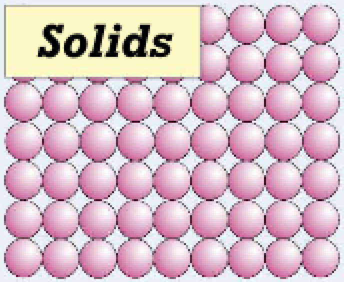
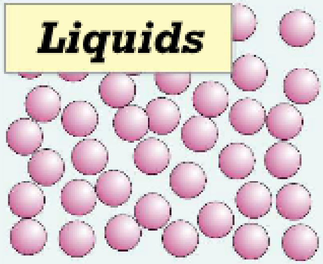
CC1a/SC1a States of Matter

Solids



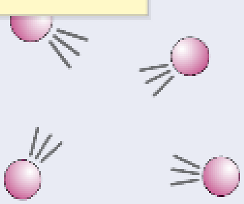
1) There are strong forces of attraction between particles, which hold them in fixed positions in a very regular lattice arrangement.
2) The particles don't move from their positions, so all solids keep a definite shape and volume, and don't flow like liquids.
3) The particles vibrate about their positions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

Liquids



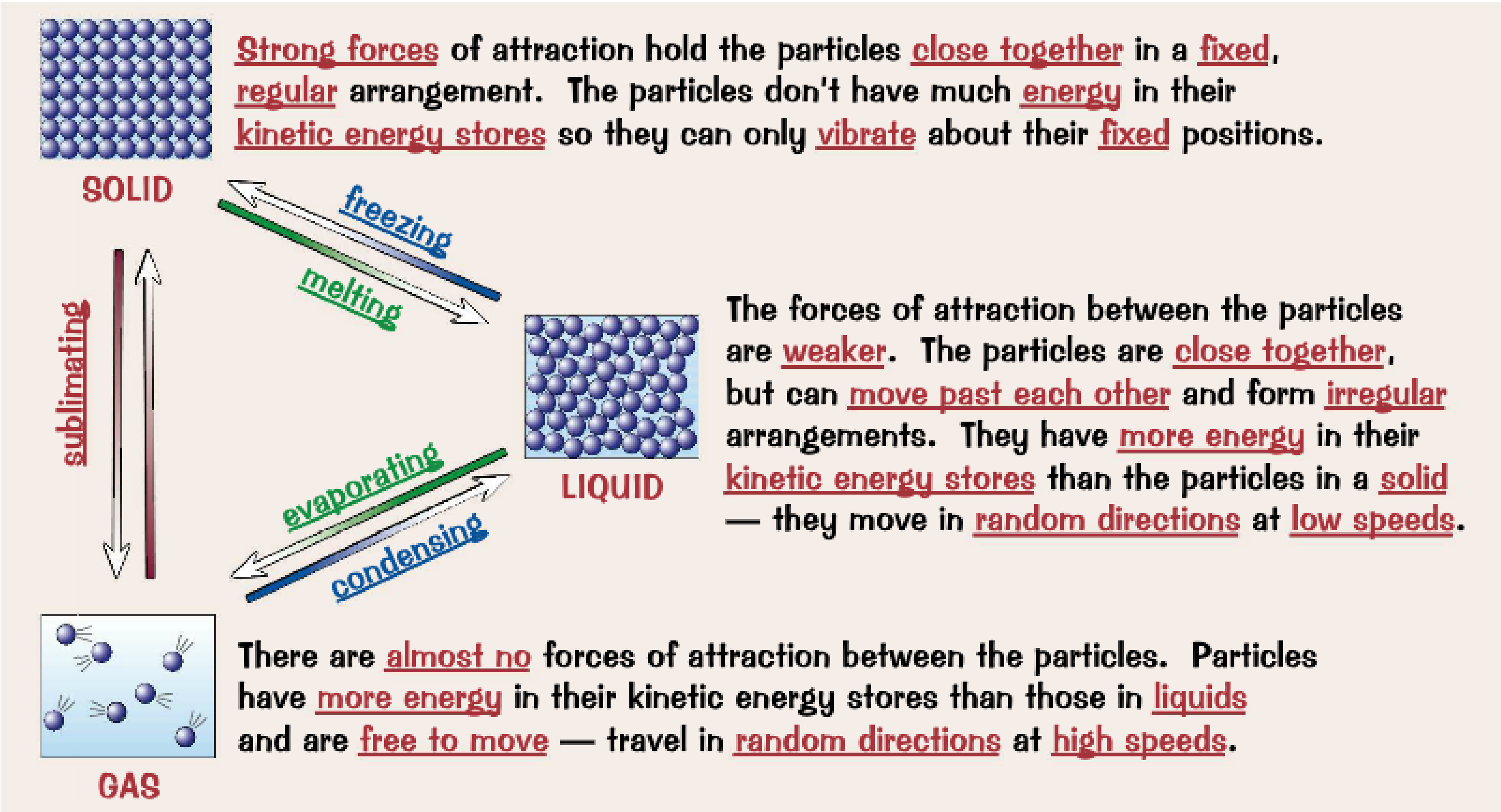
1) There is some force of attraction between the particles. They're free to move past each other, but they do tend to stick together.
2) Liquids don't keep a definite shape and will flow to fill the bottom of a container. But they do keep the same volume.
3) The particles are constantly moving with random motion. The hotter the liquid gets, the faster they move. This causes liquids to expand slightly when heated.
4) If you cool a liquid, it will freeze and become solid. If you heat a liquid enough, it evaporates (or boils) and becomes a gas.

Gases



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.
4) If you cool a gas, it will condense and become a liquid.

CC1a/SC1a States of Matter Question



Describe the differences in the movement and arrangement of particles in a solid, a liquid and a gas.

CC2a/ SC2a Mixtures

In everyday life, the word 'pure' is often used to mean 'clean' or 'natural'.

In chemistry, it's got a more specific meaning — a substance is pure if it's completely made up of a single element or compound.

If you've got more than one compound present, or different elements that aren't all part of a single compound, then you've got a mixture, not a pure substance.

So, for example, fresh air might be thought of as nice and 'pure', but it's chemically impure, because it's a mixture of nitrogen, oxygen, argon, carbon dioxide, water vapour and various other gases.

Lots of mixtures are really useful for example alloys but sometimes chemists need to obtain a pure sample of a substance .

Every pure substance has a specific melting point and boiling point.
For example, pure ice melts at 0 °C, and pure water boils at 100 °C.

So you can test the purity of a sample of a substance by comparing the actual melting or boiling point of the sample to the expected value.

If a substance is impure, the melting point will be too low.
So if some ice melts at –2 °C, it's probably got an impurity in it (e.g. salt).

The boiling point of an impure substance will be too high. For example, seawater contains salt (and other impurities). Its boiling point tends to be around 100.6 °C.

CC2a/ SC2a Mixtures

- Explain why mixtures melt over a range of temperatures but pure substances have precise melting points.
- Describe how you would separate marbles for sand .
- List the ways in which pure substances are different from mixtures.

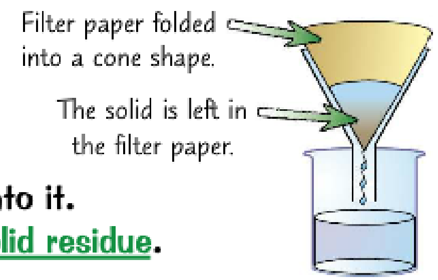
CC2b/SC2b Filtration and Crystallisation

If you've mixed a solid with a liquid, it should be pretty easy to separate them out again. Which method you'll need to use depends on whether or not the solid can dissolve in the liquid.

Filtration is Used to Separate an Insoluble Solid from a Liquid

PRACTICAL

1) If the product of a reaction is an insoluble solid, you can use filtration to separate it out from the liquid reaction mixture.
2) It can be used in purification as well. For example, solid impurities can be separated out from a reaction mixture using filtration.
3) All you do is pop some filter paper into a funnel and pour your mixture into it. The liquid part of the mixture runs through the paper, leaving behind a solid residue.

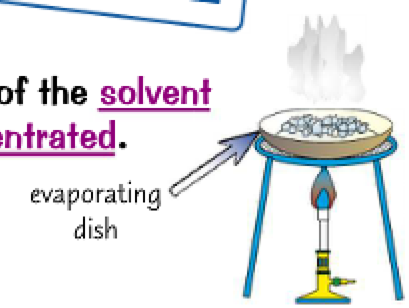


Crystallisation Separates a Soluble Solid from a Solution

PRACTICAL

Here's how you crystallise a product...

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



CC2b/SC2b Filtration and Crystallisation

- Give one example of a mixture that can be separated by filtration. Explain how this mixture is separated by filtration.
- You are given a solution that has been made by dissolving copper sulfate crystals in water.
- Describe a method you could use to extract pure copper sulfate crystals from the solution.

CC2c/SC2c Chromatography

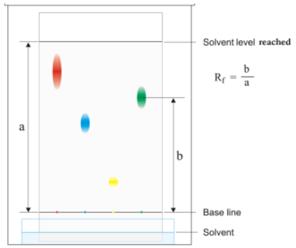
- Inks, paints and foods often contain mixtures of coloured compounds. Some coloured compounds dissolve better in a solvent than others . 'Solvent' is what substances dissolve into e.g water.
- Mixtures of coloured compounds can be separated by their solubilities through chromatography:
  – In paper chromatography, samples are placed near the bottom of a sheet of special paper ('base line' on diagram). The solvent soaks up the paper (solvent must be placed above the bottom edge of the paper but below where the samples are placed)
  – More soluble compounds in a sample are carried up the paper faster (and further) than less soluble ones, separating them. The paper with the separated components on it is called a chromatogram

Rf value:

- the distance the compound has moved up the paper divided by the distance the solvent has moved
- Further up the paper the compound has moved, greater Rf value, more soluble compound

Uses:

- The Food Standards Agency - to separate food colourings – ensures colourings used are safe
- The police - to compare a suspect's DNA sample to the DNA sample found at the crime scene
- Analysing paints and dyes – museum staff can mix exact copies of old-fashioned paints, to restore old paintings or to identify fakes



CC2c/SC2c Chromatography - Questions

- How does the process of chromatography work?
- What is the Rf value?
- What does it mean if the liquid has a higher rf value?
- How would museum staff and police use Chromatography?

CC2d/SC2d Distillation

Distillation is used to separate mixtures that contain liquids.
There are two types that you need to know about — simple and fractional.

Simple Distillation is Used to Separate Out Solutions

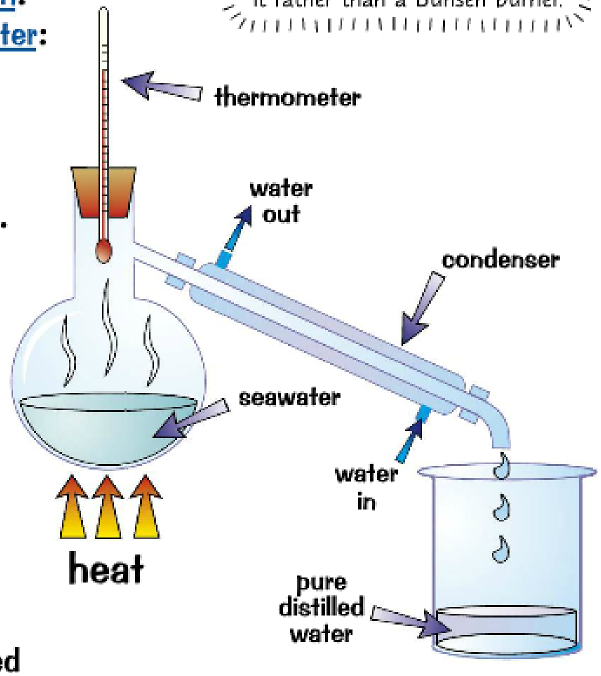
Simple distillation is used for separating out a liquid from a solution.
Here's how to use simple distillation to get pure water from seawater:

1) Pour your sample of seawater into the distillation flask.
2) Set up the apparatus as shown in the diagram. Connect the bottom end of the condenser to a cold tap using rubber tubing. Run cold water through the condenser to keep it cool.
3) Gradually heat the distillation flask. The part of the solution that has the lowest boiling point will evaporate — in this case, that's the water.
4) The water vapour passes into the condenser where it cools and condenses (turns back into a liquid). It then flows into the beaker where it is collected.
5) Eventually you'll end up with just the salt left in the flask.

The problem with simple distillation is that you can only use it to separate things with very different boiling points.

If you have a mixture of liquids with similar boiling points, you need another method to separate them out — like fractional distillation...

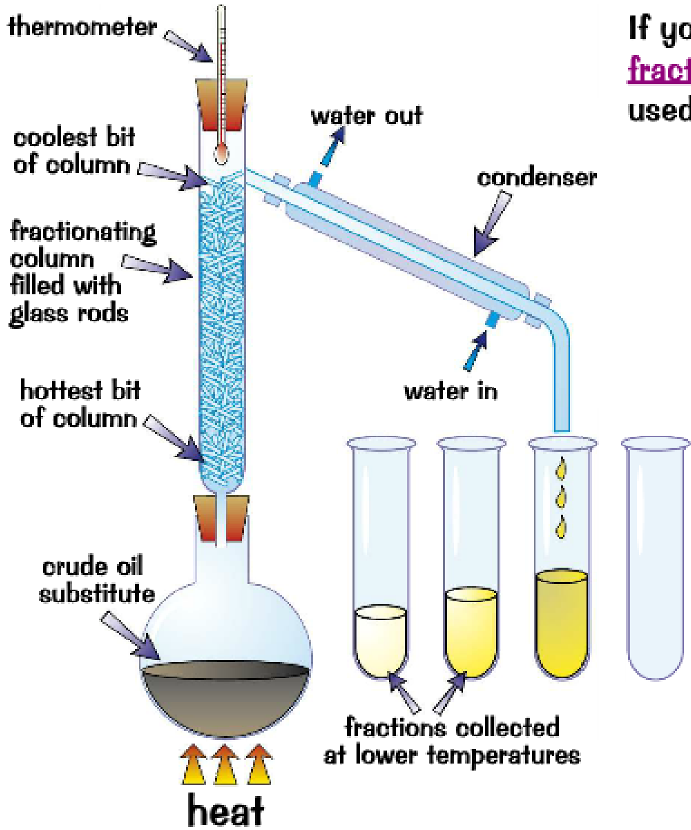
If the liquid you're heating is flammable, use an electric heater or a water bath to heat it rather than a Bunsen burner.



Explain what distillation is and how the distillation apparatus works

CC2d/SC2d Distillation

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



If you've got a mixture of liquids you can separate it using fractional distillation. Here's a lab demonstration that can be used to model fractional distillation of crude oil at a refinery:

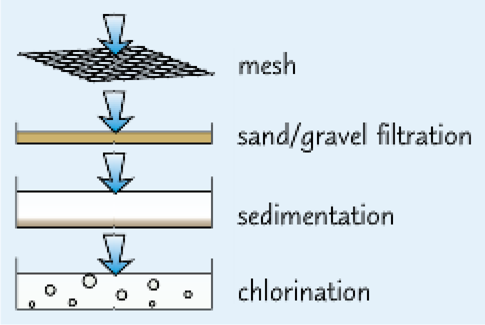
1) Put your mixture in a flask. Attach a fractionating column and condenser above the flask as shown.
2) Gradually heat the flask. The different liquids will all have different boiling points — so they will evaporate at different temperatures.
3) The liquid with the lowest boiling point evaporates first. When the temperature on the thermometer matches the boiling point of this liquid, it will reach the top of the column.
4) Liquids with higher boiling points might also start to evaporate. But the column is cooler towards the top, so they will only get part of the way up before condensing and running back down towards the flask.
5) When the first liquid has been collected, raise the temperature until the next one reaches the top.

Compare and contrast simple and fractional distillation

CC2e/SC2e Drinking Water

Water is Purified in Water Treatment Plants

How much purification the water needs depends on the source. Groundwater from aquifers is usually quite pure, but surface water needs a lot of treatment. Before we can use it, most water will be purified using the following processes:



1) Filtration — a wire mesh screens out large twigs etc., and then gravel and sand beds filter out any other solid bits.
2) Sedimentation — iron sulfate or aluminium sulfate is added to the water, which makes fine particles clump together and settle at the bottom.
3) Chlorination — chlorine gas is bubbled through to kill harmful bacteria and other microbes.

Some soluble impurities that are dissolved in the water are not removed — because they can't be filtered out. These include minerals which cause water hardness and some harmful chemicals (see below).

Tap Water Can Still Contain Impurities

The water that comes out of our taps has to meet strict safety standards, but low levels of pollutants are still found. These pollutants come from various sources:

1) Nitrate residues from excess fertiliser 'run-off' into rivers and lakes. If too many nitrates get into drinking water it can cause serious health problems, especially for young babies. Nitrates prevent the blood from carrying oxygen properly.
2) Lead compounds from old lead pipes. Lead is very poisonous, particularly in children.
3) Pesticide residues from spraying pesticides too near to rivers and lakes.

CC2e/SC2e Drinking Water

You Can Get Fresh Water by *Distilling Sea Water*

1) In some very dry countries, e.g. Kuwait, sea water is distilled to produce drinking water.
2) Distillation needs loads of energy, so it's really expensive and not practical for producing large quantities of fresh water.

- Outline how surface water is purified in a water treatment plant
- Name 3 impurities that could be present in tap water and where they come from .
- Explain why it may not be safe to drink water straight from a river .
- Describe how water is treated to deal with leaves and twigs , grit and silt and with micro organisms

CC3aSC3a Structure of the Atom

- The 3 subatomic particles are protons, neutrons, electrons.
- The nucleus contain protons and neutrons.
- The masses and charges of subatomic particles are very small and can only be measured relative to the mass of something else - the ‘relative mass’ and the ‘relative charge’
- All atoms contain the same number of protons (+) and electrons(-) so there is no overall charge
- All elements have a different number of protons in their atoms

Relative atomic mass (A_r) - atoms have very small masses - relative atomic mass is used instead of its actual mass in kilograms. A_r is the top number shown in periodic table

- A_r is the relative mass of an atom based on the mass of Carbon atoms.
  - Carbon has a relative atomic mass of 12
  - Mass of a helium atom is one third that of carbon - its relative atomic mass is 4

Isotopes - different atoms of an element with the same number of protons and electrons, but different numbers of neutrons (i.e same atomic number, different mass number)

- The presence of isotopes means that some relative atomic masses aren’t whole numbers:
  - The relative atomic mass is the average mass of all the different atoms (isotopes) of an element (relative to carbon), taking their abundance into account

Subatomic particle	Relative mass	Relative charge
Proton	1	+ 1
Neutron	1	0
Electron	Negligible (i.e 0)	- 1

CC3a/SC3a Structure of Atom - questions

- What is an atom made up of?
- What is the charge of an atom and why?
- Define relative atomic mass
- What is the relative atomic mass of a proton? A neutron? An electron?
- What is the relative charge of a proton? Neutron? Electron?
- What is an isotope?

CC3b/SC3b Atomic Number , Mass Number

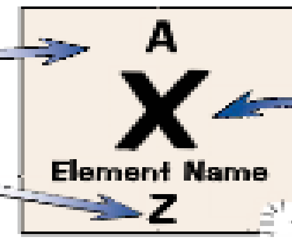
Atomic Number and Mass Number Describe an Atom

These two numbers tell you how many of each kind of particle an atom has.

In some notations and periodic tables (like the one on the data sheet in the exam), these numbers are the other way round. Just remember the bigger one is the mass number.

The Mass Number = Total no. of protons and neutrons

The Atomic Number = Number of protons



The big 'X' is the element's symbol.

1) The atomic (proton) number tells you how many protons there are.
2) Atoms of the same element all have the same number of protons — so atoms of different elements will have different numbers of protons.
3) To get the number of neutrons, just subtract the atomic number from the mass number.
4) The mass (nucleon) number is always the biggest number. On a periodic table the mass number is actually the relative atomic mass
5) Neutral atoms have no charge overall (unlike ions, see below). This is because they have the same number of protons as electrons. The charge on the electrons is the same size as the charge on the protons, but opposite — so the charges cancel out. So, the number of electrons in a neutral atom is also equal to the atomic number.

Ions have Different Numbers of Protons and Electrons

1) Ions form when atoms (or groups of atoms) gain or lose electrons
2) Negative ions form when atoms gain electrons — they have more electrons than protons. Positive ions form when atoms lose electrons — they have more protons than electrons.

- F^- — there's a single negative charge, so there must be one more electron than protons. F has an atomic number of 9, so has 9 protons. So F^- must have $9 + 1 = $ 10 electrons.
- Fe^{2+} — there's a 2+ charge, so there must be two more protons than electrons. Fe has an atomic number of 26, so has 26 protons. So Fe^{2+} must have $26 - 2 = $ 24 electrons.

CC3b/SC3b Atomic Number , Mass Number Questions

- Carbon has an atomic number of 6. How many protons does it have?
- Use a periodic table to find the number of protons and electrons in nitrogen and potassium.
- In terms of structure what do all atoms of a certain element have in common?

CC3c/SC3c Isotopes

Isotopes are different forms of the same element, which have the same number of protons but a different number of neutrons.

1) Isotopes have the same atomic number but different mass numbers.
2) If they had different atomic numbers, they'd be different elements altogether.
3) A famous example is the two main isotopes of carbon.

In the periodic table, the elements all have two numbers next to them. The bigger one is the relative atomic mass (A_r) of the element.

The relative atomic mass of an element is the average mass of one atom of the element, compared to $\frac{1}{12}$ of the mass of one atom of carbon-12.

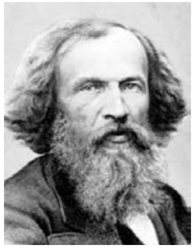
1) If an element only has one isotope (see p.85), its A_r will be the same as its mass number.
2) If an element has more than one isotope, its A_r will be the average of the mass numbers of all the different isotopes, taking into account how much there is of each one.

Example: Chlorine has two stable isotopes, chlorine-35 and chlorine-37. There's quite a lot of chlorine-35 around and not so much chlorine-37 — so chlorine's A_r works out as 35.5.

CC3c/SC3c Isotopes Questions

- What does the relative atomic mass of an element tell you?
- Describe with examples the similarities and differences between isotopes of the same element.
- State the number of protons and neutrons in an atom of uranium - 235

CC4a&b/SC4a&b History of the Periodic Table

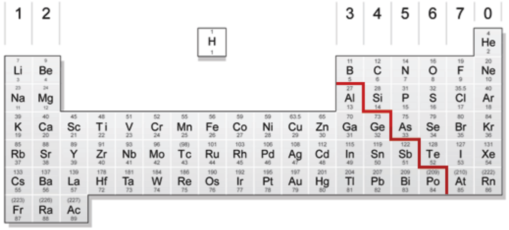


MENDELEEV - 1800s, a chemist put the properties of elements into a table:

- order of increasing atomic mass (number of protons and neutrons)
- Elements with similar properties into vertical columns (called 'groups'). Horizontal rows are called periods
- sometimes broke the 'increasing atomic mass rule'
  - he switched tellurium and iodine around so that they would be in the same groups as elements with similar properties
  - left some gaps in his table - realised from the big jumps in atomic mass that there were still some elements to discover. He predicted their properties

THE MODERN PERIODIC TABLE - elements in the modern periodic table are arranged in order of increasing atomic number rather than in order of increasing atomic mass

- In a chemical symbol the atomic number is the bottom number (the smaller one, number of protons) and the mass number is the top number (the bigger one, numbers of protons+neutrons)
- The vertical columns are called groups:
  - numbered 1-7 from left to right The group on the far right is called group 0.
- Diagram of periodic table below:
  - Elements to the left of the jagged line are metals
  - Elements to the right of the line are non-metals

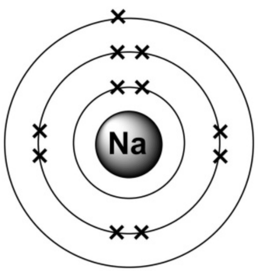


CC4a&b/SC4a&b Periodic Table - Questions

- How did Mendeleev order the elements?
- Why were elements placed in groups?
- How is the modern Table a bit different?
- Where is group 0?
- Explain how the stepped line divides the periodic table.

CC4c/SC4c Electron Configurations

- Electronic configuration - electrons are arranged in shells around the nucleus of the atom - shown as a circle drawn around the chemical symbol for the atom
- Full outer shells are stable
- Elements which have the maximum number of electrons in their outer shells are said to have 'full outer shells'. The first shell takes 2, the second 8, then 8, then 8…
- Electronic configurations can be worked out using atomic numbers e.g:
  - Atomic number of sodium is 11→it has 11 protons→11 electrons
  - 2 electrons in first shell
  - 8 electrons in second shell
  - 1 electron in third (outer) shell
  - Its electronic configuration can also be written in the form '2.8.1'. The dots separate each shell

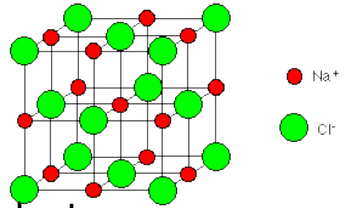


- The number of occupied shells is the same as the period number:
  - E.g magnesium is in period 3 of the periodic table…its configuration is '2.8.2'…→has 3 occupied shells
- The number of outer electrons is the same as the group number (apart from elements in group 0 which all have full outer shells):
  - E.g magnesium is in group 2 of the periodic table…its configuration is '2.8.2'…→has 2 electrons in its outer shell

CC4c/SC4cElectron Configurations - questions

- What do the electronic configurations of sodium and the other elements in group 1 have in common?
- Explain how you can tell from their electronic configuration that sodium and chlorine are in the same period.
- Describe how you can determine the atomic number, Z, of an element from its electronic configuration.
- Write the electronic configuration of phosphorus, Z=15.

CC5a,b&c / SC5a,b&c Ionic Compounds



- An ion is an atom or groups of atoms with a positive or negative charge. When electrons are gained or lost, the atoms become ions. Ionic bonds form between +ve and -ve charged ions
- Atoms of most elements have incomplete outer shells. They lose/gain electrons during chemical reactions to obtain full outer shells
- **Cations:** Metal atoms lose their outermost electrons to form positively charged ions. For elements in groups 1 and 2, the number of outer electrons lost is the same as their group number: Sodium is in group 1, (2.8.1), loses 1 electron to become a Na^{+} cation (2.8)
- **Anions:** Non-metal atoms gain electrons to form negatively charged ions. For elements in groups 6 and 7, the number of electrons they gain is 8 minus their group number: Oxygen is in group 6 (2.8.6), gains 2 electrons to become an anion (2.8.8)
- **Ionic Compounds**: form when a metal reacts with a non-metal because electrons lost by the metal are transferred to the oppositely charged non-metal
- **Names of ionic compounds:** Ionic compounds end in –ide (e.g NaCl - sodium chlor<u>ide</u>), if they contain oxygen, end in –ate (e.g $Mg(NO_3)_2$ - magnesium nitr<u>ate</u>)
- **Structure of ionic compounds:** ions in an ionic compound are tightly together and arranged in a regular lattice structure, held together by strong electrostatic forces of attraction (ionic bonds).
- **Properties of ionic compounds**: don't conduct electricity when solid (but they do when molten or in a solution) as the ions must be free to move; high boiling and melting points as ionic bonds are very strong and need a lot of energy to break

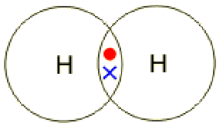
CC5a,b&c/SC5a,b&c Ionic Compounds - questions

- What is an ion?
- How do atoms gain full outer shells?
- What is the difference between an anion and a cation?
- How do you name ionic compounds?
- What is the structure of an ionic compound?
- What are the properties of ionic compounds?

CC6a/SC6a Covalent Bonds

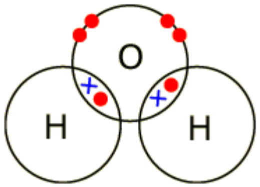
- Non-metal compounds held together by pairs of electrons in the outer shell (covalent bonds)
- Electron sharing allows both atoms to have full outer shells (more stable) = formation of molecules
- Shown by dot-cross diagrams (only outer electrons are shown because they form the bonds)

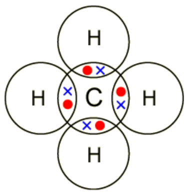
Between atoms of the same element:



- 2 hydrogen <u>atoms</u> (H) form 1 <u>molecule</u> of hydrogen gas (H_2):
  - Each hydrogen atom contributes one electron to the covalent bond. Both hydrogen atoms have full outer shells. Stable H_2 molecule is formed.
- In diagram they're shown as dots and crosses to show which atom each electron is from

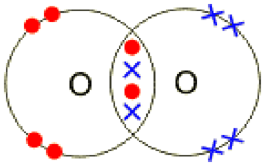
Between atoms of different elements:





- Water (H_2O):
- Methane (CH_4):

Double bonds: Atoms share more than one electron pair if needed for each atom to have a full outer shell. Two pairs of shared electrons form a double bond.



- E.g O=O (O_2):

Properties (simple molecular eg. hydrogen, methane, oxygen, water):

- Low melting and boiling points - (although there are strong covalent bonds between atoms in each molecule) there are only weak intermolecular forces between neighbouring molecules
- Poor conductors of electricity - they haven't gained or lost electrons - no ions can move around

CC6a/SC6a Covalent Bonds - Questions

- What is a covalent bond?
- Why does bonding make the atoms more stable?
- Can you draw the dot-cross diagram for water?
- Why do covalent bonds have low boiling points?
- Why are they poor conductors of electricity?

CC7a/SC7a Molecular Compounds

- All compounds contain atoms of more than one element, chemically joined together by bonds. The properties of a compound are influenced by its atoms and by its type of bonding.
- Some compounds exist as molecules-distinct groups of atoms joined by covalent bonds for example water. One molecule of water always contains one atom of oxygen covalently bonded to two atoms of hydrogen. These molecules are also known as simple molecules.
- Simple molecular substances have low melting and boiling points. They also have no overall charge so therefore cannot carry an electric current.
- Monomers are small simple molecules that can be joined in a chain to form a polymer.
- Poly(ethene) or polythene is a common polymers made of ethene molymers.

CC7a/SC7a Molecular Compounds Questions

- Given that fluorine atoms have the electronic structure 2.7, draw a dot and cross diagram to show the covalent bonding in the fluorine molecule F_2.
- Explain why water is a poor conductor of electricity.
- What is a polymer?

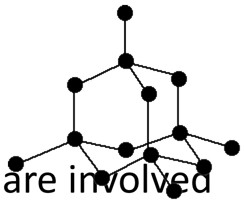
CC7b/SC7b Allotropes of Carbon

Giant molecular covalent substances (sand - silicon and oxygen atoms), diamond and graphite (both made of carbon atoms). Billions of atoms all joined together by covalent bonds.

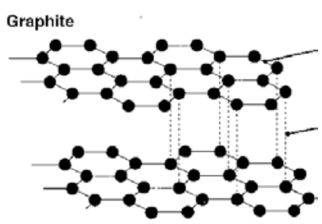
- Most of these substances have high melting and boiling points because all the atoms are joined to other atoms by strong covalent bonds (lots of heat energy is needed to break these bonds)

Properties of diamond and graphite:

- Both diamond and graphite have high melting and boiling points because of the strong covalent bonds between the carbon atoms. But they have very different properties, eg. graphite is a form of carbon in layers whereas diamond is a very compact structure.
- **Diamond:**
  - Hard because all the atoms joined with strong covalent bonds. Used to make cutting tools.Doesn't conduct electricity because there are no free ('delocalised') electrons that can move around (all four outer shell electrons in each carbon atom are involved in making bonds)



- **Graphite:** is very soft because although the covalent bonds within the layers are very strong, there are only weak forces between the layer. Graphite is used as a lubricant and for electrodes.
  Conducts electricity because only 3 outer shell electrons in each carbon atom are involved in making bonds. 1 electron from each carbon atom is free to move (electron is said to be 'delocalised'), current can flow.



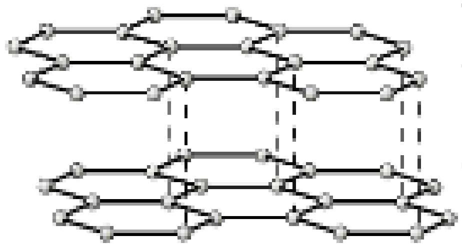
Different structural forms of the same element are called **<u>Allotropes</u>**

CC7b/SC7b Allotropes of Carbon - Questions

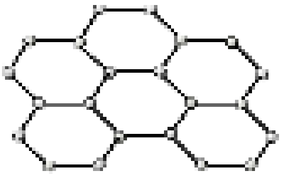
- What are the similarities and differences between graphite and diamond?
- Name a use of diamond and graphite
- Describe the structure of graphite
- Explain why graphite can conduct electricity, and why diamond can't.
- Why do they both have such a high boiling point?

CC7b/SC7b Allotropes of Carbon 2

Graphite **and** ***Graphene***



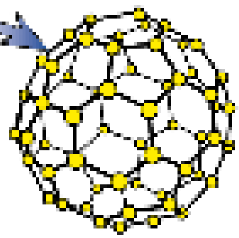
1) Graphite is black and opaque, but still kind of shiny.
2) Each carbon atom only forms three covalent bonds, creating sheets of carbon atoms which are free to slide over each other.
3) The layers are held together weakly so they are slippery and can be rubbed off onto paper to leave a black mark — that's how a pencil works. This also makes graphite ideal as a lubricating material.
4) Graphite's got a high melting point — the covalent bonds need loads of energy to break.
5) Since only three out of each carbon's four outer electrons are used in bonds, there are lots of delocalised (free) electrons that can move. This means graphite conducts electricity.



6) A single sheet of graphite is called graphene. Graphene's a bit of a wonder material — its covalent bonds make it extremely strong and a sheet of graphene is so thin that it's transparent and incredibly light. Its delocalised electrons are completely free to move about, which makes it even better at conducting electricity than graphite.

Fullerenes **are Large Carbon Molecules**

1) Fullerenes are another form of carbon. They aren't giant covalent structures, they're large molecules shaped like hollow balls or tubes. Different fullerenes contain different numbers of carbon atoms.
2) The carbon atoms in fullerenes are arranged in rings, similar to those in graphite. And like graphite, they have delocalised electrons so they can conduct electricity.
3) Their melting and boiling points aren't anything like as high as those of diamond and graphite, but they're pretty high for molecular substances because they're big molecules (and bigger molecules have more intermolecular forces).

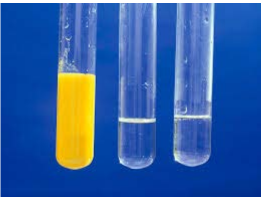


CC7b/SC7b Allotropes of Carbon 2 Questions

- Give 2 similarities and 2 differences between diamond and graphite, in terms of their structure and their properties.
- What is an allotrope?
- Explain why : a) graphite is used as electrodes.

b)Diamonds are used on cutting heads when drilling through rocks.

c)Spherical fullerenes are used in some lubricants .



CC8g/SC8g Solubility

Precipitation reactions: a reaction where an insoluble solid (precipitate) is produced from two soluble substances

Lead nitrate + potassium iodide → lead iodide + potassium nitrate

$Pb(NO_3)_{2\ (aq)} + 2KI_{\ (aq)} \rightarrow PbI_{2\ (s)} + 2KNO_{3\ (aq)}$

State symbols show that all substances are dissolved in water (aq) except for lead iodide, which is insoluble (so it's shown as a solid s)

The precipitate is separated from the unreacted ions by filtration, washed on filter paper, dried in an oven.

Diagnosing intestinal problems: patients swallow a drink - 'barium meal' - containing barium sulfate. As it passes through the patient's digestive system, x-ray photos are taken. Barium (like bone) absorbs x-rays, shows up as white on the photos, shows problems with digestive system

Most barium salts are toxic but barium sulphate is insoluble, so can't enter the patient's blood so is safe to swallow

Soluble in water	Insoluble
All common sodium, potassium and ammonium salts	
All nitrates	
Most chlorides	Silver/lead chlorides or iodides
Most sulfates	lead/barium/calcium sulfates
sodium/potassium/ammonium carbonates	Most carbonates
sodium/potassium/ammonium hydroxides	Most hydroxides

CC8g/SC8g Solubility - Questions

- What is a precipitation reaction?
- How would you recognise a precipitation reaction from the state symbols?
- Which chlorides are insoluble in water?
- Which hydroxides are soluble in water?
- What is barium meal used for?
- Why doesn't barium poison us?

CC7c/SC7c Properties of Metals

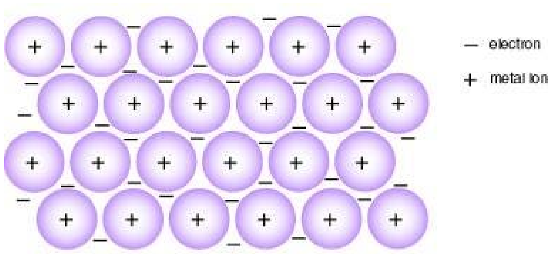
The atoms in metals are held together by metallic bonding (see below) – this gives metals different properties from other types of substances

- Metals are good conductors of electricity and heat
- Metals are solids at room temperature (→metallic bonds are strong), except for mercury, which is a liquid at room temperature
- Metals don't dissolve in water
- Metals are malleable (can be hammered into shape) metals have many uses e.g. they're used to make cars, buildings, tools

Metallic bonding:

- Metal structure is a giant lattice of positive ions surrounded by a 'sea' of outer shell electrons. The term 'sea of electrons' is used because electrons in the outer shells of metal atoms are free to move through the structure. The electrons aren't located in specific atoms so we say they are 'delocalised electrons'.
- Metals conduct electricity because delocalised electrons move around randomly in all directions between the positive ions. If a potential difference (i.e voltage) is applied across a piece of metal, all the delocalised electrons start to move in the same direction. This movement of electrons is an electric current

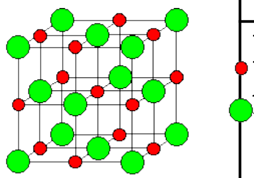
- Metals are malleable:
  - If a large force is applied, the layers of positive ions in a metal can slide over each other. The positive ions are still held together by the sea of electrons and the metal spreads out (changes shape) instead of breaking

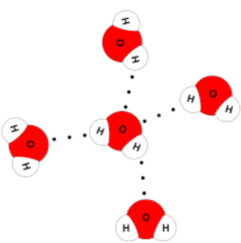


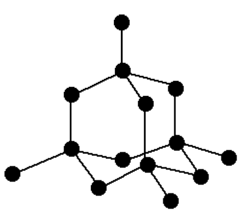
CC7c/SC7c Properties of Metals Questions

- Give four properties of metals
- What is the word that means metals can be hammered into shape?
- What is the phrase that describes the structure of metals?
- Explain why metals can conduct electricity.
- Explain why metals are malleable.

CC7d/SC7d Bonding Models





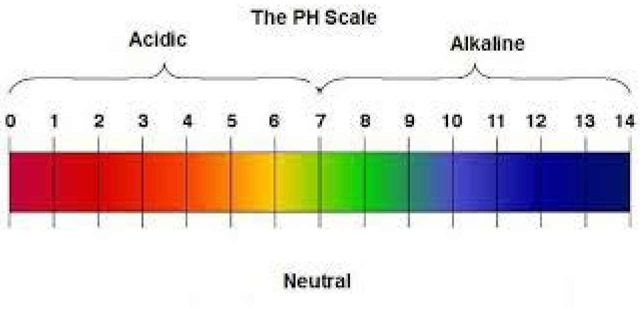


Type of structure & bonding	Giant Ionic	Simple Molecular Covalent	Giant molecular covalent
How bonds form	metal reacts with a non-metal • Metals lose electrons to become positive cations • Non-metals gain electrons to become negative anions • Oppositely charged ions attract	Between atoms of non-metal elements: Electrons are shared between atoms so they end up with stable, full outer shells	between atoms of non-metal elements Electrons are shared between atoms so they end up with a stable, full outer shells
Examples	sodium chloride	water, oxygen	diamond
Strength of bonds	strong	Strong covalent bonds between atoms, but weak forces holding separate molecules together	Strong across all atoms in a structure
Melting and boiling points	high - solids at room temperature	low – most are liquids or gases at room temperature	high – solids at room temperature
Solubility	many dissolve in water	some dissolve in water	Insoluble in water
Do they conduct electricity?	• yes when molten or in an aqueous solution • no when solid	no	no, except graphite

CC7d/SC7d Questions on Bonding Models

- How does an ionic bond form? What does it form between?
- What is a giant structure compared to a simple structure?
- Which type of structures have strong bonds?
- Describe how the melting and boiling point of structures is linked to their structure.
- Which structures are soluble?

CC8a/SC8a Acids, Alkalis and Indicators



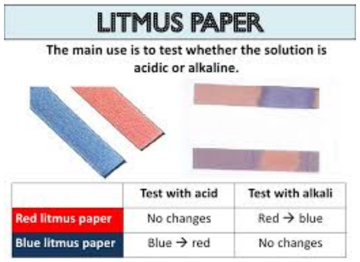
- Acids and alkalis can be described using the pH scale
  - Scale runs from pH1 (strong acid) to pH14 (strong alkali)
  - A neutral liquid has a pH of 7 (e.g. water)
  - pH5 - weak acid and pH9 - weak alkali

Note - difference between a base and an alkali

- – *Some bases are soluble and when dissolved in water, bases are called alkalis*

- Universal indicator can be used to find out if a liquid is an acid or an alkali by dipping it into the liquid and observing its colour:
  - Yellow/orange/red – acid
  - green – neutral
  - blue/purple – alkali
  - The closer to red the stronger the acid (i.e red – strong acid, yellow – weak acid)
  - The closer to purple the stronger the alkali (i.e purple – strong alkali, blue – weak alkali)

- Litmus paper (must be damp to work) is also an indicator
  - – Blue litmus paper turns red under acidic conditions (no change under alkaline or neutral conditions)
  - – Red litmus paper turns blue under alkaline conditions (no change under acidic or neutral conditions)



CC8a/SC8a Questions Acids, Alkalis and Indicators

- If the pH of a substance =6, what two words would you use to describe it?
- If the pH of a substance = 14, what two words would you use to describe it?
- Water has a pH of 7 which means it is neutral. What colour would it turn Universal Indicator?
- How is an alkali related to a base?
- What do you have to do to litmus paper for it to work?
- What colour would blue litmus turn in acidic conditions?

CC8b/SC8b Looking at Acids

Acids Can be Strong or Weak

1) Strong acids (e.g. sulfuric, hydrochloric and nitric acids) ionise almost completely in water. A large proportion of acid molecules dissociate to release H^+ ions. They tend to have low pHs (pH 0-2).
2) Weak acids (e.g. ethanoic, citric and carbonic acids) do not fully ionise in solution. Only a small proportion of acid molecules dissociate to release H^+ ions. Their pHs tend to be around 2-6.
3) The ionisation of a weak acid is a reversible reaction, which sets up an equilibrium mixture. Since only a few of the acid molecules release H^+ ions, the equilibrium lies well to the left.

Strong acid: $HCl \rightarrow H^+ + Cl^-$

Weak acid: $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$

Don't Confuse Strong Acids with Concentrated Acids

1) Acid strength (i.e. strong or weak) tells you what proportion of the acid molecules ionise in water.
2) The concentration of an acid is different. Concentration measures how much acid there is in a litre (1 dm^3) of water. Concentration is basically how watered down your acid is.
3) An acid with a high proportion of acid molecules compared to the volume of water is said to be concentrated. An acid with a low proportion of acid molecules compared to the volume of water is said to be dilute.

Concentration is measured in g/dm^3 or mol/dm^3.

4) Note that concentration describes the total number of dissolved acid molecules — not the number of molecules that produce hydrogen ions.
5) The more grams (or moles) of acid per dm^3, the more concentrated the acid is.
6) So you can have a dilute but strong acid, or a concentrated but weak acid.

CC8b/SC8b Questions Looking at Acids

Changing the Concentration of an Acid Affects its pH

If the concentration of H^+ ions increases by a factor of 10, the pH decreases by 1. So if the H^+ ion concentration increases by a factor of 100 (= 10 × 10), the pH decreases by 2 (= 1 + 1), and so on. Decreasing the H^+ ion concentration has the opposite effect — a decrease by a factor of 10 in the H^+ concentration means an increase of 1 on the pH scale.

EXAMPLE: A solution with a hydrogen ion concentration of 0.001 mol/dm³ has a pH of 4. What would happen to the pH if you increased the hydrogen ion concentration to 0.01 mol/dm³?

The H^+ concentration has increased by a factor of 10, so the pH would decrease by 1. So the new pH would be 4 – 1 = 3.

Explain the difference between a strong acid and a weak acid.

What does the pH of a solution measure?

The most acidic rain recorded in Britain had a pH of 2. How much more acidic is this than normal rainwater with a pH of 5?

CC8c/SC8c Bases and Salts

Bases are substances that neutralise acids to form a salt and water only. All metal oxides are bases

Acid + Metal oxide (base) → salt + water

Acid +Metal hydroxide (alkali) → salt + water

Acid +Metal carbonate → salt + water + carbon dioxide

The salt formed depends on the acid. In the case of indigestion remedies the acid being neutralised is HCl (as it is the one present in the stomach) so chloride salts are produced.

e.g. if indigestion tablet contains calcium carbonate

Hydrochloric acid + calcium carbonate →calcium chloride + water + carbon dioxide

$$2HCl_{(aq)} + CaCO_{3\,(s)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

Sulphuric acid and nitric acid can also be neutralised.

If sulfuric acid is neutralised→ sulphate salts are produced:

e.g. Sulfuric acid + copper oxide→ copper sulfate + water

$$H_2SO_{4\,(aq)} + CuO_{(s)} \rightarrow CuSO_{4\,(aq)} + H_2O_{(l)}$$

If nitric acid is neutralised→ nitrate salts are produced:

e.g. Nitric acid + sodium hydroxide → sodium nitrate + water

$$HNO_{3\,(aq)} + NaOH_{(s)} \rightarrow NaNO_{3\,(aq)} + H_2O_{(l)}$$

Problem	Neutraliser	Effect
Decaying Teeth	Toothpaste	Acid + Alkali = Neutral
Indigestion	Milk of Magnesia	Acid + Alkali = Neutral
Wasp Sting	Vinegar	Alkali + Acid = Neutral
Bee Sting	Baking Soda	Acid + Alkali = Neutral
Soil Decay	Adding Lime	Acid + Alkali = Neutral

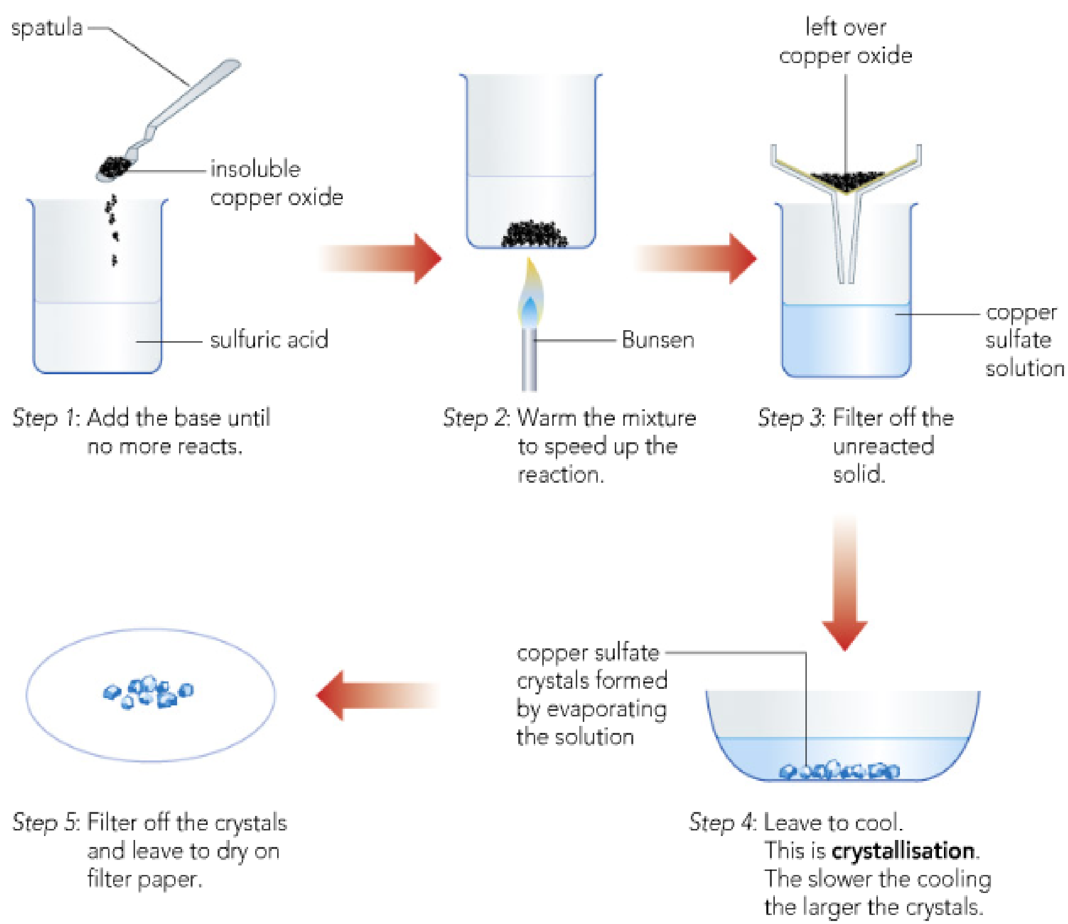
CC8c/SC8c Questions on Bases and Salts

- Finish the following general equations:
- Acid + Base →
- Acid + Metal →
- Acid + Metal carbonate →
- Name the salt formed in the following reactions:
- Hydrochloric acid + potassium oxide
- Nitric acid + Lithium hydroxide
- Sulphuric Acid + copper carbonate

CC8c/SC8c Bases and Salts

<u>Making Copper Sulphate from Copper Oxide</u>

1. React excess oxide (***insoluble***) with accurate volume acid.
2. Filter excess copper oxide and collect copper sulphate solution.
3. Evaporate solvent (water) to crystallise the salt.



CC8c/SC8c Bases and Salts Questions

- Draw a flow chart to explain the steps involved in preparing a soluble salt from an insoluble base

CC8d/SC8d Alkalis and Balancing Equations

A base is any substance that reacts with an acid to form a salt and water only. Many bases are insoluble in water. A base that can dissolve in water is called an alkali . Alkalis form alkaline solutions with pH values above 7

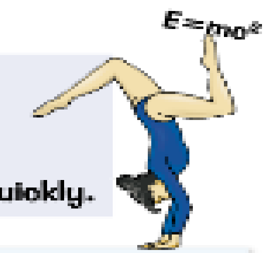
Symbol Equations **Need to be** *Balanced*

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

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

3) The formulas are all correct but the numbers of some atoms don't match up on both sides.
4) You can't change formulas like H_2SO_4 to H_2SO_5. You can only put numbers in front of them.

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



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

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

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

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

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

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

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

CC8d/SC8d / Alkalis and balancing equations - questions

- Explain why all alkalis are bases but not all bases are alkalis.
- What happens when metal hydroxides react with acids?
- Magnesium hydroxide solution reacts with stomach acid which is hydrochloric acid. Magnesium chloride solution is a product. Write a balanced equation for the reaction including state symbols.

CC8e/SC8e Alkalis and Neutralisation

Acids and Bases Neutralise Each Other

The reaction between acids and bases is called neutralisation. It produces a salt and water

$$HCl + NaOH \rightarrow NaCl + H_2O$$

acid base salt water

Neutralisation reactions in aqueous solution can also be shown as an ionic equation in terms of H^+ and OH^- ions:

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

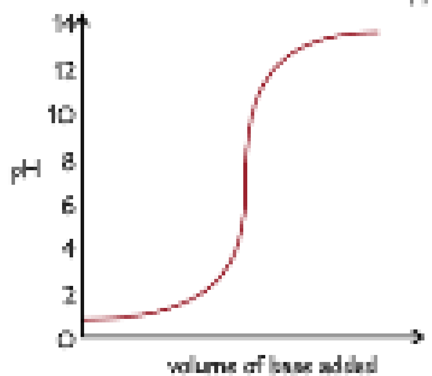
This is just the ionic equation of the reaction above.

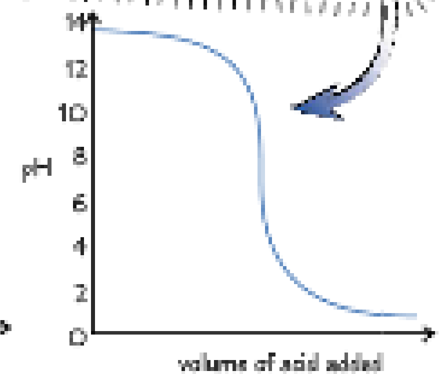
When an acid neutralises a base (or vice versa), the products are neutral, i.e. they have a pH of 7. At pH 7, the concentration of hydrogen ions equals the concentration of hydroxide ions. An indicator can be used to show that a neutralisation reaction is over (Universal indicator will go green).

Titration Curves Show pH Changes with Volume

1) Experiments called titrations are used to work out how much of an acid is used to neutralise a base of unknown concentration (or vice versa).
2) Titration curves are used to show where neutralisation happens during a titration. There's a vertical point in the curve which is where the solution is neutral (at pH 7). This is called the end point of the titration.

If you add acid to a base, the pH will decrease. If you add a base to an acid, then the pH will increase.





CC8e/SC8e Alkalis and Neutralisation Questions

- Explain how water is formed in neutralisation reactions.
- Explain in terms of ions how potassium hydroxide solution reacts with sulphuric acid to form water and potassium sulphate.
- The pH of an unknown solution is found to be 6. Is this solution acidic or alkaline?

CC8f/SC8f Reactions of acids with metals and carbonates

Many Metals React With Acids to Give Salts

Acid + Metal → Salt + Hydrogen

A salt is an ionic compound, formed as part of a neutralisation reaction.

Hydrochloric Acid Produces Chloride Salts:

$2HCl + Mg \rightarrow MgCl_2 + H_2$ (Magnesium chloride)
$6HCl + 2Al \rightarrow 2AlCl_3 + 3H_2$ (Aluminium chloride)

Sulfuric Acid Produces 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)

Nitric Acid Produces Nitrate Salts When NEUTRALISED, But...

The reaction of nitric acid with metals is more complicated — you get a nitrate salt, but instead of hydrogen gas, the other products are usually a mixture of water, NO and NO_2.

You can't really predict the balanced equation for the reaction of nitric acid with metals.

Metal Carbonates Give Salt + Water + Carbon Dioxide

Acid + Metal Carbonate → Salt + Water + Carbon Dioxide

Examples:
$2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$ (Sodium chloride)
$H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + H_2O + CO_2$ (Potassium sulfate)
$2HNO_3 + ZnCO_3 \rightarrow Zn(NO_3)_2 + H_2O + CO_2$ (Zinc nitrate)

Again, as above, hydrochloric acid produces chloride salts, sulfuric acid produces sulfate salts, and nitric acid produces nitrate salts.

CC8f/SC8f Reactions of acids with metals and carbonates - questions

- Zinc reacts with dilute hydrochloric acid. Write a) the word equation and b) the balanced equation.
- Write a balanced chemical equation for the reaction of hydrochloric acid with calcium carbonate.
- Describe a test that is to show a gas is hydrogen.

CC9a/SC9a Masses and Empirical Formulae

The true formula for a simple molecular compound is called the ‘molecular formula’ – this shows the actual number of atoms of each element in a molecule
Substances can also be represented by an empirical formula – this shows the simplest whole number ratio of atoms or ions of each element in a substance.
e.g. Ethene has the molecular formula C_2H_4 and Propene has the molecular formula C_3H_6.
For every molecule of ethene or propene, there are twice as many hydrogen atoms than carbon atoms. The molecular formulae of both ethene and propene can be simplified to CH_2 –the empirical formula.
So ethene and propene have different molecular formulae (C_2H_4 and C_3H_6) but the same empirical formula (CH_2)

Calculating the empirical formula
If you know the mass (in grams) of each element present in a compound, you can use this and the relative atomic mass of each element to calculate the empirical formula of the compound e.g. for calcium chloride:

	Ca	Cl
Mass in g	10.0	17.8
Relative atomic mass	40	35.5
Step 1. Divide the mass of each element by its relative atomic mass	10.0/40 = 0.25	17.8/35.5 = 0.5
Step 2. Divide the answers by the smallest number obtained after step 1 (in this case 0.25) to find the simplest ratio	0.25/0.25 = 1	0.5/0.25 = 2
	i.e for every molecule of calcium chloride, there are twice as many chlorine atoms than calcium atoms	
Empirical formula	$CaCl_2$	

CC9a/SC9a Masses and Empirical Formulae Questions

- What does the formula CO_2 mean?
- What is the difference between the molecular formula and the empirical formula?
- What is the empirical formula of $C_6H_{12}O_6$?
- What is the empirical formula of Na_2O_2?
- Outline the key steps to calculate the empirical formula given the masses of elements reacting.

CC9a/SC9a Masses and Empirical Formulae 2

RELATIVE MASSES

The relative atomic mass (A_r) is the mass of an atom compared to that of carbon-12

Calculating the relative formula mass (M_r)

The ‘relative formula mass’ (M_r) of a substance is the sum of the relative atomic masses (A_r) of all the atoms or ions in its formula:

e.g CO_2 – 1 carbon atom, 2 oxygen atoms:

- A_r of C = 12
- A_r of O = 16 x 2 = 32
- relative formula mass (M_r) of CO_2 = 12 + 32 = 44

e.g $Ca(NO_3)_2$ – 1 calcium atom, 2 nitrogen atoms, 6 oxygen atoms:

- A_r of Ca = 40
- A_r of N = 14 x 2 = 28
- A_r of O = 16 x 6 = 96
- relative formula mass (M_r) of $Ca(NO_3)_2$ = 40 + 28 + 96 = 164

CC9a/SC9a Questions on Masses and Emperical Formulae 2

- Helium has an atomic mass of 4. Magnesium has an atomic mass of 24. How many times heavier is an atom of Magnesium than an atom of Helium?
- Calculate the formula masses of the following:

$CaCO_3$

LiCl

CO_2

Glucose - $C_6H_{12}O_6$

CC9b/SC9b Conservation of Mass

Atoms/matter/mass are **NOT** made or destroyed in a chemical reaction. They are only **rearranged** to form new products.

total mass before = total mass after a reaction

This rearrangement of atoms means products and reactants have different physical and chemical properties

Precipitation reactions

- This is when two soluble (shown by $_{(aq)}$ state symbol) substances react together to form an insoluble (shown by $_{(s)}$ state symbol) product, called the precipitate.

e.g.

Silver nitrate + potassium bromide → potassium nitrate + silver bromide

$AgNO_{3\ (aq)} + KBr_{(aq)} \rightarrow KNO_{3\ (aq)} + AgBr_{(s)}$

AgBr is the solid precipitate formed in this example

Remember that even though a solid is formed, the total mass before and after the reaction is still the same!

CC9b/SC9b Questions on Conservation of Mass

- State the Law of Conservation of Mass.
- How would you identify a precipitation reaction?
- $AgNO_{3\,(aq)}$ + $KBr_{(aq)}$ → $KNO_{3\,(aq)}$ + $AgBr_{(s)}$

  58g + 45g → 63g + X

- Using the Law of conservation of mass what is the mass of AgBr in the equation above?

CC9b/SC9b Calculating the Masses of Reactants or Products:

During a chemical reaction, no atoms are lost or made – they are just rearranged to make new substances (the 'products'.) You can use relative masses and the balanced equation for a reaction to calculate the mass of a reactant or product.
E.g potassium nitrate (KNO_3) is decomposed to potassium nitrite (KNO_2) and oxygen (O_2).
what mass of potassium nitrate is needed to make 1.6g of oxygen?

Step 1 – balanced chemical equation: $2KNO_3 \rightarrow 2KNO_2 + O_2$
Step 2 – work out the relative masses of the substances needed in the calculation

- M_r of KNO_3 = 101
- There are two particles of KNO_3 in the balanced equation →we multiply the M_r of KNO_3 by 2…→ 101 x 2 = 202
- Relative mass (only one atom so A_r/M_r it's the same) of O_2 = 32

Step 3 – divide the answers by the smallest relative mass calculated above (in this case 32) to find the ratio:

- $2KNO_3$: 202/32 = 6.3
- O_2: 32/32 = 1
- →ratio of $2KNO_3$:O_2 is 6.3:1

Step 4 – use the ratio to find the answer!

- Ratio is 6.3:1→6.3g of potassium nitrate are needed to make 1g of oxygen
- to find the mass of potassium nitrate needed to make 1.6g of oxygen we multiply 1.6 by 6.3→1.6 x 6.3 = 10.1
- →10.1g of potassium nitrate are needed to make 1.6g of oxygen

CC9b/SC9b Questions on Masses of Reactants and Products

$Li + O_2 \rightarrow Li_2O$

How much Lithium oxide can form from 7g of Lithium if there is an excess of oxygen?

$C_2H_4 + O_2 \rightarrow 2CO_2 + 2H_2O$

What mass of ethene is needed to make 8.8g of Carbon dioxide?

CC9c/SC9c Moles

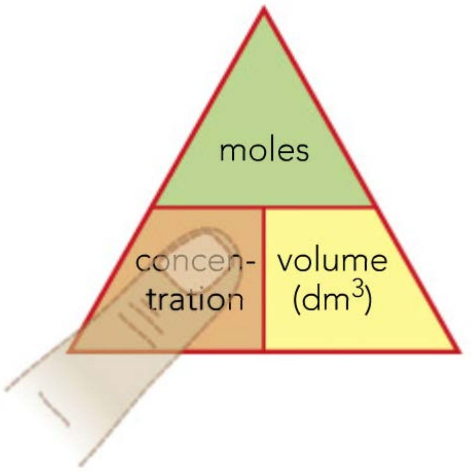
"The Mole" is Simply the Name Given to a Certain Number

1) Just like a million is this many: 1 000 000, or a billion is this many: 1 000 000 000, a mole is given by Avogadro's constant, and it's this many: 602 200 000 000 000 000 000 000 or 6.022×10^{23}.
2) But what does Avogadro's constant show? The answer is that when you get that number of atoms or molecules, of any element or compound, then, conveniently, they weigh exactly the same number of grams as the relative atomic mass, A_r (or relative formula mass, M_r) of the element or compound.

- To calculate the number of moles use the following formulas:

$$\text{number of moles of an element} = \frac{\text{mass of element in grams}}{\text{relative atomic mass}}$$

$$\text{number of moles of a compound} = \frac{\text{mass of compound in grams}}{\text{relative formula mass}}$$

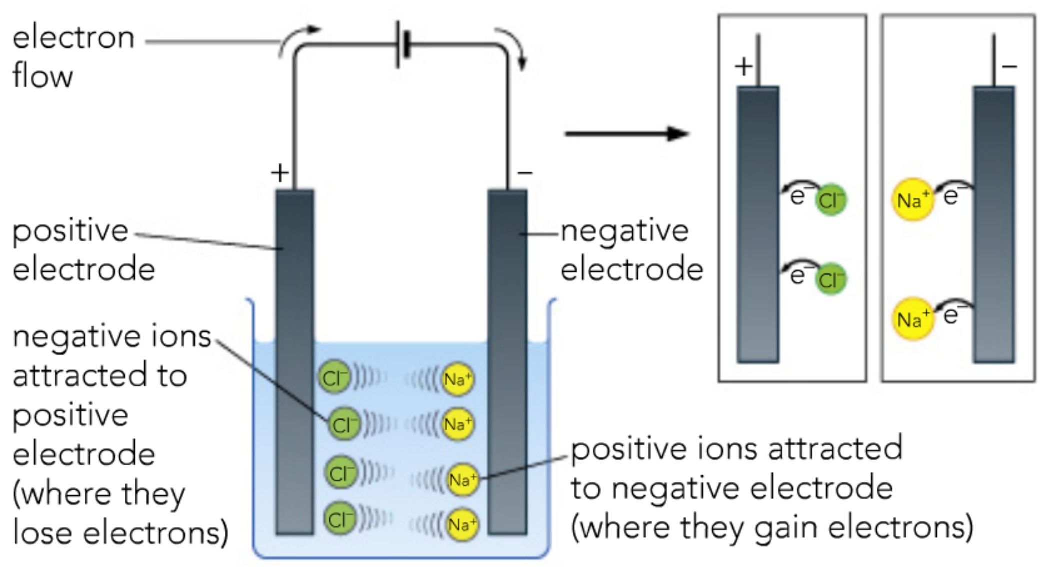


CC9c/SC9c Questions on Moles

- Calculate the mass of : a) 2mol of nitrogen molecules, N_2 b) 0.1mol of sulphur dioxide molecules, SO_2.
- Calculate the number of molecules in 90g of water, H_2O.
- Give the mass of the Avogadro constant number of a) carbon atoms b) sodium atoms .

CC10a/SC10a Electrolysis

- Electrolysis can only happen when **IONIC** substances are either **DISSOLVED** or **MOLTEN**.
- Sodium metal is produced through the electrolysis of **MOLTEN** Sodium Chloride.



CC10a/SC10a Electrolysis 2

- Oxidation is the loss of electrons and happens at the **ANODE**.
- Reduction is the gain of electrons and happens at the **CATHODE**.

Half-equations

A **half-equation** shows the change at just one of the electrodes. In the electrolysis of molten sodium chloride, the reaction at the cathode is:

$$2Na^+ + 2e^- \rightarrow 2Na$$

The reaction at the anode is:

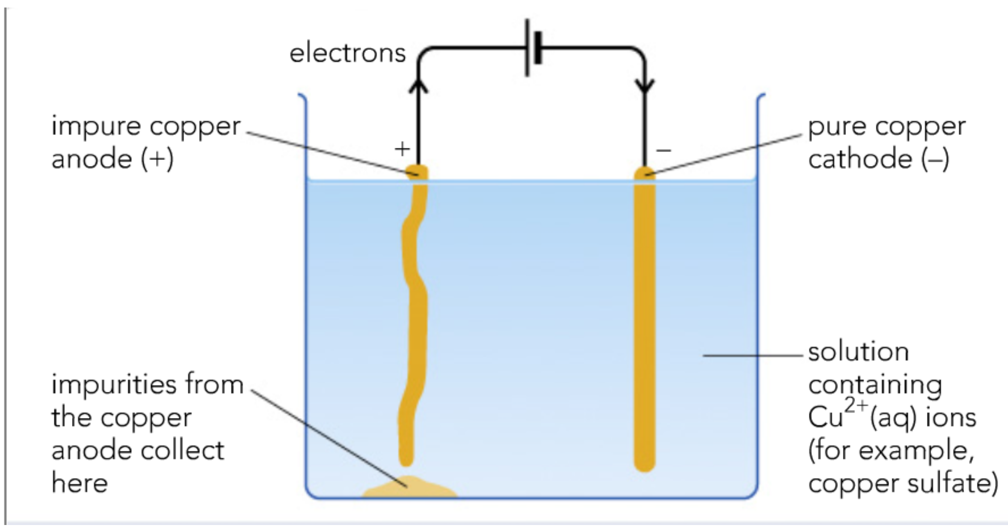
$$2Cl^- \rightarrow Cl_2 + 2e^-$$

Note that two Cl^- ions are needed to form one chlorine molecule, Cl_2.

- Sodium metal is used in **street lights** as it gives out a yellow coloured light.
- Liquid Sodium is used as a coolant in **Nuclear Reactors** as it has a high **THERMAL CONDUCTIVITY**.

CC10a/SC10a Uses of Electrolysis

- Purification of Copper
- An electrode of impure copper is used as the **ANODE**.
- Pure copper is used as the **CATHODE**.
- The **electrolyte** is Copper Sulphate solution.



The half-equation for the anode reaction is:

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$$

The half-equation for the cathode reaction is:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

CC10a/SC10a Electrolysis Questions

Name the type of copper used at the anode.

Name the type of copper used at the cathode.

Name the electrolyte in the purification of copper.

State the half equations for the purification of copper.

CC10b/SC10b Products of Electrolysis

- Electrolysis of salts in solution
  - This involves both the electrolysis of the salt and the electrolysis of water.
  - The salt splits into its two component ions.
  - The water splits into Hydrogen ions (H^+) and Hydroxide Ions (OH^-).
  - To perform electrolysis you must have **INERT** (unreactive) electrodes as some of the products can be highly corrosive.

CC10b/SC10b Questions Products of Electrolysis

Write the half equations for the electrolysis of Lead Bromide.

Write the symbol and charge for a hydrogen Ion.

Write the symbol and charge for a Hydroxide Ion.

Name the 4 aqueous ions in salt water.

What is the overall equation for the electrolysis of Sodium chloride? And the half equations?

CC10b/SC10b Electrolysis of Sea Water (NaCl)

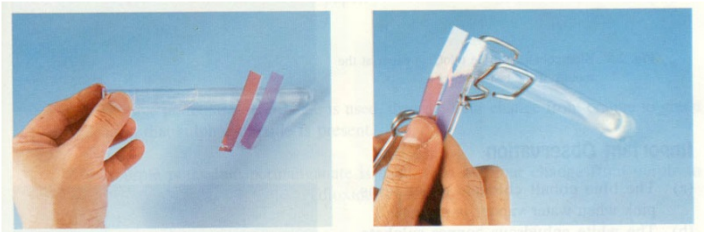
The most common dissolved substance in sea water is sodium chloride (common salt).

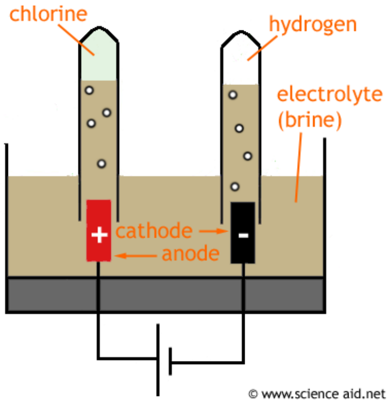
If a direct current is passed through sea water, chlorine gas is produced at one of the electrodes (hydrogen produced at the other.

Chemical test for chlorine:

Hold a piece of damp blue litmus paper in the mouth of the test tube

If the gas is chlorine the paper will first turn red and then turn white as it’s bleached





Uses of chlorine

Chlorine is a yellow-green gas with a pungent smell

Uses of chlorine:

- Kills microorganisms so is used to treat water supply (e.g. used in swimming pools)
- In manufacturing bleach and other cleaning products
- In the manufacture of plastics such as PVC

Chlorine has many uses mainly because of its high reactivity and ability to readily form compounds with other substances. However, chlorine gas is also toxic and if a gas leak occurred near a town, or a tanker transporting the gas had an accident, it could have devastating consequences

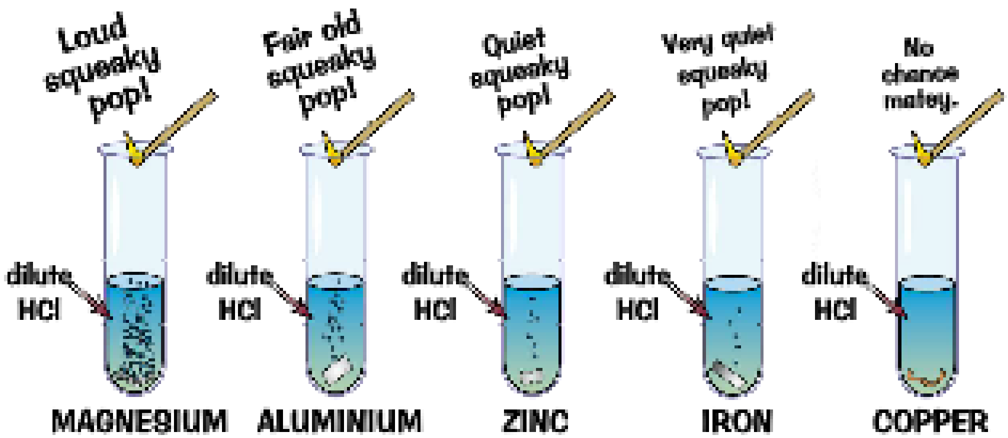
CC10b/SC10b Questions on Electrolysis of Sea Water

- What is the most abundant substance found in sea water?
- Name the two products formed during this electrolysis.
- How would you test for the gases produced?
- Give two uses for chlorine gas.
- What risks are associated with chlorine gas?

CC11a/SC11a Reactivity

How Metals React With Acids Tells You About Their Reactivity

1) The easier it is for a metal atom to lose its outer electrons and form a positive ion, the more reactive it will be.
2) Here's a classic experiment that you can do to show that some metals are more reactive than others. All you do is to place little pieces of various metals into dilute hydrochloric acid:



3) The more reactive the metal is, the faster the reaction will go.
4) Very reactive metals (e.g. magnesium) will fizz vigorously, less reactive metals (e.g. zinc) will bubble a bit, and unreactive metals (e.g. copper) will not react with dilute acids at all.
5) You can show that hydrogen is forming using the burning splint test The louder the squeaky pop, the more hydrogen has been made in the time period and the more reactive the metal is.
6) The speed of reaction is also indicated by the rate at which the bubbles of hydrogen are given off — the faster the bubbles form, the faster the reaction and the more reactive the metal.

You could also follow the rate of the reaction by using a gas syringe to measure the volume of gas given off at regular time intervals.

CC11a/SC11a Reactivity 2

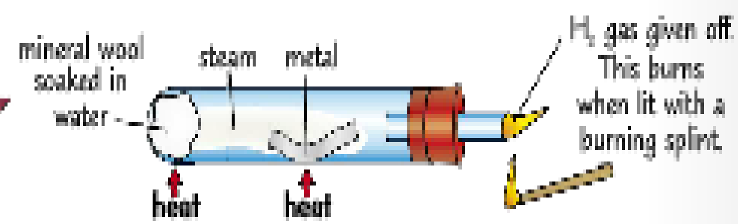
Metals Also React With Water

The reactions of metals with water also show the reactivity of metals. This is the basic reaction:

metal + water → metal hydroxide + hydrogen
(Or: less reactive metal + steam → metal oxide + hydrogen)

1) Very reactive metals like potassium, sodium, lithium and calcium will all react vigorously with water.

2) Less reactive metals like magnesium, zinc and iron won't react much with cold water, but they will react with steam. You could show this in the lab using this experiment:



3) Copper won't react with either water or steam.

CC11a/SC11a Reactivity 3

The Reactivity Series Shows How Reactive Metals Are

A reactivity series is just a table that lists metals in order of their reactivity. Here's an example:

The Reactivity Series		
Potassium	K	most reactive
Sodium	Na	
Calcium	Ca	
Magnesium	Mg	
Aluminium	Al	
Zinc	Zn	
Iron	Fe	
Copper	Cu	
Silver	Ag	least reactive

You can use displacement reactions to work out where in the reactivity series a metal should go.

Example: A student adds some metals to metal salt solutions and records whether any reactions happen. Use her table of results, below, to work out an order of reactivity for the metals.

	copper nitrate	magnesium chloride	zinc sulfate
copper	no reaction	no reaction	no reaction
magnesium	magnesium nitrate and copper formed	no reaction	magnesium sulfate and zinc formed
zinc	zinc nitrate and copper formed	no reaction	no reaction

- Magnesium displaces both copper and zinc, so it must be more reactive than both.
- Copper is displaced by both magnesium and zinc, so it must be less reactive than both.
- Zinc can displace copper, but not magnesium, so it must go between them.

The order of reactivity, from most to least, is:
magnesium, zinc, copper

CC11a/SC11a Reactivity Questions

- Write the word equation for the reaction of calcium with water.
- Write the balanced equation for the reaction of magnesium with dilute hydrochloric acid. Include state symbols.
- Tin sits between iron and copper in the reactivity series. State whether tin would displace zinc from zinc sulphate solution and explain your answer

CC11b/SC11b Ores

Some Metals are found as pure elements

e.g gold and platinum – are found naturally in the Earth's crust as elements. They are found as uncombined elements because they are ***very unreactive*** and don't form compounds. A physical method can be used to collect them such as sieving or panning.

Most Metals are found as compounds

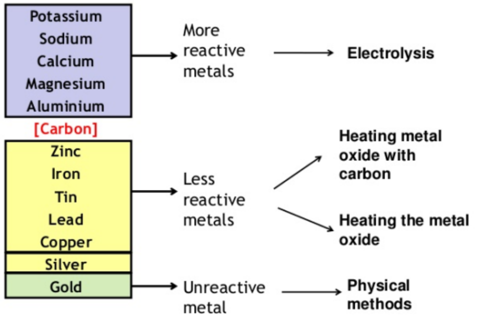
However, most metals are reactive and readily form compounds (the more reactive a metal is, the more easily it reacts with other substances to form compounds). The compounds formed are mainly metal oxides which are found in rocks called ores in the Earth's crust as metal ***ores.***

To obtain these metals they must be extracted from ores in the Earth's crust

Examples

- Haematite (iron ore) contains enough iron oxide for it to be profitable to extract iron from it
- Similarly, aluminium is extracted from bauxite (aluminium ore) and copper from malachite (copper ore)

Summary



Extracting metals from ores

Some metals can be extracted by heating their compounds with carbon:

e.g iron: iron oxide + carbon → iron + carbon dioxide

Chemical equation: $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$

Other metals are extracted from their compounds by electrolysis of a molten compound:

e.g aluminium: aluminium oxide → aluminium + oxygen

Chemical equation: $2Al_2O_3 \rightarrow 4Al + 3O_2$

The method used to extract a metal depends on its reactivity. The less reactive metals are extracted using carbon (e.g iron, zinc, lead, copper). More reactive metals are more difficult to extract because they form stronger, more stable compounds. The more reactive metals need to be extracted by electrolysis (e.g sodium, calcium, magnesium, aluminium) – a more powerful method. Metals with very low reactivity (e.g. gold) are found as elements and don't need to be extracted.

Extraction using electrolysis is more expensive than with carbon due to the cost of electricity and the more reactive the metal, the harder and more expensive it is to extract from its ore

CC11b/SC11b Questions on Ores

- Why are gold and platinum found as pure elements?
- What is an ore? Name an ore of iron and copper.
- Why can you extract zinc using carbon but to extract aluminium from its ore you have to use electrolysis?
- Write the word equation for the extraction of iron from iron oxide using carbon. Can you write the symbol equation?
- Why would scientists prefer not to use electrolysis?

CC11c/SC11c OXIDATION AND REDUCTION

Oxidation reactions: gain of oxygen

Reduction reactions: loss of oxygen

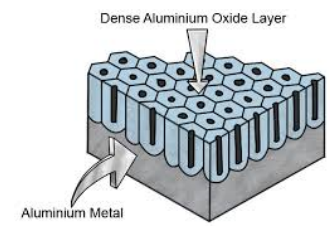
Corrosion of metals is oxidation

Corrosion happens when the surface of a metal changes by reaction with oxygen (sometimes with water)**.** e.g when iron corrodes it forms rust which is iron oxide.

Most metals corrode. The more reactive the metal, the more readily it becomes oxidised and the more rapidly it corrodes

Exception : aluminium!

- Aluminium doesn’t corrode as much as expected because upon reaction with oxygen it forms aluminium oxide, which acts as a protective layer and prevents any further corrosion.
- Less reactive metals are more resistant to oxidation and corrode less
- Very unreactive metals e.g. gold - don’t corrode at all



Metal extraction is reduction

The majority of the compounds found in ores (from which metals are extracted) are metal oxides.

In order to obtain metals from their oxides, oxygen is removed and the process of extracting metals from their ores is a reduction reaction (the metal oxides lose their oxygen and are therefore ‘reduced’)

e.g. iron oxide + carbon → iron + carbon dioxide

In this example, the iron oxide is reduced to iron (it has lost its oxygen)

CC11b/SC11b Questions on Oxidation and Reduction

- What type of reaction involves gaining oxygen?
- What type of reaction involves losing oxygen?
- Which of the above is the corrosion of metals?
- Why doesn’t aluminium corrode?
- Is extracting metals from their ores a reduction or oxidation?

CC11d/SC11d Lifecycle Assessment and Recycling

Tonnes of waste is produced every year, much of it ending up buried in the ground in landfill sites. Problems with landfill sites:

- Use up a lot of land
- Risk of pollution
- Materials can't be used again (this is a problem because we are running out of some raw materials completely e.g zinc)

Recycling is the process of taking materials out of waste before disposal and converting them into new products we can use.

Recycling can reduce demand for resources and also reduce the problem of waste disposal.

- Materials that can be recycled:Metals in drinks cans can be melted down and recycled as new drinks cans or part of a car
- Paper can be recycled as more paper or cardboard
- Plastic bottles can be recycled as fleece clothing

Life-cycle Assessments Show Total Environmental Costs

A life-cycle assessment (LCA) looks at each stage of the life of a product — from making the material from natural raw materials, making the product from the material, using the product and disposing of the product. It works out the potential environmental impact of each stage.

CC11d/SC1d Questions on Life Cycle Assessment and Recycling

- Give three problems associated with landfill sites.
- Define recycling.
- Describe two advantages of recycling.
- Give two examples of materials that can be recycled.
- Which material can be recycled to form fleece clothing?

CC11d/SC11d Life Cycle Assessments and Recycling 2

Metals can be melted down and made into something new –recycling.

Advantages of recycling

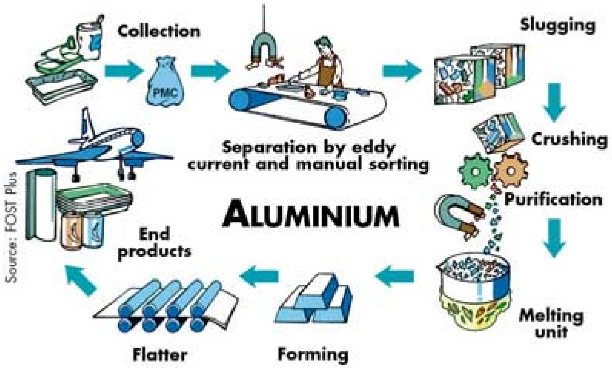
Natural reserves of metal ores will last longer

For most metals, less energy (and therefore less expense) is needed to recycle them than to extract them from their ores.

Recycling reduces the need to mine ores (mining can damage the landscape and create dust and noise pollution in the same way as limestone quarrying.

It leads to less pollution as extracting some metals produces greenhouse gases.

e.g. sulphur dioxide is formed when lead is extracted from its ore, galena



Disadvantages of recycling

Costs and energy used in collecting, sorting and transporting metals to be recycled

For some metals, it is more expensive to recycle them than to extract them from their

How metals are recycled

Iron and steel are easily separated from other metals as they're magnetic. Others are separated by hand which is labour intensive.

PROPERTIES OF METALS

Shiny, conduct heat, conduct electricity, malleable (can be hammered into shape), ductile (can be stretched into wires.) Different metals, though, have slightly different properties and have different uses.

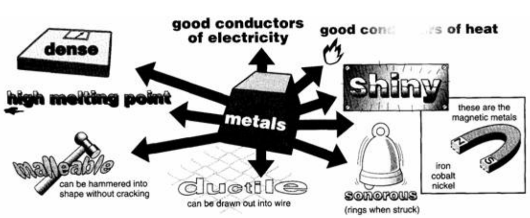
Aluminium: Has a low density and doesn't corrode as a protective layer of aluminium oxide forms quickly on its surface

It is used to make aeroplanes as it has such low density (the lighter the aircraft, the less fuel it needs to fly).

Copper: Good electrical conductor so copper is used to make electrical cables. It has low reactivity and doesn't react with water so is also used in water pipes.

Gold: Very unreactive, doesn't corrode, easily worked into shapes so can be used for jewellery. It's also a very good electrical conductor and tiny amounts are used inside most electronic devices, including mobile phones and computers.

Iron and steel: Iron is fairly cheap to extract from iron ore by heating with carbon, however, pure iron is too soft and it is often made into steel (a mixture of iron, carbon a nd other metals), which is stronger and harder.



CC11d/Sc11d Questions on Life Cycle Assessment and Recycling 2

- Define recycling.
- Name two advantages of recycling metals.
- Explain two disadvantages of recycling metals.
- How are metals separated during the recycling process?
- What do ductile and malleable mean?
- What property of aluminium makes it suitable for building aeroplanes?
- If gold is an excellent conductor why aren't all wires made of gold?

CC12a/SC12a Dynamic Equilibrium

nitrogen + hydrogen $\rightleftharpoons$ ammonia

This means that it can be considered as two reactions:

Forward: nitrogen + hydrogen $\rightarrow$ ammonia

Backward: ammonia $\rightarrow$ nitrogen + hydrogen

- When **AMMONIA** is formed it releases heat (**EXOTHERMIC**). This is the **FORWARDS** reaction.
- The reverse reaction will be the opposite which makes it **ENDOTHERMIC** (takes in heat)
- When **DYNAMIC EQUILLIBRIUM** is reached these two reactions occur at the **SAME RATE**.
- Adjusting the temperature and pressure will affect the position of the equilibrium, favour the **PRODUCT** or **REACTANT**.

CC12a/SC12a Dynamic Equilibrium 2

$$N_2 + 3\ H_2 \rightleftharpoons 2\ NH_3$$

- Reactant = 4 molecules
- Products = 2 molecules

Pressure and the Haber process

- If you increased the pressure of the reaction the equilibrium would favour the **PRODUCTS (move to the right)**. This is because the particles are being forced closer together and therefore more likely to react.

Temperature and the Haber process

- As the reaction is **EXOTHERMIC** it favours cooler conditions (it releases energy into the surrounding environment).
- An increase in temperature would move the equilibrium to the **left (favour the reactants)**.
- A low temperature would increase the yield but slow the rate of reaction.

CC12a/SC12a Dynamic Equilibrium 3

- Optimal conditions are used to ensure that the maximum possible yield is produced safely and at a sufficient rate to be economically viable.

- Temperature – approx. 450OC
- Pressure – 200atm (200 times atmospheric)
- Catalyst – Iron catalyst

- A catalyst increases the rate of reaction without ever being used in the reaction. It works by lowering the activation energy for the reaction (energy required for a successful collision)

- If the temperature or pressure is too high then there can be safety implications and too low will result in a lower yield.

CC12a/SC12a Dynamic Equilibrium Questions

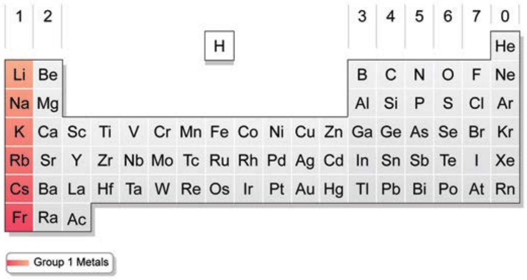
- What is happening when a reaction reaches ‘dynamic equilibrium’?
- What happens to the amount of ammonia after equilibrium is reached in the Haber process?
- Explain how pressure affects the yield in the Haber process.
- Explain how a high and a low temperature affect the yield of ammonia.
- State the temperature of the Haber process.
- State the atmospheres needed for the Haber process.
- Name the catalyst used.
- Explain the role of a catalyst.

CC13a/SC17a Group 1

The alkali metals are found in group 1 of the periodic table so they have 1 electron in their outer shell.

To gain a full outer shell they must lose their outer electron so they form ions with a charge of +1

The atoms in alkali metals are held together by metallic bonding and are solids at room temperature, but they have low melting points compared to other metals.



Alkali metals are soft metals and be cut with a knife.

All alkali metals react with water to form a metal hydroxide and hydrogen gas:

lithium + water → lithium hydroxide + hydrogen

$2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2\,(g)}$

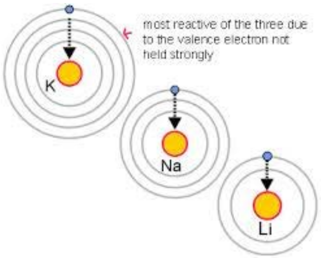
Note: all metal hydroxides are alkaline

Reactivity

If lithium is dropped in water, it just floats on the water and fizzes (that's the hydrogen being produced) until the reaction is finished. Sodium reacts more strongly with water. Sodium has a lower melting point and reaction produces enough heat to melt the metal so a molten ball of sodium that whizzes around the surface of the water (releasing hydrogen) until the reaction is finished.

Potassium reacts even more strongly and the hydrogen produced during the reaction catches fire, producing a lilac flame so the reactivity of alkali metals increases as you go down group 1.The elements at the bottom of group 1 have more electrons than the elements at the top of the group so they have more electron shells. The electron in the outer shell is further from the nucleus The attraction between the nucleus and the outer electron is weaker when charges are further apart..

- outer electron in a potassium atom (configuration: 2.8.8.1) is not held as strongly as the outer electron in a lithium atom (configuration: 2.1)
- Metals react by losing their outer electron, forming ions with a +1 charge and a full outer shell
- Potassium loses its outer electron more easily and it's more reactive



CC13a/SC17a Questions on Group 1

Explain why alkali metals become ions with a +1 charge.

Give two properties of alkali metals that are similar to other metals and two properties that are not typical.

Describe what you would see as you added Lithium, Sodium then Potassium to water.

Explain these observations using the structure of the metals.

CC13b&c/SC17b&c Group 7 and Halogen Reactivity

Group 7 so have 7 electrons in their outer shell, to form a full outer shell they gain 1 electron, form ions with a -1 charge.

At room temperature fluorine is a pale yellow gas, Chlorine is a yellow-green gas, Bromine is a brown liquid, Iodine is a grey solid.

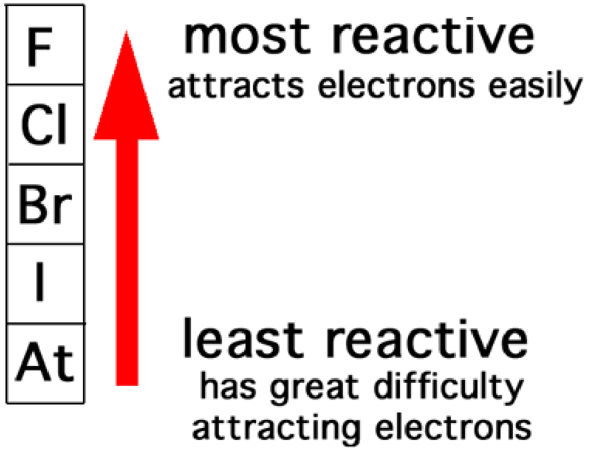
Reactivity: less reactive as you go down group 7, because they react by gaining an electron - this is easier with fewer electron shells (because outer electrons are closer to the nucleus), so fluorine is the most reactive halogen

Halogens React With metals to form metal halides e.g. sodium + chlorine → sodium chloride

Chlorine→'chloride', fluorine→'fluoride', Iodine→'iodide'

Halogens React With hydrogen: react with hydrogen gas to form hydrogen halides (they form acids when dissolved in water (when hydrogen chloride dissolves in water it forms hydrochloric acid))

Displacement reactions: more reactive halogens can 'displace' less reactive halogens from their compounds. Less reactive halogens cannot 'displace' more reactive halogens from their compounds.



CC13b&c/SC17b&c Group 7 and Halogen Reactivity Questions

- How many electrons do halogens have in their outer shell?
- What charge will the halide ion have?
- As you go down group 7 what happens to the reactivity. Explain why.
- Give a word equation for a halogen reacting with a metal and reacting with hydrogen.
- What is a displacement reaction? What would you see if the following reacted together?

$$KI + Cl_2 \rightarrow$$

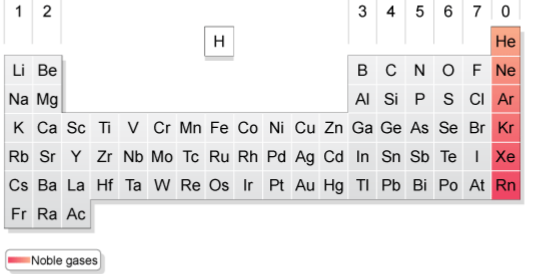
CC13d/SC17d Group 0

Group 0 or 8 of the periodic table. They already have full outer shells, compared to other elements so do not need to gain or lose electrons so do not react – they are inert (very unreactive)

Properties: All are gases at room temperature, low boiling points, increase down the group. Particles in a gas are spread far apart, low densities, decreasing down the group

Discovery: Chemists noticed that the density of pure nitrogen made in chemical reactions was less than the density of nitrogen extracted from the air. Experiments were done to find the identity of this dense gas –argon. The other noble gases were discovered soon after.

Uses: useful because unreactive. When electric current is passed through a tube filled with neon under low pressure, coloured light is produced - used in fluorescent lamps and advertising displays



CC13d/SC17d Questions on Group 0

- What group are the noble gases?
- Why are noble gases unreactive?
- Name some properties of the noble gases.
- Give a use for noble gases and explain why they can be used in this way.

CC14a7b/SC18a&b RATES OF REACTION

The rate of a chemical reaction is the speed at which it takes place – or how fast reactants are used up and products are made. In an explosion very fast chemical (exothermic) reactions release lots of heat energy and gas in a short space of time. In the rusting of iron is a slow chemical reaction occurs.

Collision theory

For a reaction to occur, particles of the reactants must collide with each other and if particles collide more frequently, rate of reaction will increase

Not all collisions lead to a reaction – particles must collide with enough energy to break the bonds in the reactants. The more energy particles have, the greater the chance of 'successful collisions'. The rate of reaction will increase if there are more high-energy collisions between particles.

The effect of temperature

The higher the temperature of the reactants, the faster the reaction - e.g eggs cook faster in boiling water than in warm water. As temperature is increased, particles have more energy, they move faster, they have a greater chance of collision and when they do collide there is a greater chance of successful collisions.

Increase in temperature →increased rate of reaction

Sometimes we cool reactions to slow them down e.g some foods are put in a fridge to slow down chemical reactions that make food go off.

The effect of concentration:

The higher the concentration of a reactant in a solution, the faster the reaction. The more concentrated a solution, the more solute particles there are in a given volume therefore it is more likely that reactant particles will collide with one another

Increase in temperature →increased rate of reaction

The effect of surface area:

The greater the surface area of a solid, the faster the rate of reaction. Increasing the surface area increases the number of particles exposed on the surface (that can collide and react) more frequent collisions between reactant particles and hence an increased rate of reaction.

To increase surface area, solids are crushed into lots of small pieces e.g. in power stations, coal is ground up into a fine powder to help it burn faster

CC14a7b/SC18a&b Questions on Rates of Reaction

- What is the collision theory?
- Do all particles that collide react?
- What is the effect on the rate of reaction of increasing temperature?
- What is the effect on rate of increasing concentration?
- Which has the biggest surface area – powdered calcium carbonate or lumps of calcium carbonate?
- What effect does decreasing the surface area have on the rate of reaction?

CC14c/SC18c Catalysts and activation energy

Catalysts are substances that speed up chemical reactions without being used up in the reactions (note: they are neither reactants nor products)

Catalysts speed up chemical reactions by increasing the probability of successful collisions by providing an alternative pathway for the reaction to occur - i.e when particles collide they don’t need as much energy to react

Many chemical processes use catalysts to increase the rate at which products are generated which means that reactions can be done at lower temperatures and pressures than they would otherwise so less energy is used to heat the reactants and it saves money.

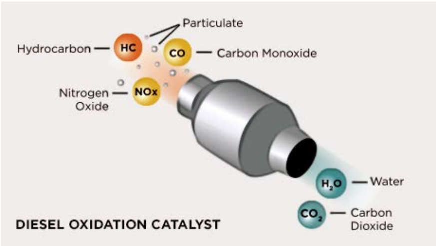
Catalytic converters

The combustion of petrol in car engines produces carbon monoxide (toxic) and unburned hydrocarbons.

Cars are now built with catalytic converters which help to combine carbon monoxide and unburned petrol with oxygen from the air to form carbon dioxide and water vapour.

This reduces pollutants in exhaust gases. Catalytic converters contain the transition metals platinum, rhodium or palladium (all very expensive), which act as catalysts. The catalysts have a high surface area (‘honeycomb structure) so the rate at which carbon monoxide and unburned petrol react with oxygen from the air to form carbon dioxide and water vapour is increased.

Catalytic converters work best at high temperatures (as particles collide more frequently and with more energy). When a car engine is first started, a catalytic converter is cool and doesn’t work well. However, the hot gases from the engine quickly heat it up.



CC14c/SC18c Questions on Catalysts and activation energy

- What is a catalyst?
- How does it work?
- Why does it save money?
- What toxic gases are emitted from a car exhaust?
- What gases are these converted into?
- Why does a large surface area (honeycomb structure) increase the efficiency of a catalyst?

CC15a/SC19a Exothermic and endothermic reactions

- In chemical reactions heat energy transfers between the reactants and the surroundings

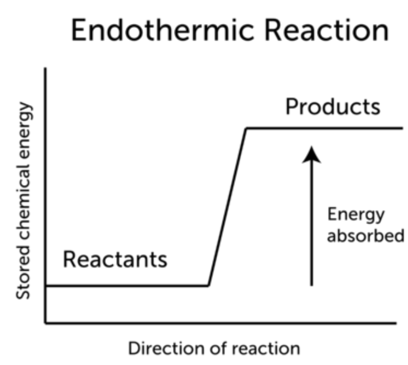
Exothermic reactions

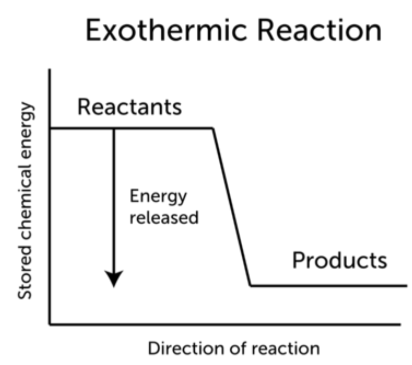
- Exothermic reactions give out heat energy to the surrounding which is shown by the temperature of the reaction mixture and its surroundings increasing.
- Most reactions are exothermic. Examples of exothermic reactions include all combustion reactions and explosions which release their heat energy quickly. e.g combustion of methane:

  methane + oxygen → carbon dioxide + water

Endothermic reactions

- Endothermic reactions take in heat energy from the surroundings so the temperature of the reaction mixture and of the surroundings decreases
- Few chemical reactions are endothermic but examples of endothermic reactions are:
  - Reaction of sodium hydrogencarbonate with hydrochloric acid (if you touch the test tube it feels cold because heat energy is taken from your hand)
  - Dissolving ammonium nitrate in water
  - Photosynthesis (takes in heat energy from the sun)





CC15a/SC19a Questions on Exothermic and endothermic reactions

- What is an exothermic reaction?
- Draw the energy diagram for an exothermic reaction?
- What will a thermometer do if it is in an exothermic reaction?
- Give an example of an exothermic reaction.
- Now answer all the questions above for an endothermic reaction.

CC15b/SC19b Energy changes in reactions

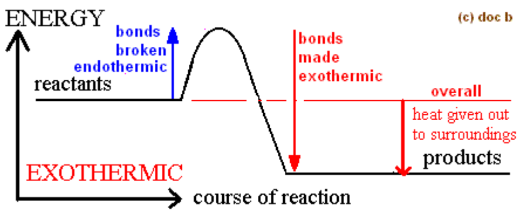
In chemical reactions, bonds in the reactants are broken and new bonds are formed to make the products.

Breaking bonds requires energy so it's an endothermic process (+)

Making bonds releases energy so it's an exothermic process (-)

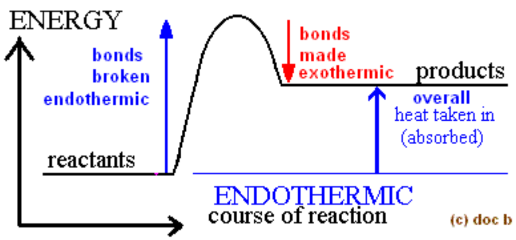
In an exothermic reaction (-)

More heat energy is released making bonds (in the products) than is required to break bonds (in the reactants) so the net result is that heat energy is released (-)



In an endothermic reaction (+)

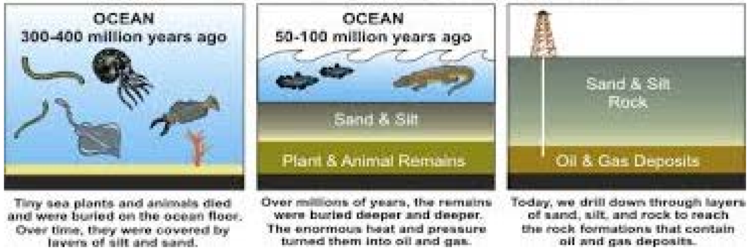
- More heat energy is required to break bonds (in the reactants) than is released making bonds (in the products) so the net result is that heat energy is taken in (+)
- These energy changes can be shown using diagrams:



CC15b/SC19b Energy changes in reactions Questions

- Chloe says in exothermic reactions bonds are formed. Is she correct?
- Does Bond Breaking require energy to go in or be taken out of the reaction?
- Does Bond Making require energy to go in or be taken out of the reaction?
- If a reaction is overall exothermic and gives out energy what does this tell you about the energy released and absorbed during the bond making and breaking process?

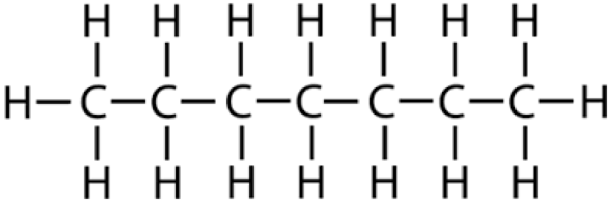
CC16a/SC20a Hydrocarbons in crude oil and natural gas



Crude oil and natural gas are fossil fuels found in some sedimentary rocks deep underground

Formation of crude oil/natural gas

- When microscopic plants and animals in the sea die, their remains fall to the bottom of the sea
- Over millions of years, their remains are buried by sediments, stopping entry of oxygen and preventing their decay
- As more sediments build up on top of the remains, the heat and pressure increases and crude oil or natural gas are gradually formed



Composition of crude oil

A hydrocarbon is a compound that contains only carbon and hydrogen atoms.

Hydrocarbons with many carbon atoms are said to have 'long' carbon chains.

Hydrocarbons with few carbon atoms are said to have 'short' carbon chains.

Crude oil is a mixture of different hydrocarbon molecules (i.e some with long carbon chains and others with short carbon chains). Crude oil can also contain some impurities like sulphur.

Crude oil is a non-renewable resource:

The rate at which we are using crude oil is much greater than that at which it is being formed underground. At the present rate of consumption, the supply of crude oil is estimated to run out in the next 40 to 50 years

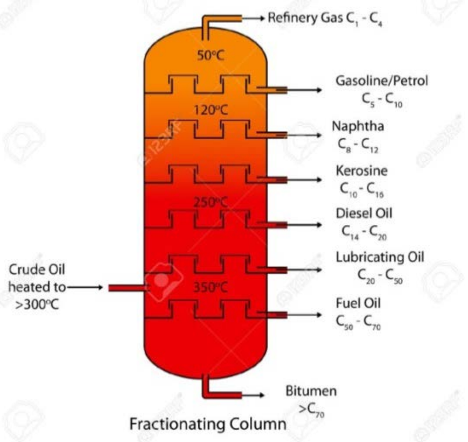
Currently most oil is obtained from drilling wells, but as demand for oil increases, surface mining of oil sands (which contain over half of the world's remaining reserves of oil) will increase.

- But is it economic to mine these sands?
- Is it worth the damage to habitats?

CC16a/SC20a Hydrocarbons in crude oil and natural gas questions

- Describe how crude oil is formed.
- Which two elements are found in a hydrocarbon?
- Is crude oil renewable or non-renewable?
- Draw the hydrocarbon that has 7 carbons and 16 Hydrogen atoms.

CC16b/SC20b Fractional distillation of crude oil



Fractional distillation

- The mixture of hydrocarbons in crude oil needs to be separated into simpler mixtures (called 'fractions') in a process called fractional distillation.
- The mixture of liquids is boiled in a fractional distillation tower and the vapour from it is condensed.
- The fractional distillation tower is hot at the bottom and cooler near the top.
- Different liquids condense at different temperatures so separates the mixture into smaller fractions of crude oil (bitumen, kerosene, petrol, diesel oil).
- Fractions with short carbon chains (e.g. gases):
  - Ignite (set alight) easily, have low boiling points and have low viscosity (they are runny) when in liquid form and condense at the top of the fractionating column
- Fractions with long carbon chains (up to 40 carbons e.g bitumen):
  Have much higher boiling points, are harder to ignite and have high viscosity (they are thick and sticky) when in liquid form and condense at the bottom of the fractionating column
- Order of fractions according to carbon chain length (from shortest to longest):

Fraction	Length of molecule	Ease of ignition	Boiling point	Viscosity
gases	short carbon chains (only a few carbon atoms)	easy	low (< 0°C)	runny
petrol	↓	↓	↓	↓
kerosene				
diesel oil				
fuel oil	long carbon chains			
bitumen	(up to 40 Carbon atoms)	difficult	high (> 350°C)	thick and sticky

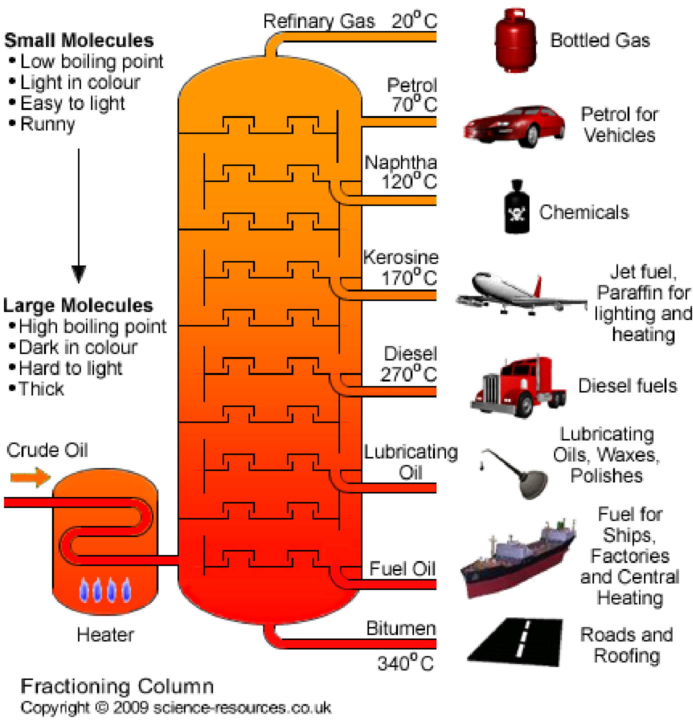
CC16b/SC20b Fractional distillation of crude oil questions

- What are fractions of crude oil?
- Is it hotter at the top or bottom of the fractionating column?
- Do shorter chains travel to the top of the column or the bottom?
- Describe the properties of shorter chain alkanes?
- Describe the properties of longer chain alkanes?

CC16b/SC20b Fractional distillation of crude oil 2

The different fractions of crude oil have different properties so have different uses.

- Gases (e.g methane): fuel for vehicles, bottled gas for camping stoves, heating and cooking in homes
- Petrol: fuel for cars
- Kerosene: fuel for aircraft
- Diesel oil: fuel for diesel engines (some cars, lorries, trains)
- Fuel oil: fuel for large ships and some power stations, fuel for heating.
- Bitumen: making roads, waterproofing flat roofs



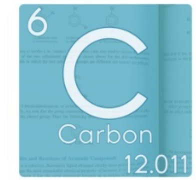
CC16b/SC20b Fractional distillation of crude oil questions 2

Give a use of:

- Gases
- Petrol
- Kerosene
- Diesel Oil
- Fuel Oil
- Bitumen

CC16c/SC20c The alkane homologous series

Carbon is in group 4 of the periodic table which means it has four electrons in the outer shell so needs four more to gain a full shell. This means each carbon forms four bonds with other atoms (either 4 single bonds, or 2 single bonds and a double bond).



Hydrogen is in group 1 so forms just one bond with other atoms.

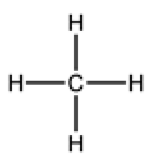
Alkanes

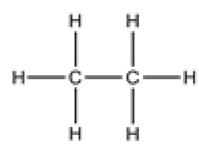
In an alkane molecule, each carbon is bonded to four other atoms with single bonds (see diagrams below).

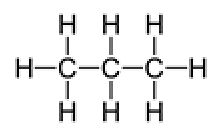
Hydrocarbons with single carbon-carbon (C-C) bonds are referred to as ‘saturated’.

Alkanes are known as saturated hydrocarbons (‘saturated’ because they can’t form bonds with any more atoms). Alkanes follow the general formula:

$$C_nH_{2n+2}$$







Methane (CH_4) is the simplest alkane, with one carbon atom joined to four hydrogen atoms

Ethane (C_2H_6) has two carbon atoms

Propane (C_3H_8) has three carbon atoms.

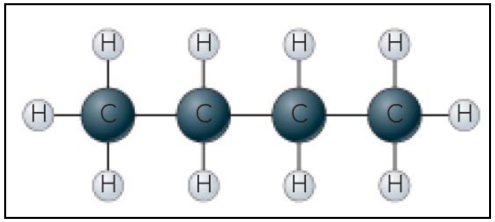
CC16c/SC20c The alkane homologous series questions

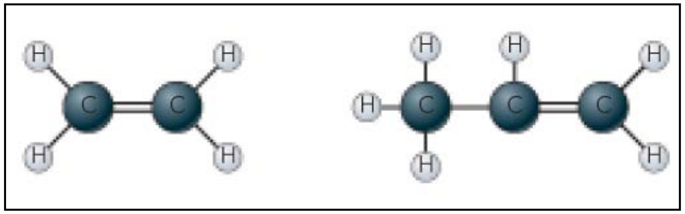
- What group is carbon in? How many electrons does it have in its outer shell? How many bonds does it try to form?
- What is the general formula for an alkane? What would the formula of hexane be if it contains 6 carbon atoms?
- What does saturated mean?
- Draw the displayed formula for octane.

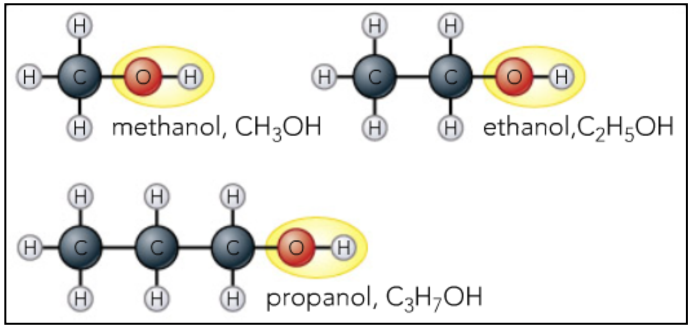
CC16c/SC20c The alkane homologous series 2

A HOMOLOGOUS series is a series of compounds that have the same general formula and similar chemical properties but have variation in boiling points.

- **Alkanes** – a hydrocarbon containing only
  C-C bonds.
- **Alkenes** – a hydrocarbon containing at least one **C=C bond**.
- **Alcohol** – a hydrocarbon containing at least one **–O-H group**.







CC16c/SC20c The alkane homologous series 2

- Define the term homologous.
- Describe the type of bonds in an alkane.
- Describe the bonds in an alkene.
- State the group found in alcohols.

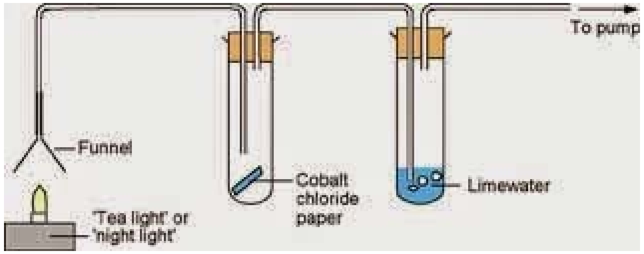
CC16d/SC20d Complete and Incomplete Combustion

When hydrocarbon fuels burn they react with oxygen and release heat and light energy. This is an oxidation reaction called combustion. When enough oxygen is present, all the hydrocarbon is used up and the only products are carbon dioxide and water. This is known as 'complete combustion'

e.g combustion of methane (the main gas in natural gas):

Methane + oxygen → carbon dioxide + water

Balanced equation: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$



Detecting products of a complete combustion reaction:

Anhydrous copper sulphate turns blue when it comes into contact with water and Limewater turns cloudy when carbon dioxide is bubbled through it. So when hydrocarbons are burnt fully in air (i.e when complete combustion occurs), anhydrous copper sulphate turns blue and limewater turns cloudy

INCOMPLETE COMBUSTION

Sometimes a burning fuel may not have enough oxygen. In this case, 'incomplete combustion' occurs which means there isn't enough oxygen for all the carbon atoms to form carbon dioxide (as each carbon must combine with 2 oxygen atoms) and some of the carbon may form carbon monoxide (CO – 1 carbon combines with 1 oxygen atom) and/or solid particles of carbon (C – no oxygen combined with carbon atom).

2 possible reactions occur when hydrocarbon fuels burn without enough oxygen.

e.g with methane:

Methane + oxygen → carbon monoxide + water — less oxygen

Methane + oxygen → carbon + water — very little oxygen

Bunsen burners and combustion

When the air hole of a Bunsen burner is open, complete combustion occurs and the flame is blue so is often called a 'clean' flame.

When the air hole of a Bunsen burner is closed there isn't enough oxygen so incomplete combustion occurs and flame is yellow. The yellow colour is caused by the hot carbon particles (soot) glowing.

Carbon monoxide problems

Carbon monoxide is an odourless, colourless toxic gas andiIt reduces the amount of oxygen that can be transported around the body in the blood.

Faulty gas boilers (in which oxygen flow is restricted) and fires produce carbon monoxide and this can lead to death by carbon monoxide poisoning so it is essential that all fuel-burning appliances must be serviced regularly and homes should be fitted with carbon monoxide detectors.

Soot (carbon) problems

- Soot produced in appliances such as boilers can clog up pipes carrying waste gases away
- Soot is also produced by vehicles. Breathing in sooty air can lead to lung disease
- Soot also leaves black marks on buildings/walls

CC16d/SC20d Questions on Complete and Incomplete Combustion

- When a hydrocarbon combusts completely what are the two products formed?
- Which tests will prove that the above products have been formed?
- During incomplete combustion which gas is less abundant than during complete combustion?
- What are the products of incomplete combustion when there is a little less oxygen and when there is a lot less oxygen?
- Which of the products is dangerous?
- How does a Bunsen burner demonstrate incomplete and complete combustion and how are the flame colours linked to the type of combustion?

CC16e/SC20e Combustible fuels and pollution

Discovering the problem

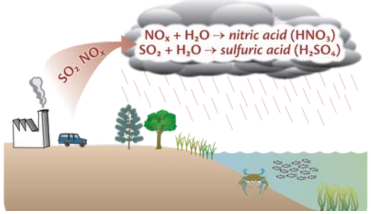
In the 1970s, the numbers of fish caught in lakes and rivers in southern Norway started to decrease and Scientists noticed that these lakes and rivers were much more acidic than those in other parts of the country that still had healthy fish.

Looking at weather patterns, scientists concluded that pollution from factories and power stations in Europe was being carried in the atmosphere, making the rainfall (more) acidic and this acid rain was killing the fish.

Causes of acid rain

Acid rain is rain that is more acidic than usual (has a pH of less than 5.2)

Hydrocarbon fuels contain sulphur impurities so when fuels are burnt, the sulphur reacts with the oxygen from the air to form sulphur dioxide gas. Sulphur dioxide dissolves in rainwater and lowers its pH, forming acid rain. Nitrogen oxides from cars are also released into the air.



Effects of acid rain

Makes rivers, lakes and soils acidic and harms organisms that live there.

- Damages trees
- Speeds up the weathering of buildings/statues made of limestone or marble and the corrosion of metal

Solutions to problem of acid rain:

- Reducing amount of sulphur in petrol, diesel and fuel oil
- Removing acidic gases from power station emissions (by neutralising them with a basic (alkaline) compound such as calcium carbonate)

CC16e/SC20e Combustible fuels and pollution questions

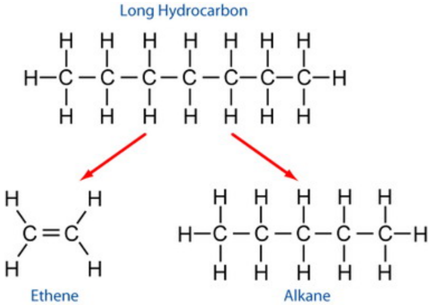
- What did Norwegian Scientists notice about the lakes in Norway?
- How did they explain what they had observed?
- How does acid rain form?
- State two effects of acid rain.
- Give details of two methods that are being used to solve the problem of acid rain.

CC16f/SC20f Breaking down hydrocarbons

Longer hydrocarbons can be broken down by heat ('thermally decomposed') into more useful shorter hydrocarbons by a process called cracking. In most cases, a long chain alkane is thermally decomposed into a shorter chain alkane and an alkene.

$C_6H_{14} \rightarrow C_4H_{10} + C_2H_4$

Note: there are still the same numbers of carbon and hydrogen atoms before and after cracking



Shorter chain alkanes produced in cracking are used as fuels (e.g petrol). Most of the short chain alkenes produced in cracking are used for making plastics. Ethene is also used for ripening fruit

Why is cracking needed?

- When crude oil is separated by fractional distillation, some fractions are present in greater amounts than others (e.g more fuel oil is present than petrol or diesel).

- The shorter fractions in crude oil (petrol, diesel) are in greater demand than the longer fractions (fuel oil, bitumen). To make supply meet demand, oil companies use cracking to break down longer molecules into more useful shorter molecules

CC16f/SC20f Breaking down hydrocarbons questions

- What happens during cracking?
- Why is this also known as a thermal decomposition?
- If $C_{12}H_{26}$ is cracked give 6 possible products that might form.
- What are the alkenes that are made used for?
- How is cracking related to supply and demand?

CC17a/SC21a THE EARLY ATMOSPHERE

The Young Earth

Evolution of life on Earth has caused the Earth's atmosphere to change. When looking for clues into the composition of the Earth's early atmosphere, scientists study planets and moons (whose atmosphere has not changed for millions of years). In particular, scientists study volcanoes because they release lots of gases

Conflicting evidence

The atmospheres of Mars and Venus are mainly carbon dioxide. The atmosphere of Titan, one of Saturn's moons, is 98% nitrogen.

Space probes have shown that Titan has an icy interior rather than a rocky one (like Earth, Mars and Venus) which makes it more likely that the Earth's early atmosphere resembled that of Mars or Venus (i.e that it contained lots of carbon dioxide).

Oxygen

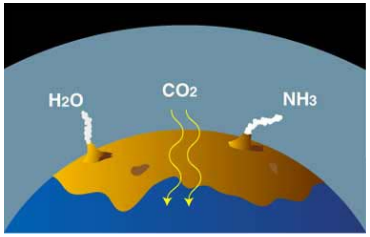
It's known what there was little or no oxygen in the Earth's early atmosphere

Evidence

The young Earth had lots of active volcanoes releasing ammonia (NH_3) and water. Volcanoes don't release oxygen. Iron compounds found in the Earth's oldest rocks could only form in the absence of oxygen.

The Earth's early atmosphere is generally considered to have lots of carbon dioxide and little oxygen.

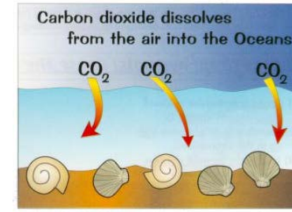
It also contained water vapour and other gases (e.g methane and ammonia).



CC17a/SC21a Questions on the Early Atmosphere

- Where do scientists look for clues about the Earth's early atmosphere?
- What fact has made scientists think that the Earth's early atmosphere was not the same as Titan's?
- How much oxygen was there in the Earth's early atmosphere?
- Where did the ammonia (NH_3)come from in the Earth's early atmosphere?

CC17b/SC21b The Changing Atmosphere



The Oceans

As the Earth became older, it cooled, and the water vapour in the atmosphere condensed into liquid water, forming the oceans.

1. Carbon dioxide dissolves in the oceans
2. Some marine organisms (e.g. coral) use carbon dioxide to make shells of calcium carbonate (over time these formed sedimentary rocks)

These mechanisms reduced the amount of carbon dioxide in the atmosphere.

Photosynthesis

- Around 1 billion years ago some organisms developed the ability to photosynthesise which involves taking in carbon dioxide and releasing oxygen into the atmosphere.
- Over time, more and more photosynthesising organisms evolved (e.g. plants) which further reduced the levels of carbon dioxide in the atmosphere and increased levels of oxygen in the atmosphere.

$$6CO_2 + 6H_2O \xrightarrow[\text{chlorophyll}]{\text{light energy}} C_6H_{12}O_6 + 6O_2$$

CC17b/SC21b Questions on the Changing Atmosphere

- How did the oceans form?
- Name the gas that dissolved in the oceans.
- What did marine organisms do with this gas?
- Which process started happening about a billion years ago?
- What effect did this process have on the early atmosphere?
- Write down the symbol equation for photosynthesis.

CC17c/SC21c THE ATMOSPHERE TODAY

The composition of the atmosphere

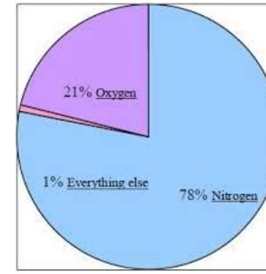
Nitrogen: 78%.

oxygen: 21%.

argon: 0.9%.

carbon dioxide: 0.04%

water vapour – variable from day to day so not included



There are also traces of other gases in the atmosphere e.g. nitrogen oxides, carbon monoxide, methane and sulphur dioxide. The amounts of these gases in the atmosphere can vary as they can be changed by natural causes or by human activities.

Natural causes

Volcanoes release carbon dioxide and sulphur dioxide

Lightning can produce nitrogen oxide

Human Activities

Deforestation: removal of trees so less photosynthesis and more carbon dioxide in the atmosphere.

Burning fossil fuels increases the amounts of carbon dioxide, carbon monoxide and sulphur dioxide in the atmosphere (harmful!)

Engines and furnaces release nitrogen oxides

Cattle and rice fields release large quantities of methane

Formation of nitrogen

Theory 1: volcanoes released nitrogen when the Earth was young (i.e. the early Earth's atmosphere already contained lots of nitrogen and resembled Titan's atmosphere. *As discussed previously, this theory is less likely*

Theory 2: nitrogen was added to the atmosphere gradually due to the reactions of nitrogen-containing compounds released from volcanoes

CC17c/SC21c The Atmosphere Today Questions

- Which gas is most abundant in today's atmosphere?
- The amount of which gases are affected by volcanic activity?
- Give two human activities which affect the amount of gases in the air.
- How did the amount of nitrogen in the air increase to the levels it is now?

CC17c/SC21c THE ATMOSPHERE TODAY 2

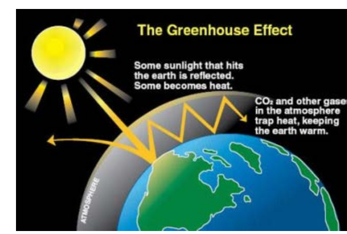
The Sun radiates energy which is either reflected back into space (by clouds, atmosphere and Earth's surface) or absorbed (by clouds, greenhouse gases in the atmosphere and Earth's surface).

The energy that is absorbed by the Earth's surface is re-radiated as infrared radiation, which can heat up the atmosphere.

For the Earth's temperature to stay the same, the energy per second absorbed by the Earth and its atmosphere must equal the energy per second radiated.

Effects of greenhouse gases on the Earth's energy balance

Greenhouse gases (such as carbon dioxide) trap heat energy so more is absorbed in the atmosphere and less is radiated back into space. This causes the temperature of the Earth to increase (global warming). To decrease the temperature of the Earth we would have to actively *remove* greenhouse gases from the atmosphere.



Strategies to stop the Earth's temperature rising

To reduce the Earth's temperature, we must reduce the amount of Sunlight that is absorbed by the Earth and its atmosphere. We can do this by increasing the amount of Sunlight that is reflected .

Possible strategies: Place huge white screens in space, about 2000km along each side or Float millions of white ping pong balls on ocean surfaces.

CC17c/SC21c THE ATMOSPHERE TODAY 2 questions

- How does the Sun’s energy get reflected back into space?
- Name three ways the sun’s energy is absorbed.
- If the Earth’s temperature is to stay the same then what must happen?
- Name a greenhouse gas and explain how it contributes to global warming.
- Suggest a possible strategy to reduce global warming.

CC17d/SC21d CLIMATE CHANGE

Some gases in the atmosphere, such as carbon dioxide, methane and water vapour, trap heat energy and help keep the Earth warm and this is known as the ***greenhouse effect*** and these global warming gases are referred to as 'greenhouse gases'.

Over the last 200 years there has been a dramatic increase in the levels of these greenhouse gases in the atmosphere, particularly CO_2 which has led to the global warming effect. Global warming effect is an increase in temperature and is likely to change weather patterns, causing climate change (also, the rising sea levels caused by melting ice caps mean flooding is an ever increasing danger to low-lying places).

The concentration of gases in the atmosphere (and the Earth's temperature) can change due to natural processes however, most scientists believe that the large increases in the levels of greenhouse gases in the atmosphere are due to human activities. The dramatic increase in CO_2 concentration in the atmosphere is thought to be due to more burning of fossil fuels and the increase in methane levels is thought to be due to increased large-scale farming.

Reducing the amount of carbon dioxide

One way of reducing the amount of carbon dioxide being added to the atmosphere is by limiting the use of fossil fuels. Chemists are currently investigating two further methods to actively reduce the level of carbon dioxide in the atmosphere.

1. Adding iron compounds to oceans – known as iron seeding -Iron is an essential nutrient for plants and is often in short supply so by adding iron encourages plants to grow. Plants use carbon dioxide for photosynthesis and when they die, the plants sink to the ocean floor and the carbonate in their shells is buried so carbon is removed from the atmosphere for a long time.
2. Converting carbon dioxide into hydrocarbons. The idea is to capture carbon dioxide from fossil-fuelled power stations and reacting it to make hydrocarbon compounds such as propane and butane.

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- Name three greenhouse gases.
- How do these gases lead to global warming?
- Which of these gases has increased the most over the last 200 years? How has this happened?
- How does global warming affect cold, icy habitats?
- Give two ways of reducing the amount of carbon dioxide in the air.