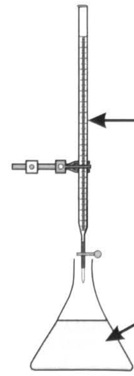
Answering Experiment Questions (i)

Science is all (well... a lot) about doing experiments carefully, and interpreting results.
And so that's what they're going to test you on when you do your exam. Among other things.

Read the question carefully

Expect at least some questions to describe experiments — a bit like the one below.

Q3 Ellen has three different bottles of citric acid: A, B and C.
The citric acid in each bottle is of a different concentration.
Ellen also has another quantity of citric acid, in the form of kitchen descaler.
Ellen wants to know if any of her three acids are the same concentration as the kitchen descaler. She plans to titrate each of the four citric acid solutions against a solution of sodium hydroxide of a known concentration, as shown.
She repeats the titration 3 times for each acid.



1. What is the independent variable in Ellen's experiment?

   The type of acid used (e.g. A, B or C).

2. What is the dependent variable?

   The quantity of acid required to neutralise the sodium hydroxide.

3. Give two variables that must be kept the same to make it a fair test.

   1. The amount of NaOH.

   2. The type and amount of indicator used.

4. Give one other precaution that Ellen should take to ensure her results are reliable.

   Wash and dry the equipment each time (to ensure no contamination).

The independent variable is the thing the experimenter changes — to see what effect the change has.

Quite often, experiments involve recording what happens over time, e.g. rate of reaction experiments. In these cases, time is always the independent variable — the experimenter isn't 'changing the time' exactly, but they do want to see what happens as the time changes.

The dependent variable is the thing the experimenter measures (every time they change the independent variable).

To make it a fair test, you've got to keep all the other variables the same (you're only changing the independent variable). That way you know that the only thing affecting the dependent variable is the independent variable.

If your experiment is being done in a lab, this should be fairly easy (though not always — e.g. you might have to keep temperature constant, which could be tricky). But it's trickier still when you don't have much control over the conditions at all — e.g. if your experiment has to be done outside (where temperature, humidity etc. can vary considerably).

Anything that might affect the results needs to be kept constant, so look at the apparatus, think what Ellen's going to be doing — and you should be able to come up with answers fairly easily.

If the equipment isn't clean, that will definitely affect the results. And if the flask's not dry, the extra water would dilute the sodium hydroxide slightly (which would affect the results). A temperature change could also be a problem (though probably a small one) — things expand as they get hotter, so Ellen could get a false reading from the burette if the temperature in the lab changes a lot between tests.

Answering Experiment Questions (ii)

5. Why did Ellen repeat the titration 3 times for each acid?

   To check for anomalous results and make the results more <u>reliable</u>.

6. The table below shows the amount of acid required in each titration.

	1st result (cm^3)	2nd result (cm^3)	3rd result (cm^3)	Mean (cm^3)
Kitchen descaler	24.1	23.9	23.7	
Acid A	23.9	23.5	24.0	23.8
Acid B	33.3	33.7	38.6	33.5
Acid C	23.7	23.9	24.1	23.9

   a) Calculate the mean amount of kitchen descaler required to neutralise the NaOH.

   Mean = (24.1 + 23.9 + 23.7) ÷ 3 = 23.9 cm^3

   b) What is the range of the quantities of kitchen descaler required?

   24.1 − 23.7 = 0.4 cm^3

7. One of the results in the table is anomalous. Circle the result and suggest why it may have occurred.

   The reading may not have been taken correctly, or the wrong quantity of NaOH may have been used.

8. Using these results, which acid can you conclude is <u>not</u> the same concentration as the kitchen descaler?

   Acid B

Sometimes you get <u>unusual results</u> — <u>repeating</u> an experiment gives you a better idea what the <u>correct result</u> should be.

When an experiment is <u>repeated</u>, the results will usually be <u>slightly different</u> each time.

To get a single <u>representative</u> value, you'd usually find the <u>mean</u> (average) of all the results.

The more times the experiment is <u>repeated</u> the <u>more reliable</u> this average will be.

To find the mean:

Add together all the data values and DIVIDE by the total number of values in the sample.

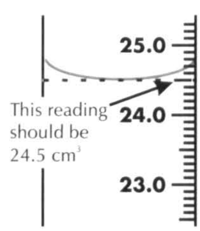
The <u>range</u> is how <u>spread out</u> the data is. You just work out the <u>difference</u> between the <u>highest</u> and <u>lowest</u> numbers.

If one result doesn't seem to fit in — it's wildly out compared with the others — then it's called an <u>anomalous</u> result. You should usually <u>ignore</u> an anomalous result (or even better — investigate it and try to work out what happened). Here, it's been <u>ignored</u> when the mean was worked out.

This one's a <u>random error</u> — one that only happens occasionally.

If you make the same mistake every time, it's a <u>systematic error</u>.

For example, if you measured the volume of a liquid using the <u>top</u> of the meniscus rather than the <u>bottom</u>, all your readings would be a little on the large side.



You have to be careful here — both Acids A and C could be the same concentration, since all experiments have a "<u>margin of error</u>" — meaning results are never absolutely spot on.

So you can say that Acid B has a different concentration — but Acids A and C <u>could</u> be the same.

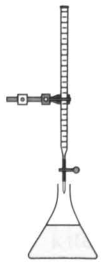
You can believe me — I'm a scientist...

This is a question all about making results <u>trustworthy</u> (a not-very-scientific way of saying <u>reliable</u>, <u>accurate</u> and <u>precise</u> — see next page). So you need to make the experiment fair, and you need to double-check results to make sure nothing weird happened. It's the same for all scientists.

Answering Experiment Questions (iii)

*Use **sensible measurements** for your **variables***

Charley has four bottles of acid labelled A to D, which he has been told are of different concentrations. He does a titration experiment to find out what volume of each acid is needed to neutralise 25 cm^3 of alkali.



He measures 25 cm^3 of the alkali into a flask, with some indicator solution, and sets up a burette filled with 50 cm^3 of acid A. He gradually adds acid to the alkali, and when the indicator shows that the alkali is almost neutralised, he adds the acid very slowly, drop by drop. He stops when the alkali has been neutralised, and writes down the volume of acid that was needed.

He repeats this experiment for the other three acids.

Before starting, Charley did a trial run, adding each acid very quickly, so he knows that between 15 cm^3 and 25 cm^3 of the acids are needed to neutralise 25 cm^3 of the alkali.

1. What kind of variable is the volume of acid needed?
   - **A** A continuous variable ✓
   - **B** A categoric variable ☐
   - **C** An ordered variable ☐
   - **D** A discrete variable ☐

Continuous data is numerical data that can take any value in a range — e.g. length, volume, temperature, time.

Note: You can't measure the exact value of continuous data. Say you measure a height as 5.6 cm to the nearest mm. It's not exact — you get a more precise value if you measure to the nearest 0.1 mm or 0.01 mm, etc.

Categoric variables are variables that can't be related to size or quantity — they're types (categories) of things. E.g. names of metals or types of fertiliser.

Ordered variables are things like small, medium and large lumps, or warm, very warm, and hot.

Discrete data is the type that can be counted in chunks, where there's no in-between value. E.g. number of people is discrete not continuous because you can't have half a person.

2. Charley should initially add the acid...
   - **A** 0.01 cm^3 at a time. ☐
   - **B** 10 cm^3 at a time. ☐
   - **C** 1 cm^3 at a time. ✓
   - **D** 5 cm^3 at a time. ☐

It's important to use sensible values for variables.

It's no good adding loads of acid at once, as you might shoot past the neutralisation point. But on the other hand, adding it in tiny amounts like 0.01 cm^3 could take ages.

3. The burette used to measure the volume of acid should be capable of measuring...
   - **A** to the nearest 0.1 cm^3. ✓
   - **B** to the nearest cm^3. ☐
   - **C** to the nearest 10 cm^3. ☐
   - **D** to the nearest 0.5 cm^3. ☐

A burette measuring only in whole cm^3 or more, would not be sensitive enough — when the alkali is nearly neutralised, Charley adds acid drop by drop, so he needs to measure to the nearest 0.1 cm^3 to get precise results.

The sensitivity of an instrument is the smallest change it can detect. E.g. some balances measure to the nearest gram, but really sensitive ones measure to the nearest hundredth of a gram. For measuring tiny changes — like from 2.00 g to 1.92 g — a sensitive balance is needed.

You also have to think about the precision and accuracy of your results.

Measurements (of the same thing) that are very precise will be close together. Really accurate measurements are those that have an average value that's really close to the true answer. So it's possible for results to be precise but not very accurate, e.g. a fancy piece of lab equipment might give results that are precise, but if it's not calibrated properly those results won't be accurate.

Answering Experiment Questions (iv)

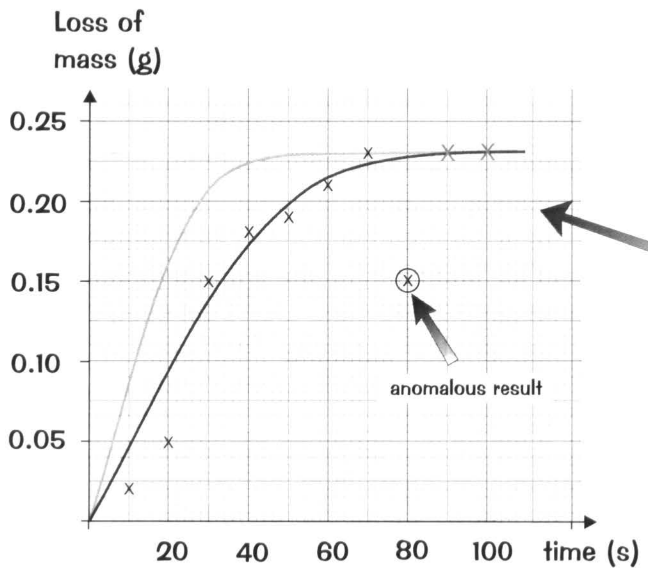
Once you've collected all your data together, you need to analyse it to find any relationships between the variables. The easiest way to do this is to draw a graph, then describe what you see...

Graphs are used to show relationships

Melissa did an experiment on rate of reaction, mixing magnesium and dilute hydrochloric acid. She measured how much mass was 'lost' from the flask of reactants as a gas was given off. These were her results.

Time (s)	10	20	30	40	50	60	70	80	90	100
Loss in mass (g)	0.02	0.05	0.15	0.18	0.19	0.21	0.23	0.15	0.23	0.23

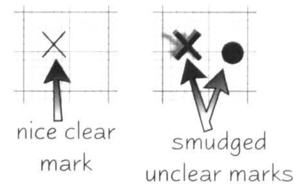
1. a) Eight of the points are plotted below. Plot the remaining **two** points on the graph.
   b) Draw the line or curve which best fits the points.



2. Explain what your graph shows about how the rate of reaction changes with time.
   The curve is steep for the first 30 s, showing that the reaction is quickest at first, but as the reactants are gradually used up, the reaction slows down and the curve flattens off.

3. Sketch a graph to show the results you would expect to see if Melissa repeated the experiment using more concentrated acid, but keeping everything else the same.

To plot the points, use a sharp pencil and make a neat little cross.



If your points lie roughly in a line, draw a line of best fit. If your points make a curve, draw a smooth curve. Whatever else you do, don't just join the dots.

When you're drawing your curve (or line), make it go as close to as many points as possible. It doesn't have to go through them — you want a smooth curve (or a straight line) not a wiggly one. In this case, the curve has to go through the origin (0, 0) as you know there'd be no gas given off if you hadn't mixed the reactants together yet.

You might have some anomalous results — usually when you've done something daft, like reading a scale wrongly.

You can ignore these anomalous results when you're drawing your curve (or line).

You'd expect that, with more concentrated acid, the reaction would go faster — so the curve would be steeper at the beginning, and reach its highest value sooner. But there's the same mass of the other reactant (magnesium) as before, so the same total amount of gas will be given off — the final loss in mass will be the same as before.

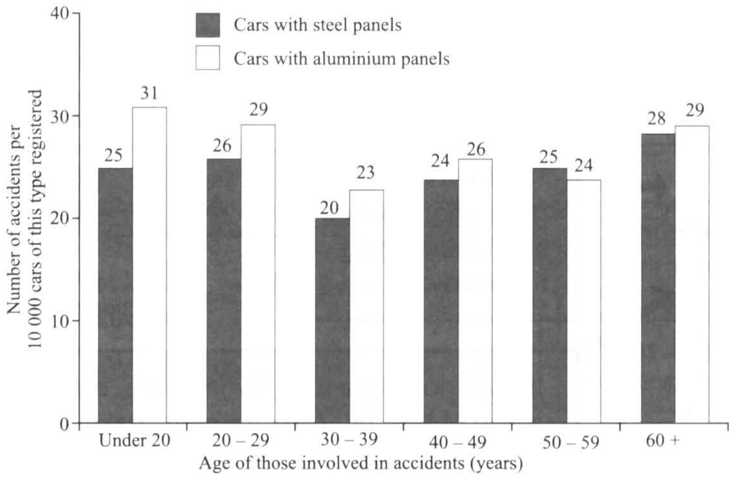
Answering Experiment Questions (v)

Not all experiments can be carefully controlled in a laboratory. Some have to be done in the real world.

Relationships do NOT always tell you the cause

Most car bodies are made from strong steel panels, but engineers are looking for innovative materials which will improve efficiency and safety. 'Alucars' has released a new car with body panels made of aluminium.

The bar chart below shows how many accidents there were involving cars with aluminium bodies, and how many involving cars with steel bodies, in one year.



In large studies done outside a lab it's really difficult to keep all the variables the same and to make sure the control group are kept in the same conditions.

In this study the control group are the people in 'normal' cars, with steel panels.

This is a bar chart. It contains a key to tell you what colour bars relate to which group.

1. There are approximately 5000 aluminium cars registered in the county of Wessex. Use the bar chart to estimate how many under-20s will be involved in accidents in one year, in aluminium-panelled cars.

   $31 \div 2 = 15.5 \approx 16$ people.

They're asking you the number of injuries you'd expect for 5000 cars — the graph tells you injuries per 10 000 cars. Don't get caught out, read the question carefully. (And don't write something daft with half a person in it.)

2. What conclusion can you draw from the results?

   There are proportionately more accidents involving cars with aluminium bodies than cars with steel bodies.

When describing the data and drawing conclusions it's really important that you don't say that having an aluminium-panelled car causes accidents. The graph only shows that there's a positive relationship between the two.

In studies like these where you can't control everything, it's possible a third variable is causing the relationship. E.g. aluminium-bodied cars would be lighter than steel-bodied cars, so they might appeal to people who like driving fast, and driving faster causes more accidents.

3. Suggest how the accident data may have been collected.

   e.g. from police records.

Use your common sense here.

Try to suggest a method to get reliable results. For example, it's very unlikely that the data would have been collected by a telephone survey or an internet search.

A relationship doesn't necessarily mean cause and effect

Just looking at numbers (like here) doesn't mean you can say that one variable changing causes the other to change too. 'Ice cream sales' and 'cases of heatstroke' probably rise and fall together — but you can't say that ice cream causes heatstroke. (They'd more likely both be caused by a heatwave.)

Pages 12-13

Warm-Up Questions

1) protons and electrons
2) Mass number is the sum of the number of protons and the number of neutrons in an atom. Atomic number is the number of protons (or electrons) in an atom.
3) It either expands, or its pressure increases.
4) E.g. copper, iron (any solid element).
5) $C_6H_{12}O_6$ + **6**O_2 → **6**CO_2 + **6**H_2O
6) The $_2$ in H_2SO_4 refers to 2 atoms (of H), while the 2 in 2NaOH refers to 2 molecules of NaOH.

Exam Questions

1 B ***(1 mark)***

2 D ***(1 mark)***

3 (a) (i) The particles are free to move past each other ***(1 mark)*** but there is some force of attraction between them so they tend to stick together ***(1 mark)***.

(ii) The particles are free to move / have virtually no force of attraction between them ***(1 mark)*** so they move in straight lines until they collide with each other or with the sides of the container ***(1 mark)***.

(iii) The particles have strong forces of attraction between them / are not free to move ***(1 mark)***, so they stay in a regular arrangement ***(1 mark)***.

(b) (i) Particles at the surface of a liquid overcome the forces of attraction from other particles and escape ***(1 mark)***.

(ii) Perfumes have to evaporate easily so they can reach our smell receptors quickly ***(1 mark)***.

4 (a) In a compound, different types of atoms are bonded together chemically. In a mixture they are not. ***(1 mark)***

(b) (i) noble gases / group 0 ***(1 mark)***

(ii) 40 ***(1 mark)***

(c) e.g. carbon dioxide / methane ***(1 mark)***

(d) $2Mg + O_2 \rightarrow 2MgO$ ***(1 mark for correct products and reactants, 1 mark for correctly balancing the equation)***

5 B ***(1 mark)***

6 C ***(1 mark)***

7 (a) sulfuric acid + ammonia → ammonium sulfate ***(1 mark)***

(b) $H_2SO_4 + 2NH_3 \rightarrow (NH_4)_2SO_4$ ***(1 mark for correct products and reactants, 1 mark for correctly balancing the equation)***

(c) 15 ***(1 mark)***
There are eight atoms of hydrogen, one atom of sulfur, four atoms of oxygen, and two atoms of nitrogen.

Page 14

Revision Summary for Section One

9) 2 sodium atoms, 1 carbon atom, 3 oxygen atoms

10) a) Bottom-left
b) top-right
c) top-left
d) bottom-right

12) calcium

14) a) $CaCO_3 + 2\,HCl \rightarrow CaCl_2 + H_2O + CO_2$
b) $Ca + 2\,H_2O \rightarrow Ca(OH)_2 + H_2$

15) $2K + 2H_2O \rightarrow 2KOH + H_2$

Pages 21-22

Warm-Up Questions

1) Any three of, e.g. as a building material / making glass / making cement / making slaked lime.
2) Any one of, e.g. destroys habitats / uses land / causes noise / causes pollution / leads to unsightly tips.
3) Cement is limestone that has been heated with clay. Mortar is cement mixed with sand and water.
4) chalcopyrite
5) Copper extracted by reduction with carbon is impure and doesn't conduct electricity well enough. Electrolysis produces very pure copper which is a much better conductor.
6) Any one of, e.g. zinc / iron / tin / copper.

Exam Questions

1 (a) $CaCO_3 \rightarrow CaO + CO_2$ ***(1 mark)***

(b) (i) calcium hydroxide ***(1 mark)***, $Ca(OH)_2$ ***(1 mark)***

(ii) e.g. neutralising acid soils ***(1 mark)***

2 B ***(1 mark)***

3 (a) Any two of, e.g. useful products can be made / provides jobs / brings money into the local area ***(1 mark each)***.

(b) Any two of, e.g. causes noise / scars landscape / loss of habitats / abandoned mine shafts can be dangerous ***(1 mark each)***.

4 (a) E.g. potassium, sodium, calcium, magnesium, aluminium (any metal above carbon in the reactivity series). ***(1 mark)***

(b) (i) removal of oxygen (accept gain of electrons) ***(1 mark)***

(ii) zinc oxide + carbon → zinc + carbon dioxide ***(1 mark)***

(c) $Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$ ***(1 mark for correct products and reactants, 1 mark for correctly balancing the equation, 1 mark for correct state symbols)***

(d) C ***(1 mark)***
Zinc is more reactive than copper, so zinc will displace copper.

5 (a) e.g. electrical wires, plumbing ***(2 marks)***

(b) Any one of, e.g. the supply of copper-rich ores is limited / demand for copper is growing ***(1 mark)***.

(c) (i) E.g. other methods need a lot of energy / other methods release sulfur dioxide gas, which causes acid rain. ***(1 mark)***

(ii) E.g. using bacteria is slower than other methods. ***(1 mark)***

Pages 29-30

Warm-Up Questions

1) Any two of e.g. iron, zinc, copper (any two transition metals).
2) Any three of, hard / strong / good electrical conductors / good conductors of heat / malleable / ductile / flexible.
3) metallic bonding
4) A mixture of metals, or a mixture of a metal and a non-metal, e.g. bronze — sculpture, medals / cupronickel — coins / solder — joining wires.
5) solvent, binding medium and pigment
6) Oil-based paints have something that dissolve oil as the solvent. Water-based paints have water as the solvent.

Exam Questions

1 (a) A colloid is a mixture of tiny particles of one kind dispersed (but not dissolved) in another substance. ***(1 mark)***

(b) Because the dispersed particles are so small that they don't settle at the bottom. ***(1 mark)***

(c) (i) Gloss paint, because it is harder-wearing. ***(1 mark)***

(ii) Emulsion paint, because it dries quickly and produces only low levels of harmful fumes. ***(1 mark)***
You don't need to memorise all the different properties, just be able to interpret information about them.

2 D ***(1 mark)***

3 B ***(1 mark)***

4 (a) It has a shape memory property ***(1 mark)***.

(b) E.g. glasses frames ***(1 mark)***

(c) E.g. metal fatigue in smart alloys is worse than in normal alloys; smart alloys are more expensive than normal alloys. ***(2 marks)***

5 A — 2 ***(1 mark)***
B — 3 ***(1 mark)***
C — 4 ***(1 mark)***
D — 1 ***(1 mark)***

6 (a) The outer electrons of each atom can move freely, creating a sea of free electrons ***(1 mark)***. These can carry electric current through the material ***(1 mark)***.

(b) Any two of, e.g. good conductor of heat / strong / can be bent/hammered into shape ***(1 mark for both properties)***.

Pages 40-41

Warm-Up Questions

1) Compounds made from carbon and hydrogen *only*.
2) Any three of, e.g. LPG / petrol / naphtha / kerosene / diesel / oil / bitumen.
3) They contain carbon-carbon double bonds.
4) Long-chain hydrocarbons are cracked to make more useful products / because there's more demand for short-chain fractions.
5) High temperature and a catalyst.
6) Any three of, e.g. transport / electricity generation / making plastics / heating / making medicines / making paints and dyes.

Exam Questions

1 A ***(1 mark)***

2 (a) (i) There should be an M in the bottom box ***(1 mark)***.

(ii) There should be a B in the top box ***(1 mark)***.

Fractions with bigger molecules have a higher boiling point, so condense at the higher temperatures at the bottom of the column. Fractions with smaller molecules have a lower boiling point, so don't condense until they reach the top of the column.

(b) Any one of, e.g. jet engine fuel, domestic heating, paint solvent ***(1 mark)***.

(c) The explanation should contain three of the following points:
- the fractions have different boiling points
- the crude oil is heated
- the fractions boil
- fractions condense at different heights in the fractionating column
- fractions are tapped off where they condense

(1 mark per point; maximum 3 marks)

3 C ***(1 mark)***

4 C ***(1 mark)***

5 (a) (i) $2C_2H_6 + 7O_2 \rightarrow 6H_2O + 4CO_2$

(1 mark for correct products and reactants, 1 mark for correctly balancing the equation)

This is a bit of a tricky one — don't panic, just take one element at a time and keep pencilling in numbers till it all balances.

(ii) Test its boiling point: 100 °C **OR** dip dry blue cobalt chloride paper in it: it turns pink **OR** add to white anhydrous copper sulfate crystals: they turn blue. ***(1 mark)***. The limewater goes cloudy, indicating the presence of CO_2 ***(1 mark)***.

(b) (i) An insufficient oxygen supply ***(1 mark)***.

(ii) carbon monoxide ***(1 mark)*** and carbon ***(1 mark)***

6 A ***(1 mark)***

Page 45

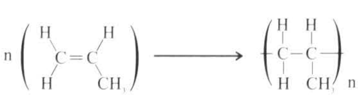
Warm-Up Questions

1) Monomers are small molecules which can be joined together to give much larger molecules called polymers.
2) The higher the melting point, the stronger the forces holding the polymer chains together.
3) covalent bonds
4) E.g. for making kettles.

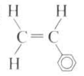
Exam Questions

1 biodegradable, burnt, toxic, recycle, expensive ***(5 marks)***

The majority of polymers aren't biodegradable, and this has significant environmental consequences. Chemists are currently working on biodegradable polymers though.

2 (a) 

(1 mark)

(b) 

(1 mark)

(c) (i) Any one of, e.g. window frames/CDs ***(1 mark)***

(ii) Any one of, e.g. clothing/synthetic leather ***(1 mark)***

Page 46

Revision Summary for Section Three

8) e.g. $2C_2H_6 + 5O_2 \rightarrow CO_2 + 6H_2O + C + 2CO$ (+ energy)

Page 53

Warm-Up Questions

1) E.g. it kills microbes / destroys the poisons present in some raw foods.
2) The protein molecules change shape irreversibly when heated (this is known as denaturing).
3) E.g. Thermochromatic pigments that change colour faster the warmer they get. This is used to tell if food's been warm long enough for microbes to grow.
4) Antioxidants are chemicals that stop fat and oil reacting with oxygen.

Exam Questions

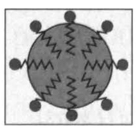
1 C ***(1 mark)***

Potatoes are plants, so each cell is surrounded by a cellulose cell wall, which humans can't digest.

2 Extract the colour from the food sample by placing it in a test tube with a few drops of solvent ***(1 mark)***. Put a spot of the coloured solution on a pencil baseline on some filter paper ***(1 mark)***. Stand the filter paper in a beaker with some solvent. ***(1 mark)***

3 (a) hydrophilic  hydrophobic ***(1 mark)***

(b) Hydrophobic means that it does not mix with water (it doesn't 'like' water) ***(1 mark)***. Hydrophilic means that it is attracted to water (it 'likes' water) ***(1 mark)***.

(c) 

(1 mark)

(d) They prevent the mayonnaise emulsion separating into its component liquids / they keep the oil and water mixed well together. ***(1 mark)***

Pages 61-62

Warm-Up Questions

1) The fruits or seeds are crushed. The oil is separated from the crushed plant material (by a centrifuge or using solvents). The oil is then distilled to refine it.
2) A nickel catalyst, at about 60°C.
3) ethanol
4) Distillation produces more concentrated alcohol.
5) (10%) ethanol mixed with (90%) petrol. It is used extensively in Brazil.
6) To make sure they are safe to use.

Exam Questions

1 A — 4 ***(1 mark)***

B — 2 ***(1 mark)***

C — 3 ***(1 mark)***

D — 1 ***(1 mark)***

2 A — 1 ***(1 mark)***

B — 2 ***(1 mark)***

C — 4 ***(1 mark)***

D — 3 ***(1 mark)***

3 (a) esterification ***(1 mark)***

(b) carboxylic acid + alcohol → ester + water ***(1 mark)***

(c) Any five of, evaporate easily, non-toxic, don't react with water, don't irritate the skin, insoluble in water, have a pleasant smell.
(1 mark for each correct property, maximum of 5 marks)

4 (a) Crude oil is non-renewable. As it runs out it will become more expensive ***(1 mark)***.

(b) (i) $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$ ***(1 mark for correct reactants and products, 1 mark for correctly balancing the equation)***

(ii) Ethanol can be produced by fermenting crops (e.g. sugar cane) ***(1 mark)***. The crops absorb carbon dioxide as they grow (by photosynthesis) ***(1 mark)***.
Photosynthesis takes carbon dioxide out of the air and produces oxygen.

5 (a) vegetable oils ***(1 mark)***

(b) (i) Biodiesel comes from recently grown plants which took in carbon dioxide from the air when they were alive ***(1 mark)***. This is released again when the biodiesel is burnt, so the net increase in carbon dioxide in the atmosphere is nil ***(1 mark)***.

(ii) Any two from, e.g. produces less sulfur dioxide than diesel or petrol / doesn't release as many particulates as diesel or petrol / it's biodegradable / it's less toxic than regular diesel / it's made from a renewable resource ***(1 mark each)***.

(c) E.g. it's more expensive than ordinary diesel / we're unable to make enough biodiesel to replace regular diesel ***(2 marks)***.

Pages 68-69

Warm-Up Questions

1) volcano / mountain chain
Earthquakes occur at plate boundaries but they aren't a geological feature.

2) Any one of, e.g. bulging of the ground near the volcano / mini-earthquakes.

3) A (theoretical) single land mass / supercontinent made from the all present continents joined together.

4) Layers of sediment laid down in lakes or seas; layers get buried under more layers, and the water is squeezed out; fluids flowing through the pores deposit natural mineral cement.

5) steps / buildings

Exam Questions

1 C ***(1 mark)***

2 B ***(1 mark)***

3 (a) From the action of heat and pressure ***(1 mark)*** over long periods of time ***(1 mark)***.

(b) (i) Limestone ***(1 mark)***

(ii) Any two from, e.g. marble has smaller crystals / a more even texture / is harder ***(1 mark each)***

4 (a) The diagram should be labelled:
A – crust ***(1 mark)***
B – mantle ***(1 mark)***
C – inner core ***(1 mark)***
D – outer core ***(1 mark)***

(b) nickel ***(1 mark)***, iron ***(1 mark)***

5 (a) Igneous rocks form when molten magma ***(1 mark)*** pushes up into the crust before cooling and solidifying. ***(1 mark)***

(b) Extrusive igneous rocks cool quickly above ground, forming small crystals, while intrusive igneous rocks cool slowly underground, forming big crystals ***(1 mark)***.
Examples:
Extrusive — any one of, e.g. basalt, rhyolite ***(1 mark)***
Intrusive — any one of, e.g. granite, gabbro ***(1 mark)***

6 B ***(1 mark)***

7 (a) Any two of, there is a jigsaw fit between the continents / identical fossils of the same age have been found in rocks in different continents, which suggests the continents were joined once upon a time / certain rock layers of similar ages on different continents show similarity / there are various living creatures found in both America and Africa that couldn't have crossed the Atlantic Ocean. ***(1 mark each)***.

(b) This theory wasn't accepted because: any one of, e.g. he didn't explain why it happened / he wasn't a geologist ***(1 mark)***.

Page 73

Warm-Up Questions

1) Nitrogen, oxygen

2) Any one of, e.g. it allowed complex organisms to evolve / it produced the ozone layer.

3) Any one of, e.g. aerosol propellants / fridge coolants.

Exam Questions

1 A — 3 ***(1 mark)***
B — 2 ***(1 mark)***
C — 4 ***(1 mark)***
D — 1 ***(1 mark)***

2 (a) There have been large variations in temperature and CO_2 concentration over the last 250 000 years ***(1 mark)***. There is a positive correlation between CO_2 concentration and temperature ***(1 mark)***.
The question's worth two marks, so you have to make two points.

(b) The X should be drawn at 25 000 years ago ***(1 mark)***.

3 (a) (i) Antarctica ***(1 mark)***
the Arctic ***(1 mark)***

(ii) They break it down ***(1 mark)***.

(b) The answer should contain three of the following points:
- ozone protects against harmful UV radiation
- ozone levels have fallen
- incidence of skin cancer has increased
- but other factors, e.g. more holidays, may play a part

(1 mark per point; maximum 3 marks)
If the question says 'discuss', try to put across more than one point of view.

Pages 80-81

Warm-Up Questions

1) It absorbs heat radiated from Earth, preventing some of the heat being lost into space.

2) Any one of, e.g. improved sanitation / modern medicine / improved farming methods.

3) We are using up finite resources (crude oil, metals, etc.) more quickly and adding more and more pollution to our environment.

4) Any two of, e.g. nitrogen oxides / sulfur dioxide / carbon monoxide / water vapour.

5) Any two of, e.g. lakes become acidic / kills fish and trees / damages limestone buildings/statues.

Exam Questions

1 C ***(1 mark)***

2 (a) The diagram should be labelled:
A – respiration ***(1 mark)***
B – compounds in plants ***(1 mark)***
C – photosynthesis ***(1 mark)***
D – burning ***(1 mark)***

(b) Large scale deforestation increases the amount of CO_2 in the atmosphere ***(1 mark)*** because: CO_2 is released into the atmosphere when trees are burnt to clear land, microorganisms feeding on dead trees release CO_2 through respiration, fewer trees mean that less CO_2 is absorbed from the atmosphere in photosynthesis ***(2 marks for any two of these points)***.

3 C ***(1 mark)***

4 A — 3 ***(1 mark)***
B — 1 ***(1 mark)***
C — 4 ***(1 mark)***
D — 2 ***(1 mark)***

5 (a) (i)Fertilisers can get washed into rivers and pollute water ***(1 mark)***.

(ii) E.g. sewage / toxic chemicals from industry ***(1 mark each)***

(b) Household waste gets dumped in landfill sites ***(1 mark)***.

(c) Any one of, e.g. pesticides / herbicides / nuclear waste ***(1 mark)***.

6 B ***(1 mark)***

Page 86

Warm-Up Questions

1) The only product when it burns in air is water.
2) Any four of, e.g. energy value / availability / cost / ease of storage / toxicity / how polluting it is.
3) Any two of, e.g. fossil fuels (oil/natural gas/coal) / uranium / stone (e.g. limestone) / metals (e.g. aluminium).
4) Any one of, e.g. collecting / sorting / processing / transporting material.
5) E.g. meeting the needs of today's population without harming the ability of future generations to meet their needs.

Exam Questions

1 (a) Making new materials (e.g. by mining, extracting and producing) uses a lot of energy ***(1 mark)***. Recycling glass and some metals uses a fraction of this energy ***(1 mark)***.

(b) Any two of, e.g. recycling uses less resources than making new things / recycling costs less than making new things / recycling cuts down on the amount of rubbish that goes into landfull ***(1 mark each)***.

(c) Any one of, e.g. collecting waste material / transporting waste material / sorting waste material / processing waste material ***(1 mark)***.

2 (a) (i) methane and CO_2 ***(1 mark)***

(ii) By microorganisms digesting waste ***(1 mark)***.
It can be burnt to heat water/generate electricity ***(1 mark)***.

(iii)Advantage: any one of, e.g. cheap / carbon neutral / renewable / uses waste products ***(1 mark)***.

Disadvantage: e.g. production is slow in cold weather ***(1 mark)***.

(b) (i) by electrolysing water ***(1 mark)***

(ii) It's highly explosive, so it has to be kept in very secure containers ***(1 mark)***.

(iii)Any one of, e.g. you would need a specially designed, expensive engine / hydrogen isn't widely and cheaply available / large, strong gas containers are heavy and would increase fuel consumption ***(1 mark)***.

Page 94

Top Tip

1) a) K^+ (2, 8, 8)
b) Al^{3+} (2,8)
c) Be^{2+} (2)
d) S^{2-} (2,8,8)
e) F^- (2,8)

Pages 95-96

Warm-Up Questions

1) 2
2) 2
3) A high boiling point.
4) cations
5) $Al(OH)_3$

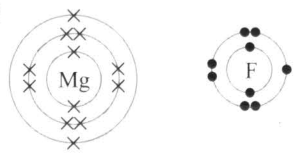
Exam Questions

1 (a) 11 ***(1 mark)***

(b) Group 1 ***(1 mark)***, because it has one electron in its outer shell / because it's got 11 electrons, so it's sodium which is in Group 1. ***(1 mark)***.

(c) 1 ***(1 mark)***

2 (a)



(1 mark) ***(1 mark)***

(b) Mg^{2+} ***(1 mark)*** and F^- ***(1 mark)***

(c) MgF_2 ***(1 mark)***

(d) The positively charged magnesium ions are attracted to the negatively charged fluoride ions ***(1 mark)***.

(e) (i) There are strong forces of attraction between the ions ***(1 mark)*** so a large amount of energy is needed to melt the compound ***(1 mark)***.

(ii) When the magnesium fluoride is molten the ions can move about and carry charge (i.e. conduct a current) through the liquid ***(1 mark)***.

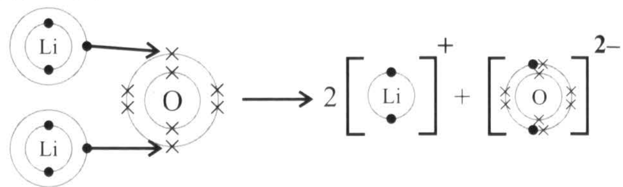
3 (a) 2+ in iron(II) oxide and 3+ in iron(III) oxide ***(1 mark)***

(b) iron(II) oxide = FeO ***(1 mark)***, iron(III) oxide = Fe_2O_3 ***(1 mark)***

(c) Any two of, e.g. copper, manganese, chromium ***(1 mark)***

4 (a) lithium oxide ***(1 mark)***

(b) (i) and (ii)



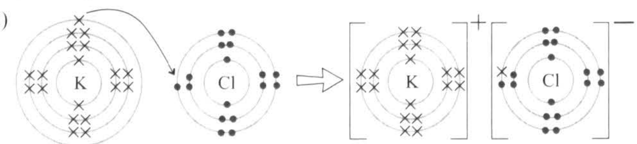
(1 mark for arrows shown correctly, 1 mark for correct electron arrangement and charge on lithium ion, 1 mark for correct electron arrangement and charge on oxygen ion).

5 (a)

	Potassium atom, K	Potassium ion, K^+	Chlorine atom, Cl	Chloride ion, Cl^-
Number of electrons	19	**18**	**17**	**18**
Electron arrangement	2, 8, 8, 1	**2, 8, 8**	**2, 8, 7**	**2, 8, 8**

(1 mark for each correct column, maximum 3 marks)

(b)



(2 marks — 1 mark for correct electron arrangements, 1 mark for correct arrow and charges on ions)

Page 101

Warm-Up Questions

1) it increases
2) The elements go from gas to liquid to solid (at room temperature).
3) hydrogen
4) E.g. silver bromide, zinc chloride (any metal halide).

Exam Questions

1 A — 2 ***(1 mark)***
B — 3 ***(1 mark)***
C — 1 ***(1 mark)***
D — 4 ***(1 mark)***

2 (a) Fluorine — gas ***(1 mark)***
Chlorine — gas ***(1 mark)***
Bromine — liquid ***(1 mark)***
Iodine — solid ***(1 mark)***

(b) Arrow should be pointing upwards. ***(1 mark)***

(c) (i) displacement ***(1 mark)***
Chlorine is displacing iodine.

(ii) iodine/I/I_2 ***(1 mark)***

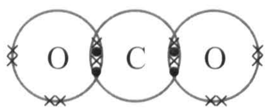
Page 108

Warm-Up Questions

1) In a covalent bond, the atoms share electrons. In an ionic bond, one of the atoms donates electrons to the other atom.
2) Each of the atoms in the covalent bond contributes two electrons to the bond instead of just one, so four electrons are shared altogether.
3) Any two of:
   Diamond is very hard and graphite is fairly soft.
   Graphite conducts electricity and diamond doesn't.
   Diamond is clear/transparent and graphite is opaque.
   Diamond is colourless and graphite is black.
4) E.g. silicon dioxide/silica
5) Because the intermolecular forces between the chlorine molecules are very weak.

Exam Questions

1



(1 mark for double bonds shown correctly, 1 mark for remaining electrons shown correctly)

2 (a) (i) giant covalent ***(1 mark)***
(ii) giant covalent ***(1 mark)***
(iii) simple molecular covalent ***(1 mark)***
(b) It doesn't contain any free electrons or ions to carry the charge ***(1 mark)***.
(c) It contains free electrons able to carry the charge ***(1 mark)***.
(d) (i) All of the atoms in silicon dioxide and in graphite are held together by strong covalent bonds ***(1 mark)***. In bromine, each molecule is held together with a strong covalent bond but the forces between these molecules are weak ***(1 mark)***.
In order to melt, a substance has to overcome the forces holding its particles tightly together in the rigid structure of a solid. If the forces between the particles are weak, this is easy to do and doesn't take much energy at all. But if the forces are really strong, like in a giant covalent structure, you have to provide loads of heat to give the particles enough energy to break free.
(ii) Iodine has bigger molecules than bromine ***(1 mark)***, and the bigger the halogen molecules, the greater the intermolecular forces are between them ***(1 mark)***.

Page 114

Warm-Up Questions

1) Their properties make them very hard to observe — they are colourless gases and are very unreactive.
2) Helium — used in balloons/airships.
   Neon — used in neon lighting.
   Both helium and neon are also used in the helium-neon laser.
3) The transition metals.
4) Metals have giant structures which include a 'sea' of free electrons. These electrons are able to move and carry the charge, allowing a current to flow through the metal.
5) 1–100 nanometres.

Exam Questions

1 (a) 2, 8, 8 ***(1 mark)***
(b) Each argon atom already has a full outer shell ***(1 mark)***, so they have no need to pair up and share electrons ***(1 mark)***.
(c) Argon takes the place of oxygen and provides an inert atmosphere ***(1 mark)*** so that the filament can't burn away even when it gets very hot ***(1 mark)***.

2 (a) 'Superconducting' means that the substance has no electrical resistance at all when a current flows through it ***(1 mark)***, meaning that none of the electrical energy is wasted as heat ***(1 mark)***.
(b) Metal oxide ceramic 2, because it's easier to cool substances to this temperature and keep them there than it is to reach lower temperatures ***(1 mark)***.
In fact, none of these substances are likely to be a lot of use in the real world. Keeping an electromagnet or power line at this temperature all the time would be really hard to do. Room temperature (or close to it) superconductors are what's needed, which may sound impossible — but think of all the stuff science has come up with that must have seemed impossible once...
(c) Any two of, power cables that carry electricity without losing any power / very fast electrical circuits/ very strong electromagnets that can work without a constant power source ***(1 mark each)***

3 In carbon nanotubes with a structure like that in the diagram, each carbon atom only forms three bonds ***(1 mark)***. This leaves each atom with a spare electron ***(1 mark)***, and if a current is applied these electrons are free to move and carry the charge, so the tube conducts ***(1 mark)***.

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Revision Summary for Section Six

26) A: metallic, B: giant molecular, C: ionic

Page 119

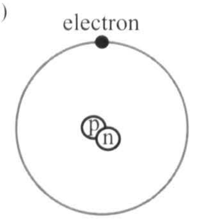
Warm-Up Questions

1) Isotopes.
2) The mass number.
3) The relative atomic mass.
4) The relative formula mass.
5) Neutron.

Exam Questions

1 (a) (i) Relative atomic mass ***(1 mark)***.
(ii) Boron-11 has one more neutron in its nucleus than boron-10 ***(1 mark)***.
(iii) Boron-11 must be the most abundant ***(1 mark)***. The A_r value takes into account the relative abundance of each isotope, and in the case of boron it is closer to 11 than to 10 ***(1 mark)***.
(b) (i) M_r of $BF_3 = 11 + (19 \times 3) = 68$ ***(1 mark)***
(ii) M_r of $B(OH)_3 = 11 + (17 \times 3) = 62$ ***(1 mark)***

2 (a)



(1 mark for showing the correct particles in the nucleus, 1 mark for showing the single orbiting electron correctly).
(b) (i) $M_r = (2 \times 2) + 16 = 20$ ***(1 mark)***
(ii) Because the mass of water molecules containing deuterium (20) is greater than the mass of ordinary water molecules (18) ***(1 mark)***.

Page 121

Top Tip

1) a) 30%
   b) 88.9%
   c) 48%
   d) 65.3%
2) CH_4

Page 122

Top Tip

1) 21.4 g
2) 38.0 g

Page 125

Warm-Up Questions

1) One mole of atoms or molecules of any substance will have a mass in grams equal to the relative formula mass for that substance.
2) One mole of O_2 weighs 16 × 2 = 32 g.
3) moles = mass ÷ M_r
   = 4 ÷ (23 + 16 + 1)
   = 0.1 moles
4) mass = moles × M_r
   = 2 × (12 + 32)
   = 88 g
5) 2 M
6) Moles = volume (l) × molarity
   = 0.5 × 0.5
   = 0.25 moles

Exam Questions

1 (a) % mass of N in $CO(NH_2)_2$ = [(A_r × no. of atoms) ÷ M_r] × 100
= [(14 × 2) ÷ (12 + 16 + 32)] × 100
= 47%

(1 mark for correct working, 1 mark for correct answer)

% mass of N in KNO_3 = [14 ÷ (39 + 14 + 48)] × 100
= 14%

(1 mark for correct working, 1 mark for correct answer)

% mass of N in NH_4NO_3 = [(14 × 2) ÷ (28 + 4 + 48)] × 100
= 35%

(1 mark for correct working, 1 mark for correct answer)

Calculations are often worth more than one mark. It can be tempting just to scribble down enough working out to get you to the answer. But it's worth bearing in mind that if you get the final answer wrong, you could still get some marks for the working. So if you put down each step clearly it could pay off.

(b) Urea ***(1 mark)***. It contains the greatest percentage mass of nitrogen, so would provide more nitrogen for plant growth per kg spread on the soil ***(1 mark)***.

2 (a) 100 – 60 = 40% ***(1 mark)***

(b) 40 g of sulfur combine with 60 g of oxygen.

S = 40	O = 60
40 ÷ 32	60 ÷ 16
= 1.25	= 3.75
1.25 ÷ 1.25 = 1	3.75 ÷ 1.25 = 3

Therefore, the formula of the oxide is $\mathbf{SO_3}$
(2 marks — 1 mark for correct working)

3 (a) 100g .... reacts to give ... 56 g
1 g .... reacts to give ... 56 ÷ 100 = 0.56 g
2 g .... reacts to give ... 0.56 × 2 = **1.12 g*(1 mark)***

(b) E.g. When transferring the $CaCO_3$ from the weighing apparatus to the test tube, or the CaO from the test tube to the weighing apparatus some of the solid may be left behind ***(1 mark)***.

Page 130

Warm-Up Questions

1) Waste by-products decrease the atom economy of a reaction.
2) 100%.
3) Because they use up resources very quickly and produce a lot of waste. This might be polluting and it has to be disposed of (e.g. in landfill sites or in the sea).
4) The raw materials would cost a lot and the waste would be expensive to dispose of. There would also be less product to sell.
5) Percentage yield = (4 ÷ 5) × 100
   = 80%

Exam Questions

1 M_r of ethanol = (12 × 2) + 6 + 16 = 46
M_r of ethene = (12 × 2) + 4 = 28 ***(1 mark)***
Atom economy = (28 ÷ 46) × 100 ***(1 mark)***
= 61 % ***(1 mark)***

2 (a) From the equation, 4 moles of CuO → 4 moles of Cu
so 1 mole CuO → 1 mole Cu ***(1 mark)***
63.5 + 16 = 79.5 g CuO → 63.5 g Cu ***(1 mark)***
1 g CuO → 63.5 ÷ 79.5 = 0.8 g (1 d.p.)
4 g CuO → 0.8 × 4 = 3.2 g ***(1 mark)***

(b) Percentage yield = (2.8 ÷ 3.2) × 100 ***(1 mark)***
= 87.5% ***(1 mark)***

(c) Any three of:

There may have been an incomplete reaction — some copper oxide was not reduced ***(1 mark)***.

There may have been unexpected reactions (which produced different products) due to impurities in the reactants ***(1 mark)***.

Some of the copper may have been left behind when it was scraped out into the beaker ***(1 mark)***.

Some of the copper may have been left on the filter paper ***(1 mark)***.

Page 131

Revision Summary for Section Seven

2) a) 40 b) 108 c) 44 d) 84
   e) 78 f) 81 g) 106 h) 58.5
5) 20.18
6) a) 75% b) 8.7% c) 27%
7) a) 57% b) 35% c) 73%
8) MgSO4
9) CaF2
10) 186.8 g
11) 80.3 g
12) 20.1 g
13) 2 moles
14) 142 g
16) 0.1 moles
17) 45.9%
20) 66%

Pages 137–138

Warm-Up Questions

1) Neutralisation.
2) A salt and hydrogen gas.
3) Copper nitrate and water.
4) Add the insoluble base to an acid until all the acid is neutralised and the excess base can be seen on the bottom of the flask. Then filter out the excess base and evaporate off the water to leave a pure, dry sample.
5) Iron is less reactive than aluminium, so it would not displace the aluminium from the salt.

Exam Questions

1 (a) A — red
B — pH 7
D — pH 8/9
E — purple
(2 marks if all correct, 1 mark for 2 or 3 correct)

(b) C *(1 mark)*

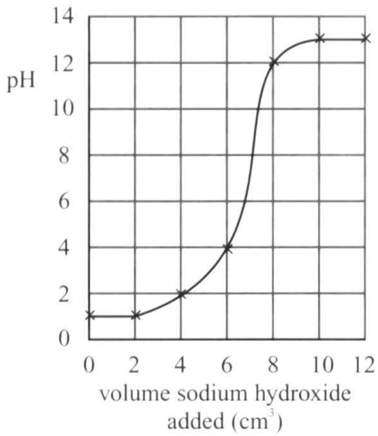
(c) E *(1 mark)*

(d) B *(1 mark)*

(e) A *(1 mark)*

With questions like this, always have a guess if you're not sure. Remember, the examiners can't take marks off you (even for a really silly answer) and if you're stuck between two possibilities you're much more likely to get a mark if you go for one of them than if you put nothing at all.

2 (a)



(1 mark for points plotted correctly, 1 mark for best fit curve)

(b) 7 cm^3 (also accept 6 or 8, depending on best fit curve at pH 7 *(1 mark)*.

(c) Because the starting pH is pH 1 *(1 mark)*.

This is the kind of question that can somehow trip you up, even if it seems obvious once you know the answer. So it's lucky you've come across it now rather than in the exam, isn't it? The pH before any alkali is added has to be the pH of the acid, and a pH of 1 means a very strong acid. See? Obvious.

(d) sodium sulfate *(1 mark)*

3 (a) sodium chloride *(1 mark)*

(b) potassium hydroxide and ammonia *(1 mark)*

(c) zinc oxide *(1 mark)*

(d) (i) $Mg + 2HCl \rightarrow MgCl_2 + H_2$ *(1 mark)*

(ii) $NH_3 + HCl \rightarrow NH_4Cl$ *(1 mark)*

(iii) $ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$ *(1 mark)*

(iv) $KOH + HCl \rightarrow KCl + H_2O$ *(1 mark)*

4 (a) It must be insoluble *(1 mark)*.

(b) silver nitrate + hydrochloric acid → silver chloride + nitric acid *(1 mark)*.

(c) First, filter the solution to remove the salt which has precipitated out *(1 mark)*. Then wash the insoluble salt (with distilled water) *(1 mark)* and then leave it to dry/dry it in an oven *(1 mark)*.

Pages 142–143

Warm-Up Questions

1) hydration
2) thermal decomposition
3) oxidation
4) precipitation
5) dehydration

Exam Questions

1 (a) Because the ethanol breaks down on heating to give two simpler products *(1 mark)*.

(b) Because water is removed from the ethanol in the reaction to make a new product *(1 mark)*.

2 (a) A *(1 mark)* and C *(1 mark)*

(b) C *(1 mark)* and E *(1 mark)*

(c) B *(1 mark)*

(d) E *(1 mark)*

3 (a) iron + oxygen + water → hydrated iron(III) oxide *(1 mark)*

(b) Because the iron atoms lose electrons *(1 mark)*.

(c) By heating to remove the water *(1 mark)*. (Also accept by using a dehydrating agent such as concentrated sulfuric acid.)

(d) B *(1 mark)* and C *(1 mark)*

4 (a) sodium hydrogencarbonate → sodium carbonate + carbon dioxide + water *(1 mark)*

(b) thermal decomposition *(1 mark)*

(c) The carbon dioxide gas produced in the reaction helps the cake to rise *(1 mark)*.

5

Reaction	Colour of precipitate when NaOH added	Name of precipitate
iron + sulfuric acid	**dark grey-green**	iron(II) hydroxide
iron + nitric acid	orange-red	**iron(III) hydroxide**

(1 mark each)

Page 146

Warm-Up Questions

1) It must be molten or dissolved in water.
2) It makes the copper a better (electrical) conductor.
3) At the cathode/negative electrode.
4) aluminium/magnesium/calcium/sodium/lithium/potassium
5) hydrogen, chlorine and sodium hydroxide

Exam Questions

1 (a) The bulb lit up because there was a flow of electrons around the circuit *(1 mark)*.

(b) At the anode/positive electrode *(1 mark)*.

(c) The positive lead ions move to the negative cathode where they gain electrons to give lead atoms *(1 mark)*. The lead metal is denser than the lead bromide so it sinks to the bottom *(1 mark)*.

(d) (i) $Pb^{2+} + 2e^- \rightarrow Pb$ *(1 mark)*

(ii) $2\,Br^- \rightarrow Br_2 + 2e^-$ *(1 mark)*

2 (a) (i) hydrogen *(1 mark)*

(ii) H^+ *(1 mark)*

(iii) $2H^+ + 2e^- \rightarrow H_2$ *(1 mark)*

(b) (i) chlorine *(1 mark)*

(ii) The negative chloride ions move to the anode and give up their electrons *(1 mark)* to form chlorine gas molecules, Cl_2 *(1 mark)*.

Page 147

Revision Summary for Section Eight

8) a) magnesium chloride: $Mg + 2HCl \rightarrow MgCl_2 + H_2$

b) aluminium sulfate: $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$

11) a) e.g. hydrochloric acid and copper(II) oxide

$2HCl + CuO \rightarrow CuCl_2 + H_2O$

b) e.g. nitric acid and calcium oxide

$2HNO_3 + CaO \rightarrow Ca(NO_3)_2 + H_2O$

c) e.g. sulfuric acid and zinc oxide

$H_2SO_4 + ZnO \rightarrow ZnSO_4 + H_2O$

Page 151

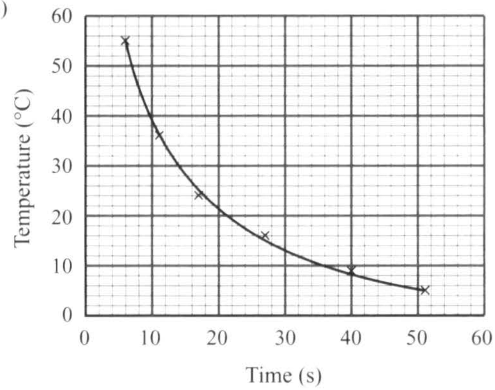
Warm-Up Questions

1) Increase the temperature (of the acid).
Use smaller pieces of/powdered magnesium.
Increase the acid concentration.
2) It would increase the time taken (i.e. reduce the rate of reaction).
3) By keeping the milk cool/storing it in a fridge.
4) The corrosion of iron is a reaction that happens very slowly.
Explosions are very fast reactions. (Other answers possible.)
5) Measure the volume of gas given off by collecting it in a gas syringe/monitor the mass of a reaction flask from which the gas escapes.

Exam Questions

1 (a) Any two of, the volume of sodium thiosulfate/hydrochloric acid / the concentration of sodium thiosulfate/hydrochloric acid / the person judging when the black cross is obscured / the black cross used (size, darkness etc.) ***(1 mark each)***.
Judging when a cross is completely obscured is quite subjective — two people might not agree on exactly when it happens. You can try to limit this problem by using the same person each time, but you can't remove the problem completely. The person might have changed their mind slightly by the time they do the next experiment — or be looking at it from a different angle, be a bit more bored, etc.

(b)



(1 mark for all points plotted correctly, 1 mark for best-fit curve)

(c) As the temperature increases the time decreases, meaning that the reaction is happening faster ***(1 mark)***. An increase in temperature causes an increase in the rate of a reaction because the particles have more energy ***(1 mark)***.

(d) Each of the reactions would happen more slowly ***(1 mark)***, although they would still vary with temperature in the same way ***(1 mark)***.

2 (a) A gas/carbon dioxide is produced and leaves the flask ***(1 mark)***.

(b) The same amount and concentration of acid was used each time, with excess marble ***(1 mark)***.

(c) (i) The marble chips were smaller/the temperature was higher ***(1 mark)***.

(ii) Smaller chips give an increased surface area, increasing the rate/ a higher temperature means the particles have more energy, increasing the rate ***(1 mark)***.

(d) The concentration of the acid is greatest at this point, before it starts being converted into products ***(1 mark)***.

Page 156

Warm-Up Questions

1) They must collide with enough energy.
2) Heat the reaction mixture.
3) The particles are squashed more closely together and so collide more often.
4) A catalyst is a substance which changes the speed of a reaction, without being changed or used up in the reaction.
5) The activation energy is the minimum amount of energy needed for the reaction to happen.

Exam Questions

1 (a) Heating makes the hydrogen peroxide particles move faster ***(1 mark)***, so they collide more often ***(1 mark)*** and with greater energy ***(1 mark)***.

(b) The surface area of the catalyst is increased ***(1 mark)***, giving the particles a greater area that they can stick to and react ***(1 mark)***.

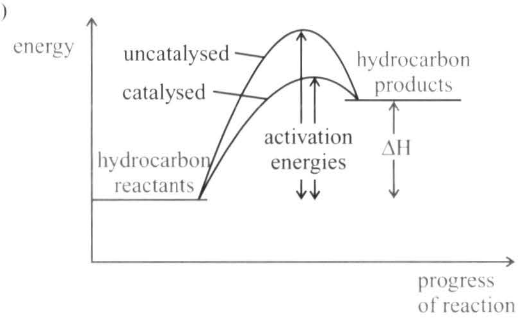
(c) By increasing the hydrogen peroxide concentration ***(1 mark)***.

2 (a) This gives the catalyst a greater surface area ***(1 mark)***.

(b) This reduces the amount of energy needed for heating ***(1 mark)***, meaning that the energy costs are reduced and the product is cheaper to make (and using less energy is also better for the environment) ***(1 mark)***.

(c) Any two of, they can be very expensive to buy / they may need to be removed from the product and cleaned, which could be costly and wastes time / they can be poisoned by impurities and need replacing ***(1 mark each)***

(d)



(1 mark for showing two curves of the correct shape that peak above the 'products' energy level. 1 mark for showing that the catalysed activation energy is lower. 1 mark for showing both the activation energies clearly.)

Pages 163–164

Warm-Up Questions

1) exothermic
2) The temperature will decrease.
3) Energy is given out — it's exothermic.
4) ΔH
5) This is the amount of energy needed to raise the temperature of 1 g of water by 1 °C.

Exam Questions

1 (a) The methane burns / gives out heat/light ***(1 mark)***.

(b) (i) C—H and O = O ***(1 mark)***

(ii) O = O, because it is a double bond ***(1 mark)***.

(c) (i) $(4 \times 414) + (2 \times 494) = 2644$ kJ/mol ***(1 mark)***

(ii) $(2 \times 800) + (4 \times 459) = 3436$ kJ/mol ***(1 mark)***

(iii) $3436 - 2644 = 792$ kJ/mole ***(1 mark)***

(d) The energy released when the new bonds are formed is greater than the energy needed to break the original bonds, so overall energy is given out ***(1 mark)***.

2 (a) x = activation energy ***(1 mark)***. y = ΔH ***(1 mark)***.

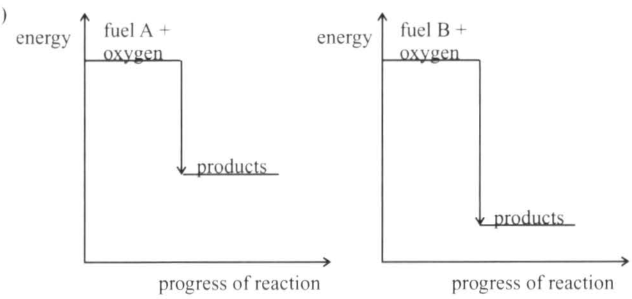
(b) Endothermic ***(1 mark)***, because the products have more energy than the reactants, so energy must be taken in during the reaction ***(1 mark)***.

(c) Because the activation energy is too high ***(1 mark)***.

3 (a) To make it a fair test/so that the temperature rise is proportional to the amount of heat produced ***(1 mark)***.

(b) $100 \times 4.2 \times 21$ ***(1 mark)*** = 8820 J/8.82 kJ ***(1 mark — units needed)***.
This gives you the energy transferred (in J) and normally you would then have to divide this by the mass of fuel burned (in g) to find the heat energy transferred per gram of fuel. But in this case only 1 g of fuel was burned anyway. So you're done.

(c)



(1 mark for showing products lower than reactants, 1 mark if products are shown at a lower level for fuel B than for A.)

Page 168

Warm-Up Questions

1) Add a catalyst.
2) They are the same.

3) It increases the amount/concentration/yield of products / shifts the position of equilibrium to the right.
4) iron/Fe
5) They are recycled and used to produce more product.

Exam Questions

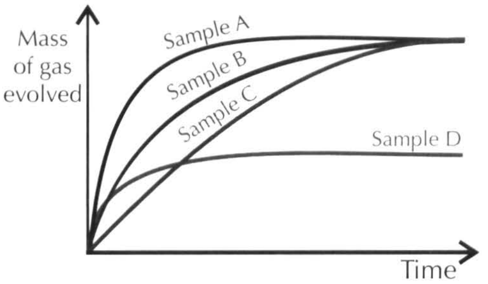
1 (a) Because the CO_2 gas that's produced would escape if the system wasn't closed. This would cause the equilibrium to shift ***(1 mark)***.
(b) Increasing the temperature ***(1 mark)*** and reducing the pressure ***(1 mark)***.

2 (a) (i) crude oil/methane/water ***(1 mark)***
(ii) air ***(1 mark)***
(b) $3H_{2(g)} + N_{2(g)} \rightleftharpoons 2NH_{3(g)}$
(1 mark for correct formula, 1 mark if correctly balanced, 1 mark for correct state symbols).
(c) E.g. fertilisers/explosives ***(1 mark)***.
(d) A high temperature reduces the equilibrium yield but increases the rate of the reaction ***(1 mark)***. If the temperature was any lower, the product would be formed too slowly ***(1 mark)***.
Remember, it's better to get a yield of 10% after 20 seconds than a yield of 20% after 60 seconds.

Page 169

Revision Summary for Section Nine

3) b)



12) a) 478 kJ/mol
b) This is an exothermic reaction.
13) 18 720 J (or 18.720 kJ)

Pages 176–177

Warm-Up Questions

1) Because ammonia is very soluble in water.
2) E.g. carbon dioxide/chlorine (any gas that is heavier than air).
3) Use damp red litmus paper — it will turn blue. (Or smell it.)
4) Oxygen.
5) Blue.

Exam Questions

1 (a) carbon dioxide / CO_2 ***(1 mark)***
(b) hydrogen ***(1 mark)***
(c) chlorine ***(1 mark)***
(d) SO_3^{2-}(s) + $\mathbf{2H^+}$(aq) ***(1 mark)*** → $\mathbf{SO_2}$(g) ***(1 mark)*** + H_2O(l)

2

Flame colour	Metal ion
blue-green	Cu^{2+}
lilac	K^+
orange-yellow	Na^+
(brick-)red	Ca^{2+}

(1 mark each)

3 Add nitric acid and then silver nitrate solution ***(1 mark)***.
A white precipitate indicates chloride ions ***(1 mark)***.
A cream precipitate indicates bromide ions ***(1 mark)***.
A yellow precipitate indicates iodide ions ***(1 mark)***.

4 Al^{3+} ***(1 mark)***
NH_4^+ ***(1 mark)***
Cu^{2+} ***(1 mark)***
Ca^{2+} ***(1 mark)***

5 (a) Add sodium hydroxide solution ***(1 mark)***. A (sludgy) green precipitate indicates iron(II) ions ***(1 mark)***.
(b) Add dilute hydrochloric acid and then barium chloride solution ***(1 mark)***. A white precipitate indicates sulfate ions ***(1 mark)***.

6 (a) Copper(II) chloride, $CuCl_2$ ***(1 mark for copper, 1 mark for chloride)***.
(b) Potassium nitrate, KNO_3 ***(1 mark for potassium, 1 mark for nitrate)***.

Pages 184–185

Warm-Up Questions

1) hydrogen ion / H^+
2) blue
3) hydrogen gas / H_2
4) The surface would char/blacken.
5) e.g. mass spectrometry

Exam Questions

1 (a) carbon dioxide / CO_2 ***(1 mark)***
(b) Bubble the gas through limewater ***(1 mark)*** and it will turn cloudy/milky ***(1 mark)***.
(c) CO_3^{2-}(s) + $\mathbf{2H^+}$ ***(1 mark)*** (aq) → $\mathbf{CO_2}$ ***(1 mark)*** (g) + H_2O(l)

2 (a) ammonia / NH_3 ***(1 mark)***
(b) NH_4^+ (aq) + $\mathbf{OH^-}$ ***(1 mark)*** (aq) → $\mathbf{NH_3}$ ***(1 mark)*** (g) + H_2O (l)
(c) It turns damp red litmus paper blue ***(1 mark)***.

3 (a) Proportion of C in CO_2 = 12 ÷ 44 ≈ 0.27
mass of C in compound = 4.4 × 0.27 ≈ 1.2 g ***(1 mark)***
moles of C = 1.2 ÷ 12 = 0.1 mol ***(1 mark)***
(b) Proportion of H in H_2O = 2 ÷ 18 ≈ 0.11
mass of H in compound = 1.8 × 0.11 ≈ 0.2 g ***(1 mark)***
moles of H = 0.2 ÷ 1 = 0.2 mol ***(1 mark)***
(c) Ratio of C:H = 0.1:0.2 = 1:2
So the empirical formula is CH_2 ***(1 mark)***.

4 (a) Mass of C in compound = 1.1 × (12 ÷ 44) = 0.3 g
Moles of C = 0.3 ÷ 12 = 0.025 mol ***(1 mark)***
Mass of H in compound = 0.675 × (2 ÷ 18) = 0.075 g
Moles of H = 0.075 ÷ 1 = 0.075 mol ***(1 mark)***
Ratio C:H = 0.025:0.075 = 1:3
So empirical formula is CH_3 ***(1 mark)***.
(b) Shake with bromine water ***(1 mark)***. The bromine water will not decolourise / will stay coloured/brown/orange ***(1 mark)***.

5 (a) C ***(1 mark)***
(b) Any two of, e.g. they're much quicker / they're more accurate / they can be carried out by technicians — you don't have to pay trained chemists to do everything / the tests can be automated / even very tiny amounts of chemical can be detected and analysed
(1 mark for each correct answer, maximum 2 marks).

6 (a) Any one of, e.g. methyl orange — turns red / blue litmus — turns red / universal indicator — turns red/orange / phenolphthalein — turns colourless ***(1 mark)***
(b) (i) hydrogen ***(1 mark)***
(ii) $2H^+$(aq) + $2e^-$ → H_2(g) ***(1 mark)***
(c) Any two of, infrared spectroscopy / ultraviolet spectroscopy / nuclear magnetic resonance (NMR) spectroscopy / mass spectrometry
(1 mark each, maximum 2 marks).

Page 186

Revision Summary for Section Ten

19) CH_2

Page 191

Warm-Up Questions

1) By evaporation.
2) Water molecules are polar.
3) calcium carbonate
4) Any one of, e.g. calcium sulfate / magnesium sulfate.
5) Ion-exchange columns and washing soda/sodium carbonate/Na_2CO_3.

Exam Questions

1 (a) 1 — evaporation ***(1 mark)***
2 — condensation ***(1 mark)***
3 — clouds ***(1 mark)***
(b) Rainwater falls onto land where it dissolves soluble substances, e.g. from rocks/fertilisers/waste ***(1 mark)*** and it then runs back to the sea (in rivers) ***(1 mark)***.

2 (a) As the temperature increases, so does the solubility ***(1 mark)***.
(b) It suggests that it is ionic bonding ***(1 mark)***.

3 (a) Any two of, e.g. Forms scum with soap. / Requires more soap for cleaning. / Forms limescale on heating systems/kettles, etc. / May block pipes. ***(1 mark each)***.
(b) $Ca(HCO_3)_2(aq) \rightarrow$ **$CaCO_3$(s)** + **H_2O(l)** + **CO_2(g)** ***(1 mark each)***

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Warm-Up Questions

1) To remove solids.
2) To kill microorganisms.
3) By distillation (boiling water and condensing the steam).
4) The higher the temperature, the lower the solubility of gases.
5) enzymes

Exam Questions

1 (a) Because above this temperature the water becomes a gas — 100 °C is the boiling point of water ***(1 mark)***.
And, fairly obviously, you can't dissolve a solid in a gas...
(b) 5 g (accept answers between 4 g and 6 g) ***(1 mark)***
(c) Solubility is 15 g per 100 g of water at 80 °C ***(1 mark)***
250 g of water will dissolve (250 ÷ 100) × 15 g = 37.5 g ***(1 mark)***
(Accept answers between 35 g and 40 g)
(d) At 90 °C, 20 g dissolved. At 10 °C, 4 g dissolved ***(1 mark)***.
20 – 4 = 16 g will crystallise out ***(1 mark)***.
(Accept answers between 14 g and 18 g.)

2 (a) Detergent molecules have a hydrophobic tail and a hydrophilic head ***(1 mark)***. The hydrophobic tail is attracted to the oil and surrounds it ***(1 mark)*** and the hydrophilic head is attracted to the water, pulling the oil away from the fabric ***(1 mark)***.
Examiners would also have to accept an answer based on surfactant properties of detergents and the lowering of surface tension here — which is all getting a bit too fancy-pants in my opinion, but it just goes to show that there's often more than one correct way to explain things.
(b) (i) sodium hydroxide (NaOH) ***(1 mark)***
Can't accept potassium hydroxide here I'm afraid — it is sometimes used to make detergents, but the question asked about sodium salts.
(ii) It's a neutralisation reaction ***(1 mark)***.
(c) Dry cleaning uses a solvent other than water / an organic solvent ***(1 mark)***. The molecules in the dry cleaning solvent are strongly attracted to the molecules in the stain, pulling the stain apart ***(1 mark)***.

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Warm-Up Questions

1) The equilibrium will move to the right/towards products.
2) The equilibrium will move to the right/towards products.
3) It would have no effect (because there are equal numbers of gas molecules on both sides).
4) The forward reaction will be faster than the reverse reaction, producing more NH_3 to try to replace the NH_3 that's been removed.

Exam Questions

1 (a) (i) The yield decreases ***(1 mark)***.
That's because the forward reaction is exothermic, and the equilibrium always shifts to try and oppose any change.
(ii) The yield decreases ***(1 mark)***, because there are more gas moles on the left/fewer gas moles on the right ***(1 mark)*** and so the equilibrium moves to the left to try and increase the pressure again ***(1 mark)***.
(b) More product (SO_3) will form because the reverse/backward reaction won't occur/will be much slower ***(1 mark)***.

2 (a) 20% ***(1 mark)*** (Accept 18–22%)
(b) As the pressure increases, the equilibrium moves to the left/towards the reactants ***(1 mark)***, because there are fewer gas moles on the left/more gas moles on the right ***(1 mark)*** and the equilibrium shifts to oppose the change ***(1 mark)***.
(c) (i) As the temperature increases the yield increases ***(1 mark)***, so the equilibrium has moved to the right/towards products ***(1 mark)***.
(ii) endothermic ***(1 mark)***

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Warm-Up Questions

1) H^+ and SO_4^{2-}
2) A proton donor.
3) An acid that only partially ionises in solution.
4) $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$
5) A strong acid is one that almost completely ionises in water. A concentrated acid is one with a large number of molecules (or ions) in a particular unit volume of it.

Exam Questions

1 (a) $NH_3(g) + H_2O(l)$ ***(1 mark)*** $\rightarrow NH_4^+(aq)$ ***(1 mark)*** $+ OH^-(aq)$
(b) Ammonia is a proton acceptor ***(1 mark)*** — it doesn't need to be in solution to do this ***(1 mark)***.
(c) Weak alkalis only partially ionise when dissolved in water ***(1 mark)***. Strong alkalis (almost) completely ionise when dissolved in water ***(1 mark)***.

2 (a) Any one of, e.g. to make the experiment a fair test / to ensure that concentration was a controlled variable in the investigation ***(1 mark)***.
(b) Hydrochloric acid, because it is a stronger acid ***(1 mark)*** and so it has a greater concentration of ions in solution ***(1 mark)***.
The moving charges carry the current, so the more of them there are, the better the solution will conduct.
(c) (i) hydrochloric acid ***(1 mark)***
(ii) The weak acid is in equilibrium with its ions at first ***(1 mark)*** but as the H^+ ions react they are used up, so their concentration decreases ***(1 mark)*** and the equilibrium moves to the right to replace them ***(1 mark)***.

Revision Summary for Section Eleven

3) lead nitrate, ammonium chloride, potassium sulfate, Ag_2SO_4, $CuSO_4$, barium sulfate, $Ba(NO_3)_2$

11) a) 75 g
b) 35 °C
c) 95 – 75 = 20 g

26) 7 cm^3

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Warm-Up Questions

1) 0.15 × 1000 = 150 cm^3
2) n = c × V = 0.1 × (25/1000) = 0.0025 mol

3) $M_r = (2 \times 23) + 12 + (3 \times 16) = 106$
Concentration = $2.65 \div 106 = 0.025$ mol/dm³

4) $280 \div 24\,000 = 0.0117$ mol

5) The point at which the indicator changes colour/the solution is just neutralised.

Exam Questions

1 (a) (i) $4.9 \times (1000/250) = 19.6$ g/dm³ ***(1 mark)***
(ii) $M_r = (3 \times 1) + 31 + (4 \times 16) = 98$ ***(1 mark)***
Concentration = $19.6 \div 98 = 0.2$ mol/dm³ ***(1 mark)***
(b) Mass = n × M_r
mass = $0.1 \times 98 = 9.8$ g ***(1 mark)***

2 (a) M_r of $CuCO_3$ = 124 and M_r of CO_2 = 44 ***(1 mark)***
Mass of $CO_2 = 37.2 \times (44/124) = 13.2$ g ***(1 mark)***
(b) n = $13.2 \div 44 = 0.3$ ***(1 mark)***
volume = $0.3 \times 24\,000 = 7200$ cm³ or 7.2 dm³ ***(1 mark)***

3 Measure 25 cm³ of the hydrochloric acid using a pipette and place the acid in a conical flask ***(1 mark)***. Add a few drops of indicator/methyl orange/phenolphthalein ***(1 mark)***. Fill a burette with the sodium hydroxide solution (1 ***mark)***. Add the sodium hydroxide slowly until the indicator changes colour ***(1 mark)***. Repeat the experiment and find the average volume required ***(1 mark)***. Use this volume to calculate the number of moles of each reactant used in the titration and use this to calculate the concentration of the hydrochloric acid ***(1 mark)***.

(Any five for a maximum of 5 marks — note that it is acceptable to have the sodium hydroxide in the conical flask and add the acid.)

4 (a) n = c × V ***(1 mark)***
n = $0.1 \times (27.5/1000) = 0.00275$ mol ***(1 mark)***
(b) HCl and NaOH react in a 1:1 ratio, so number of moles of HCl = 0.00275 mol ***(1 mark)***
(c) c = n ÷ V ***(1 mark)***
V = 25/1000 = 0.025
c = $0.00275 \div 0.025 = 0.11$ mol/dm³ ***(1 mark)***

5 (a) n = c × V ***(1 mark)***
n = $1.0 \times (30.3/1000) = 0.0303$ mol ***(1 mark)***
(b) H_2SO_4 and NaOH react in a 1:2 ratio, so number of moles of H_2SO_4 = $0.0303 \div 2 = 0.01515$ mol ***(1 mark)***
(c) c = n ÷ V ***(1 mark)***
V = 25/1000 = 0.025
c = $0.01515 \div 0.025 = 0.606$ mol/dm³ ***(1 mark)***

6 (a) Any one of, e.g. it was an anomalous result/an outlier / the first titration is often a 'rough' titration and its result is not accurate ***(1 mark)***.
(b) Moles of NaOH = c × V
= $0.1 \times (9.0/1000) = 0.0009$ mol ***(1 mark)***
HA and NaOH react in a 1:1 ratio, so moles of HA = 0.0009 moles ***(1 mark)***
Concentration of HA = n ÷ V
= $0.0009 \div 0.025 = 0.036$ mol/dm³ ***(1 mark)***
(c) The acid used in lemonade is a weak acid (e.g. citric acid) ***(1 mark)***. There will actually be very few free H^+ ions in the lemonade/the acid will be only partially ionised ***(1 mark)***.

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Warm-Up Questions

1) Oxidation is loss of electrons and reduction is gain of electrons.
2) It should be placed at the cathode (the negative electrode).
3) Oxygen forms at the anode.
4) A faraday is the charge contained in 1 mole of electrons.
1 F = 96 000 C.
5) t = Q ÷ I = 200 000 ÷ 2.5 = 80 000 seconds

Exam Questions

1 (a) K^+, SO_4^{2-}, H^+, OH^-
(2 marks if all four are correct, 1 mark if there is only one error.)
(b) (i) $2H^+ + 2e^- \rightarrow H_2$ ***(1 mark)***
(ii) $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
or: $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$ ***(1 mark)***
(c) It doesn't react with any of the reactants or products ***(1 mark)***.

2 (a) Oxidation — the copper is losing electrons ***(1 mark)***.
(b) (i) Q = I × t ***(1 mark)***
Q = 4 × 7200 = 28 800 C ***(1 mark)***
(ii) Moles = 28 800 ÷ 96 000 = 0.3 mol ***(1 mark)***
Remember, one faraday (96 000 C) is equivalent to one mole of electrons.
(iii) 1 mole of copper ions, Cu^{2+}, needs 2 moles of electrons, and so:
moles of copper = 0.3 ÷ 2 = 0.15 mol ***(1 mark)***
(iv) m = n × A_r = 0.15 × 64 = 9.6 g ***(1 mark)***

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Revision Summary for Section Twelve

1) 2.5 mol/dm³
2) 0.1 moles
3) a) 7.5 g/dm³ b) 3.75 g/dm³
4) 120 g/dm³
5) 4 g/dm³
7) a) 144 dm³ b) 0.15 moles
8) a) 44 g b) 17 g c) 40 g d) 58.5 g e) 98 g
9) a) 16 g b) 3 g c) 14 g
10) 3 dm³
12) b) 0.1 mol/dm³
14) a) 0.167 mol/dm³ b) 10.5 g/dm³
19) 6 C
22) a) 1.8 g b) 0.675 dm³

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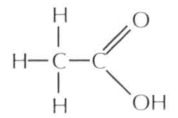
Warm-Up Questions

1) Any two of, e.g. start-up costs are lower / it's cost-effective for producing small quantities of substances such as pharmaceutical drugs / it allows several different chemicals to be made using the same equipment / any problem with contamination can easily be traced to a specific batch.
2) Conditions (such as temperature and pressure) that give a high rate of reaction may not necessarily give a high percentage yield, and so intermediate conditions may need to be used. There may also be a compromise between cost and yield.
3) $C_nH_{2n-1}OH$
4) It's used as a cleaning fluid and as a fuel (other answers possible).
5) Ethanoic acid — flavouring / preservative
Citric acid — kettle descaler / flavouring
Aspirin (acetylsalicylic acid) — analgesic / painkiller.
(Other answers are possible.)

Exam Questions

1 (a) O–H group ***(1 mark)***
(b) (i) $C_2H_5OH \rightarrow C_2H_4$ ***(1 mark)*** + H_2O ***(1 mark)***.
(ii) Any two of, e.g. plant/equipment costs / rate of production for each method / amount of energy required / how labour-intensive each process is ***(1 mark each)***.
(iii) Any two of, e.g. only one product is being made / the process can be automated / production never has to stop / it allows for large-scale production / it is a relatively simple process ***(1 mark each)***.

2 (a)



(1 mark for showing the –COOH functional group correctly, 1 mark for showing two carbons in total and correct hydrogens.)

(b) (i) ethyl ethanoate ***(1 mark)***

(ii) methyl propanoate ***(1 mark)***

(iii) water ***(1 mark)***

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Warm-Up Questions

1) It allows lots of similar compounds to be made quickly (for comparison of their properties and so that testing is cost-effective).
2) Market research, research and development, trialling, marketing, and manufacturing (including energy, labour and raw material costs).
3) A painkiller/painkilling drug.
4) By reacting it with sodium hydroxide or sodium carbonate.
5) vanadium pentoxide (V_2O_5)

Exam Questions

1 (a) (i) E.g. all have a benzene ring ***(1 mark)***.

(ii) Both have a $-COCH_3$ group ***(1 mark)***.

(iii) $C_8H_9O_2N$ ***(1 mark)***

(b) Crush the plant material ***(1 mark)***.
Dissolve the compound in a suitable solvent ***(1 mark)***.
Separate out the compound using chromatography ***(1 mark)***.

2 (a) The Contact process ***(1 mark)***.

(b) (i) A higher temperature increases the rate of reaction ***(1 mark)***.

(ii) 1–2 atmospheres/atmospheric pressure ***(1 mark)***.
There are fewer moles of product than there are of reactant so a high pressure would give more product ***(1 mark)***, but high pressures are expensive to maintain / would liquefy the SO_2 / are not really needed, as the equilibrium is already on the right ***(1 mark)***.

(c) $SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$
(1 mark for reactants, 1 mark for product).

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Warm-Up Questions

1) An electrical cell that uses the reaction between a fuel (e.g. hydrogen) and oxygen (from air, for example) to generate electricity.
2) potassium hydroxide
3) 18
4) Any three of, e.g. more than one ion (with different charges) / coloured compounds / make useful catalysts / good conductors of heat/electricity / dense / strong / shiny / hard / have high melting points.

Exam Questions

1 (a) (i) Reduction, because the reactants are gaining electrons ***(1 mark)***.

(ii) $2H_2 + O_2 \rightarrow 2H_2O$
(1 mark for the equation, 1 mark if correctly balanced.)

(b) E.g. the only product is water, so they are practical, safe and provide a useful additional source of water for long space journeys ***(1 mark)***.

(c) Any one of, e.g. conventional fuel (petrol or diesel) supplies are non-renewable and will run out in the future / petrol and diesel are very polluting and consumers/governments are increasingly seeking other options ***(1 mark)***.

2 (a) C and E ***(1 mark each)***.
C is scandium and E is nickel, if you're interested.

(b) (i) Any two of, e.g. iron has a higher melting point / iron is stronger / iron is harder / iron is denser / ***(1 mark each)***.

(ii) Any one of, e.g. iron is less reactive / iron forms more than one ion (Fe^{2+}, Fe^{3+}) ***(1 mark)***.

(c) (i) Cu^{2+} / copper(II) ***(1 mark)***

(ii) It is coloured/blue ***(1 mark)***.

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Warm-Up Questions

1) Cheshire/Teeside
2) chlorine, hydrogen and sodium hydroxide
3) hydrogen and oxygen
4) An atom or molecule with an unpaired electron.
5) They are unreactive and can exist in the atmosphere for long periods before starting to react.

Exam Questions

1 (a) $2H^+$ ***(1 mark)*** $+ 2e^-$ ***(1 mark)*** $\rightarrow H_2$

(b) $2Cl^-$ ***(1 mark)*** $- 2e^- \rightarrow Cl_2$ ***(1 mark)***

(c) (i) Any one of, e.g. Haber process/making ammonia / making margarine ***(1 mark)***.

(ii) Any one of, e.g. making disinfectants / killing bacteria / making bleach / making plastics / making hydrochloric acid / making insecticides ***(1 mark)***.

(d) sodium ***(1 mark)*** and chlorine ***(1 mark)***

2 (a) They are unreactive/inert, but the strong ultraviolet light in the stratosphere has enough energy to break up the molecules ***(1 mark)***.

(b) (i) $O_3 + Cl\bullet \rightarrow O_2$ ***(1 mark)*** $+ ClO\bullet$ ***(1 mark)***

(ii) $ClO\bullet$ ***(1 mark)*** $+ O_3 \rightarrow 2O_2 + Cl\bullet$ ***(1 mark)***

(c) It is a chain reaction ***(1 mark)***. The Cl• free radical is a product/is recycled and can react with more ozone ***(1 mark)***.

(d) alkanes ***(1 mark)*** and HFCs/hydrofluorocarbons ***(1 mark)***

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Revision Summary for Section Thirteen

8) $7 \times 6 \times 12 = 504$

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