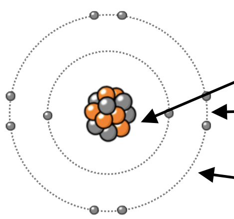
GCSE Chemistry Atomic structure and periodic table part 1

Atoms, elements and compounds

Atom	*The smallest part of an element that can exist*	Have a radius of around 0.1 nanometres and have no charge (0).
Element	*Contains only one type of atom*	Around 100 different elements each one is represented by a symbol e.g. O, Na, Br.
Compound	*Two or more elements chemically combined*	Compounds can only be separated into elements by chemical reactions.



Central nucleus	Contains protons and neutrons
Electron shells	Contains electrons

Electronic structures

Electronic shell	Max number of electrons
1	2
2	8
3	8
4	2

Relative electrical charges of subatomic particles

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

7
Li
3

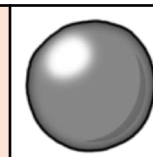
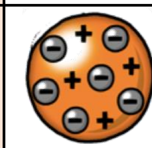
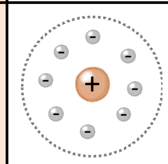
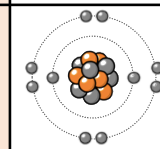
Mass number	*The sum of the protons and neutrons in the nucleus*	
Atomic number	*The number of protons in the atom*	Number of electrons = number of protons

Mixtures

Mixtures	*Two or more elements or compounds not chemically combined together*	Can be separated by physical processes.

Method	Description	Example
Filtration	*Separating an insoluble solid from a liquid*	To get sand from a mixture of sand, salt and water.
Crystallisation	*To separate a solid from a solution*	To obtain pure crystals of sodium chloride from salt water.
Simple distillation	*To separate a solvent from a solution*	To get pure water from salt water.
Fractional distillation	*Separating a mixture of liquids each with different boiling points*	To separate the different compounds in crude oil.
Chromatography	*Separating substances that move at different rates through a medium*	To separate out the dyes in food colouring.

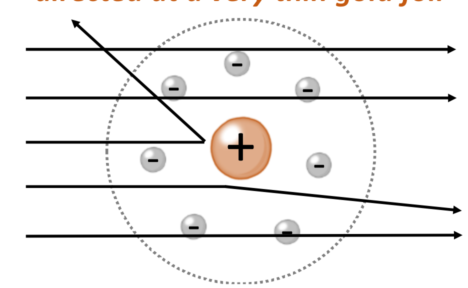
The development of the model of the atom

Pre 1900		*Tiny solid spheres that could not be divided*	Before the discovery of the electron, John Dalton said the solid sphere made up the different elements.
1897 'plum pudding'		*A ball of positive charge with negative electrons embedded in it*	JJ Thompson 's experiments showed that showed that an atom must contain small negative charges (discovery of electrons).
1909 nuclear model		*Positively charge nucleus at the centre surrounded negative electrons*	Ernest Rutherford's alpha particle scattering experiment showed that the mass was concentrated at the centre of the atom.
1913 Bohr model		*Electrons orbit the nucleus at specific distances*	Niels Bohr proposed that electrons orbited in fixed shells; this was supported by experimental observations.

James Chadwick	*Provided the evidence to show the existence of neutrons within the nucleus*

Rutherford's scattering experiment

A beam of alpha particles are directed at a very thin gold foil



Most of the alpha particles passed right through.
A few (+) alpha particles were deflected by the positive nucleus.
A tiny number of particles reflected back from the nucleus.

Chemical equations

Chemical equations	*Show chemical reactions - need reactant(s) and product(s) energy always involves and energy change*	Law of conservation of mass states the total mass of products = the total mass of reactants.
Word equations	*Uses words to show reaction* **reactants → products** magnesium + oxygen → magnesium oxide	Does not show what is happening to the atoms or the number of atoms.
Symbol equations	*Uses symbols to show reaction* **reactants → products** $2Mg + O_2 \rightarrow 2MgO$	Shows the number of atoms and molecules in the reaction, these need to be balanced.

Relative atomic mass

Isotopes	*Atoms of the same element with the same number of protons and different numbers of neutrons*	**^{35}Cl (75%) and ^{37}Cl (25%)** Relative abundance = (% isotope 1 x mass isotope 1) + (% isotope 2 x mass isotope 2) ÷ 100 **e.g. (25 x 37) + (75x 35) ÷ 100 = 35.5**

GCSE Chemistry Atomic structure and periodic table part 2

The Periodic table

Alkali metals

Halogens

Noble gases

Transition metals

1	2											3	4	5	6	7	0
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?						

Metals to the left of this line, non metals to the right

Elements arranged in order of atomic number	*Elements with similar properties are in columns called groups*	Elements in the same group have the same number of outer shell electrons and elements in the same period (row) have the same number of electron shells.

Development of the Periodic table

Before discovery of protons, neutrons and electrons	*Elements arranged in order of atomic weight*	Early periodic tables were incomplete, some elements were placed in inappropriate groups if the strict order atomic weights was followed.
Mendeleev	*Left gaps for elements that hadn't been discovered yet*	Elements with properties predicted by Mendeleev were discovered and filled in the gaps. Knowledge of isotopes explained why order based on atomic weights was not always correct.

Metals and non metals

Metals	*To the left of the Periodic table*	Form positive ions. Conductors, high melting and boiling points, ductile, malleable.
Non metals	*To the right of the Periodic table*	Form negative ions. Insulators, low melting and boiling points.

Group 1

Alkali metals	*Very reactive with oxygen, water and chlorine*	Only have one electron in their outer shell. Form +1 ions.
	Reactivity increases down the group	Negative outer electron is further away from the positive nucleus so is more easily lost.

With oxygen	Forms a metal oxide	Metal + oxygen → metal oxide	e.g. $4Na + O_2 \rightarrow 2Na_2O$
With water	Forms a metal hydroxide and hydrogen	Metal + water → metal hydroxide + hydrogen	e.g. $2Na + 2H_2O \rightarrow 2NaOH + H_2$
With chlorine	Forms a metal chloride	Metal + chlorine → metal chloride	e.g. $2Na + Cl_2 \rightarrow 2NaCl$

Group 7

Halogens	*Consist of molecules made of a pair of atoms*	Have seven electrons in their outer shell. Form -1 ions.
	Melting and boiling points increase down the group (gas → liquid → solid)	Increasing atomic mass number.
	Reactivity decreases down the group	Increasing proton number means an electron is more easily gained

With metals	Forms a metal halide	Metal + halogen → metal halide e.g. Sodium + chlorine → sodium chloride	e.g. NaCl metal atom loses outer shell electrons and halogen gains an outer shell electron
With hydrogen	Forms a hydrogen halide	Hydrogen + halogen → hydrogen halide e.g. Hydrogen + bromine → hydrogen bromide	e.g. $Cl_2 + H_2 \rightarrow 2HCl$
With aqueous solution of a halide salt	A more reactive halogen will displace the less reactive halogen from the salt	Chlorine + potassium bromide → potassium chloride + bromine	e.g. $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$

Group 0

Noble gases	*Unreactive, do not form molecules*	This is due to having full outer shells of electrons.
	Boiling points increase down the group	Increasing atomic number.

Transition metals (Chemistry only)

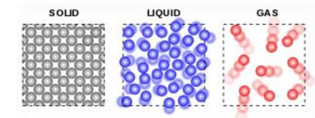
Compared to group 1	• *Less reactive* • *Harder* • *Denser* • *Higher melting points*	• *Cu^{2+} is blue* • *Ni^{2+} is pale green, used in the manufacture of margarine* • *Fe^{2+} is green, used in the Haber process* • *Fe^{3+} is reddish-brown* • *Mn^{2+} is pale pink*
Typical properties	• *Many have different ion possibilities with different charges* • *Used as catalysts* • *Form coloured compounds*	

GCSE Chemistry Topic 2: Structure and Bonding 1

Chemical bonds

Ionic	*Particles are oppositely charged ions*	Occurs in compounds formed from metals combined with non metals.
Covalent	*Particles are atoms that share pairs of electrons*	Occurs in most non metallic elements and in compounds of non metals.
Metallic	*Particles are atoms which share delocalised electrons*	Occurs in metallic elements and alloys.

The three states of matter

Solid, liquid , gas	*Melting and freezing happen at melting point, boiling and condensing happen at boiling point.*	 The amount of energy needed for a state change depends on the strength of forces between particles in the substance.	(HT only) Limitations of simple model: • There are no forces in the model • All particles are shown as spheres • Spheres are solid

s	solid
l	liquid
g	gas

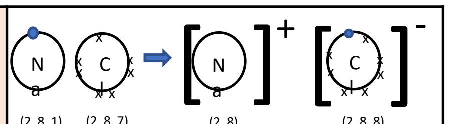
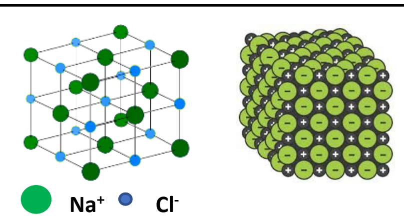
Properties of ionic compounds

High melting and boiling points	Large amounts of energy needed to break the bonds.
Do not conduct electricity when solid	Ions are held in a fixed position in the lattice and cannot move.
Do conduct electricity when molten or dissolved	Lattice breaks apart and the ions are free to move.

Ionic bonding

Electrons are transferred so that all atoms have a noble gas configuration (full outer shells).	*Metal atoms lose electrons and become positively charged ions*	Group 1 metals form +1 ions Group 2 metals form +2 ions
	Non metals atoms gain electrons to become negatively charged ions	Group 6 non metals form -2 ions Group 7 non metals form -1 ions

Ionic compounds

Dot and cross diagram	
Giant structure	

Structure	• *Held together by strong electrostatic forces of attraction between oppositely charged ions* • *Forces act in all directions in the lattice*

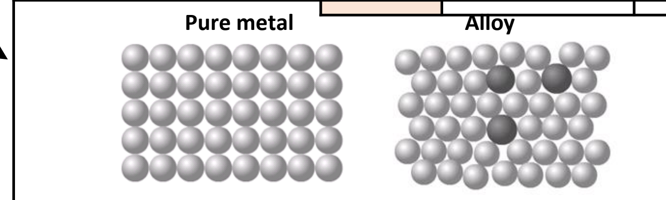
Metals as conductors

Good conductors of electricity	Delocalised electrons carry electrical charge through the metal.
Good conductors of thermal energy	Energy is transferred by the delocalised electrons.

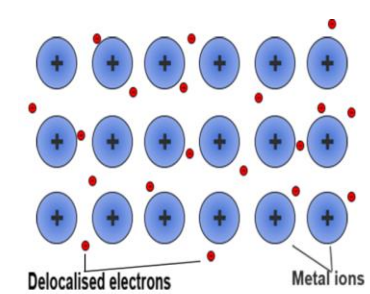
Properties of metals and alloys

High melting and boiling points	This is due to the strong metallic bonds.
Pure metals can be bent and shaped	Atoms are arranged in layers that can slide over each other.

Alloys	*Mixture of two or more elements at least one of which is a metal*	Harder than pure metals because atoms of different sizes disrupt the layers so they cannot slide over each other.

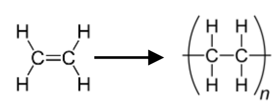


Metallic bonding

Giant structure of atoms arranged in a regular pattern		Electrons in the outer shell of metal atoms are delocalised and free to move through the whole structure. This sharing of electrons leads to strong metallic bonds.

GCSE Chemistry Topic 2: Structure and Bonding 2

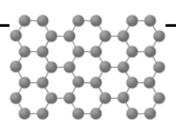
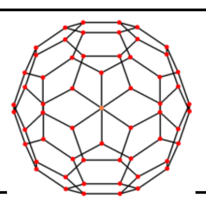
Polymers

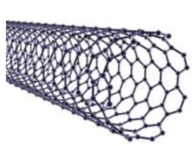
Very large molecules	*Solids at room temperature*	Atoms are linked by strong covalent bonds.	

Properties of small molecules

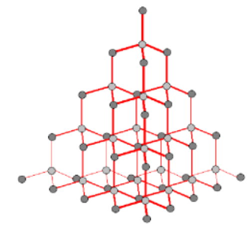
Usually gases or liquids	*Covalent bonds in the molecule are strong but forces between molecules (intermolecular) are weak*	Low melting and boiling points.	Due to having weak intermolecular forces that easily broken.
		Do not conduct electricity.	Due to them molecules not having an overall electrical charge.
		Larger molecules have higher melting and boiling points.	Intermolecular forces increase with the size of the molecules.

Graphene and fullerenes

Graphene	 *Single layer of graphite one atom thick*	Excellent conductor.	Contains delocalised electrons.
		Very strong.	Contains strong covalent bonds.
Fullerenes		Buckminsterfullerene, C_{60} First fullerene to be discovered.	Hexagonal rings of carbon atoms with hollow shapes. Can also have rings of five (pentagonal) or seven (heptagonal) carbon atoms.

Carbon nanotubes		*Very thin and long cylindrical fullerenes*	Very conductive.	Used in electronics industry.
			High tensile strength.	Reinforcing composite materials.
			Large surface area to volume ratio.	Catalysts and lubricants.

Diamond

Each carbon atom is bonded to four others		Very hard.	Rigid structure.
		Very high melting point.	Strong covalent bonds.
		Does not conduct electricity.	No delocalised electrons.

Giant covalent structures

Diamond, graphite, silicon dioxide	*Very high melting points*	Lots of energy needed to break strong, covalent bonds.

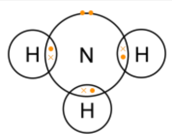
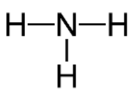
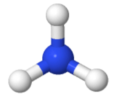
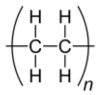
Size of particles and their properties (Chemistry only)

Nanoparticles	*Between 1 and 100 nanometres (nm) in size*	1 nanometre (1 nm) = 1 x 10^{-9} metres (0.000 000 001m or a billionth of a metre).

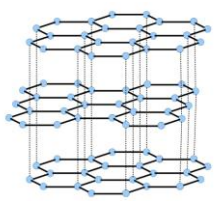
Use of nanoparticles

Healthcare, cosmetics, sun cream, catalysts, deodorants, electronics.	Nanoparticles may be toxic to people. They may be able to enter the brain from the bloodstream and cause harm.

Covalent bonding

Atoms share pairs of electrons	*Can be small molecules e.g. ammonia*		Dot and cross : + Show which atom the electrons in the bonds come from - All electrons are identical
			2D with bonds: + Show which atoms are bonded together - It shows the H-C-H bond incorrectly at 90°
			3D ball and stick model: + Attempts to show the H-C-H bond angle is 109.5°
	Can be giant covalent structures e.g. polymers		

Graphite

Each carbon atom is bonded to three others forming layers of hexagonal rings with no covalent bonds between the layers		Slippery.	Layers can slide over each other.
		Very high melting point.	Strong covalent bonds.
		Does conduct electricity.	Delocalised electrons between layers.

GCSE Chemistry – Topic 3 QUANTITATIVE CHEMISTRY 1

Conservation of mass and balanced symbol equations

M_r	*The sum of the relative atomic masses of the atoms in the numbers shown in the formula. The Mr is the mass of 1 mole of the formula in grams.*	The sum of the M_r of the reactants in the quantities shown equals the sum of the M_r of the products in the quantities shown.

$2Mg + O_2 \rightarrow 2MgO$

48g + 32g = 80g

80g = 80g

Balanced symbol equations	*Represent chemical reactions and have the same number of atoms of each element on both sides of the equation*	$H_2 + Cl_2 \rightarrow 2HCl$ Subscript (small number at the bottom, after the element) numbers show the number of atoms of the element to its left. Normal script (in front of the formula) numbers show the number of molecules.

Conservation of mass	*No atoms are lost or made during a chemical reaction*	Mass of the products equals the mass of the reactants.

Mass changes when a reactant or product is a gas

Mass appears to increase during a reaction	*One of the reactants is a gas*	Magnesium + oxygen → magnesium oxide
Mass appears to decrease during a reaction	*One of the products is a gas and has escaped*	Calcium carbonate → carbon dioxide + calcium oxide

Relative formula mass (M_r)

Limiting reactants (HT only)

The reactant that is completely used up	*Limits the amount of product that is made*	Less moles of product are made.

Chemical measurements

Whenever a measurement is taken, there is always some uncertainty about the result obtained	*Can determine whether the mean value falls within the range of uncertainty of the result*	1. Calculate the mean 2. Calculate the range of the results 3. Estimate of uncertainty in mean would be half the range

Concentration of solutions

Measured in mass per given volume of solution (g/dm^3)	*Conc. = $\frac{\text{mass (g)}}{\text{volume (dm}^3\text{)}}$*	Greater mass = higher concentration. Greater volume = lower concentration.

Moles (HT only)

Chemical amounts are measured in moles (mol)	*Mass of one mole of a substance in grams = relative formula mass*	One mole of H_2O = 18g (1 + 1 + 16) One mole of Mg = 24g
Avogadro constant	*One mole of any substance will contain the same number of particles, atoms, molecules or ions.*	6.02×10^{23} per mole One mole of H_2O will contain 6.02×10^{23} molecules One mole of NaCl will contain 6.02×10^{23} Na^+ ions
	Number of moles = $\frac{\text{mass (g)}}{A_r}$ or $\frac{\text{mass (g)}}{M_r}$	How many moles of sulfuric acid molecules are there in 4.7g of sulfuric acid (H_2SO_4)? Give your answer to 1 significant figure. $\frac{4.7}{98}$ = 0.05 mol (M_r of H_2SO_4)

Using moles to balance equations (HT only)

The balancing numbers in a symbol equation can be calculated from the masses of reactants and products	*Convert the masses in grams to amounts in moles and convert the number of moles to simple whole number ratios.*

Amounts of substances in equations (HT only)

Chemical equations show the number of moles reacting and the number of moles made	*$Mg + 2HCl \rightarrow MgCl_2 + H_2$* *One mole of magnesium reacts with two moles of hydrochloric acid to make one mole of magnesium chloride and one mole of hydrogen*	If you have a 60g of Mg, what mass of HCl do you need to convert it to $MgCl_2$? A_r : Mg =24 so mass of 1 mole of Mg = 24g M_r : HCl (1 + 35.5) so mass of 1 mole of HCl = 36.5g So 60g of Mg is 60/24 = 2.5 moles Balanced symbol equation tells us that for every one mole of Mg, you need two moles of HCl to react with it. So you need 2.5x2 = 5 moles of HCl You will need 5 x 36.5g of HCl= 182.5g

GCSE Chemistry – Topic 3 QUANTITATIVE CHEMISTRY 2

Atom economy

A measure of the amount of starting materials that end up as useful products

Atom economy = $\frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula mass of all reactants from equation}} \times 100$

High atom economy is important or sustainable development and economic reasons

Calculate the atom economy for making hydrogen by reacting zinc with hydrochloric acid:

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

M_r of H_2 = 1 + 1 = 2
M_r of Zn + 2HCl = 65 + 1 + 1 + 35.5 + 35.5 = 138

Atom economy = $^2/_{138} \times 100$
= $^2/_{138} \times 100$ = 1.45%

This method is unlikely to be chosen as it has a low atom economy.

Percentage yield

Yield is the amount of product obtained

It is not always possible to obtain the calculated amount of a product

- The reaction may not go to completion because it is reversible.
- Some of the product may be lost when it is separated from the reaction mixture.
- Some of the reactants may react in ways different to the expected reaction.

Percentage yield is comparing the amount of product obtained as a percentage of the maximum theoretical amount

% Yield = $\frac{\text{Mass of product made}}{\text{Max. theoretical mass}} \times 100$

A piece of sodium metal is heated in chlorine gas. A maximum theoretical mass of 10g for sodium chloride was calculated, but the actual yield was only 8g. *Calculate the percentage yield.*

Percentage yield = 8/10 x 100 =80%

Using concentrations of solutions in mol/dm³ (HT only, chemistry only)

Concentration of a solution is the amount of solute per volume of solution

Concentration (mol/dm³) = $\frac{\text{moles (mol)}}{\text{volume (dm}^3\text{)}}$

Concentration (g/dm³) = $\frac{\text{mass (g)}}{\text{volume (dm}^3\text{)}}$

g/dm³ $\xrightarrow{\div M_r}$ *mol/dm³*; *mol/dm³* $\xrightarrow{\times M_r}$ *g/dm³*

What is the concentration of a solution that has 35.0g of solute in 0.5dm³ of solution?

35/0.5 = 70 g/dm³

Titration

If the volumes of two solutions that react completely are known and the concentrations of one solution is known, the concentration of the other solution can be calculated.

$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

It takes 12.20cm³ of sulfuric acid to neutralise 24.00cm³ of sodium hydroxide solution, which has a concentration of 0.50mol/dm³.

Calculate the concentration of the sulfuric acid in mol/dm³:

0.5 mol/dm³ x (24/1000) dm³ = 0.012 mol of NaOH
The equation shows that 2 mol of NaOH reacts with 1 mol of H_2SO_4, so the number of moles in 12.20cm³ of sulfuric acid is (0.012/2) = 0.006 mol of sulfuric acid

Calculate the concentration of sulfuric acid in mol/ dm³
0.006 mol x (1000/12.2) dm³ =0.49mol/dm³

Use of amount of substance in relation to volumes of gases (HT only, chemistry only)

Equal amounts of moles or gases occupy the same volume under the same conditions of temperature and pressure

The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmospheric pressure) is ***24 dm³***

No. of moles of gas = $\frac{\text{vol of gas (dm}^3\text{)}}{24\text{dm}^3}$

What is the volume of 11.6 g of butane (C_4H_{10}) gas at RTP?

M_r : (4 x 12) + (10 x 1) = 58

11.6/58 = 0.20 mol

Volume = 0.20 x 24 = 4.8 dm³

6g of a hydrocarbon gas had a volume of 4.8 dm³. Calculate its molecular mass.

1 mole = 24 dm³, so 4.8/24 = 0.2 mol

M_r = 6 / 0.2 = 30

If 6g = 0.2 mol, 1 mol equals 30 g

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Oxidation **I**s **L**oss (of electrons) **R**eduction **I**s **G**ain (of electrons)

HT ONLY: Reactions between metals and acids are redox reactions as the metal donates electrons to the hydrogen ions. This displaces hydrogen as a gas while the metal ions are left in the solution.

GCSE Chemistry Topic 4 Chemical Changes 1

Oxidation and reduction in terms of electrons (HT ONLY)

Ionic half equations (HT only)		
For displacement reactions	*Ionic half equations show what happens to each of the reactants during reactions*	For example: The ionic equation for the reaction between iron and copper (II) ions is: **$Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$** The half-equation for iron (II) is: **$Fe \rightarrow Fe^{2+} + 2e^-$** The half-equation for copper (II) ions is: **$Cu^{2+} + 2e^- \rightarrow Cu$**

Reactions of acids

Reactions of acids and metals

Acids react with some metals to produce salts and hydrogen.

Reactions with acids	*metal + acid → salt + hydrogen*	**magnesium + hydrochloric acid → magnesium chloride + hydrogen** **zinc + sulfuric acid → zinc sulfate + hydrogen**

Neutralisation of acids and salt production

Acid name	Salt name
Hydrochloric acid	Chloride
Sulfuric acid	Sulfate
Nitric acid	Nitrate

sodium hydroxide + hydrochloric acid → sodium chloride + water

calcium carbonate + sulfuric acid → calcium sulfate, + carbon dioxide + water

Neutralisation	*Acids can be neutralised by alkalis and bases*	An **alkali** is a soluble base e.g. metal hydroxide. A **base** is a substance that neutralises an acid e.g. a soluble metal hydroxide or a metal oxide.

Extraction of metals and reduction

Unreactive metals, such as gold, are found in the Earth as the metal itself. They can be mined from the ground.

Extraction using carbon	
Metals less reactive than carbon can be extracted from their oxides by reduction.	For example: **zinc oxide + carbon → zinc + carbon dioxide**

Reactivity of metals

	Reactions with water	Reactions with acid
Group 1 metals	*Reactions get more vigorous as you go down the group*	*Reactions get more vigorous as you go down the group*
Group 2 metals	*Do not react with water*	*Observable reactions include fizzing and temperature increases*
Zinc, iron and copper	*Do not react with water*	*Zinc and iron react slowly with acid. Copper does not react with acid.*

Metal oxides

Metals and oxygen	*Metals react with oxygen to form metal oxides*	**magnesium + oxygen → magnesium oxide** $2Mg + O_2 \rightarrow 2MgO$
Reduction	*This is when oxygen is removed from a compound during a reaction*	e.g. metal oxides reacting with hydrogen, extracting low reactivity metals
Oxidation	*This is when oxygen is gained by a compound during a reaction*	e.g. metals reacting with oxygen, rusting of iron

The reactivity series

Metals form positive ions when they react	*The reactivity of a metal is related to its tendency to form positive ions*	The reactivity series arranges metals in order of their reactivity (their tendency to form positive ions).
Carbon and hydrogen	*Carbon and hydrogen are non-metals but are included in the reactivity series*	These two non-metals are included in the reactivity series as they can be used to extract some metals from their ores, depending on their reactivity.
Displacement	*A more reactive metal can displace a less reactive metal from a compound.*	Silver nitrate + Sodium chloride → Sodium nitrate + Silver chloride

potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

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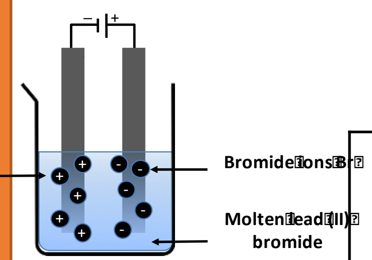
GCSE Chemistry Topic 4 Chemical Changes 2

Electrolysis

Process of electrolysis	*Splitting up using electricity*	When an ionic compound is melted or dissolved in water, the ions are free to move. These are then able to conduct electricity and are called electrolytes. Passing an electric current though electrolytes causes the ions to move to the electrodes.
Electrode	*Anode* *Cathode*	The positive electrode is called the anode. The negative electrode is called the cathode.
Where do the ions go?	*Cations* *Anions*	Cations are positive ions and they move to the negative cathode. Anions are negative ions and they move to the positive anode.

Higher tier: Half equations, for example:
At the cathode: $Pb^{2+} + 2e^- \rightarrow Pb$
At the anode: $2Br^- \rightarrow Br_2 + 2e^-$



Extracting metals using electrolysis

Metals can be extracted from molten compounds using electrolysis.
This process is used when the metal is too reactive to be extracted by reduction with carbon.
The process is expensive due to large amounts of energy needed to produce the electrical current. Example: aluminium is extracted in this way.
Aluminium extraction uses CRYOLITE to lower the melting point of aluminium oxide. Electrodes have to be replaced regularly as they are made of GRAPHITE (CARBON) which react with the oxygen produced.

Electrolysis of aqueous solutions

The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.

At the negative electrode	Metal will be produced on the electrode if it is less reactive than hydrogen. Hydrogen will be produced if the metal is more reactive than hydrogen.
At the positive electrode	Oxygen is formed at positive electrode. If you have a halide ion (Cl^-, I^-, Br^-) then you will get chlorine, bromine or iodine formed at that electrode.

Strong and weak acids (HT ONLY)

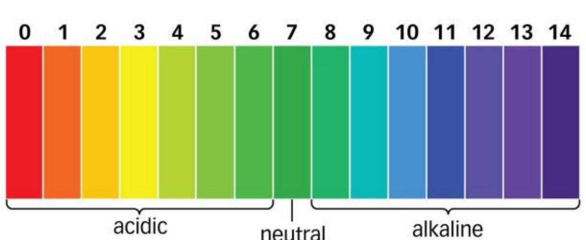
Strong acids	*Completely ionised in aqueous solutions e.g. hydrochloric, nitric and sulfuric acids.*
Weak acids	*Only partially ionised in aqueous solutions e.g. ethanoic acid, citric acid.*
Hydrogen ion concentration	*As the pH decreases by one unit (becoming a stronger acid), the hydrogen ion concentration increases by a factor of 10.*

Reactions of acids

Soluble salts

Soluble salts	*Soluble salts can be made from reacting acids with solid insoluble substances (e.g. metals, metal oxides, hydroxides and carbonates).*
Production of soluble salts	*Add the solid to the acid until no more dissolves. Filter off excess solid and then crystallise to produce solid salts.*

The pH scale and neutralisation



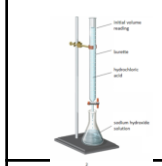
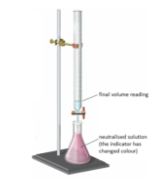
You can use universal indicator or a pH probe to measure the acidity or alkalinity of a solution against the pH scale.

Acids	*Acids contain hydrogen ions (H^+) in aqueous solutions.*
Alkalis	*Aqueous solutions of alkalis contain hydroxide ions (OH^-).*

In neutralisation reactions, hydrogen ions react with hydroxide ions to produce water:
$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Titrations (Chemistry only)

Titrations are used to work out the precise volumes of acid and alkali solutions that react with each other.

	1. Use the pipette to add 25 cm^3 of alkali to a conical flask and add a few drops of indicator. (a **pipette** is used for **fixed volumes** only)
	2. 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. (a **burette** is used for **variable volumes**)
	3. 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.

Calculating the chemical quantities in titrations involving concentrations in mol/dm³ and in g/dm³ (HT ONLY):

$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

It takes 12.20cm^3 of sulfuric acid to neutralise 24.00cm^3 of sodium hydroxide solution, which has a concentration of 0.50mol/dm^3.

Calculate the concentration of the sulfuric acid in g/dm³

0.5 mol/dm^3 x (24/1000) dm^3 = 0.012 mol of NaOH

The equation shows that 2 mol of NaOH reacts with 1 mol of H_2SO_4, so the number of moles in 12.20cm^3 of sulfuric acid is (0.012/2) = 0.006 mol of sulfuric acid

Calculate the concentration of sulfuric acid in mol/ dm³

0.006 mol x (1000/12.2) dm^3 =0.49mol/dm^3

Calculate the concentration of sulfuric acid in g/ dm³

H_2SO_4 = (2x1) + 32 + (4x16) = 98g
0.49 x 98g = 48.2g/dm^3

GCSE Chemistry – Topic 5 Energy changes

Types of reaction

Endothermic	*Energy is taken in from the surroundings so the temperature of the surroundings decreases*	• Thermal decomposition • The reaction of citric acid and sodium hydrogencarbonate	• Sports injury packs
Exothermic	*Energy is transferred to the surroundings so the temperature of the surroundings increases*	• Combustion • Many oxidation reactions • Neutralisation	• Hand warmers • Self-heating cans

Fuel cells (Chemistry only)

Hydrogen fuel cells	*Word equation:* *hydrogen + oxygen → water*	Symbol equation: $2H_2 + O_2 \rightarrow 2H_2O$
	Advantages: • No pollutants produced • Can be a range of sizes	Disadvantages: • Hydrogen is highly flammable • Hydrogen is difficult to store
Ionic half equations	Negative electrode: $2H_2\,(g) + 4OH^-\,(aq) \rightarrow 4H_2O\,(l) + 4e^-$	Positive electrode: $O_2\,(g) + 2H_2O\,(l) + 4e^- \rightarrow 4OH^-\,(aq)$

The energy change of reactions (HT only)

Breaking bonds in reactants	*Endothermic process*
Making bonds in products	*Exothermic process*

Overall energy change of a reaction	*Exothermic*	Energy released making new bonds is greater than the energy taken in breaking existing bonds.
	Endothermi c	Energy needed to break existing bonds is greater than the energy released making new bonds.

Bond energy calculation

Calculate the overall energy change for the forward reaction

$N_2 + 3H_2\ \ 2NH_3$

Bond energies (in kJ/mol): H-H 436, H-N 391, NN 945

Bond breaking: 945 + (3 x 436) = 945 + 1308 = 2253 kJ/mol

Bond making: 6 x 391 = 2346 kJ/mol

Overall energy change = 2253 - 2346 = -93kJ/mol

Therefore reaction is exothermic overall.

Cells and batteries (Chemistry only)

Simple cell	*Make a simple cell by connecting two different metals in contact with an electrolyte*	Increase the voltage by increasing the reactivity difference between the two metals.
Batteries	*Consist of two or more cells connected together in series to provide a greater voltage.*	

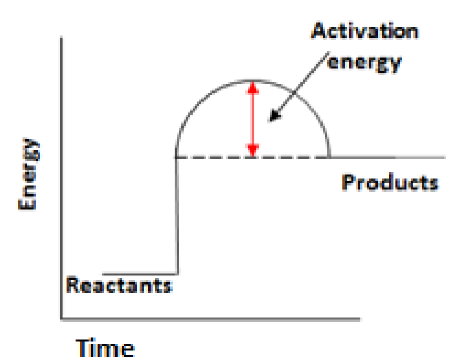
Non-rechargeable cells	*Stop when one of the reactants has been used up*	Alkaline batteries
Rechargeable cells	*Can be recharged because the chemical reactions are reversed when an external electrical current is supplied*	Rechargeable batteries

Reaction profiles

Reaction profiles	*Show the overall energy change of a reaction*

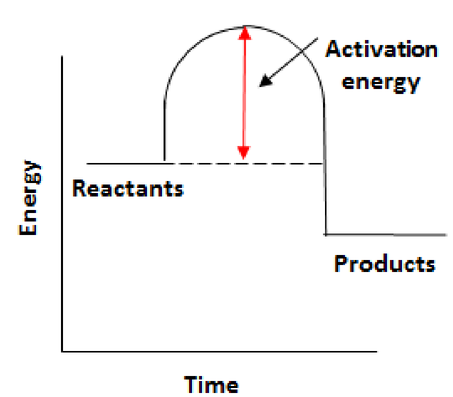
Activation energy	*Chemical reactions only happen when particles collide with sufficient energy*	The minimum amount of energy that colliding particles must have in order to react is called the activation energy.

Endothermic



Products are at a higher energy level than the reactants. As the reactants form products, energy is transferred from the surroundings to the reaction mixture. The temperature of the surroundings decreases because energy is taken in during the reaction.

Exothermic



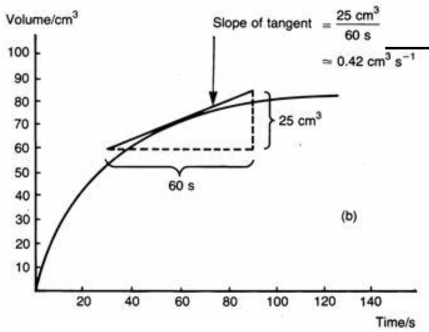
Products are at a lower energy level than the reactants. When the reactants form products, energy is transferred to the surroundings. The temperature of the surroundings increases because energy is released during the reaction.

Chemistry Topic 6 :The rate and extent of chemical change Higher

Rate of reaction

Calculating rates of reactions

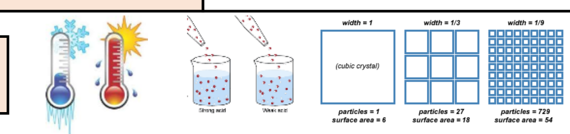
Rate of chemical reaction	*This can be calculated by measuring the quantity of reactant used or product formed in a given time.*	Rate = $\frac{\text{quantity of reactant used}}{\text{time taken}}$ Rate = $\frac{\text{quantity of product formed}}{\text{time taken}}$



Factors affecting rates

Factors affecting the rate of reaction	
Temperature	*The higher the temperature, the quicker the rate of reaction.*
Concentration	*The higher the concentration, the quicker the rate of reaction.*
Surface area	*The larger the surface area of a reactant solid, the quicker the rate of reaction.*
Pressure (of gases)	*When gases react, the higher the pressure upon them, the quicker the rate of reaction.*

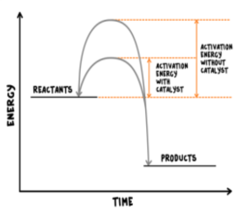
Collision theory and activation energy



Collision theory	*Chemical reactions can only occur when reacting particles collide with each other with sufficient energy.*	Increasing the temperature increases the frequency of collisions and makes the collisions more energetic, therefore increasing the rate of reaction. Increasing the concentration, pressure (gases) and surface area (solids) of reactions increases the frequency of collisions, therefore increasing the rate of reaction.
Activation energy	*This is the minimum amount of energy colliding particles in a reaction need in order to react.*	

Catalysts

Catalyst	A catalyst changes the rate of a chemical reaction but is not used in the reaction.
Enzymes	These are biological catalysts.
How do they work?	Catalysts provide a different reaction pathway WITH A LOWER ACTIVATION ENERGY.



If a catalyst is used in a reaction, it is not shown in the word equation.

Reversible reactions and dynamic equilibrium

Equilibrium

Reversible reactions

Reversible reactions	In some chemical reactions, the products can react again to re-form the reactants.
Representing reversible reactions	A + B $\rightleftharpoons$ C + D
The direction	The direction of reversible reactions can be changed by changing conditions: A + B $\underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}}$ C + D

If one direction of a reversible reaction is exothermic, the opposite direction is endothermic. The same amount of energy is transferred in each case.

Energy changes and reversible reactions

For example:
Hydrated copper sulfate (blue) $\underset{\text{exothermic}}{\overset{\text{endothermic}}{\rightleftharpoons}}$ Anhydrous copper sulfate (white) + Water

Ammonium chloride $\rightleftharpoons$ ammonia + hydrogen chloride

Equilibrium in reversible reactions	When a reversible reaction occurs equilibrium is reached when the forward and reverse reactions occur exactly at the same rate IN A CLOSED SYSTEM..

Changing conditions and equilibrium (HT)

The relative amounts of reactants and products at equilibrium depend on the conditions of the reaction.

Le Chatelier's Principles	States that when a system experiences a disturbance (change in condition), it will respond to restore a new equilibrium state.
Changing concentration	If the concentration of a reactant is increased, more products will be formed . If the concentration of a product is decreased, more reactants will react.
Changing temperature	If the temperature of a system at equilibrium is increased: - Exothermic reaction = products decrease - Endothermic reaction = products increase
Changing pressure (gaseous reactions)	For a gaseous system at equilibrium: - Pressure increase = equilibrium position shifts to side of equation with smaller number of molecules. - Pressure decrease = equilibrium position shifts to side of equation with larger number of molecules.