

Year	Chemistry
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Slides with this icon are for separate or triple science students only

Keyword	Definition
Substance	A material with particular properties
Matter	Anything that has mass and takes up space.
Solute	Substance that dissolves in a liquid to make a solution
Solvent	The liquid in which a solute dissolves to make a solution
Solution	Formed when a substance has dissolved in a liquid eg salt (solute) dissolves in water (solvent) forming salt water (solution).
Insoluble	Describes a substance that cannot be dissolved in a certain liquid eg nail varnish is insoluble in water.
Compound	A substance that can be split into simpler substances because it contains two or more different elements chemically joined together.
Mixture	Two or more substances jumbled together but not chemically joined together. Many mixtures can be easily separated. Melt over a range of temperatures eg 4 – 9 °C
Element	A substance made of only atoms with the same number of protons in the nucleus.
Pure	A single substance, with a fixed composition, that does not have anything else mixed with it eg a gold bar. Melt at a specific temperature eg 25 °C. As opposed to daily use of pure often used to mean natural or clean.
Impure	A mixture of different substances eg air contains, water, oxygen, carbon dioxide, nitrogen etc

### States of matter and changes of state

- See 'How to be a scientist' knowledge organiser. You need to know the state names, names of state changes and properties of each state.
- · Gas to solid or solid to gas is called sublimation.
- The properties of each state of matter depend on the forces between molecules.
- · When substances change state it is a physical change (easily reversed by heating or cooling and no new substances are created).
- · A chemical change results in the formation of a new substance.
- · On a graph showing temperature and time horizontal lines show a change of state for a pure substance (ie temperature does not change during a change of state for pure substances).

### Worked example - explain what happens when a gas is cooled.

- · When cooled particles in the gas lose energy.
- · The particles vibrate less strengthening the bonds between them making the gas' volume decrease.
- At a certain temperature the bonds between particles become strong enough to hold them together. This is called condensing and the gas changes to a liquid.
- · As the liquid is cooled the bonds continue to strengthen, the volume continues to decrease and particle energy continues ot decreases.
- · At a certain temperature particles have lost enough energy for the particles to be held together even more tightly fixing them to a particular point. This is called freezing; the liquid is now a solid.
- · As cooling continues particles continue to lose energy and vibrate less and less.

### Predicting state from data

- · Look at the given temperature.
- . If this temperature is less than the melting point (MP) of a substance it will be a solid.
- · If this temperature is between the melting and boiling points (BP) of a substance it will be a liquid.
- If this temperature is above the boiling point the substance will be a gas.

Substance	MP/°C	BP/°C	State at 30 °C
Α	450	700	solid
В	-200	-10	gas
С	7	85	liquid



Separation method	Used to separate	Example mixture	Notes on method
Simple distillation	a solvent and solute from a solution	Water and salt from salt water	Can only be used to separate substances with very different boiling points
Fractional distillation	a mixture of liquids with different boiling points	Fractions from crude oil	Substance with lowest BP collected first, highest BP collected last
Filtration	a mixture of a liquid and solid	Sand from rock salt	Insoluble solid impurities can be removed from solutions this way
Crystallisation	a solute from a solution (solvent is lost)	Copper sulfate from copper sulfate solution	The solvent will evaporate leaving crystals of the solute. Using an oven will speed this up.
Paper chromatography	and identify soluble substances from a mixture	Identify dyes used to make an ink	Mobile phase – solvent in which molecules move eg water Stationary phase – molecules cannot move eg paper
		<del>-</del>	

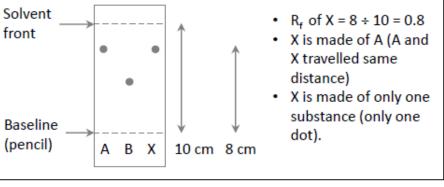
### Chromatography

each substance. It depends on how soluble they are in the mobile phase and how strongly they are attracted to the stationary phase. · Very soluble substances that are weakly attracted to the

The rate a substance travels up the stationary phase differs for

- stationary phase will travel furthest up the paper. · Chromatography can be used to see if a substance is pure or if
- it is a mixture.

### Worked example



- R<sub>f</sub> factors can be used to identify what substances are in a mixture by comparison to a known substances R<sub>f</sub> value. R<sub>f</sub> values are always less than 1.
- $R_f = \frac{distance\ travelled\ by\ solute}{...}$
- distance travelled by solvent

### Water treatment and analysis

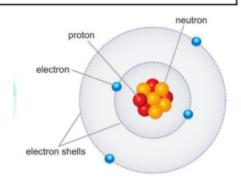
- · Waste and ground water can be made potable (drinkable) using the following processes:
- 1. Filtration large solids (sand, twigs etc) are removed by wire mesh and gravel and sand beds
- 2. Sedimentation the water is held in large tanks to let smaller particles settle to the bottom
- 3. Chlorination Chlorine gas bubbled through water to kill bacteria
- and other microbes Distillation can be used to make seawater drinkable. It requires a lot of energy to do and is therefore expensive.
- Water used for chemical analysis must be pure. Tap water contains dissolved substances that can affect experiments and give false results if used in chemical analysis.

	Word	Meaning
;	atom	Atoms are small particles from which all substances are made. They are the smallest neutral part of an
,		element that can take part in chemical reactions.
5	electron	Tiny particle with a negative charge that is found in shells around the nucleus of an atom.
	electron shell	
	electron shell	Area around a nucleus that can be occupied by
)		electrons, usually drawn as a circle (in 'target
)		diagrams'). Also called an electron energy level or an 'orbit'.
	element	A simple substance made up of only one type of
)		atom.
	neutron	Electrically neutral subatomic particle found in the
		nucleus of most atoms.
	nucleus	The positively charged centre of an atom.
	proton	A positively charge subatomic particle in the nucleus of all atoms.
	relative charge	The electric charge of a subatomic particle compared
		to the charge on a proton.
	relative mass	The mass of a subatomic particle compared to the
		mass of a proton.
	subatomic particles	The smaller particles that make up atoms - protons,
		neutrons and electrons.

Dalton's model of	-All matter made of atoms which:			
atoms	- Are tiny hard spheres			
	- Can't be broken down			
	- Can't be created or destroyed			
	- Atoms of an element are identical			
	- Different elements have different atoms			
IJ Thompson	Discovered that atoms contain even smaller particles called subatomic particles. He discovered electrons			

Atoms in elements have no overall charge because they have equal numbers of positive protons and negative electrons, so the charges cancel each other out

all of its mass



Subatomic particle	Relative charge	Relative mass	
proton	+1 (positive)	1	
electron	−1 (negative)	1/1835 (negligible)	
neutron	0 (no charge)	1	

The nucleus of an mass number (A) (protons + neutrons) atom contains nearly atom symbol atomic number (Z) (protons only)

Word	Meaning		
chemical property	How a substance reacts with other substances.		
periodic table	An ordered list of all known elements.		
physical property A description of how a material behaves and res			
	to forces and energy. Hardness is a physical property.		
prediction	What you think will happen in an experiment and why		
	you think this.		
relative atomic	The mean mass of an atom relative to the mass of		
mass, A <sub>r</sub>	one-twelfth of an atom of carbon-12, which is assigned		
	a mass of 12. The A <sub>r</sub> of an element is the mean		
	relative mass of the isotopes in the element.		

Word Meaning					
atomic number	The number of protons in the nucleus of an atom				
	(symbol Z). Also known as the proton number.				
mass number	The total number of protons and neutrons in the				
	nucleus of an atom (symbol A). Also known as the				
	nucleon number.				
periodic table	Chart in which the elements are arranged in order				
	of increasing atomic number.				



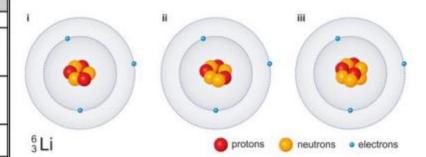
Word	Meaning			
A <sub>r</sub>	Symbol for relative atomic mass (RAM).			
isotopes	Atoms of an element with the same number of protons (atomic number) but different mass numbers due to different numbers of neutrons.			
mean	An average calculated by adding up the values of a set of measurements and dividing by the number of measurements in the set.			
nuclear fission	The reaction in which the nucleus of a large atom, such as uranium, splits into two smaller nuclei.			
relative atomic mass (RAM)	The mean mass of an atom relative to the mass of an atom of carbon-12, which is assigned a mass of 12. The RAM of an element is the mean relative mass of the isotopes in the element.			

Word	Meaning  Tiny particle with a negative charge that is found in shells around the nucleus of an atom.		
electron			
electron shell	Areas around a nucleus that can be occupied by electrons, usually drawn as circles. Also called an electron energy level.		
electronic configuration	The arrangement of electrons in shells around the nucleus of an atom.		

Each electron shell can hold a maximum number of electrons: First shell = 2 electrons

Second and third – maximum of 8 electrons

1	2	3	4	5	6	7	0
H							He
	Be	(B)	<b>©</b>	(N)	<b>(</b>	(F)	Ne
Na	Mg	(A)	(Si)	<b>(D)</b>	(S)		(Ar)
		Electron configuration of elements 1-20 Written form is e.g. 2.8.1 for K					



Isotopes of Lithium have different numbers of neutrons

Calculating relative atomic mass

Relative atomic mass can be calculate using: Mass number of an elements isotopes The abundance ( how much) of each isotope

### Example:

Chlorine naturally exists as two isotopes,  $_{17}^{35}Cl$  (chlorine-35) and  $_{17}^{37}Cl$  (chlorine-37). The abundance of chlorine-35 is 75% and the abundance of chlorine-37 is 25%. In other words, in every 100 chlorine atoms, 75 atoms have a mass number of 35, and 25 atoms have a mass number of 37.

To calculate the relative atomic mass,  $A_r$ , of chlorine:

$$\mathbf{A}_r = \frac{total\; mass\; of\; atoms}{total\; number\; of\; atoms} = \frac{(35\times75) + (37\times25)}{(75+25)}$$

$$A_r = \frac{2625 + 925}{100} = \frac{3550}{100}$$

$$A_r = 35.5$$
 (to 1 decimal place)

Notice that the answer is closer to 35 than it is to 37. This is because the chlorine-35 isotope is much more abundant than the chlorine-37 isotope.



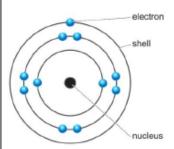
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Word	Meaning
atomic number	The number of protons in the nucleus of an atom
	(symbol Z). Also known as the proton number.
group	A vertical column of elements in the periodic table.
	Elements in the same group generally have similar
	properties.
inert	Does not react.
period	A horizontal row in the periodic table.
relative atomic mass	The mean mass of an atom compared to 1/12th the
	mass of an atom of carbon-12. (One atom of carbon-
	12 has been assigned a mass of 12.)
X-ray	Electromagnetic radiation that has a shorter
	wavelength than UV but longer than gamma rays.

Remember from KS3 – the periodic table hasn't always looked like the one we see today

Electron configuration is related to an elements position on the periodic table:

Number of occupied shells = period number Number of electrons in the outer shell = group number



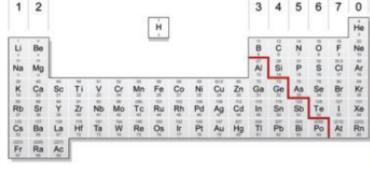
Sodium Atamia A

Atomic Number 11 Electron configuration 2.8.1

Group = 1 Period = 3

Dmitri Mendeleev	Russian chemist, developed the
	periodic table.
Mendeleev's periodic	Ordered by increasing A <sub>r</sub> , some
table	elements switched according to
	their chemical properties.
Gaps in Mendeleev's	Mendeleev left gaps where no
periodic table	known element fitted and
	predicted these would be filled
	with newly discovered elements.
Eka-aluminium	An element that Mendeleev
	thought would fill a gap. He
	predicted its properties, which
	matched gallium when discovered.

Knowledge of isotopes made it possible to explain why the order based on atomic weights was not always correct, because some elements have a higher mass than others when isotopes are taken into account, but a lower one if you only look at one specific isotope.



The red line shows the divide between metals and nonmetals in the periodic table

- Metals = elements that react to form positive ions.
- o Majority of elements are metals.
- o Found to the left and towards the bottom of the periodic table., because they lose electron(s) in order to form these positive ions, forming an electronic structure that is stable, like that of a noble gas
- Non-metals = elements that do not form positive ions.
- o Found towards the right and top of the periodic table, because they gain electron(s) in order to form these negative ions, forming an electronic structure that is stable, like that of a noble gas

Pair reversals	Elements (like Ar and K) that are not in order of increasing
	mass.
Explaining pair	It means elements should be order elements by increasing
reversals	atomic number instead.

Keyword	Definition	
Intermolecular force	A weak force of attraction between molecules	
Allotropes	Different structural forms of the same element	
Covalent, giant structure	3D lattice of carbon atoms linked by covalent bonds	
Lattice structure	Regular grid-like repeating arrangement of particles such as atoms, molecules or ions	
Polymer	A long-chain molecule made by joining many smaller molecules (monomers) together.	

### Allotropes of carbon

 Carbon can take several different structures each with differing physical properties.

### Graphite

- Sheets of carbon atoms joined by weak intermolecular forces between sheets. Each atom is covalently bonded to 3 others.
- Each atom has one delocalised electron that can move throughout the structure making graphite electrically conductive. As it is also chemically non-reactive it is used to make electrodes.
- The weak intermolecular forces between sheets means they can easily slide past each other meaning graphite can be used as a lubricant.

### Diamond

- A giant tetrahedral arrangement of carbon atoms. Each atom is covalently bonded to 4 others.
- No delocalised electrons result in diamond being an electrical insulator.
- As the entire structure is joined through strong covalent bonds diamond is extremely hard with a high melting and boiling point. It is therefore used in cutting tools.

### Allotropes of carbon (continued)

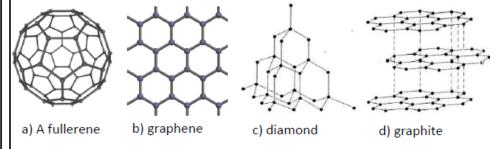
### **Fullerenes**

- Simple molecules of carbon atoms joined covalently into a ball (eg C<sub>60</sub>). Each carbon atom makes 3 covalent bonds.
- Delocalised electrons exist in fullerenes but can only move within a single molecule so fullerenes are electrical insulators.
- Weak intermolecular forces between molecules mean a low melting point and make fullerenes soft and slippery.

### Graphene

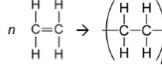
- A sheet of carbon atoms one atom thick. Each atom is covalently bonded to 3 others.
- Each atom has one delocalised electron making graphene electrically conductive.
- · The covalent bonds make graphene extremely strong.

### Ball and stick models of:



### **Polymers**

- Polymers are long molecules made of a chain of covalently bonded carbon atoms which other atoms are joined to.
- They are formed when many small, identical units (monomers) join together.
- An example of a polymer is poly(ethene); made from monomers of ethene:





	lonic	Simple molecular	Giant covalent	Metallic
Structure	Lattice structure made of billions of ions	Small, separate groups of covalently bonded atoms	Lattice structure made of billions of atoms	Lattice structure made of billions of cations in layers in a 'sea' of delocalised electrons
Melting / boiling points	High	Low	High	High
Explanation	Strong electrostatic forces between ions must be broken. Therefore a lot of energy is needed so MP/BP are high.	Only the weak intermolecular forces between molecules are broken. Therefore little energy is needed so MP/BP are low.	Strong covalent bonds between atoms must be broken. Therefore a lot of energy is needed so MP/BP are high.	Strong electrostatic forces between ions and delocalised electrons must be broken. Therefore a lot of energy is needed so MP/BP are high.
Electrical properties	Conductors when molten (liquid) or in solution (dissolved in water).	Insulators	Most are insulators (graphite and graphene are conductive).	Conductors when solid or liquid.
Explanation	Only electrically conductive when ions can move (ie when molten or in solution).	No ions or delocalised electrons are present so electricity cannot be conducted.	No ions or delocalised electrons are present so electricity cannot be conducted.	Delocalised electrons can move through the structure.
Solubility	Many are soluble in water	Few are soluble in water	Insoluble in water	Insoluble in water

### Metals and non-metals

- · Metals Generally, shiny solids at room temperature with high melting/boiling points, high density, good electrical conductors, malleable.
- Non-metals Generally, low melting/boiling points, poor electrical conductors, brittle.

### Malleability of metals

- Metals can be hammered or rolled into shape (they are malleable).
- The layers of ions in a metal can slide over each other and the 'sea' of delocalised electrons will hold them together.

### Problems with bonding models

- · Ball and stick models show which atoms are joined together and shape of structure but show atoms too far apart and sticks that do not exist or hold atoms together.
- · Dot and cross diagrams show how electrons are shared in covalent bonds but suggest electrons in different atoms are different and does not show formed structures.
- 3D space filling models show relative size of atoms/ions but do not show how electrons have been shared/transferred.
- No printed model can show vibrations of ions/atoms or show the movement of delocalised electrons.

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Keyword	Definition	
lon	An atom or group of atoms with a positive or negative charge (ie unequal numbers of protons and electrons)	
Anion	A negatively charged ion (a molecule with more electrons than protons)	
Cation	A positively charged ion (a molecule with fewer electrons that protons)	
Bond	A force that holds some atoms tightly together	
Electrostatic force Force of attraction between oppositely charged particles and force of repulsion between particles with the same charge		
Ionic bond	Strong electrostatic force of attraction between oppositely charged ions	
Covalent bond	The bond formed when pair of electrons is shared between two atoms	
Metallic bond	Bonding found in metals. Can be thought of as positively charged ions in a 'sea' of negatively charged delocalised electrons	
Delocalised electrons	An electron that is free to move (ie not bound to one nucleus) and can carry an electrical charge	
Lattice structure	Regular grid-like repeating arrangement of particles such as atoms, molecules or ions	

### Ionic compound structures

· Ionic compounds form a lattice with a regular, repeated arrangement of oppositely charged ions held together by the strong electrostatic forces of attraction between them.

Ionic compound formulas show the ratio of ions in the

compound. Cl-

 $Na^+$ 

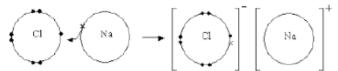
A ball and stick model showing the lattice structure of sodium chloride.

### Ionic bonding

· Ionic bonds are formed when electrons are transferred from one atom to another producing a cation and an anion. They typically form between a metal and a non-metal.

### Dot and cross diagrams

- · Only electrons in the outermost shells need to be shown.
- The number of electrons in the outer shell of an atom is the same as the group number it is in. (eg Cl group 7, 7 electrons in outer shell).
- · An atom of sodium transfers one electron to the outer shell of chlorine creating an anion and a cation.



### Formation of ions

· When forming ions, atoms will gain or lose electrons until they have a full outer shell - a stable arrangement of electrons.

Group	Outer electrons	Electrons gained or lost	Charge on ion
1	1	1 lost	+1
2	2	2 lost	+2
6	6	2 gained	-2
7	7	1 gained	-1

- Mass number = the number of protons + the number of neutrons
- Atomic number = the number of protons
- · The number of electrons, neutrons and protons can be calculated if given the formula of an ion.
- For example: Li<sup>+</sup>
  - · Lithium: mass number = 7; atomic number = 3
  - Therefore Li has 4 neutrons (7-3 = 4) and 3 protons.
  - For an atom the number of electrons = the number of protons
  - · Li\* has a charge of +1 and therefore has lost 1 electron compared to an atom of lithium leaving 2 electrons.

### Names of compounds

- A name ending in –ide indicates an element has gained one or more electrons eg a chlorine anion is called a chloride ion.
- A name ending in –ate indicates the presence of oxygen eg sodium carbonate is Na<sub>2</sub>CO<sub>3</sub>

### Calculating an ionic compound's formula

- Ionic compounds are always neutral. This means that the charges
  of the ions that make up the compound add together to make 0.
- The formula of an ionic compound can therefore be calculated if the ions that are in it are known.

### Worked example

- · Deduce the formula of sodium carbonate.
- Sodium is in group 1 and therefore forms an ion with a charge of +1.
- Carbonate ions have the formula CO<sub>3</sub><sup>2-</sup> and therefore have a charge of -2.
- 3. A ratio of 1 Na<sup>+</sup>: 1 CO<sub>3</sub><sup>2-</sup> gives: (+1) + (-2) = -1
- 4. To reach 0 we must add one more ion with a charge of +1
- 5. A ration of 2 Na<sup>+</sup>:  $1 CO_3^{2-}$  gives: 2(+1) + (-2) = 0
- 6. The charge is 0 so the formula is Na<sub>2</sub>CO<sub>3</sub>.

Common ion name	Formula of ion
Nitrate	NO <sub>3</sub> -
Carbonate	CO <sub>3</sub> <sup>2-</sup>
Sulfate	SO <sub>4</sub> <sup>2-</sup>
Hydroxide	OH-

### Covalent bonding

- · A covalent bond is pair of electrons shared between two atoms.
- · They typically form between non-metal atoms.
- Covalent bonds cause molecules to form.
- Atoms and small molecules (eg H<sub>2</sub>O) are typically about 10<sup>-10</sup> m across

### Drawing covalent bonds

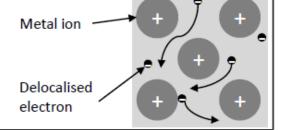
Atom	Electrons in outer shell	Extra electrons needed to fill outer shell	Number of covalent bonds formed
Group 7 (eg F)	7	1	1
Group 6 (eg O)	6	2	2
Group 5 (eg N)	5	3	3
Group 4 (eg C)	4	4	4
Hydrogen	1	1	1

- By matching the number of bonds formed by an element a formula for a covalent molecule can be figured out:
- · Carbon dioxide is made form carbon and oxygen atoms.
- · Carbon will form 4 covalent bonds; oxygen 2.
- Therefore 2 oxygen atoms will bond with 1 carbon atom:

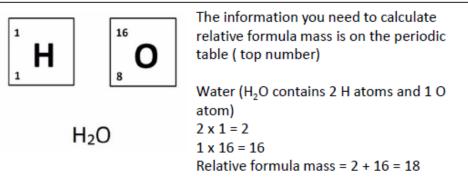


### Metallic bonds

- Metals form structures with a giant lattice structure.
- The electrons in the outer shell of metals are delocalised.
- There is a strong electrostatic force of attraction between delocalised electrons and the metal cations – a metallic bond.
- All metals (including alloys) are held together by metallic bonds.



Word	Meaning
empirical formula	The formula showing the simplest whole number ratio
	of atoms of each element in a compound.
molecular formula	The formula showing the actual number of atoms of
	each element in a molecule of a compound.
relative formula mass	The sum of the relative atomic masses of all the
	atoms in a formula.



The concentration of a solution can be calculated using: the mass of **dissolved** solute in grams, g the volume of solution (or solvent) in cubic decimetres, dm<sup>3</sup>

concentration = 
$$\frac{\text{mass of solute in g}}{\text{volume of solution in dm}^3}$$

Volumes used in concentration calculations must be in dm<sup>3</sup>, not in cm<sup>3</sup> or ml. 1 dm<sup>3</sup> = 1000 cm<sup>3</sup>.

This means:

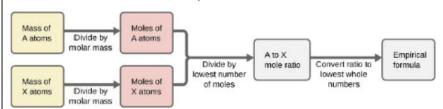
- · divide by 1000 to convert from cm3 to dm3
- multiply by 1000 to convert from dm³ to cm³

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concentration	The amount of a solute dissolved in a certain volume	
	of solvent.	
solute	A substance that dissolves in a liquid to make a	
	solution.	
solvent	Describes the liquid in which a substance dissolves to	
	make a solution.	

### 

We can divide each number in  $C_6H_{12}O_6$  by 6 to make a simpler whole number ratio – the empirical formula



### WORKED EXAMPLE

GCSE pod

Symbol for element	Ca	Cl
Mass (g)	10.0	17.8
Relative atomic mass, $A_{\rm r}$	40	35.5
Divide the mass of each element by its relative atomic mass	$\frac{10.0}{40} = 0.25$	$\frac{17.8}{35.5} = 0.5$
Divide the answers by the smallest number to find the simplest ratio	$\frac{0.25}{0.25} = 1$	$\frac{0.5}{0.25} = 2$
Empirical formula	CaCl.	

By weighing before and after heating ( to constant mass) you can practically find empirical formula



To find a molecular	- Calculate M <sub>r</sub> for the empirical formula
formula from an	- Divide the M <sub>r</sub> of the molecular formula by
empirical formula	this number
	- Multiply the empirical formula by your
	answer









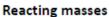
Mass does not change in a reaction-the mass of product will always equal the mass of reactants

n insoluble substance that is formed when two luble substances react together in solution. Then substances cannot enter or leave an	
hen substances cannot enter or leave an	
oserved environment, e.g. a stoppered test	
be and a precipitation reaction.	
ne idea that mass is never lost or gained during	
chemical reaction or physical change.	
stem When substances can enter or leave an	
served environment . For example, a	
rbonate reacting with acid producing CO <sub>2</sub>	
ibbles: the mass appears to decrease because	
ou can't weigh the gas that goes into the air,	
owever it is still there.	

Avogadro constant	This is the number of particles in one mole of a	
	substance (6.02 ×10 <sup>23</sup> mol <sup>-1</sup> ).	
limiting reactant	The reactant that determines the amount of product	
	formed in a chemical reaction. Any other reactants	
	will be present in excess.	
mole	The mass of a mole of a substance is the relative	
	formula mass expressed in grams.	
stoichiometry	The molar ratio of the reactants and products in a	
	chemical reaction.	

1 mol of a substance contains the Avogadro constant number of particles

1 mol of magnesium has a mass of 24g and contains  $6.02 \times 10^{23}$  atoms number of moles of substance =  $\frac{\text{mass of substance (g)}}{A_c \text{ or } M_c}$ 



Used to calculate mass of a product or reactant from a balanced equation

### Worked example

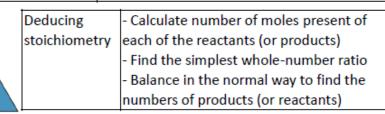
Calculate the mass of chlorine needed to make 53.4g of aluminium chloride.

Calculate ratio of masses (multiply  $M_r$  values by the balancing numbers shown in the equation).

3Cl₂ makes 2AlCl₃ so 3 × 71 = 213 g Cl₂ makes 2 × 133.5 = 267 g AlCl,■

Work out the mass for 1g of reactant or product. (Here we want 1g of the product because that's the mass we know already.)

Calculating moles	Quantity in moles =		
from a number of	number of particles		
particles	6.02 x 10 <sup>23</sup>	3	
Calculating the	Number of particles =		
number of particles	mass	X	6.02 x 10 <sup>23</sup>
from a mass of	relative formula mass)		
substance			



Word	Meaning	
actual yield	The actual amount of product obtained from a	
	chemical reaction.	
percentage yield	The actual yield divided by the theoretical yield, as a	
	percentage.	
side reactions	A reaction which takes place at the same time as	
	another main reaction.	
theoretical yield	The maximum calculated amount of a product that	
	could be formed from a given amount of reactants.	
yield	The amount of product formed in a reaction.	
atom economy	The percentage, by mass, of reactants that are	
	converted into useful products.	
by-product	Any product formed in a reaction in addition to the	
	required product.	

relative formula mass  $(M_r)$  of the useful product atom economy =  $\frac{1}{\text{sum of relative formula masses of all the reactants}}$ 

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

$$(M_r = 180)$$
  $(M_r = 46)$ 

$$\frac{2\times46}{180}\times100=51.1\%$$

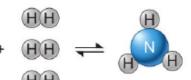
Atom economy shows how efficient a reaction is in making use of all of the atoms in a reaction.

Not all product s may be useful

1 product = 100% atom economy

Industry could sell by products, or use in another process







100% atom economy equilibrium reached Loss of product (in transfer)

Yield is less than expected due to:

Unwanted side reactions – reactions making different products

High % yield is useful. Less waste, more profit

Incomplete reaction - not left long enough to react, or

percentage yield = 
$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Disadvantage – does not indicate amount of waste product.

If a reaction has high yield but low atom economy - lots of waste products.

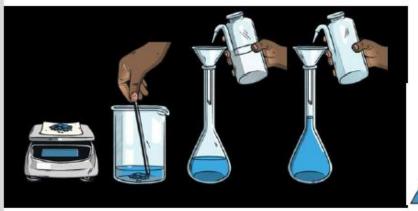
Waste reduction is important

Energy consumption, rate of reaction, raw materials and conditions all need to be considered to produce a high yield if an equilibrium is reached.

	Fermentation	Reaction of ethene with steam
raw materials	carbohydrates (e.g. from sugar cane, sugar beet, maize) – these are renewable	ethene obtained from crude oil – this is non-renewable
temperature	30–40°C	300°C
pressure	atmospheric pressure	high pressure (60–70 atm)
catalyst	enzymes in yeast	concentrated phosphoric acid
rate of reaction	slow	fast
purity	impure – must be fractionally distilled	pure

(g dm<sup>-2</sup>

Word	Meaning
calibrated	Marked with a scale for accurate readings.
volumetric flask	A flask which is accurately calibrated to hold a given volume of solution.



Acid-alkali titration are used to find the exact volume of acid that neutralises a specific volume of alkali

1. Use the pipette to add 25 cm3 of alkali to a conical flask

1	and add a few drops of indicator.
	Fill the burette with acid and note the starting volume.     Slowly add the acid from the burette to the alkali in the conical flask, swirling to mix.
1	<ol> <li>Stop adding the acid when the end-point is reached (the appropriate colour change in the indicator happens). Note the final volume reading. Repeat steps 1 to 3 until you get consistent readings.</li> </ol>

concentration in g dm<sup>-3</sup> =  $\frac{\text{mass of solute in g}}{\text{volume of solution in dm}^3}$ 

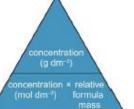
Word Meaning
mole ratio The ratio between the amounts, in moles, of substances in a balanced equation.

Solutions with accurate concentrations are made in volumetric flasks.

A known amount of solute is dissolved in distilled water, then made up to the graduation,. The flask is stoppered and shaken.

Volume is measured from the BOTTOM of the meniscus.

concentration in mol dm<sup>-3</sup> =  $\frac{\text{concentration in g dm}^{-3}}{\text{relative formula mass of solute}}$ 



concentration in mol dm<sup>-3</sup> =  $\frac{\text{number of moles of solute}}{\text{volume of solution in dm}^3}$ 

Worked example

 $25.0\,\text{cm}^3$  of 0.200 mol  $dm^{\text{-}3}$  sodium hydroxide solution was neutralised by 0.150 mol  $dm^{\text{-}3}$  sulfuric acid. Calculate the volume of sulfuric acid needed for the neutralisation.

2NaOH + H<sub>2</sub>SO<sub>2</sub> 
$$\rightarrow$$
 Na<sub>2</sub>SO<sub>2</sub> + 2H<sub>2</sub>O mole ratio NaOH: H<sub>2</sub>SO<sub>2</sub> = 2:1

Step 1: work out moles of NaOH (from given information) number of moles of NaOH = concentration of NaOH × volume used (mol dm<sup>-3</sup>) (dm<sup>3</sup>) =  $0.200 \times \frac{25.0}{1000} = 0.00500 \text{ mol}$ 

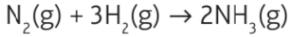
2 mol of NaOH reacts with 1 mol of H<sub>2</sub>SO<sub>4</sub>,

so 0.00500 mol of NaOH reacts with  $\frac{0.00500}{2} = 0.00250$  mol of H<sub>2</sub>SO<sub>4</sub>

Step 3:Calculate volume of acid needed.

volume in dm<sup>3</sup> =  $\frac{\text{number of moles of acid}}{\text{concentration in mol dm}^{-3}}$ =  $\frac{0.00250}{0.150}$  = 0.0167 dm<sup>3</sup> or 16.7 cm<sup>3</sup>

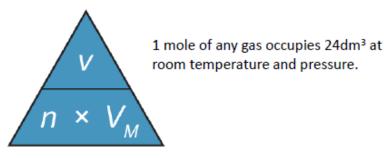
Word	Meaning
Avogadro's law	The law that states that, if the temperature and pressure are the same, equal volumes of different gases contain an equal number of molecules.
molar gas volume	The volume occupied by one mole of molecules of any gas. It is 24 dm <sup>3</sup> , or 24 000 cm <sup>3</sup> , at room temperature and pressure.



1 mole of nitrogen reacts with 3 moles of hydrogen to form 2 moles of ammonia.

So:

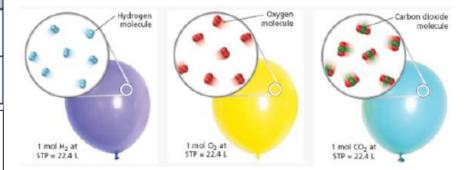
100cm<sup>3</sup> of nitrogen reacts with 300cm<sup>3</sup> of hydrogen to produce 200cm<sup>3</sup> of ammonia at the same temperature and pressure.



n = amount in mol, v = volume of gas,

 $V_m$  = molar volume

volume of gas = amount of gas (mol) × molar volume



Volume of gas does not depend on relative formula mass.

Temperature, pressure and number of particles determines volume

### Remember:

- Balance equation
- Calculate number of moles
- Use equation to work out moles of unknown
- Calculate unknown amount

Calcium carbonate decomposes to form calcium oxide and carbon dioxide when it is heated:

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

Calculate the volume of carbon dioxide produced from completely decomposing 1.0 g of calcium carbonate. ( $M_r$  of CaCO<sub>3</sub> = 100, molar volume = 24,000 cm<sup>3</sup>)

Amount of calcium carbonate = 1.0/100 = 0.010 mol

Amount of carbon dioxide = 0.010 mol (because the mole ratio is 1:1)

Volume of carbon dioxide =  $0.010 \times 24,000$ 

= 240 cm<sup>3</sup>



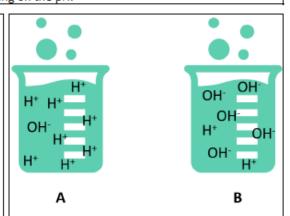
Word	Meaning	
acid	A solution with a pH of less than 7 and that contains an	
	excess of hydrogen (H+) ions. Acids turn litmus red.	
acidic	Containing or having the properties of an acid. (adjective)	
acidity	The amount of acid in a solution.	
alkali	A solution with a pH of more than 7 and that contains an	
	excess of hydroxide (OH-) ions. Alkalis turn litmus blue.	
alkaline	Having a pH of more than 7.	
alkalinity	The amount of alkali in a solution.	
aqueous solution	A solution with water as the solvent.	
concentration	A measure of how much solute is dissolved in a solvent	
	such as water. (Units g dm <sup>-3</sup> or mol dm <sup>-3</sup> )	
indicator	A substance that changes colour depending on the pH of a	
	solution.	
neutral	A substance that is neither an acid nor an alkali. Neutral	
	solutions have a pH of 7 and the same concentrations of	
	hydrogen (H <sup>+</sup> ) and (OH <sup>-</sup> ) ions.	
pH scale	A numerical scale up to 14 that measures the acidity or	
	alkalinity of a solution based on the concentrations of	
	hydrogen (H+) and (OH-) ions.	
universal indicator	An indicator, containing a mixture of different pH	
	indicators, designed to produce a range of colours	
	depending on the pH.	

Explosives	Oxidiser	Flammables
Corrosives	Toxic	Irritant
Compressed Gas	Environmentally Hazardous	Health Hazard

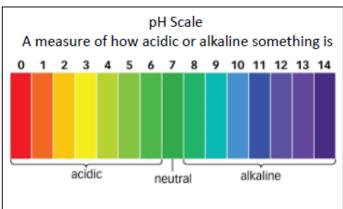
Many substances are

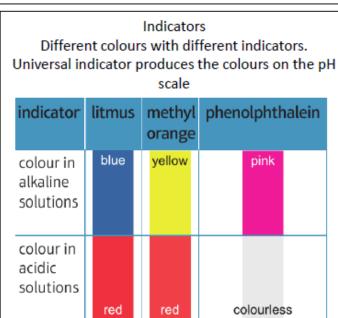
HAZARDOUS and are identified

by a universal symbol so they can be used safely



- Acids produce excess H+ ions in solution
- Alkalis produce excess OH- ions in solution





### HIGHER

More H+ ions= higher concentration, more acidic, lower pH

More OH- ions = more alkaline, higher pH

Neutral H<sup>+</sup> = OH<sup>-</sup> ion concentration

Word	Meaning						
concentrated	Containin volume o	T 12 T 3	amount of	solute diss	olved in a	small	
dilute	Containin volume o	The second second second	amount of	solute diss	olved in a	large	
dissociate	For exam	ple, acid i	molecules	fferent par dissociate i olve in wat	nto H <sup>+</sup> ion	s and	
pH meter	Electronic	c device u	sed to mea	sure the p	H of a solu	ition.	
strong acid	An acidic it dissolve		at dissociat	es comple	tely into io	ns when	
weak acid	100000000000000000000000000000000000000	solute than		t dissociate	complete	ly into	9
рН	0	1	2	3	4	5	
difference in	4	14	14	14	14	14	

× 10

Common acids	Formula
hydrochloric acid	HCI
sulfuric acid	H <sub>2</sub> SO <sub>4</sub>
nitric acid	HNO <sub>3</sub>
Common alkalis	Formula
sodium hydroxide	NaOH
potassium hydroxide	кон
calcium hydroxide	Ca(OH),



An acid with a pH of 0 IS 10 x 10 x 10 x 10 = 100000 times more acidic than an acid of pH4

× 10

 $Concentration = \frac{amount\ dissolved}{volume\ of\ solution}$ 

concentration of

H+ ions

Units: grams per decimetre cubed, g dm-3

For example, if 4 g is dissolved in 50 cm3

Concentration =  $\frac{4}{0.05}$  = 80 g dm<sup>-3</sup>

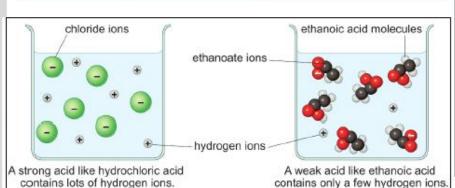
Note: the minus sign shows that g is divided by dm<sup>3</sup>. 1 dm<sup>3</sup> is the same volume as 1 litre or 1000 cm<sup>3</sup>.

× 10

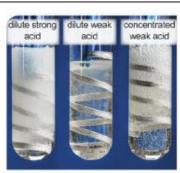
× 10

× 10

Note: divide cm<sup>3</sup> by 1000 to change into dm<sup>3</sup>.

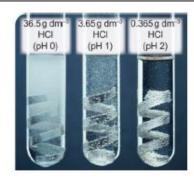


× 10



 $\times 10$ 

A dilute strong acid and a concentrated weak acid can have a similar pH and reactivity due to similar concentrations of H<sup>+</sup> ions



Lower pH = more H+ ions Higher pH = less H+ ions Lower pH react faster with metals such as Mg



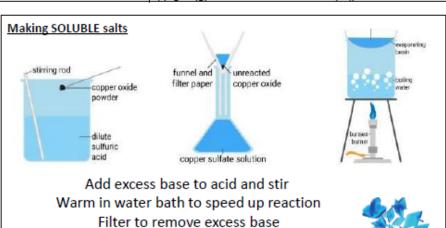
Word	Meaning
base	Any substance, soluble or insoluble, that neutralises an
	acid, forming a salt and water only.
crystallisation	The process of forming crystals.
filter (verb)	To remove or separate a solid from a liquid by passing the
	mixture through a porous material.
neutralise (verb)	To make a solution neither acidic nor alkaline. During
	neutralisation a base reacts with an acid, forming a salt and
	water.
salt	An ionic compound produced by a neutralisation reaction.
state symbols	Standard set of symbols written after chemical formulae to
	indicate the state of a substance. These are: solid (s), liquid
	(I), gas (g) and dissolved in water (aq)

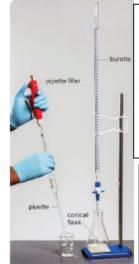
Acid	Salt formed
hydrochloric acid	chloride
sulfuric acid	sulfate
nitric acid	nitrate

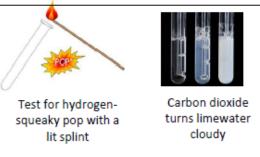
The type of salt formed comes from the acid used

### Neutralisation reactions

ACID + BASE	$\longrightarrow$	SALT + WATER
ACID + ALKALI	$\longrightarrow$	SALT + WATER
ACID + METAL	$\rightarrow$	SALT + HYDROGEN
ACID + CARBONATE	$\rightarrow$	SALT + WATER + CARBON
		DIOVIDE







Titration can be used to make salts from alkalis. This ensures you mix an acid and alkali in the correct proportions so you end up with a salt and water <u>ONLY</u>

	You can monitor changes in pH be adding small amounts of base to acid, and take a pH reading after each addition
pl	l can be
	easured
W	ith a pH

meter

Heat over water bath to concentrate solution

Leave to crystallise

Word	Meaning
burette	Apparatus used to accurately measure the volume of
	solution that has been added during a titration.
end-point	In a titration, when just enough solution has been added
	from the burette to react with all the solution in the flask.
pipette	Apparatus used to accurately measure a set volume of a
	solution, which can be used in a titration.
titration	Method used to mix acids and alkalis in the correct
	proportions to produce a solution containing only salt and
	water. It can be used to find the concentration of an acid
	or an alkali.

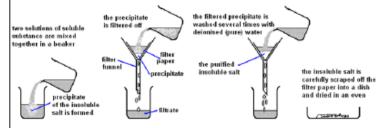


Soluble in water	Insoluble in water
all common sodium, potassium and ammonium salts	
all nitrates	
most chlorides	silver, lead chlorides
most sulfates	lead, barium, calcium sulfates
sodium, potassium and ammonium carbonates	most carbonates
sodium, potassium and ammonium hydroxides	most hydroxides

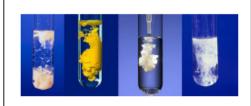
copper sulfate + potassium carbonate → copper carbonate + potassium sulfate

In precipitation reactions, the ions from the salts swap, The solubility rules can help to decide whether a precipitate will form

To make an insoluble salt, two soluble salts need to react together in a precipitation reaction



Mix solutions, filter, rinse solid with deionised water, leave to dry in a warm place



### GCSE Pod



### Useful websites

https://www.bbc.co.uk/bitesize/topics/zsgcv9q

https://app.senecalearning.com/classroom/course/1fd02c70 -10e2-11e8-92d6-f7c3df19f7bc/section/d092e750-159c-11e8-946f-958618255889/section-overview

Word	Meaning
effervescence	Fizzing or a stream of bubbles produced during a
	reaction.
half equation	A balanced equation, including electrons, that shows
	what happens to one substance during a redox
	reaction.
ionic equation	A balanced equation that only shows the ions that
	react together. The spectator ions are not included in
	the equation.
oxidation	A reaction in which a substance gains oxygen or in
	which an atom or ion loses electrons.
reactivity series	A list of metals in order of reactivity with the most
1	reactive at the top.
reduction	A reaction in which a substance loses oxygen or in
	which an atom or ion gains electrons.
spectator ions	These are ions that do not change during a reaction.
precipitate	An insoluble product formed when solutions of two
	soluble reactants are mixed.
precipitation	A reaction in which an insoluble product is formed
]	from two soluble reactants in solution.



### HIGHER

Ionic equations show the ions that react together in a precipitation reaction- the ions that do not change are not included — **spectator ions** 

 $Pb^{2+}(aq) + 2Cl^{-}(aq) \rightarrow PbCl_{2}(s)$ 

NA All	
.k /	
^ \	

Keyword	Definition
Electrolysis	The composition (breakdown) of a compound into its elements using a direct electric current.
Electrode	A conductor used to make contact with the electrolyte. The negatively charged electrode is called the cathode. The positively charged electrode is called the anode.
Electrolyte	A molten or aqueous ionic compound
lon	An atom that has lost electrons to become positively charged (AKA a cation) or gained electrons to become negatively charged (AKA an anion).
Aqueous	Dissolved in water
Oxidation	Loss of electrons
Reduction	Gain of electrons
Dissociation	The break up of an ionic compound into individual ions when dissolved in water.
Metal	An element from the left hand side of the periodic table that always forms a cation (a positively charged ion).
Inert	Non-reactive. Does not participate in a chemical reaction.

### Naming electrodes

Electrodes are named after the ion that they attract. Positive cations are attracted to the negative cathode. Negative anions are attracted to the positive anode.

PANIC - Positive Anode Negative Is Cathode.

### Electrolysis of molten compounds

The only source of ions are from the molten compound. Therefore a metal is always formed at the cathode and the nonmetal always forms at the anode.

EG electrolysis of KBr will form  $K + Br_2$ 

### Electrolysis of aqueous compounds

Some water molecules will dissociate to form H<sup>+</sup> and OH<sup>-</sup> ions. Therefore four ions will be present during electrolysis.

### Cathode

If the metal present is less reactive than hydrogen (ie copper, silver or gold) that metal will form at the cathode. If the metal is more reactive than hydrogen, hydrogen gas will form at the cathode.

### Anode

If there is a halogen present (F, Cl, Br or I) molecules of that gas will form at the anode, eg Cl<sub>2</sub>
If no halogen is present molecules of oxygen will form at the anode ie O<sub>2</sub>

### Half-equations (HIGHER ONLY)

- · Half-equations show us what is happening at each electrode during electrolysis
- Half-equations are balanced by adding or removing a number of electrons (e<sup>-</sup>) equal to the total charge on the ion in the equation.
- For electrolysis to the left of the arrow there is always an ion
- For electrolysis to the right of the arrow there is always an element (uncharged) [sometimes a molecule]
- · Electrons left of the arrow are added to the ion
- Electrons right of the arrow have been removed from the ion EG Half equations for electrolysis of  $Al_2O_3$ :  $Al^{3+} + 3e^- \rightarrow Al$

 $20^{2-} \rightarrow 0_2 + 4e^{-}$ 

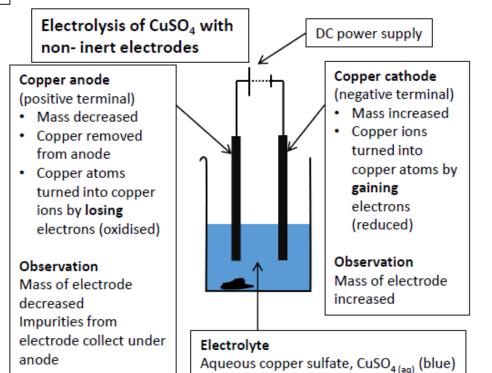
### Electrolytes

Electrolysis can only take place when ions are free to move. Therefore covalent compounds (no ions present) and solid ionic compounds (ions cannot move) cannot undergo electrolysis.

# ST IVO ACADEMY Astrea Academy Trust Astrea Academy Trust

### Electrolysis of CuSO<sub>4</sub> DC power supply with inert electrodes Graphite cathode Graphite anode (negative terminal) (positive terminal) Positive cations · Negative anions (copper ions) (hydroxide ions) attracted attracted · Cations GAIN Anions LOSE electrons electrons Therefore Therefore reduced oxidised Observation Observation Metal forms on Bubbles of gas form electrode (oxygen)

### **Electrolyte** Aqueous copper sulfate, CuSO<sub>4 (aq)</sub> (blue)



Keyword	Definition
Oxidation	A reaction in which a substance gains oxygen or loses electrons.
Reduction	A reaction in which a substance loses oxygen or gains electrons.
Ore	A rock that contains a high concentration of metal or metal compound.
Extraction	A process in which a metal in obtained from its ore.
Recycling	Converting waste products into new products.

### Reactions of metals with acids

- The more reactive the metal the more vigorously and faster hydrogen bubbles will be generated.
- Reactive metals like magnesium produce lots of hydrogen quickly (giving a loud squeaky pop when ignited).
- Relatively unreactive metals like copper will not react with dilute acids at all.

### Reactions of metals with water

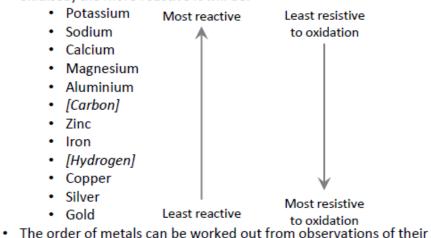
- metal + water → metal hydroxide + hydrogen
- less reactive metal + steam → metal oxide + hydrogen
- Highly reactive metals like potassium, sodium and lithium will react vigorously with cold water.
- Less reactive metals like magnesium and zinc will react with steam and slowly with cold water.
- · Copper will not react with water or steam.

### Displacement reactions are redox reactions - (HIGHER ONLY)

- Reduction and oxidation happen simultaneously during a displacement reaction - they are an example of a redox reaction.
- The more reactive metal is oxidised and the less reactive metal is reduced.

### The reactivity series

- Metals can be sorted by their reactivity. Carbon and hydrogen are commonly included.
- The more easily a metal loses an electron (the more easily it is oxidised) the more reactive it will be.



### Displacement reactions

reactions with dilute acids and water

- A more reactive metal will displace a less reactive metal from a salt. eg copper sulfate + iron → copper + iron sulfate
- A less reactive metal cannot displace a more reactive metal from a salt. eg zinc sulfate + iron → no reaction
- A series of displacement reactions can be used to sort metals into a reactivity series.
- Count the total number of time a metal displaces another. The more displacements the more reactive that metal is.

### Extraction of metals with carbon

- Ores containing metals below carbon in the reactivity series can be reduced using carbon.
  - iron oxide + carbon → iron + carbon dioxide
  - 2Fe<sub>2</sub>O<sub>3</sub> + 3C → 4Fe + 3CO<sub>2</sub>

### Ores and extraction of metals

- Many metals are found in ores in the Earth's crust chemically combined to other elements.
- Some unreactive metals (eg gold, platinum) are found as uncombined elements and only require refining before use.
- The ores are mined and the metals extracted from them. Metals that are chemically combined with other elements must be reduced to form pure metals.
- · Generally the less reactive the metal the cheaper it is to extract.

### Extraction of metals by electrolysis

- Ores containing metals above carbon in the reactivity series can be reduced using electrolysis. (See 'Electrolysis' KO)
- Electrolysis costs more than carbon extraction due to higher energy costs from the required electricity and energy needed to melt ores.
- · Aluminium is extracted by electrolysis
  - Overall reaction: 2Al<sub>2</sub>O<sub>3(I)</sub> → 4Al<sub>(I)</sub> + 3O<sub>2(g)</sub>
  - At the cathode: Al<sup>3+</sup> + 3e<sup>-</sup> → Al
     At the anode: 20<sup>2-</sup> → 0<sub>2</sub> + 4e<sup>-</sup>

### Life cycle assessment (LCA)

- A LCAs can be used to determine the potential environmental impact of a product at every stage of its 'life' from cradle to grave.
- Choice of material: eg metals require mining and extracting which uses a lot of energy and can create pollution
- Manufacture: different methods can result in more or less energy and water usage and waste production.
- Product use: some products give off more pollution than others during use eg solvent based paints are more polluting during use than water based paints.
- Disposal: some products are easily recycled (eg aluminium drinks cans) reducing landfill usage and pollution and others are harder to recycle (eg blended plastics).

### Extraction of metals by biological methods - (HIGHER ONLY)

- As the limited supply of some ores (eg copper) begin to run low new extraction methods are being developed to use ores that are uneconomically viable using traditional methods:
- Bioleaching: bacteria are used to separate metals from their ores (eg separating copper from copper sulfide). A solution from the process (the leachate) contains the metal ions that are then reduced using electrolysis or a displacement reaction.
- 2. Phytoextraction: Plants are grown in soil containing metal compounds. These compounds are absorbed and concentrations become higher in the leaves than the soil. The plants are harvested, dried, and burnt to leave the metal compounds. Metals are extracted using electrolysis or a displacement reaction.
- Advantage less environmentally damaging than traditional mining. Disadvantage – slower than traditional mining.

### Recycling

- Recycling conserves finite resources (both ores and fossil fuels that form a large source of the energy needed to extract and refine raw materials).
- Recycling uses a tiny fraction of the energy needed to extract and refine a raw material – therefore saving money.
- Raw materials commonly come from mines that damage the environment and destroy habitats. Recycling reduces the number of mines needed.
- Recycling reduces the amount of waste sent to landfills which take up space and can pollute their surroundings.
- Recycling creates more jobs than disposing of waste in landfills.
   Recyclable materials need transporting to and processing at recycling centres and reprocessing into new saleable products.
- Recycling aluminium saves: ~95% of the energy needed to mine and extract the same mass of aluminium; 4 kg of aluminium ore and associated waste.

Keyword	Definition
Dynamic equilibrium	When the forwards and backwards reactions of a revisable reaction are occurring at the same rate.
Concentration	The amount of a solute dissolved in a certain volume of a solvent. Typical units mol / dm³ (NB 1 dm³ = 1 L).
Closed system	When substances cannot enter or leave an observed environment eg a stoppered test tube.
Le Chatelier's Principle	A change of temperature, concentration or pressure to a dynamic equilibrium will result in the point of equilibrium changing to minimise the change.
Dynamic equilibrium vs equilibrium	There are different types of equilibriums. A dynamic equilibrium is specifically an equilibrium in which reactions are constantly taking place but the rates of reactions are identical. You may see both terms being used to mean the same thing during your GCSEs.

### Equilibrium

- Take the general reversible reaction A + B ⇌ C + D
- As A and B react their concentrations decrease and so the rate of the forward reaction decreases. At the same time the concentrations of C and D will increase and so the rate of the backwards reaction will increase.
- Eventually the rates of the forward and backward reactions will be the same – equilibrium has been reached.
- At equilibrium both reactions are still happening but there is no change in concentration of reactants or products.
- Equilibrium can only be reached in closed systems. Equilibriums do not usually end with equal concentrations of products.

### Reversible reactions & the Haber process

- A reversible reaction is a chemical reaction that can work in both directions. The products of one reaction are the reactants for another.
- The Haber process is a reversible reaction that can reach dynamic equilibrium:

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

- Nitrogen is taken from the air and hydrogen is obtained from natural gas.
- · The reaction conditions for the Haber process are:

1. Temperature: 450 °C

2. Pressure: 200 atm (atmospheres)

3. Iron catalyst present

### Position of Equilibrium

- At equilibrium there will typically be either more reactants (A and B - the equilibrium lies to the left) or more products (C and D - the equilibrium lies to the right).
- The exact point of equilibrium depends on the reaction occurring and the following conditions:
  - 1. Temperature
  - 2. Pressure (when gases are involved)
  - 3. Concentrations

### HIGHER ONLY - Le Chatelier's principle

• You can predict how the point of equilibrium will change when the conditions are changed. Collectively these rules form what is called Le Chatelier's principle.

Condition	Change	Equilibrium point changes to
Temperature	increased	endothermic reaction rate increases (heat absorbed from the environment).
	decreased	exothermic reaction rate increases (heat released to the environment).
Pressure of a gas	increased	rate of reaction that produces less moles of gas increases.
	decreased	rate of reaction that produces more moles of gas increases.
Concentration of a substance	increased	rate of reaction that uses up the added substance increases.
	decreased	rate of reaction that forms the removed substance increases.

### Worked example

Condition	Change	Equilibrium point changes to
Temperature	increased	endothermic reaction rate increases – more ${\rm N_2}$ and ${\rm H_2}$ produced
	decreased	exothermic reaction rate increases – more NH <sub>3</sub> produced
Pressure	increased	rate of reaction that produces less moles of gas increases – more NH <sub>3</sub> produced
	decreased	rate of reaction that produces more moles of gas increases – more $\mathrm{N}_2$ and $\mathrm{H}_2$ produced
Concentration of NH <sub>3</sub>	increased	rate of reaction that uses up the added substance increases - more $\rm N_2$ and $\rm H_2$ produced
	decreased	rate of reaction that forms the removed substance increases - more $\mathrm{NH_3}$ produced

Keyword	Definition
Batch process	A method of manufacture in which a substance is made in separate stages rather than continuously. Eg making a bottle of bleach at a time.
Continuous process	A method of manufacture in which a substance is made continuously. Eg making a continuous stream of bleach.
Fertiliser	Water soluble compounds added to soils to replace the minerals used by plants.
Yield	The amount of useful product you can get from something
Refining	The removal of impurities or unwanted elements from a substance.
Equilibrium	When the forwards and backwards reactions of a revisable reaction are occurring at the same rate. In a closed system the concentration of each compound would be constant.

### **Fertilisers**

- Nitrogen, phosphorus and potassium (N, P, K) are the three essential elements in fertilisers.
- Plants absorb these elements from the soil gradually reducing their concentration.
- Without enough of these elements a plants growth and life process are negatively affected.
- Fertilisers add (or increase the amount of) these elements to a soil increasing crop yields as plants can grow faster and larger.
- Ammonia based fertilisers have advantages over traditional fertilisers such as manure. The composition of ammonia based fertilisers can be carefully controlled as well as the amount produced.

### Conditions of industrial manufacturing - (HIGHER ONLY)

- The aim of chemical industrial manufacturing is to produce a substance at an acceptable yield and timeframe in an economically viable way (ie a profit can be made).
- Conditions (temperature, pressure, presence of a catalyst) are controlled to balance costs, yields and rates of production.
- Higher temperatures and pressures generally cost more to maintain due to higher energy demands.
- The cost of extraction and refining of raw materials also impacts the economic viability of a process.

### Predicting rate of reaching equilibrium - (HIGHER ONLY)

- Increasing the temperature and pressure (or concentration) of reactants increase the rate at which equilibrium is reached.
- Depending on the reaction being carried out these conditions could result in a higher or lower yield.
- A catalyst will increase the rate reaction (and so equilibrium will be reached faster) without affecting the yield or equilibrium point of a reaction.

### Conditions of the Haber process

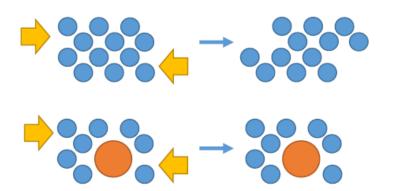
- •The Haber process is used to produce ammonia industrially
  - $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} (+ heat)$
- •Higher pressures favour the forward reaction (4 moles of gas: 2)
- •Lower temperatures favour the forwards reaction (forwards is exothermic) but slow the rate of reaction.
- •The pressure is maintained at ~200 atmospheres an acceptable compromise between costs and yield.
- •Temperature is maintained at ~450 °C an acceptable compromise between yield and rate of reaction.
- •An iron catalyst is used to reach equilibrium faster.



Keyword	Definition
Transition metal	A metal element in the block between groups 2 and 3 of the periodic table.
Alloy	A metal with one or more other elements added (usually metals) to improve its properties (eg hardness).
Corrosion	The gradual deterioration of a substance when it reacts with substances in the environment. (eg when a metal oxidises in the air).
Sacrificial protection	Using a more reactive metal to protect another metal from corrosion (eg attaching zinc blocks to iron to prevent the iron rusting).
Compound	A substance that can be split into simpler substances because it contains atoms of two or more elements chemically joined together.

### Alloys

- Pure metals are malleable due to the regular arrangement of ions in layers that can easily slide over each other
- · Some pure metals are not strong enough for certain uses
- Adding another element to a pure metal that has a different size disrupts the regular arrangement of metal ions and makes it difficult for the ions to slide over each other - creating a harder material



### Transition metals

- · The majority of metals are transition metals
- A transition metal typically has the following properties: high melting point, high density, forms coloured compounds (eg Fe<sup>2+</sup> is green, Fe<sup>3+</sup> is orange-brown) and can be used as a catalyst (eg Fe is the catalyst for the Haber process for making ammonia)

### Alloys of aluminium

- Aluminium has a low density which is suitable for aircraft
  manufacturing but is not strong enough when pure. Alloys can
  increase the strength of aluminium without increasing its density.
- Magnalium (aluminium + magnesium) at 5% magnesium magnalium is suitable for aircraft manufacturing. It is stronger, less dense and more corrosion resistant than pure aluminium.

### Alloys of copper

- Bronze (copper + tin) harder than copper. Used for decorative items such as medals, ornaments and statues
- Brass (copper + zinc) more malleable than bronze. Used where low friction is needed such as in water taps and door fittings

### Alloys of gold

- Gold is used to make jewellery when pure it is very soft and malleable.
- Metals such as zinc, copper and silver are used to make gold alloys that are harder than pure gold and have a different colour.

### Alloys of iron

- Alloys of iron, called steels, have different properties depending on what is added to the iron.
- · Adding carbon to iron produces steel. It is harder than iron.
- Adding carbon and another metal produces a steel alloy. Adding chromium (and sometimes nickel) produces stainless steel. Unlike iron, stainless steel is resistant to corrosion.

### Corrosion

- Most metals will corrode when exposed to oxygen and water to form metal oxides.
- Corrosion of metals is a redox reaction. The metal is oxidised (it loses electrons) and oxygen is reduced (it gains electrons).

### Corrosion - Prevention of rusting

- Rusting is the corrosion of iron. Both oxygen and water are required for rusting to happen. Preventing either of these from contacting the iron will stop rusting.
- Coating the iron with a barrier will prevent it from rusting. Paint
  prevents rusting for non-moving parts and oil/grease can be used
  for moving parts like bike chains.
- A more reactive metal than iron (eg zinc) can be used as a sacrificial protection by attaching it to the iron. Electrons are taken from the zinc in preference to the iron preventing rusting.
- Galvanising is an example of sacrificial protection. A barrier of a more reactive metal is coated over the iron (eg steel buckets are often galvanised).

### Corrosion prevention - Electroplating

- Electroplating is process of coating the surface of a metal object with another metal using electrolysis.
- Electroplating is typically carried out to improve corrosion resistance or the appearance of the metal object
- Gold and silver plated objects (eg jewellery) can be made this way.
  - The cathode (negatively charged) is the object to be plated and the anode (positively charged) is a bar of the metal that will form the plate. The electrolyte contains ions of the metal that will form the plate.

### GCSE

**CHEMISTRY: Fertilisers and the Haber process** 

### Ammonia based fertilisers

- Ammonia is used to produce nitrogen containing compounds that are used as fertilisers.
- Reacting ammonia with different chemicals will produce different fertilising compounds.
- Reactions of ammonia with water and oxygen are used to produce nitric acid (HNO<sub>3</sub>) which is then used to produce a fertiliser.
- nitric acid + ammonia → ammonium nitrate (a fertiliser)
   HNO<sub>3(aq)</sub> + NH<sub>3(aq)</sub> → NH<sub>4</sub>NO<sub>3(aq)</sub>
- sulfuric acid + ammonia → ammonium sulfate
   H<sub>2</sub>SO<sub>4(aq)</sub> + NH<sub>3(aq)</sub> → (NH<sub>4</sub>)<sub>2</sub>SO<sub>4(aq)</sub>

Production of ammonia sulfate	Small scale (lab)	Large scale (industrial)
Starting materials	Ammonia solution and dilute sulfuric acid	Raw materials of ammonia and sulfuric acid
Stages	Titration – methyl orange indicator, ammonia solution in flask, sulfuric acid in burette (see 'Acids' KO for method) Crystallisation	Several stages
Type of process	Batch	Continuous











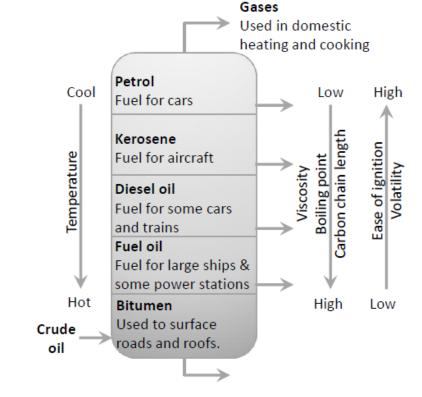
Keyword	Definition
Crude oil	A liquid fossil fuel. Formed from remains of dead plants and animals that lived in the sea. It is a complex mixture of hydrocarbons. It is a finite resource.
Finite resource	Something useful that is no longer made or that is being made very slowly.
Hydrocarbon	A compound containing <b>only</b> hydrogen and carbon.
Viscosity	How thick or runny a liquid is. High viscosity is thick eg honey, low viscosity is runny eg water.
Homologous series	A family of compounds with similar chemical properties and a gradual variation in physical properties (eg boiling points). They have the same general formula with successive members differing in their molecular formulae by CH <sub>2</sub> .
General formula	A type of formula that represents the composition of any member of an homologous series eg Alkanes – $C_nH_{2n+2}$
Saturated (relating to hydrocarbons)	A molecule containing only single bonds between carbon atoms in a chain. Alkanes.
Unsaturated (relating to hydrocarbons)	A molecule that contains one or more double bonds between carbon atoms in a chain. Alkenes.

### Hydrogen as a fossil fuel alternative

- · Hydrogen can be used to fuel vehicles
- Advantages: very clean (only product of combustion is water). Obtained from water (a renewable resource).
- Disadvantages: Special (expensive) engine needed (a fuel cell). It needs manufacturing; expensive and energy often comes from fossil fuels. Difficult to store and not widely available.

### Crude oil and fractional distillation

- Crude oil contains a mixture of hydrocarbons where the carbon atoms form rings or chains. It is an important source of useful substances (fuels and feedstock for the petrochemical industry).
- Fractional distillation is used to separate crude oil into fractions (simpler mixtures of hydrocarbons with similar chain lengths) using the differing boiling points of the fractions.
- · Crude oil is heated until it is mostly gas (only bitumen remains liquid). The lower the boiling point of a fraction the higher up the tower it travels before it is cool enough to condense to a liquid and be tapped off.



- · Good Penguins Keep Diving For Bass
- Gas Petrol Kerosene Diesel Fuel oil Bitumen



## ACADEMY TASK

### Combustion - complete and incomplete

- · Burning hydrocarbons releases a lot of energy.
- Complete combustion occurs when hydrocarbons are burnt in excess oxygen. Only water and carbon dioxide is produced:
  - hydrocarbon + oxygen → carbon dioxide + water
- Incomplete combustion occurs when hydrocarbons are burnt in a limited supply of oxygen:
  - hydrocarbon + oxygen → carbon dioxide + water
     + carbon + carbon monoxide
- Carbon from incomplete combustion is called soot. It is a black dust that makes buildings look dirty and can worsen respiratory problems like asthma.
- Carbon monoxide is a toxic gas. It combines with haemoglobin in red blood cells and prevents oxygen being transported around the body. In severe cases this leads to death.

### Acid rain

- When burnt sulfur impurities in fossil fuels form sulfur dioxide (SO<sub>2</sub>).
- Sulfur dioxide will react with water in the atmosphere forming sulfuric acid – aka acid rain.
- Acid rain can cause lake and rivers to become acidic; killing plants and animals. Acid rain also damages limestone buildings and corrodes metals.

### Nitrogen oxides

- Nitrogen oxides (NO<sub>x</sub>) form from oxygen and nitrogen in the air at high temperatures when hydrocarbons are burnt eg in car engines – they are pollutants.
- NO<sub>x</sub> contribute to acid rain and photochemical smog at ground level (a type of air pollution that can cause breathing difficulties and other issues).

### Cracking

- Cracking turns unwanted long alkanes (saturated) into shorter, more useful alkanes and alkenes (unsaturated – can be made into plastics).
- Cracking helps match supply and demand eg more bitumen is produced than is needed but less petrol is produced than is needed.
- Cracking is a form of thermal decomposition. One molecule breaks down into two or more smaller molecules when heated.
   A catalyst is used to increase the rate of the reaction.
- The total number of each element is the same before and after cracking eg
   C<sub>10</sub>H<sub>22</sub> → C<sub>8</sub>H<sub>18</sub> + C<sub>2</sub>H<sub>4</sub>

### Properties of hydrocarbons

- Alkanes (saturated, C<sub>n</sub>H<sub>2n+2</sub> eg C<sub>2</sub>H<sub>6</sub>) and alkenes (unsaturated, C<sub>n</sub>H<sub>2n</sub> eg C<sub>2</sub>H<sub>4</sub>) are two different homologous series found in hydrocarbons.
- All molecules within a homologous series have similar chemical properties.
- The physical properties change gradually with the number of carbon atoms in the molecule.
- Shorter molecules have weaker intermolecular forces that need less energy to overcome so are less viscous, easier to ignite and have lower boiling points.
- Longer molecules have stronger intermolecular forces that need more energy to overcome so are more viscous, harder to ignite and have higher boiling points.

Can you define homologous series?

Keyword	Definition
combustion	A chemical reaction in which a compound reacts with oxygen.
incomplete combustion	When a substance reacts only partially with oxygen, such as when carbon burns in air producing carbon monoxide and soot (unburnt carbon).
oxidation	A reaction in which oxygen is added to a substance; loss of electrons by an atom or negative ion
toxic	Poisonous.
acid rain	Rainwater that is more acidic than usual due to air pollution, usually caused by sulfur dioxide and nitrogen oxides.
oxide of nitrogen	Any one of a variety of gaseous compounds consisting of only nitrogen and oxygen atoms. Together they are often represented as NO <sub>x</sub> .
pollutant	A substance that harms living organisms when released into the environment.
weathering	When rocks are broken up by physical, chemical or biological processes.
greenhouse gas	A gas that helps to trap 'heat' in the atmosphere. Carbon dioxide, methane and water vapour are greenhouse gases.
atmosphere	The layer of gases that surrounds the Earth.
climate change	Changes that happen to the global weather patterns as a result of global warming.

### Fuel + oxygen → carbon dioxide + water

Combustion is an oxidation reaction. Energy is also released (Exothermic) . Complete combustion occurs when there is plenty of oxygen for the reaction. Incomplete combustion happens when there is not enough oxygen to allow a fuel to burn completely:

Fuel + oxygen → carbon dioxide + carbon monoxide + carbon + water

Product	Problem
Carbon dioxide	Greenhouse gas
Carbon monoxide	Toxic – binds irreversibly to haemoglobin so it can't carry oxygen, reduces oxygen in blood: tiredness, coma and death.
Carbon (soot)	can block flues, makes buildings dirty, respiratory problems and reduced air quality
Water	Damages decorations

When buring fossil fuels there are also impurities Sulphur + oxygen → sulphur dioxide (causes acid rain) Nitrogen + oxygen → oxides of Nitrogen (acid rain and photochemical smog – a type of air pollution that causes breathing difficulties)

### Acid rain

Causes: acidic gases which dissolve in rain water

Examples: sulfur dioxide, oxides of nitrogen

Problems: makes lakes and rivers acidic, which harms fish and other life. It also damages trees, buildings made of limestone or marble, and makes metal corrode.

Solutions: 1) removing sulfur from car fuels; 2) by removing sulfur dioxide from power station waste gases.

### Climate change

Causes: greenhouse gases trap heat (Infrared radiation) that is radiated from the earth, keeping earth warm Activities that release greenhouse gases: burning fuel, cows (digestive processes)

and rice fields

Increased human population means an increased demand for energy, greater deforestation (demand for space, fewer trees taking CO<sub>2</sub> from atmosphere) Examples: carbon dioxide, methane and water vapour

Effects: Global warming

Solutions: Reduce use of fossil fuels and

be more energy efficient

### The Evidence

The evidence for climate change shows a correlation between human activities such as burning fossil fuels and an increase of temperature. However there are some uncertainties caused by the location where these measurements were taken and historical accuracy. We can estimate by looking at: fossils, tree rings, gas bubbles in ice. (Less precise)



### History of the Atmosphere Stage Gases Large amounts of Carbon dioxide, ammonia, nitrogen and Volcanoes (Hot) water vapour - no oxygen. Earth cooled Water vapour condense to make oceans. Carbon dioxide dissolved in oceans Photosynthesised, absorbed carbon dioxide, gave out Marine plants evolved oxygen Some carbon dioxide taken in to make shells, some shells Marine organisms form sedimentary carbonate rocks evolved

### Current Atmosphere

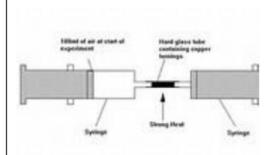
Our atmosphere 78% Nitrogen, 21% oxygen, 1% Argon, 0.04% carbon dioxide Processes that affect the amount of gasses in our atmosphere:

		L
Take in CO₂	Release CO <sub>2</sub>	
Photosynthesis	Respiration, burning fossil fuels, deforestation,	
		Г

### Measuring the amount of Oxygen in the air:

Various experiments exist to demonstrate the amount of oxygen in the air by calculating the percentage of air used when a substance is burned.

Here is one example:



Heat a metal in air, push / pull syringes to move air

- volume of air in syringes decreases as oxygen removed from air as metal reacts with oxygen, metal oxide formed • continue until no further change and allow (apparatus) to cool
- final reading on gas syringe will decrease 21 cm<sup>3</sup> if you start with 100cm<sup>3</sup> as oxygen has reacted, shows it's 21%



Keyword	Definition
Polymer	A substance of high average relative molecular mass made up of small repeating units called monomers.
Functional group	An atom or group of atoms that are mainly responsible for a molecule's chemical properties
Biodegradable	A substance that can be broken down by microorganisms
Organic compound	A compound with a central framework of carbon atoms onto which hydrogen and other molecules are attached. $CH_4$ organic. $CO_2$ is not as it contains no hydrogens.
Homologous series	A series of compounds that fit a general formula, have similar chemical properties and differ in structure by a number of CH <sub>2</sub> units
Hydrocarbons	A compound containing only carbon and hydrogen atoms.
Carbohydrates	A group of compounds made of C, H and O eg sugars like glucose.

### Recycling polymers

### Advantages

- · Reduces amount of non-biodegradable waste in landfill
- · Reduces emissions of greenhouse and toxic gases released through combustion
- · Generally uses less water, energy and crude oil than making 'new' polymers from raw materials
- · Generally saves money and creates jobs.

### Disadvantages

- Polymers must be separated by type before being recycled - can be difficult and expensive
- If polymers are mixed the final recycled polymer quality can be reduced
- · Melting polymers can release toxic gases
- Polymers can only be recycled a finite number of times.

### Hydrocarbon combustion

- · Hydrocarbons undergo combustion reactions with oxygen.
- · See 'Fuels' knowledge organiser for details.

### Covalent bonds

Lines between element symbols in a displayed formula represents a covalent bond (See 'Bonding' KO)

	No. of C		Alkane (C <sub>n</sub> H <sub>2n+2</sub> )		Alkene (C <sub>n</sub> H <sub>2n</sub> )			
		Prefix	Name	Molecular formula	Displayed formula	Name	Molecular formula	Displayed formula
	1	Meth-	Methane	CH <sub>4</sub>	Н  - Н—С—Н  - Н	-	-	-
	2	Eth-	Ethane	C <sub>2</sub> H <sub>6</sub>	H H     H-C-C-H     H H	Ethene	C <sub>2</sub> H <sub>4</sub>	H H C=C H H
	3	Prop-	Propane	C <sub>3</sub> H <sub>8</sub>	H H H H-C-C-C-H H H H	Propene	C <sub>3</sub> H <sub>6</sub>	H H H H-C-C=C H H
	4	But-	Butane	C <sub>4</sub> H <sub>10</sub>	H H H H H-C-C-C-C-H	Butene	C <sub>4</sub> H <sub>8</sub> C=C can be at the end or centre of molecule	H H H H H-C-C-C=C

### Hydrocarbons - saturated and unsaturated

- · Alkanes are saturated hydrocarbons - they contain only single bonds between carbon atoms
- · Alkenes are unsaturated hydrocarbons - they contain one or more double bonds between carbon atoms
- See 'Fuels' knowledge organiser for details.



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### Testing for alkenes - Bromine water

- Adding bromine water to an alkene and shaking will result in a colour change. Orange → colourless.
- This is an addition reaction (bromine is 'added' to the alkene).

### Addition polymerisation

- Molecules containing a C=C bond can open up this bond and join together to form a polymer.
- They are named by adding 'poly' in front of the monomer name in brackets.

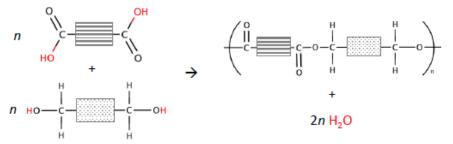
- If given a monomer you need to be able to draw the polymer. If given a polymer you need to be able to draw a monomer.

### Condensation polymerisation - (HIGHER ONLY)

 Typically two different monomers react to form a polymer and a small molecule (eg water) is formed for every bond formed in the polymer chain.

### **Polyesters**

 Formed between dicarboxylic acid and diol (compound containing two alcohol functional groups) monomers.



### Problems with polymers

- Starting materials are commonly from crude oil – (see 'Fuel' Knowledge organiser for issues with this).
- Most are non-biodegradable. Filling landfills and taking up valuable land.
- Disposal by combustion (to avoid landfill and produce energy) often releases toxic gases like HCl and greenhouse gases like CO<sub>2</sub>.

### Naturally occurring polymers

- DNA is a polymer made from four different monomers called nucleotides.
- Starch is a polymer formed from glucose monomers.
- Proteins are polymers based on amino acid monomers.

Properties and uses of polymers - Each polymer has its own properties making them suitable for different uses

Polymer	Properties	Uses
Poly(ethene)	Cheap, electrical insulator, flexible	Wire insulation, bags, bottles
Poly(propene)	Mouldable, flexible, strong, tough	Crates, furniture, pipes
Poly(chloroethane) (aka PVC)	Cheap, tough	Window frames, water pipes
Poly(tetrafluoroethene) (aka PTFE)	Tough, unreactive, non-stick	Non-stick pans, waterproof clothing

	Carboxylic acids				
Name	Molecular formula	Displayed formula			
Methanoic acid	CH₃OOH	H-C,0			
Ethanoic acid	CH₃COOH	H-C-C-0-H			
Propanoic acid	C₂H₅COOH	н-с-с-с			
Butanoic acid	C₃H <sub>7</sub> COOH	H H H H O O H			

### Alcohols as fuels

- When burnt, alcohols release energy and so can be used as fuels.
- Some alcohols are better fuels than others – they release more energy per gram of alcohol burnt.
- To compare different alcohols as fuels you can heat a known volume of water using different alcohols. See 'Chemistry practicals' KO for details.
- The longer the carbon chain in an alcohol the more efficient it will be as a fuel.

### Alcohols

- Alcohols are a homologous series. containing the functional group –OH.
- Alcohols can be dehydrated (ie have water removed from them) to form alkenes.

### Homologous series

- Members of a homologous series will react in similar ways to each other because they all contain the same functional group
- If you know how one member of a series reacts you can predict how all members of that series react

### Carboxylic acids

- Carboxylic acids are a homologous series containing the function group –COOH.
- They can be formed from the oxidation of certain alcohols.
  - Oxidising methanol makes methanoic acid
  - Oxidising ethanol makes ethanoic acid
  - etc
- They have typical properties of acids eg form solutions with pH <7</li>

### Production of ethanol by fermentation

- Fermentation uses yeast (containing natural enzymes) to convert carbohydrates to ethanol (and CO<sub>2</sub>).
- Mix yeast and carbohydrate solution in sterile container. Seal container.
- Store at 30 40 °C. Lower temperatures slow reaction, higher denatures enzyme stopping reaction.
- The container must be sealed to maintain an anaerobic atmosphere (no oxygen present). If oxygen entered the ethanol would oxidise to form ethanoic acid.
- When the ethanol concentration reaches 10-20 % the yeast will be killed by the ethanol.
- The yeast will settle to the bottom and the alcohol solution can be collected.

### Concentrating ethanol

- Increasing ethanol concentration to >20% can be achieved by fractional distillation of the fermentation mixture.
- Ethanol has a lower boiling point than water so will evaporate and be collected before the water evaporates.
- See 'States of matter and mixtures' KO for more on fractional distillation.

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### **Development of the Periodic Table**

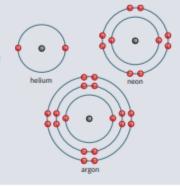
The Periodic Table has changed over time as scientists have organised it differently. Mendeleev was able to accurately predict the properties of undiscovered elements based on the gaps in the table.

	First lists of elements	Mendeleev's Periodic Table	Modern Periodic Table
How are elements ordered?	by atomic mass	normally by atomic mass but some elements were swapped around	by atomic number
Are there gaps?	по даре	gaps left for undiscovered elements	no gaps – all elements up to a certain atomic number have been discovered
How are elements grouped?	not grouped	grouped by chemical properties	grouped by the number of electrons in the outer shells
Metals and non-metals	no clear distinction	no clear distinction	metals to the left, non-metals to the right
Problems	some elements grouped inappropriately	incomplete, with no explanation for why some elements had to be swapped to fit in the appropriate groups	_

### Group 0

Elements in **Group 0** are called the **noble gases**. Noble gases have the following properties:

- full outer shells with eight electrons, so do not need to lose or gain electrons
- are very unreactive (inert) so exist as single atoms as they do not bond to form molecules
- · boiling points that increase down the group.



### (P) Key terms

Make sure you can write a definition for these key terms.

alkali metals chemical properties displacement groups halogens inert isotopes noble gas organised Periodic Table reactivity undiscovered unreactive

### **Group 1 elements**

Group 1 elements react with oxygen, chlorine, and water, for example:

lithium + oxygen → lithium oxide

lithium + chlorine → lithium chloride

lithium + water → lithium hydroxide + hydrogen

Group 1 elements are called **alkali metals** because they react with water to form an alkali (a solution of their metal hydroxide).

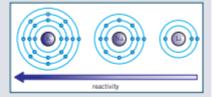


### **Group 1 properties**

Group 1 elements all have one electron in their outer shell.

Reactivity increases down Group 1 because as you move down the group:

- · the atoms increase in size
- the outer electron is further away from the nucleus, and there are more shells shielding the outer electron from the nucleus
- the electrostatic attraction between the nucleus and the outer electron is weaker so it is easier to lose the one outer electron
- · the melting point and boiling point decreases down Group 1.



### Group 7 elements

Group 7 elements are called the halogens. They are non-metals that exist as molecules made up of pairs of atoms.

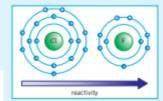
	Name	Formula	State at room temperature	Melting point and boiling point	Reactivity
Г	fluorine	F <sub>2</sub>	gas		
Г	chlorine	Cl <sub>2</sub>	gas		d
Г	bromine	Br <sub>2</sub>	liquid	increases down the group	decreases down the group
Г	iodine	I <sub>2</sub>	solid		



### Group 7 reactivity

Reactivity decreases down Group 7 because as you move down the group:

- the atoms increase in size
- the outer shell is further away from the nucleus, and there are more shells between the nucleus and the outer shell
- the electrostatic attraction from the nucleus to the outer shell is weaker so it is harder to gain one electron to fill the outer shell.



### Group 7 displacement

More reactive Group 7 elements can take the place of less reactive ones in a compound. This is called **displacement**.

For example, fluorine displaces chlorine as it is more reactive:

fluorine + potassium chloride → potassium fluoride + chlorine

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Word	Meaning
alkali metals	A group of very reactive metals found in group 1 of the periodic table.
group (chemistry)	A vertical column of elements in the periodic table. Elements in the same group generally have similar properties.
periodic table	The chart in which the elements are arranged in order of increasing atomic number.
reactivity	A description of how quickly or vigorously something reacts.

### Alkali Metals

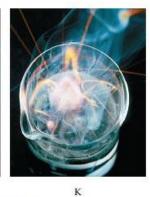
- · React with oxygen to form oxides
- React with water to form the hydroxide and hydrogen
- · React with chlorine to form chlorides



reactivity



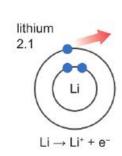
Na

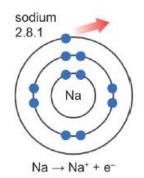


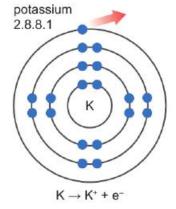
4K(s)	+	O <sub>2</sub> (g)	$\rightarrow$	2K <sub>2</sub> O(s)
potassium	+	oxygen	$\rightarrow$	potassium oxide

$$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$$
  
lithium + water  $\rightarrow$  lithium hydroxide + hydrogen

	lithium + water	bubbles fiercely on the surface
	sodium + water	melts into a ball and fizzes about the surface
7	potassium + water	bursts into flames and flies about the surface



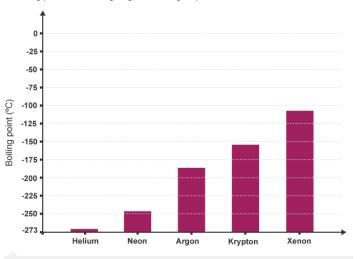




### **Boiling points**

The noble gases all have low boiling points:

- helium, at the top of group 0, has the lowest boiling point of any element
- boiling point increases going down the group



The boiling points of group 0 elements

This is because, going down group 0:

- the atoms become larger
- the intermolecular forces between the atoms become stronger
- more <u>energy</u> is needed to overcome these forces

### Properties and uses of noble gases

The main **properties** of the noble gases include:

- they have low **densities**
- they are inert, so they are not **flammable**

Many uses of the noble gases are linked to one or more of these properties.

### Helium

Helium is used as a lifting gas in party balloons and airships. Helium is:

- less dense than air, so balloons and airships rise
- non-flammable so the helium cannot set on fire

### Argon

Argon is used as a 'shield gas' when **welding** pieces of metal together. Argon is:

- denser than air, so it stops air getting to the metal
- inert, so the hot metal cannot oxidise and spoil the weld

### Question

Filament lamps contain thin metal wires. These become very hot and glow brightly when an <u>electric current</u> is passed through them. Explain why argon, krypton or xenon are used in these lamps.



The alkali metals share similar **<u>physical</u>** and **<u>chemical properties</u>**.

### Physical properties

The alkali metals:

- are soft (they can be cut with a knife)
- have relatively low melting points

The alkali metals also have low <u>densities</u>. They are low enough for the first three (lithium, sodium and potassium) to float on water.

### Example

The table shows the melting points of five alkali metals. Use this information to describe how the melting point changes in group 1

Element	Melting point
Lithium, Li	180°C
Sodium, Na	98°C
Potassium, K	63°C
Rubidium, Rb	39°C
Caesium, Cs	28°C

Word	Meaning
bleach	To take the colour out of something.
diatomic	Two atoms chemically bonded together.
disinfectant	Something that destroys or neutralises disease-
	carrying microorganisms.
halide	A compound formed between a halogen and another
	element such as a metal or hydrogen.
halogen	An element in group 7 of the periodic table.
salt	A compound formed by neutralisation of an acid by a
	base.



### Physical properties

The halogens exist as simple <u>molecules</u>. Each molecule contains two halogen <u>atoms</u> joined by a single <u>covalent bond</u>. The table shows the colour and physical <u>states</u> of chlorine, bromine and iodine at room temperature.

Element	Colour	State
Chlorine, Cl <sub>2</sub>	Pale green	Gas
Bromine, Br <sub>2</sub>	Brown	Liquid
lodine, I <sub>2</sub>	Purple-black	Solid

The **melting points** and **boiling points** of the halogens increase going down group 7. This is because, going down group 7:

- the molecules become larger
- the intermolecular forces become stronger
- more energy is needed to overcome these forces

### Chemical test for chlorine

Damp <u>litmus paper</u> is bleached white when it is placed in chlorine. If damp blue litmus paper is used, the paper turns red then white.

### Chemical properties of the halogens

Atoms of group 7 elements all have seven electrons in their outer shell. This means that the halogens all have similar chemical properties.

### Reactions with metals

The halogens react with metals to produce <u>salts</u> (the word 'halogen' means 'salt former'). For example, chlorine reacts with sodium:

Sodium + chlorine → sodium chloride

 $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$ 

Sodium and chlorine react vigorously when heated, giving an orange flame and clouds of white sodium chloride.

The halogens become less <u>reactive</u> going down group 7. The table describes what is seen when halogens react with iron wool.

Halogen	Reaction
Fluorine	Cold iron wool burns to produce white iron(III) fluoride
Chlorine	Hot iron wool burns vigorously to produce orange-brown iron(III) chloride
Bromine	Hot iron wool burns quickly to produce red-brown iron(III) bromide
lodine	Hot iron wool reacts slowly in iodine vapour to produce grey iron(II) iodide

### Reactions with hydrogen

The halogens react with hydrogen to produce <u>compounds</u> called hydrogen halides. For example, chlorine reacts with hydrogen:

Hydrogen + chlorine  $\rightarrow$  hydrogen chloride

 $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ 

The hydrogen halides are gases at room temperature. They dissolve in water to produce acidic solutions. Hydrogen chloride dissolves in water to produce hydrochloric acid, HCl(aq).

The table describes what is seen when halogens react with hydrogen.

Halogen	Reaction
Fluorine	Explodes in the cold and dark, forming hydrogen fluoride
Chlorine	Explodes with a flame or in sunlight, forming hydrogen chloride
Bromine	Vigorous reaction with burning hydrogen, forming hydrogen bromide
lodine	Very slow reaction when heated strongly, forming some hydrogen iodide

### Halogen displacement reactions

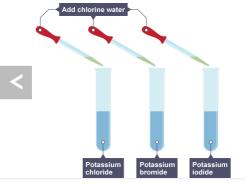
A more <u>reactive halogen</u> can <u>displace</u> a less reactive halogen from <u>solutions</u> of its <u>salts</u>. For example, chlorine is more reactive than iodine. A solution of chlorine can displace iodine from potassium iodide solution:

chlorine + potassium iodide → potassium chloride + iodine

 $Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$ 

The reaction mixture turns darker and iodine solution forms.

The slideshow shows what happens when solutions of chlorine, bromine and iodine are added to various potassium **halide** salts.



### Determining a reactivity series

A <u>reactivity series</u> can be produced by attempting some <u>displacement reactions</u>. Different combinations of halogen solution and salt solution are tested. The table shows the results of these tests. A halogen cannot displace itself from a solution of one of its salts, so these three tests were not done.

In the table below, the following formula names are used:

- KCl potassium chloride solution
- KBr potassium bromide solution
- KI Potassium iodide solution

	KCl solution	KBr solution	Kl solution	Reactions
Chlorine	Not done	Solution darkens	Solution darkens	2
Bromine	No visible reaction	Not done	Solution darkens	1
lodine	No visible reaction	No visible reaction	Not done	0

### Example

Use the results in the table to deduce an order of reactivity, starting with the most reactive halogen.

The order of reactivity is chlorine > bromine > iodine. This is because chlorine could displace bromine and iodine, bromine could only displace iodine, but iodine could not displace chlorine or bromine.

### Explaining the trend

When an element in group 7 takes part in a reaction, its <u>atom's</u> outer shells gain an <u>electron</u> and form negatively charged <u>ions</u>, called <u>anions</u>. The less easily these anions form, the less reactive the halogen.

Going down group 7:

- the atoms become larger
- the outer shell becomes further from the nucleus
- the force of attraction between the nucleus and the outer shell decreases
- an outer electron is gained less easily
- the halogen becomes less reactive

### Displacement reactions as redox reactions - Higher

A <u>balanced equation</u> for the reaction between chlorine and potassium bromide solution can be written in terms of the ions involved:

$$Cl_2(aq) + 2K^+(aq) + 2Br^-(aq) \rightarrow 2K^+(aq) + 2Cl^-(aq) + Br_2(aq)$$

Notice that potassium ions,  $K^+$ , appear on both sides of the equation. They do not take part in the reaction and are called **spectator ions**. The equation can be rewritten without them:

$$Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$$

This equation is an example of a balanced **ionic equation**. It can be split into two **half equations**:

$$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$$
, (reduction)

$$2Br^{-}(aq) \rightarrow Br_{2}(aq) + 2e^{-}$$
, (oxidation)

Notice that:

- chlorine atoms (in chlorine molecules) gain electrons they are reduced
- bromide ions lose electrons and form molecules they are oxidised

**Reduction** and **oxidation** happen at the same time, so the reactions are called **redox reactions**.

Displacement reactions are just one example of redox reactions. **Electrolysis** reactions are also redox reactions.



Meaning	
A substance formed in a reaction.	$\Box$
How quickly something happens.	П
A substance used up in a chemical reaction.	П
A factor that can change.	$\Box$
The minimum amount of energy needed to start a reaction.	
	A substance formed in a reaction.  How quickly something happens.  A substance used up in a chemical reaction.  A factor that can change.  The minimum amount of energy needed to start a

Rate of reaction is the speed at which reactants are turned into products

Some reactions are fast (like explosions)

Some reactions are slow (like rusting)

If we alter variables in a reaction, we can control the rate

Faster reaction — steeper part of graph

Curve gets less steep as reaction slows

Reaction stops when curve flattens

For chemical reactions to occur, particles must COLLIDE with enough energy to react

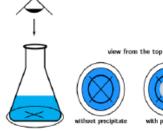
Effect of concentration on rate	Increasing the concentration increases the rate because there are more particles so there are more collisions and more reactions.
Effect of surface area on rate	Increasing the surface area (by decreasing particle sizes) in creases the rate by exposing more particles to collisions leading to more collisions and more reactions.
Effect of pressure on rate	Increasing the pressure increases the rate because particles are pushed closer together so they collide more often.
Effect of temperature on rate	Increasing the temperature increases the rate because particles move faster so they collide more, and collide with more energy to a greater proportion of collisions lead to reactions.

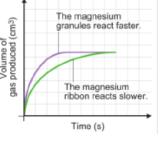
Measuring rates –	- Collect gas in a gas syringe and measure the volume every 30 secs.	1
reactions that produce gas	- Collect gas over water (up-turned measuring cylinder full of water) and measure volume every 30 secs.	
	- Do reaction on a balance and record the change in mass every 30 secs.	ŀ
Measuring rates –	Do the reaction in a beaker placed on piece of paper with a cross marked on it. Looking down through the beaker,	1
reactions that form a	time how it takes for the cross to disappear.	
precipitate ( go cloudy)		











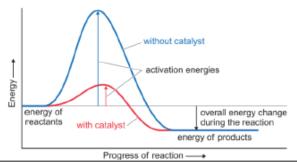
Smaller pieces of solid reactant-more gas is produced

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Meaning The space in an enzyme where the substrate fits
The enace in an enzyme where the substrate fits
The space in an enzyme where the substrate his
during an enzyme-catalysed reaction.
A substance that increases the rate of a reaction
without itself being used up.
An enzyme in which the shape of the active site has
changed so much that its substrate no longer fits and
the reaction can no longer happen.
A protein produced by living organisms that acts as a
catalyst to increase the rate of a reaction.
A polymer made up of amino acids.
A diagram to show how the energy stored in
substances changes during a chemical reaction.
A substance that is changed during a reaction.

Catalysts provide an alternative reaction route that requires less activation energy.. As less energy is needed to start the reaction, more reactant molecules have enough energy y = more collisions = faster reaction

Reaction profiles show the energy changes. Overall energy does not change



exothermic reaction

Progress of reaction

endothermic reaction

Progress of reaction

energy given out products

products

energy taken in

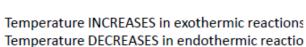
reactants

Energy is transferred between surroundings and reactants in chemical reactions- most often by heat

Word	Meaning	1
endothermic	A type of reaction in which energy from the surroundings is transferred to the products.	
	The products have more stored energy than	
	the reactants have.	
exothermic	A type of reaction in which energy is	I
	transferred to the surroundings from the	
	reactants. The products have less stored energy than the reactants have.	

Measuring	-Sit a polystyrene beaker inside a glass beaker
energy changes	(insulation)
	- Measure the starting temperature of the reactants.
	- Mix the reactants in the polystyrene beaker
	- Cover with lid fitted with a thermometer
	- Monitor and record the lowest temperature.

Endothermic Reaction Exothermic Reaction Activation nergy f Energy of reactants Energy of reactant Direction of reaction Direction of reaction



Bond breaking- Endothermic

Bond making- Exothermic

Exothermic reaction - more energy given out making bonds than needed to break bonds Endothermic reaction - more energy needed to break bonds than is given out when bonds made

Calculating	The energy change in a reaction is the difference between the energy required to
energy	break the old bonds and the energy released by making the new ones.
changes from	Add up the total strength of old bonds broken and subtract the total strength of
bond	new bonds made. You will be given the data to use
strengths (H)	A negative answer is exothermic.

Endothermic	- Dissolving (most) salts
reaction	- Some precipitation
examples	- Photosynthesis
Exothermic	- Neutralisation
reaction	- Displacement
examples	- Combustion
	- Some precipitation
	- Respiration

reactants

Keyword	Definition		
lon	An atom that has lost electrons to become positively charged (AKA a cation) or gained electrons to become negatively charged (AKA an anion).		
Precipitate	An insoluble product formed when solutions of two soluble reactants are mixed.		
Flame photometer	A machine used to identify metal ions in solution and determine their concentration.		
Standard solution	A solution with a known concentration of a substance.		
Halide	A compound containing an element from group 7 (ie F, Cl, Br or I)		
Calibration curve	1		
Emission spectrum	The spectrum of light emitted by an atom, ion or molecule.		
Aqueous (aq)	Dissolved in water		

### Testing for ions

Each test is unique. An observed result only indicates one ion is present. A test with a result that indicate one or more different ions is not a useful test.

### Instrumental analysis

These tend to be faster, more accurate and have a higher sensitivity (can detect smaller amount of ions) than flame tests or simple chemical tests.



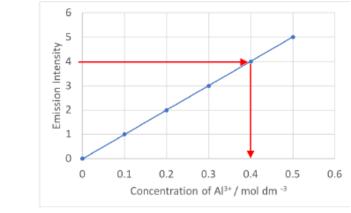
### Flame photometry

Allows the identification of ions in dilute solutions by comparison of data with reference data.

Can be used on mixtures of ions (unlike flame tests).

Concentrations of ions can be calculated using intensity of light and a calibration curve.

The calibration curve below shows that at an emission intensity of 4 the concentration of Al<sup>3+</sup> ions is 0.4 mol dm<sup>-3</sup>



### Flame test method

- Clean a nichrome wire loop by dipping in acid then rinse in distilled water.
- 2. Dip the wire loop in a sample of the metal compound
- Put the wire loop in the blue flame of a Bunsen burner and record the colour observed

### Anion tests - addition of acid

Acid is added to all anion tests to remove any compounds that could give a misleading result.

For example, when testing for halides, acid is added before the silver nitrate solution so that any carbonate ions in solution are removed otherwise they form a precipitate that would invalidate the test.

Cation		Flame test	Add sodium hydroxide	
Lithium ions	Li+	Red flame	-	
Sodium ions	Na <sup>+</sup>	Yellow flame	S	
Potassium ions	K+	Lilac flame	-	
Calcium ions	Ca <sup>2+</sup>	Orange-red flame	White precipitate (does not dissolve when more NaOH added) $Ca^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Ca(OH)_{2(s)}$	
Copper ions	Cu <sup>2+</sup>	Blue-green flame	Blue precipitate $Cu^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Cu(OH)_{2(s)}$	
Aluminium ions	Al³+	-	White precipitate (dissolves when more NaOH added) $Al^{3+}_{(aq)} + 3OH^{-}_{(aq)} \rightarrow Al(OH)_{3(s)}$	
Iron (II) ions	Fe <sup>2+</sup>	-	Green precipitate $Fe^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Fe(OH)_{2(s)}$	
Iron (III) ions	Fe³+	-	Brown precipitate $Fe^{3+}_{(aq)} + 3OH^{-}_{(aq)} \rightarrow Fe(OH)_{3(s)}$	
Ammonium	NH <sub>4</sub> <sup>+</sup>	-	Ammonia gas released – will turn damp red litmus paper blue	

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Anion		Add dilute acid	Add barium chloride (aq) with hydrochloric acid	Add silver nitrate (aq) with nitric acid
Carbonate ions	CO <sub>3</sub> <sup>2-</sup>	Bubbles of gas that turns limewater cloudy	(Bubbles of gas formed when acid added)	(Bubbles of gas formed when acid added)
Sulfate ions	SO <sub>4</sub> <sup>2-</sup>	-	White precipitate $Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \rightarrow BaSO_{4(s)}$	-
Chloride ions	Cl-	-	-	White precipitate $Ag^+_{(aq)} + Cl^{(aq)} \rightarrow AgCl_{(s)}$
Bromide ions	Br	-	-	Cream precipitate $Ag^{+}_{(aq)} + Br^{-}_{(aq)} \rightarrow AgBr_{(s)}$
Iodide ions	ŀ	-	-	Yellow precipitate $Ag^{+}_{(aq)} + I^{-}_{(aq)} \rightarrow AgI_{(s)}$

Definition	
Piece of a material consisting of a few hundred atoms, and between 1 nm and 100 nm in size.	
A hard, durable, non-metallic material that is generally unaffected by heat eg china and glass	
The total amount of surface area of an object divided by its volume	
A material that breaks before it deforms (changes shape)	
A material that resists both deformation (changing shape) and failure (eg snapping)	
A material that resists dents, scratches etc	
Resists deformation (changing shape)	
A material that can be hammered or pressed into shape without breaking	

### **Polymers**

- A wide range of polymers exist with differing physical properties such as strength, stretchiness, density, ease of moulding etc
- The properties of a polymer determine its use.
- Strong, rigid polymers (eg high-density poly(ethene), HDPE) are used to make water pipes.
- Light, stretchy polymers (eg low-density poly(ethene), LDPE) are used to make carrier bags.
- · Poly(styrene) foam is used for packaging and thermal insulation
- Heat resistant polymers (eg poly(propene), PP) are used to make kettles.

### Ceramics

- Ceramics are stiff, brittle, strong and hardwearing materials that are electrical and thermal insulators. They are resistant to corrosion and do not degrade overtime like some other materials.
- They are produced by heating substances to high temperatures such as clay to produce bricks and china, and sand to produce glass.
- Glass is different to other ceramics as the atoms inside it are not arranged regularly and so glass is transparent.

### Metals

- Metals are good thermal and electrical conductors. They are malleable, strong, have high melting temperatures and typically have a high density and are less brittle than ceramics or polymers.
- Some metals corrode easily (eg iron) but corrosion resistant metals can last a very long time (eg copper, gold).
- Metals can be mixed with other elements to form alloys (such as steel) that can have improved properties such increased hardness or corrosion resistance over individual metals.

### Composites

- Composites are made from two or more materials with different properties. The properties of the composite depend on the materials used to make it.
- Composites (eg reinforced concrete) are made from a reinforcing material (iron bars) embedded in a matrix/binder (concrete)
- The component materials can be changed to obtain a material with the properties needed for a specific use.
- The cost of producing composites is their main disadvantage.



### **Nanoparticles**

- · Nanoparticles contain only a few hundred atoms. They are larger than atoms ( $\sim 0.1 - 0.5$  nm) and simple molecules (eg O<sub>2</sub>) but are smaller than nearly everything else.
- A nanoparticle has very different properties than the bulk material – eg fullerenes (carbon nanoparticles) have different properties than a lump of carbon. Gold nanoparticles will change colour from red to purple depending on the size of the particles.
- Nanoparticles have extremely high surface area to volume ratios which means a greater proportion of their atoms can interact with substances compared to their bulk material.

### Nanoparticles and health

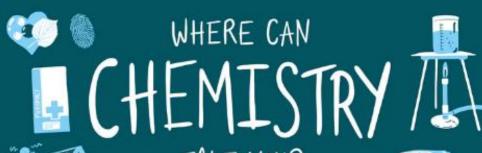
- The full impact of nanoparticles on health is not yet fully understood. New products need to be tested to minimise risks to health.
- Some people worry that products containing nanoparticles are available without full investigations into possible side-effects or long-term impacts of their use on human health or the environment. The higher surface area to volume ratio could mean they are more toxic compared to their bulk materials.
- For example, some nanoparticles used in medicines do not break down easily so could build up in cells. Breathing nanoparticles in could lead to breathing issues such as inflamed lungs.

### Uses of nanoparticles

- The high surface area to volume ratios means nanoparticles can make excellent catalysts as reactions can take place on the surface of the nanoparticles.
- · Some sunscreens use certain nanoparticles as they absorb UV light without leaving white marks on the skin.
- Nanomedicines using fullerenes that contain drugs are being developed. Fullerenes are more easily absorbed into the body than most particles and could deliver drugs into cells where they are needed.
- · Carbon nanotubes conduct electricity and can be used to make smaller parts for computers.
- Certain nanoparticles can be added to plastics in sporting equipment to make them stronger without adding much mass
- Silver nanoparticles have antibacterial properties and can be added to polymer fibres used to make surgical masks and wound dressings.









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### **ENERGY & UTILITIES**

Today, about 500,000 people work in the energy sector. But with the demand for green energy growing, by 2020 half a million people could be working in renewables alone.

Coreer paths: Geochemist, mudlogger, renewable energy engineer



### **ENGINEERING**

The proportion of young engineers has dropped over the last decade. This means there will be a high demand for younger workers in the coming years.

Career paths: Chemical engineer, civil engineer, nuclear engineer



### FAST CONSUMER GOODS

Online grocery shopping is expected to increase by 68% between 2016 and 2021. This means that new e-commerce jobs will be created to match the growing demand.

Career paths: Food scientist, market research, quality controller



### MANUFACTURING

The manufacturing sector employs around three million people and accounts for 9% of employment in the UK. That's a lot of lobs!

Career paths: Manufacturing manager, stock control manager



### MEDICINE & HEALTHCARE

The UK healthcare industry employs over four million people, making it one of the biggest employment. sectors. Four of the five highest average graduate salaries are in fields related to medicine.

Career paths: Doctor, nurse, opticion:



### RECRUITMENT & HR

The Recruitment & HR industry employs around 100,000 people and this is only expected to grow in the future.

Career paths: HR officer, recruitment consultant, training manager



### SCIENCE & RESEARCH

Between 2016 and 2023, jobs in science and research will grow at twice the rate of other industries. creating 142,000 new jobs. One in every six jobs will be in science and research.

Career paths: Lab scientist, phermacologist, research and development.







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