The Perfect Answer Revision Guide To...

CHEMISTRY

By SwH Learning

CAMBRIDGE IGCSE

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Core & Extended Assessment

2nd Edition

HAZEL LINDSEY &
MARTIN BAILEY

SwH

About SwH Learning

SwH Learning is comprised of a small team of like-minded, highly-educated individuals who have a shared passion for making first-class educational materials accessible for students worldwide. We're proud to offer one-to-one tuition for any subject, taught by specialist, enthusiastic and experienced tutors, as well as online revision classes, revision guides and workbooks.

If you are interested in arranging tuition with one of our subject specialists, visit us at www.swhlearning.co.uk

About The Authors

Hazel Lindsey

Hazel is the creator of the Science with Hazel YouTube channel and a professional science tutor. She specialises in GCSE and IGCSE biology, chemistry and physics across all exam boards.

A qualified secondary school science teacher, Hazel graduated with a degree in Pre-Clinical Veterinary Sciences from St. John's College, University of Cambridge, where she subsequently completed her PGCE (Post-Graduate Certificate of Education).

Through her tuition and YouTube channel, Hazel is responsible for helping hundreds of thousands of students from around the globe achieve their dream results in their GCSEs, IGCSEs, A Levels and IB examinations.

Martin Bailey

Martin is the co-founder of SwH Learning. Together, he and Hazel create and edit numerous science revision videos, write and publish The Perfect Answer GCSE, IGCSE and A-level revision guides & workbooks, and run the SwH Learning revision courses.

Martin graduated with a degree in Natural Sciences from Trinity College, University of Cambridge. He has thousands of hours of experience tutoring students of all abilities, from GCSE up to degree level.

Whilst at Cambridge, Martin specialised in pathology and biological anthropology. He also studied chemistry, maths, geology and history & philosophy of science.

Contents

NOTE: Core content is given in this format (Grades 1-5). *Extended content is given in italics (grades 5-9).* A copy of the periodic table will be provided in all papers

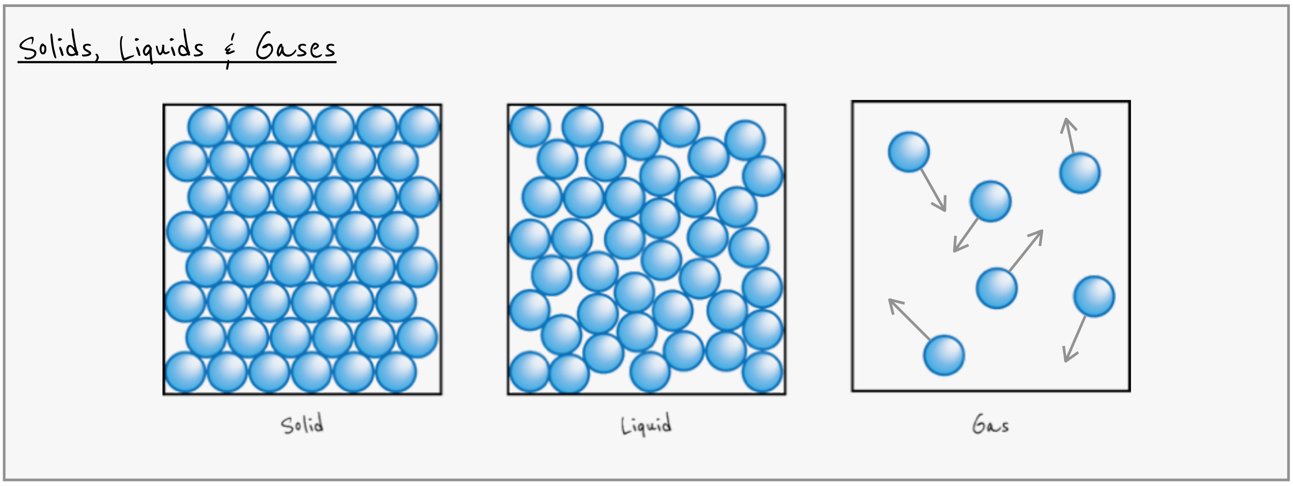
1. STATES OF MATTER

1.1 Solids, liquids and gases

Outline the properties and structure of solids, liquids and gases

	Properties	Structure
Solids	- Have a fixed shape - Do not flow - High density	- Particles arranged regularly and packed closely together - Vibrate in fixed positions - Strong forces between particles - Small amount of kinetic energy
Liquids	- Flow easily - Fixed volume but will flow to fit shape of container	- Particles are mostly touching with some gaps - Particles move about at random - Medium forces between particles - Moderate amount of kinetic energy
Gases	- No fixed volume or shape - Less dense than same volume of solid or liquid	- Particles move at random and quickly - Particles are far apart - Weak forces between particles - High amount of kinetic energy - Collide with each other and sides of container

Draw the structure of a solid, liquid and a gas



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Give the name for the following conversions:
- Solid → liquid = melting
- Liquid → gas = boiling / evaporating
- Gas → liquid = condensing
- Liquid → solid = freezing

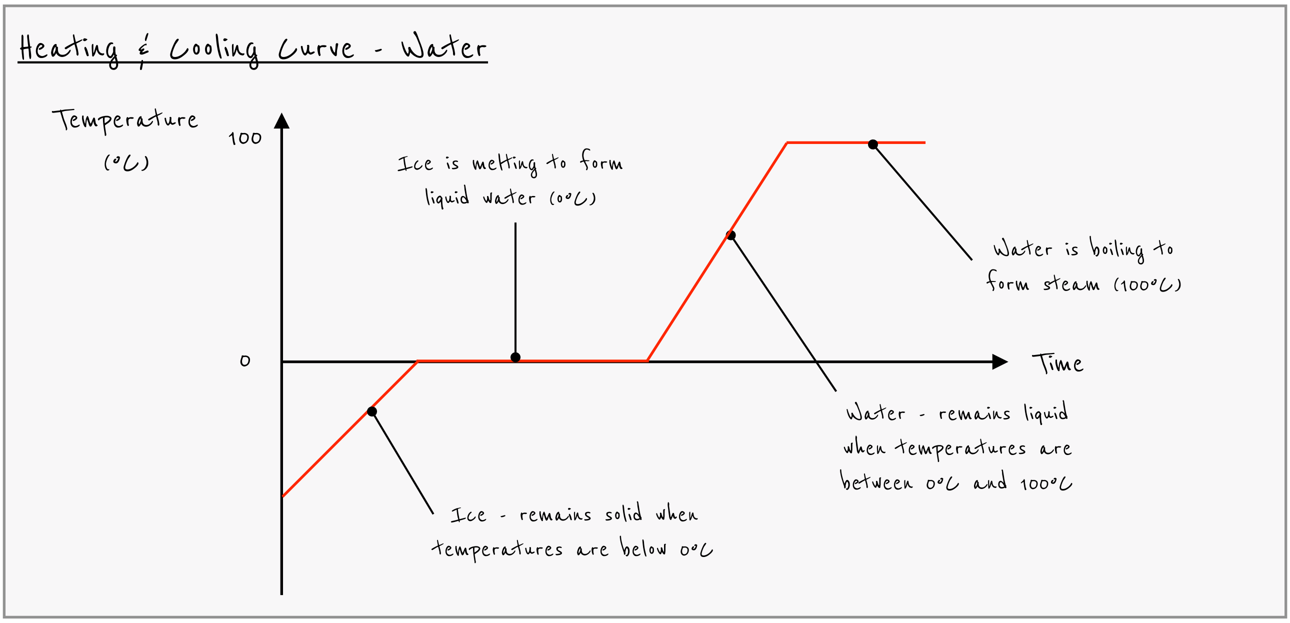
Describe what happens when water vapour cools to form liquid water
- *Particles lose kinetic energy*
- *Particles move closer together*
- *Particles move slower and less randomly*

Describe what happens when liquid water boils to form water vapour
- *Particles gain kinetic energy*
- *Particles move further apart*
- *Particles move quicker and more randomly*

Describe what happens when liquid water cools to form ice
- *Particles lose kinetic energy*
- *Particles move closer together*
- *Particles move slower and less randomly*

Describe what happens when ice melts to form liquid water
- *Particles gain kinetic energy and vibrate more*
- *Particles move further apart as forces of attraction are overcome*
- *Particles move quicker and more randomly*

How does evaporation occur?
- *Particles in liquid have differing amounts of energy*
- *Particles with the greatest amount of kinetic energy break away from surface of liquid*
- *Average kinetic energy of remaining particles is lowered*
- *In a closed container both evaporation and condensation occur simultaneously*



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Describe what happens to the volume of a gas when its temperature increases
- Volume of gas increases

Use kinetic theory to explain why the volume of a gas increases when temperature increases
- *Particles gain kinetic energy*
- *Particles move quicker and more randomly*
- *Particles move further apart*

Describe what happens to the pressure of a gas in a container when its temperature increases
- Pressure increases

Use kinetic theory to explain why pressure in a container increases when temperature increases
- *Particles have more kinetic energy → move faster*
  - *Each particle collides with greater force → increases pressure*
    - *As pressure = force ÷ area*
- *Collide more frequently with container walls → increases pressure*

Describe what happens to the pressure of a gas in a container when the volume decreases
- Pressure increases

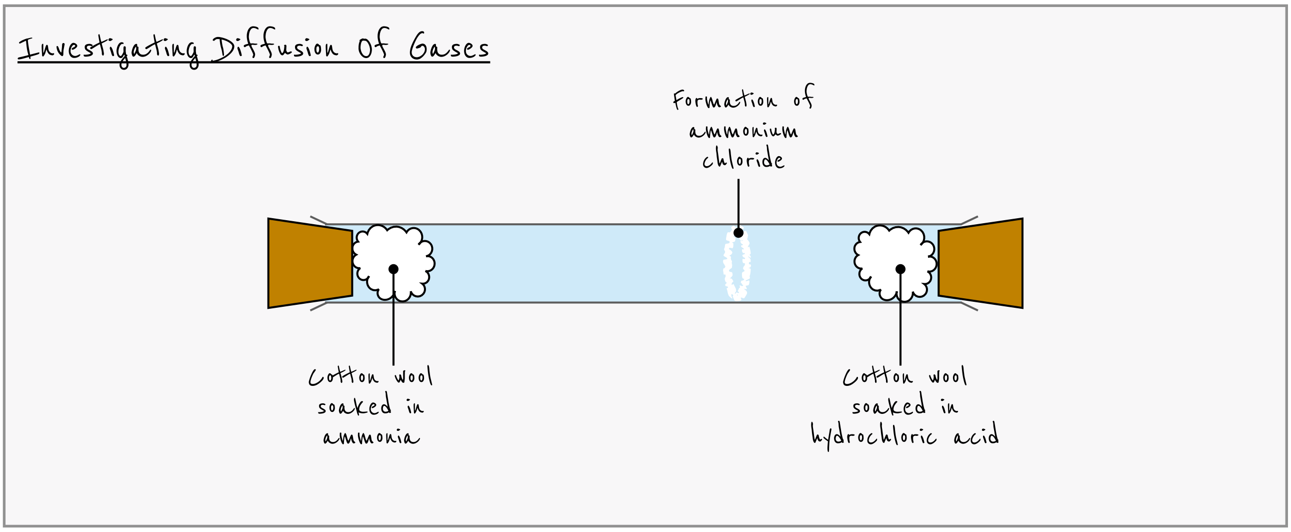
Use kinetic theory to explain why pressure in a container increases when volume decreases
- *Temperature remains constant → average kinetic energy of gas particles stays the same*
- *However container is smaller → particles collide with walls more often*
- *Increased frequency of collisions → increase in pressure*

1.2 Diffusion

Define diffusion
- Net movement of particles
- From an area of high concentration to low concentration

Which factors affect the rate of diffusion?
- Molecular mass (M_r)
  - The lower the mass of its particles, the faster a gas diffuses
- Temperature
  - The greater the temperate, the greater the kinetic energy of the particles
  - Rate of diffusion increases

How can diffusion be demonstrated experimentally?
- *Use cotton wool balls soaked in ammonia (NH_3) and hydrochloric acid (HCl)*
- *Place cotton wool at either end of a sealed tube*
- *White ring of ammonium chloride forms closer to HCl end*
- *Tells you that NH_3 diffuses faster*
  - *NH_3 has a lower M_r, so has lower mass*
  - *Diffuses more quickly than HCl*



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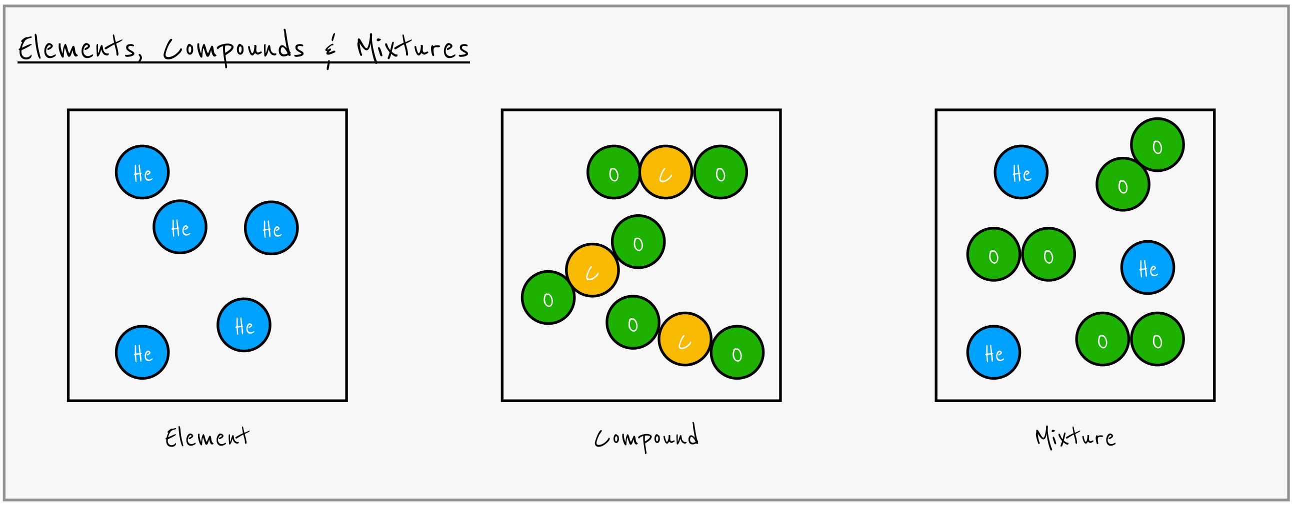
2. ATOMS, ELEMENTS AND COMPOUNDS

2.1 Elements, compounds and mixtures

What is an element?
- A substance which contains one TYPE of atom only
- Cannot be split into anything simpler by any chemical means

What is a compound?
- A substance made up of two or more elements chemically combined

What is a mixture?
- A substance made up of two or more elements NOT chemically bonded together



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Examples of elements, compounds and mixtures:

Element	Compound	Mixture
Iron	Calcium carbonate	Honey
Lead	Ammonia	Air
Sulfur	Carbon dioxide	Sea water
Nitrogen	Water	Blood
Oxygen	Iron sulfide	Soup

What is a pure substance?
- Contains one type of material only
  - e.g. one type of element or molecule

Describe the melting and boiling points of pure substances
- Fixed
  - e.g. boiling point of pure water is exactly 100°C
  - e.g. melting point of pure water is exactly 0°C

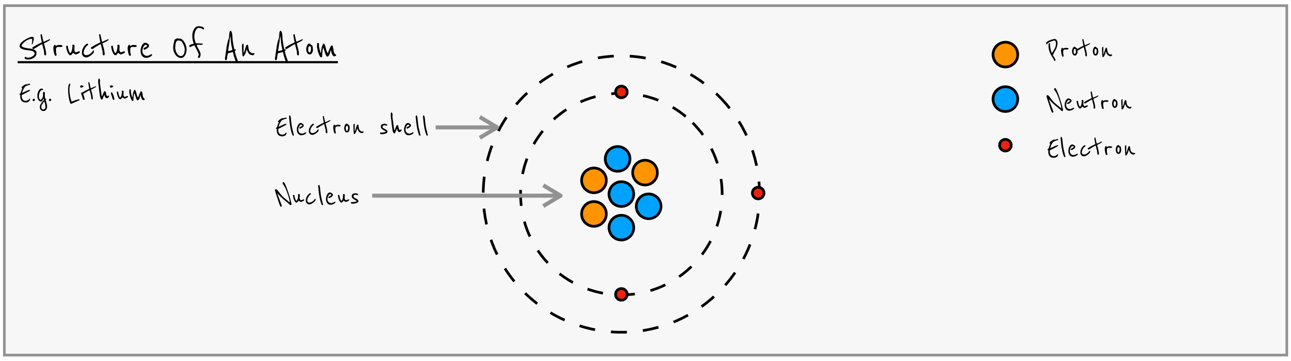
Describe the melting and boiling points of mixtures
- Melt over a range of temperatures
- Boil over a range of temperatures

2.2 Atomic structure and the Periodic Table

What is an atom?
- Smallest particle of a chemical element that can exist

What is a molecule?
- Two or more atoms chemically bonded together

Describe the structure of an atom
- Central nucleus containing neutrons and protons
- Surrounded by shells of electrons



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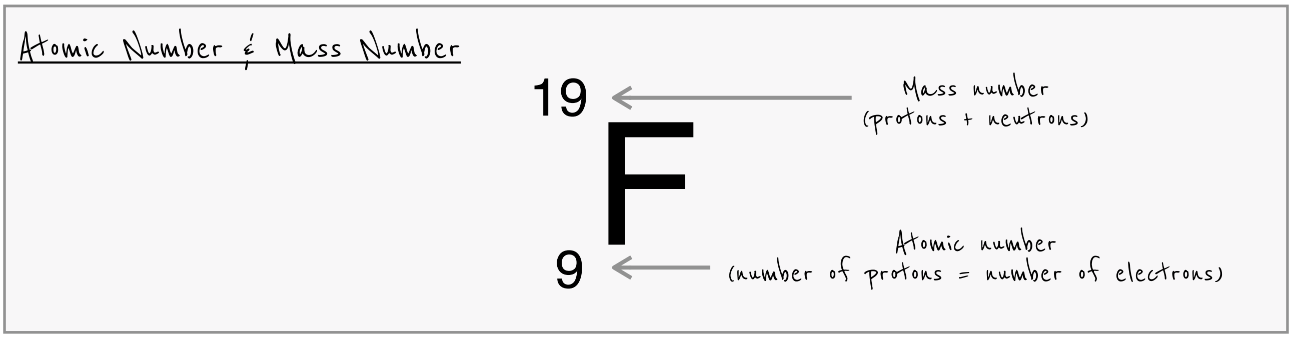
Give the relative charge and mass of a proton, neutron and electron

	Proton	Neutron	Electron
Relative charge	1	0	-1
Relative mass	1	1	1/1836

What is the atomic number?
- The number of protons

What is the mass number?
- The total number of protons and neutrons

What is the nucleon number?
- Same as mass number
  - i.e. total number of protons and neutrons



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How is the number of protons in an atom calculated?
- Same as atomic number

How is the number of electrons in an atom calculated?
- Same as atomic/proton number

How is the number of neutrons in an atom calculated?
- Mass number – atomic number

How many electrons can fit in each electron shell?
- 2 in the first shell
- 8 in all other shells

Give the electron configuration of a) Carbon, b) Magnesium, c) Potassium
- Carbon: 2.4
- Magnesium: 2.8.2
- Potassium: 2.8.8.1

What is the group number of the periodic table?
- Tells you the number of electrons in the outer shell
  - e.g. F has 7 electrons in its outer shell and is therefore in group VII

Why do elements in the same group have similar chemical properties?
- Same number of electrons in outer shell

Why are noble gases (group VIII) unreactive?
- They have a full outer shell of electrons
- Stable

What is the period number of the periodic table?
- Tells you the number of shells of electrons
  - e.g. Ca (2.8.8.2) has four shells of electrons and is therefore in period 4

2.3 Isotopes

What is an isotope?
- Atoms of the same element with the same number of protons but different number of neutrons

What is relative atomic mass (A_r)?
- The average mass of the isotopes of an element compared with 1/12th of the mass of an atom of ^{12}C

Why do isotopes of the same element have the same chemical properties?
- Have same number of electrons
- Have same electron configuration

How do you calculate the A_r of an element from the relative abundance of a particular isotope?
- $\frac{(\% \text{ of isotope 1} \times \text{mass of isotope 1}) + (\% \text{ of isotope 2} \times \text{mass of isotope 2}) + (\ldots)}{100}$

Worked Example 1: Relative Abundance

A naturally occurring sample of the element chlorine contains 75% of the Cl-35 isotope and 25% of the Cl-37 isotope. Calculate the relative atomic mass of chlorine.

Answer: $\frac{(75 \times 35) + (25 \times 37)}{100}$

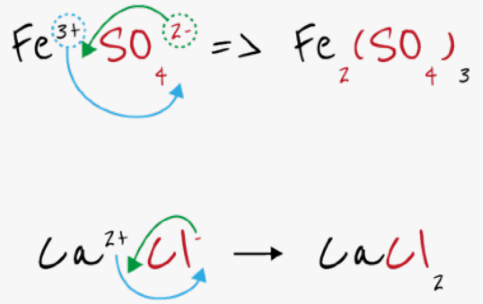
$= 35.5$

2.4 Ions and ionic bonds

How is an ion formed?
- An atom loses or gains electrons
  - Loses electron → forms positive ion (cation)
  - Gains electron → forms negative ion (anion)

When working out the charge on an ion, remember:
- For groups 1 to 3, the charge on the ion is the same as the group number
  - e.g. Mg is in group 2 and therefore forms Mg^{2+}
- For groups 5,6 and 7, the charge on the ion is (8 – group number)
  - e.g. N is in group 5. 8 – 5 = 3, therefore N^{3-}

Ionic Compound Formulae - The Drag & Drop Method



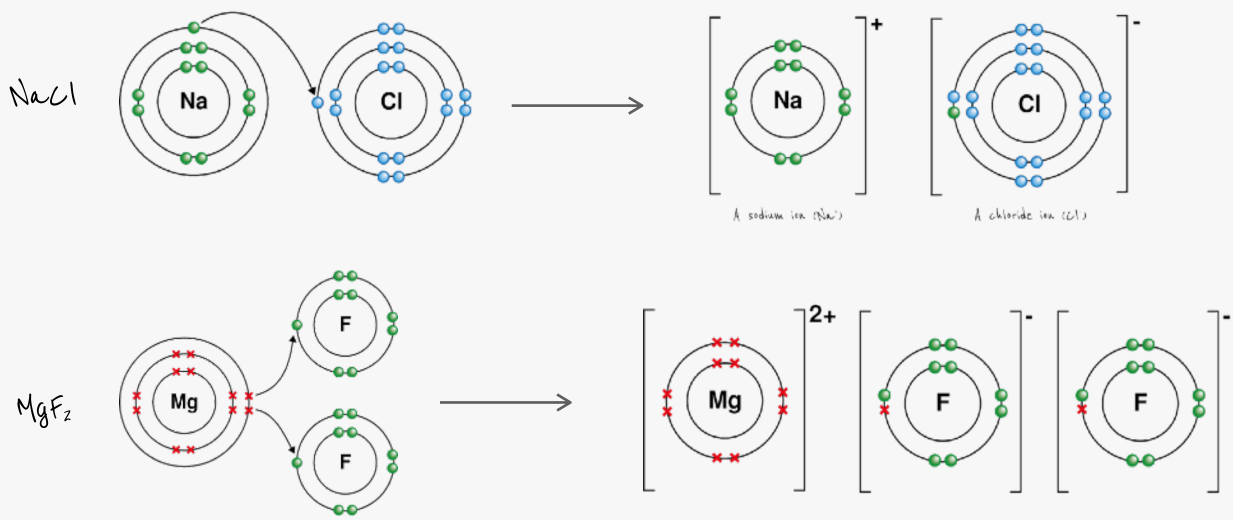
Learn the following ions off by heart

Positive ions (cations)
- H^+
- Ag^+
- Cu^{2+}
- Fe^{2+}
- Fe^{3+}
- Pb^{2+}
- Zn^{2+}
- NH_4^+

Negative ions (anions)
- OH^-
- NO_3^-
- CO_3^{2-}
- SO_4^{2-}

What is an ionic bond?
- Strong electrostatic force of attraction between oppositely charged ions
  - Form between a metal and a non-metal

Dot-And-Cross Diagrams - Ionic Bonding



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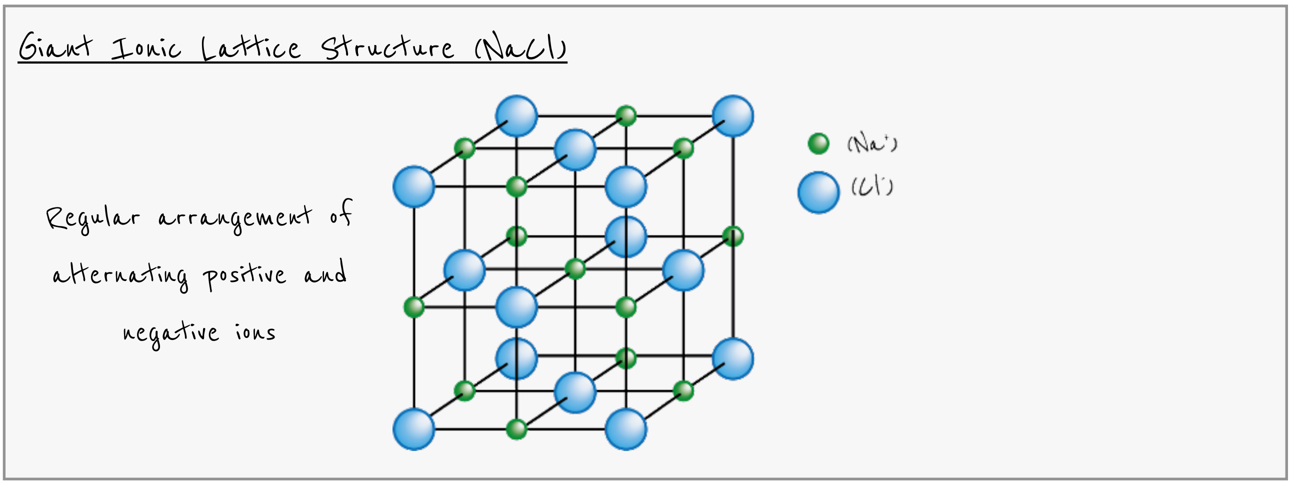
Give the properties of giant ionic lattices
- High melting and boiling points
- Conduct electricity when molten/dissolved in aqueous solution
- Brittle

Why do ionic structures have high melting and boiling points?
- *Form giant ionic lattices*
- *Strong electrostatic forces of attraction between oppositely charged ions*
- *Requires lots of energy to break*



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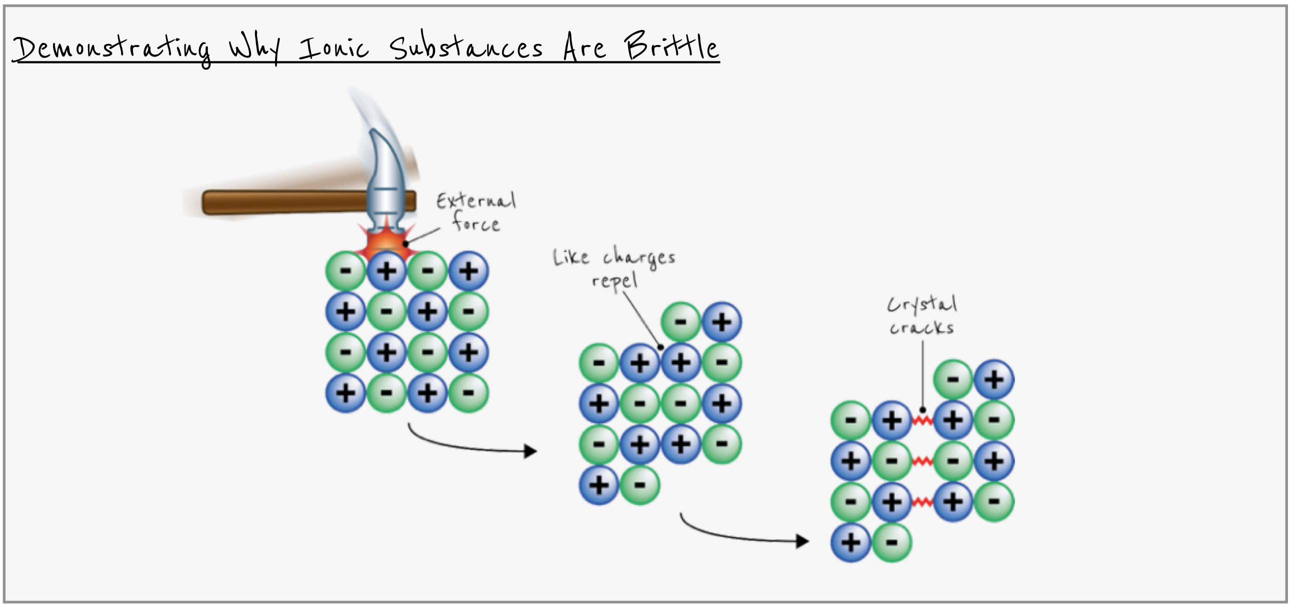
Why don't ionic substances conduct electricity when solid?
- *Ions are held tightly in fixed positions*
- *Not free to move*

Why do ionic substances conduct electricity when molten/dissolved?
- *Ions are free to move*
- *Can carry electric charge*

Explain why ionic substances are brittle
- *Applying force causes ions to move*
- *Like charges align and repel*
- *Lattice structure breaks apart*



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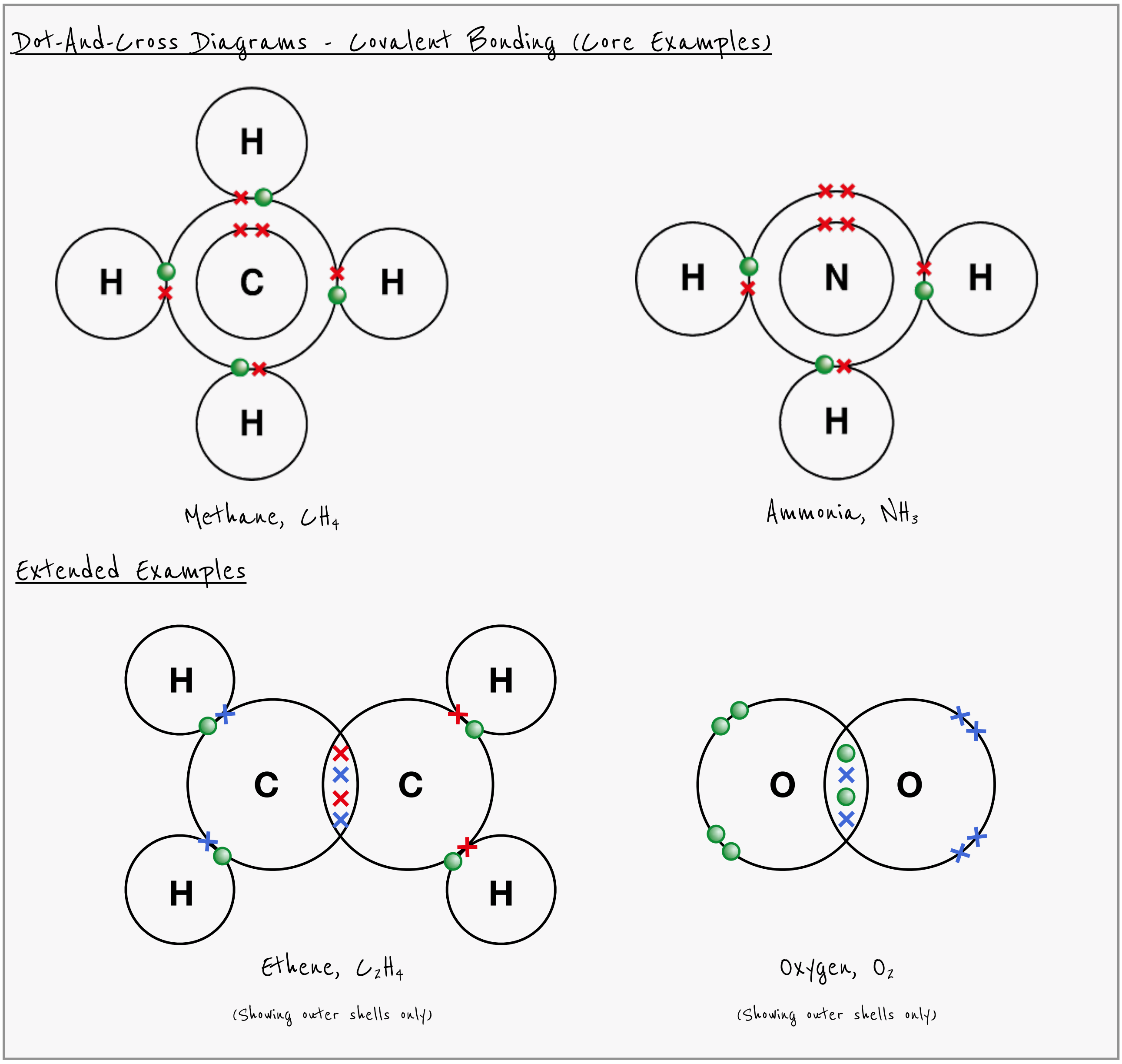
2.5 Simple molecules and covalent bonds

What is a covalent bond?
- Basic definition: a pair of electrons shared between two atoms
- Detailed definition: strong electrostatic forces of attraction between nuclei (positively charged) and shared pair of electrons (negatively charged)
- Form between two non-metals
- Result in noble gas electron configurations (full outer shell of electrons)



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What is a simple molecular substance?
- Small, covalently bonded molecule
  - e.g. H_2O, CO_2

Describe the properties of simple molecular compounds
- Low melting and boiling points
- Poor electrical conductivity

Define intermolecular force
- *Temporary weak attraction between different molecules*

Why do simple molecular substances have low melting points?
- *Weak intermolecular forces of attraction*
- *Do not require a lot of energy to overcome*

Why does the boiling point of simple molecular substances increase with increasing relative molecular mass?
- *Boiling overcomes the intermolecular forces of attraction between molecules*
- *Greater M_r = greater intermolecular forces of attraction to be overcome*
- *More heat energy needed to overcome these forces*

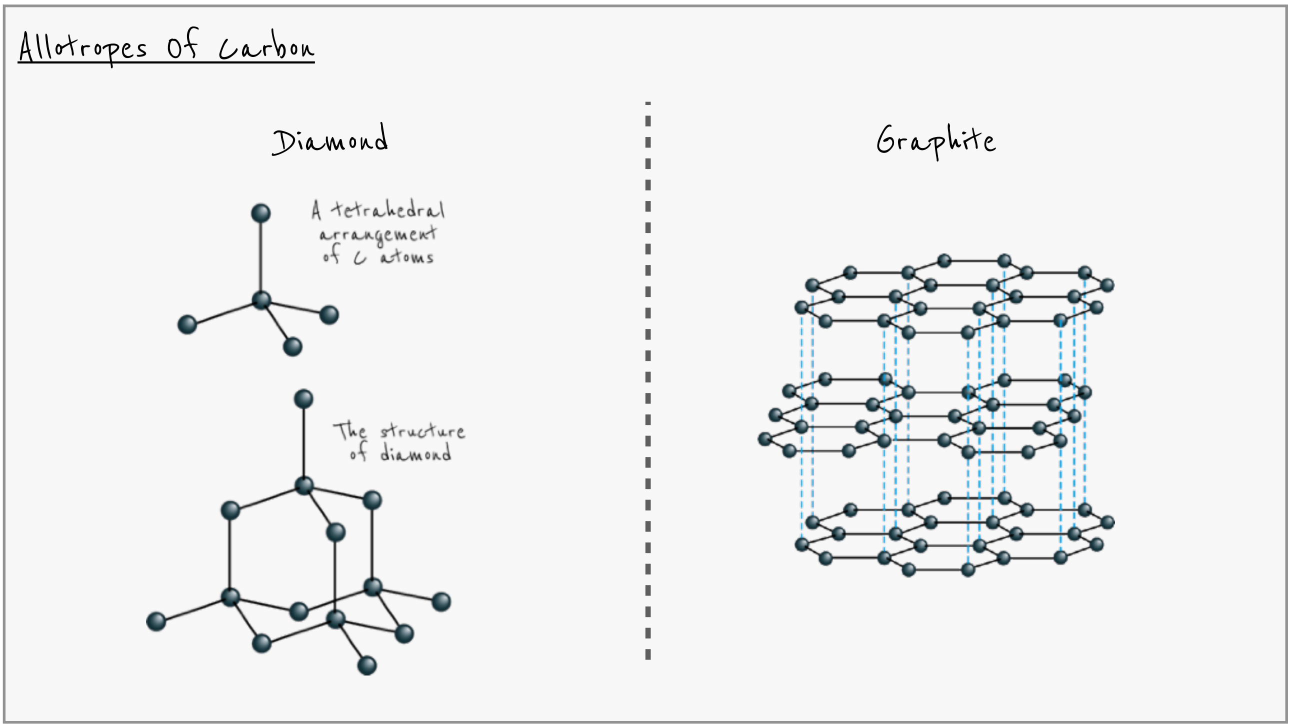
Why don't simple molecular substances conduct electricity?
- *No overall electric charge*
- *No free electrons*

2.6 Giant covalent structures

Define giant covalent structure
- Large lattice of covalently bonded atoms (macromolecules)
  - e.g. graphite, diamond, *silicon dioxide*

What is an allotrope?
- Different forms of the same element

Give 2 allotropes of carbon
- Diamond
- Graphite



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Why does diamond have such a high melting point?
- Giant covalent structure
- Each carbon atom covalently bonded to 4 others
- Many strong covalent bonds
- Require lots of energy to break

Why is diamond used in cutting tools?
- Very hard and very high melting point
  - Every carbon atom forms four strong covalent bonds to neighbouring carbon atoms
  - No free electrons and no ions
  - No weak intermolecular forces between molecules

Why don't most covalent substances conduct electricity?
- No free electrons
- Each electron in outer shell is bonded

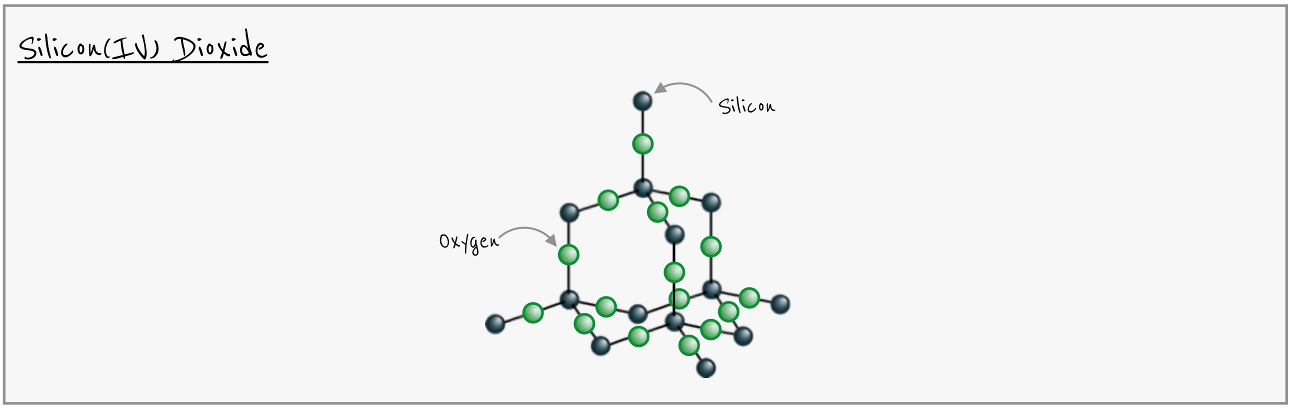
Why doesn't diamond conduct electricity?
- No free electrons
- Each electron in outer shell is bonded

Why does graphite have such a high melting point?
- Many strong covalent bonds
- Require lots of energy to break

Why does graphite conduct electricity?
- Each carbon atom is only bonded to 3 others
- 4th electron free to move and carry charge
- Graphite can therefore be used as an electrode

Why is graphite used as lubricant?
- Carbon atoms are arranged in layers
- Layers held together by weak intermolecular forces
  - Do not require a lot of energy to break
- Layers slide over each other

Describe the structure of silicon(IV) oxide (silicon dioxide)
- *Each silicon atom bonds covalently to 4 oxygen atoms*
- *Each oxygen atom bonds to 2 silicon atoms*
- *Giant structure*



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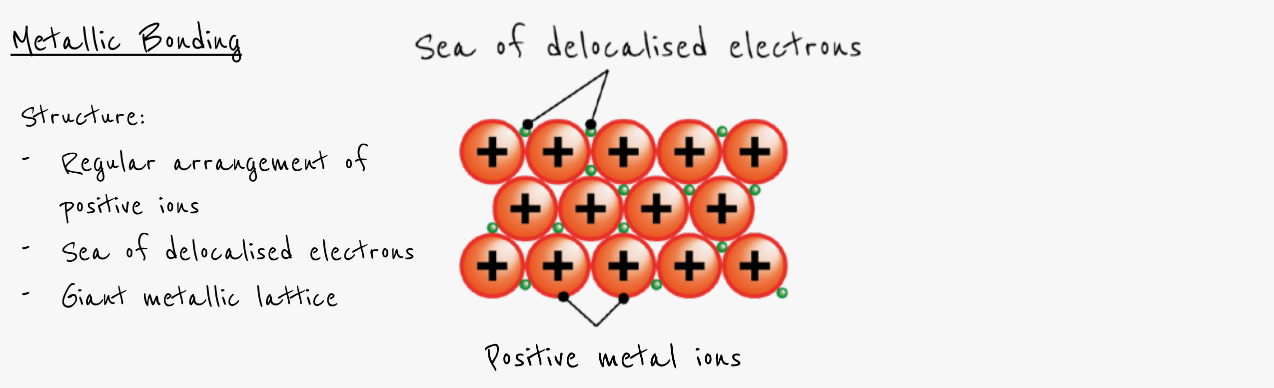
What are the similarities in properties between diamond and silicon(IV) oxide?
- *Both giant covalent structures*
- *Both have very high melting points*
  - *Many strong covalent bonds*
  - *Require lots of energy to break*
- *Both are hard structure*
  - *Each carbon (diamond) and silicon (silicon(IV) oxide) atom forms 4 strong covalent bonds*
- *Neither can conduct electricity*
  - *No free electrons or ions to carry a charge*

2.7 Metallic bonding

What is a metallic bond?

- *Electrostatic forces of attraction between positive metal ions and 'sea' of delocalised electrons*



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Why are metals good conductors of heat?

- *Delocalised electrons*
- *As electrons are free to move, heat energy is transferred throughout structure*

Why do metals conduct electricity?

- *Delocalised electrons*
- *Free to move*

Why do metals have high melting and boiling points?

- *Strong electrostatic forces of attraction*
- *Between positive metal ions and delocalised electrons*
- *Require a lot energy to break*

Why are metals malleable?

- *Layers of ions can slide over each other*

Why are metals ductile?

- *Layers of ions can slide over each other*

Bonding Summary

Structure	Simple Covalent (Molecular)	Giant Covalent	Giant Ionic Lattice	Giant Metallic Lattice
Made Up Of	Non-metals	Non-metals	Metal + Non-metal	Metals
Key Examples	Methane Ammonia Ethene	Diamond Graphite Silicon dioxide	Sodium chloride Magnesium oxide	Iron Zinc Aluminium
Melting / Boiling Point	Low	High	High	High
Conducts As Solid?	No	No (except graphite)	No	Yes
Conducts As Liquid?	No	No	Yes	Yes

3. STOICHIOMETRY

3.1 Formulae

What does molecular formula mean?
- The exact number and type of different atoms in one molecule

What are the state symbols?
- Gas = (g)
- Liquid = (l)
- Solid = (s)
- Aqueous (in water) = (aq)

Worked Example 2: Balancing Equations

Give the balanced symbol equation for the reaction between copper and oxygen

copper + oxygen → copper oxide

Remember: When balancing equations, only change the large number IN FRONT of each element/compound. Do not change the little numbers!

Step 1: Write out your symbols (include state symbols if possible)

$$\_Cu\ (s) + \_O_2\ (g) \rightarrow \_CuO\ (s)$$

Step 2: Create a tally chart for the number of each atom on both sides of the equation

Cu	I		Cu	I
O	II		O	I

Step 3: We need an extra oxygen on the right hand side, so add an extra CuO

Cu	I		Cu	II
O	II		O	II

Step 4: We now need an extra Cu on the left hand side, so add an extra Cu

Cu	II		Cu	II
O	II		O	II

Step 5: Check that everything balances, and write out the balanced symbol equation using the number of each element/compound you've used (there's no need to write a '1' in front of the O_2)

$$2\ Cu\ (s) + O_2\ (g) \rightarrow 2\ CuO\ (s)$$

What does empirical formula mean?
- *The simplest whole number ratio of the different atoms or ions present in a compound*

3.2 Relative masses of atoms and molecules

What is relative atomic mass (A_r)?
- The average mass of the isotopes of an element compared with 1/12th of the mass of an atom of ^{12}C

What is relative molecular mass (M_r)?
- The sum of the relative atomic masses in a molecule

Worked Example 3: Calculating M_r From A_r

Find the M_r of $MgCO_3$

Use the periodic table to find the mass number of each of the elements present

Answer: 24 + 12 + (3 x 16)
= 84

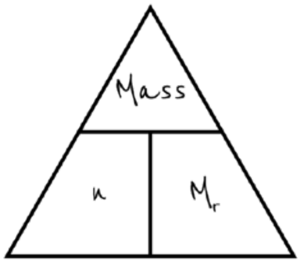
3.3 The mole and the Avogadro constant

Give the units of concentration
- mol/dm^3
- g/dm^3

What is the Avogadro constant?
- *The number of elements/molecules in a mole*
- *6.02×10^{23}*

Define the mole
- *Amount of a substance that contains 6.02×10^{23} particles*

How do you calculate the number of moles in a given mass?
- *Number of moles = $\frac{\text{mass (in grams)}}{\text{molar mass } (M_r)}$*



Worked Example 4 & 5: Calculations Involving Moles, Mass & M_r

Example 4: Find the number of moles in 54g of H_2O

Answer: *Moles = Mass / M_r*
Moles = 54 / 18
Moles = 3

Example 5: Find the mass of 0.2 moles of $CaCO_3$

Answer: *Mass = Moles x M_r*
Mass = 0.2 x (40 + 12 + (3x16))
= 0.2 x 100
= 20g

Worked Example 6: Calculating Reacting Masses

4.5g of hydrochloric acid, HCl, reacted with calcium hydroxide, $Ca(OH)_2$. Calculate the mass of Calcium Chloride, $CaCl_2$, formed.

Step 1: Balance your equation and set out a table like this:

	2HCl +	$Ca(OH)_2$ →	$CaCl_2$ +	H_2O
Mass				
M_r				
Moles				

Step 2: Fill in the information that you know (or can work out from the periodic table)

	2HCl +	$Ca(OH)_2$ →	$CaCl_2$ +	H_2O
Mass	4.5		You are working out this value	
M_r	1 + 35.5 = 36.5		40+(35.5x2) = 111	
Moles	4.5 / 36.5 = 0.123287…			

Step 3: Use the stoichiometry (big numbers in front of the compounds) to calculate the unknown moles of $CaCl_2$. There is a '2' in front of HCl and an (invisible) '1' in front of $CaCl_2$ so there are half as many moles of $CaCl_2$ produced as there are moles of HCl being used

	2HCl +	$Ca(OH)_2$ →	$CaCl_2$ +	H_2O
Mass	4.5		You are working out this value	
M_r	36.5		111	
Moles	0.123287		0.123287 / 2 = 0.061643	

Step 4: Calculate the mass of $CaCl_2$ produced

	2HCl +	$Ca(OH)_2$ →	$CaCl_2$ +	H_2O
Mass	4.5		111 x 0.061643 = 6.84g	
M_r	36.5		111	
Moles	0.123287		0.061643	

Answer = 6.84g

Worked Example 7: Determining Limiting Reagents

Copper reacts with concentrated nitric acid according to the equation

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2H_2O + 2NO_2$$

3.2g of copper is reacted with 0.4mol concentrated nitric acid. Work out which reagent is limiting and which is in excess

Step 1: Set out your table and fill in the known information

	Cu +	$4HNO_3$ →	$Cu(NO_3)_2$	+ $2H_2O$ + $2NO_2$
Mass	3.2			
M_r	63.5			
Moles	3.2 / 63.5 = 0.05…	Theoretical number of moles required		

Step 2: Calculate the theoretical number of moles of concentrated nitric acid required using the stoichiometry

	Cu +	$4HNO_3$ →	$Cu(NO_3)_2$	+ $2H_2O$ + $2NO_2$
Mass	3.2			
M_r	63.5			
Moles	0.05…	0.05 x 4 = 0.2 moles		

According to the question, we have 0.4 moles of concentrated nitric acid, but we only need 0.2 moles to react with the 3.2g of copper

Therefore the concentrated nitric acid is in excess (and the copper is limiting)

Give the formula that links moles, concentration and volume

- *Moles = Concentration x Volume*
  - *Make sure to convert volume to dm^3 if necessary*
  - *1 dm^3 = 1000 cm^3*

n = Number of moles
c = Concentration (mol dm⁻³)
v = Volume (dm³)

Worked Example 8: Moles, Concentration & Volume

Calculate the number of moles in 25cm³ of 2 mol/dm³ HCl

Answer:

$25cm^3 = 0.025dm^3$

moles = concentration x volume

moles = 2 x 0.025

moles = 0.05

What is the volume that one mole of gas occupies?

- 24 dm^3 (at room temperature and pressure)
  - 24 dm^3 = 24000 cm^3

Worked Example 9: Gas Volumes

Find the volume of CO_2 produced when 60g of ethane is burnt completely

$$C_2H_6 + 3.5O_2 \rightarrow 2CO_2 + 3H_2O$$

Answer:

Moles of ethane = mass / M_r

Moles of ethane = 60/30

Moles of ethane = 2

Moles of CO_2 produced = 2 x 2 (using stoichiometry)

Moles of CO_2 produced = 4

Volume of CO_2 produced = 24 x 4

Volume of CO_2 produced = 96 dm^3

Worked Example 10: Calculating Empirical Formula

Find the empirical formula of a compound which contained 5.85g K, 2.10g N and 4.8g O

Step 1: Calculate the moles of each element in the compound

	K	N	O
Mass	5.85	2.1	4.8
M_r	39	14	16
Moles	0.15	0.15	0.3

Step 2: Divide all the moles by the smallest value (0.15 in this example)

	K	N	O
Divide by smallest number	0.15/0.15 = 1	0.15/0.15 = 1	0.3/0.15 = 2
Ratio	1	1	2

The ratio gives the number of atoms of each element. (You may need to multiply them all to give a whole number)

Answer = KNO_2

Worked Example 11: Titration Calculations

20.0 cm^3 hydrochloric acid were used to neutralise 25.0cm^3 of 0.10mol/dm^3 calcium hydroxide

$$2HCl\ (aq) + Ca(OH)_2\ (aq) \rightarrow CaCl_2\ (aq) + 2H_2O\ (l)$$

Calculate the concentration of the hydrochloric acid

Answer:
Step 1: Balance the equation (if necessary) and set out a table like this with the information provided

	$2HCl$	+ $Ca(OH)_2$ →	$CaCl_2$ +	$2H_2O$
Concentration (moles/dm^3)	Conc. of HCl required	0.10		
Volume (dm^3)	20/1000 = 0.02	25/1000 = 0.025		
Moles				

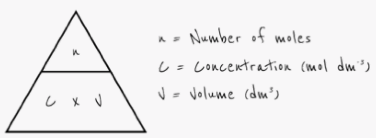
Step 2: Calculate the moles of $Ca(OH)_2$ in the solution

	$2HCl$	+ $Ca(OH)_2$ →	$CaCl_2$	+ $2H_2O$
Concentration (moles/dm^3)	Conc. of HCl required	0.10		
Volume (dm^3)	0.02	0.025		
Moles		0.10 x 0.025 = 0.0025		

Step 3: Use stoichiometry to calculate the amount of moles of HCl required (1 mole of $Ca(OH)_2$ reacts with 2 moles of HCl)

	$2HCl$	+ $Ca(OH)_2$ →	$CaCl_2$	+ $2H_2O$
Concentration (moles/dm^3)	Conc. of HCl required	0.10		
Volume (dm^3)	0.02	0.025		
Moles	2 x 0.0025 = 0.005	0.0025		

Step 4: Use concentration = moles / volume to calculate the concentration of HCl



Concentration = 0.005 / 0.02
= 0.25 mol/dm^3

How is percentage yield calculated?

- *Percentage yield =* $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$

Worked Example 12: Calculating Percentage Yield

Copper oxide is reacted with sulfuric acid to make copper sulfate and water. 1.6 g of dry copper sulfate crystals are made. If the theoretical yield is 2.0 g, calculate the percentage yield of copper sulfate.

Answer:

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

$$= \frac{1.6}{2.0} \times 100$$

$$= 80\%$$

How do you calculate the percentage by mass of element in a compound?

- *% by mass =* $\frac{\text{total mass of element}}{\text{mass of compound}} \times 100$

Worked Example 13: Calculating Percentage By Mass

Find the percentage by mass of magnesium (Mg) in magnesium oxide (MgO)

Answer:

$$\% \text{ by mass} = \frac{\text{total mass of element}}{\text{mass of compound}} \times 100$$

$$= \frac{24}{(24+16)} \times 100$$

$$= 60\%$$

4. ELECTROCHEMISTRY

4.1 Electrolysis

Why don't most covalent substances conduct electricity?
- No free electrons
- Each electron in outer shell is bonded

Why don't ionic substances conduct electricity when solid?
- Ions are held tightly in fixed positions
- Not free to move

Why do ionic substances conduct electricity when molten/dissolved?
- Ions are free to move
- Can carry electric charge

What is an ion?
- Charged particle
  - An atom which has either gained or lost electrons

What is an anion?
- Negative ion (gained electrons)

What is a cation?
- Positive ion (lost electrons)

Define electrolysis
- Decomposition of a molten or aqueous ionic compound
- By the passage of an electric current

Define anode
- Positive electrode

Define cathode
- Negative electrode

Use PANC to help you:
Positive Anode
Negative Cathode

Outline the process of electrolysis
- Positively charged ions move to negative electrode
- Negatively charged ions move to positive electrode
  - Ions discharged at electrodes, producing elements

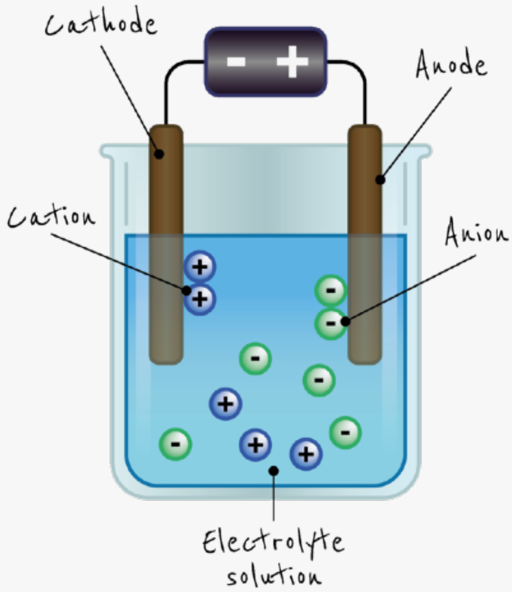
What sort of substances undergo electrolysis?
- Giant ionic structures

Why does the electrolyte need to be molten?
- So ions are free to move

What are the electrodes usually made out of?
- Inert substances e.g. graphite

What does oxidation and reduction mean? (Use OIL RIG to help you)
- Oxidation is loss (of electrons)
- Reduction is gain (of electrons)

Simple Electrolysis Apparatus



Credit: Tom Morley for SwH Learning

How to write half equations:

- *Make sure all charges and number of atoms balance*
- *At the negative electrode (cathode)…*
  - *Positive ions (cations) are attracted*
  - *They gain electrons to form elements*
    - *e.g.* $Ag^{+} + e^{-} \rightarrow Ag$
- *At the positive electrode (anode)*
  - *Negative ions (anions) are attracted*
  - *They lose electrons to form elements/compounds*
    - *e.g.* $2Br^{-} - 2e^{-} \rightarrow Br_2$
    - *e.g.* $4OH^{-} - 4e^{-} \rightarrow 2H_2O + O_2$
    - *Note: These can also be written as* $2Br^{-} \rightarrow Br_2 + 2e^{-}$ *and* $4OH^{-} \rightarrow 2H_2O + O_2 + 4e^{-}$

Electrolysis Rules: Molten Ionic Compounds

For electrolysis using a molten ionic compound…

Liquid electrolyte contains:

- Ions that make up the ionic compound

At the negative cathode:

- Solid metal forms from cations in the ionic compound
  - e.g. $Cu^{2+} + 2e^{-} \rightarrow Cu$

At the positive anode:

- Negative ion in ionic compound forms its element
  - e.g. $2Br^{-} - 2e^{-} \rightarrow Br_2$

Worked Example 14: Electrolysis Of A Molten Ionic Compound

Predict the products at the cathode and anode of the electrolysis of molten lead (ii) bromide using inert platinum electrodes

Answer: At the cathode: $Pb^{2+} + 2e^- \rightarrow Pb$ (electrons gained = reduction)

At the anode: $2Br^- - 2e^- \rightarrow Br_2$ (electrons lost = oxidation)

Electrolysis Rules: Ionic Solutions (i.e. An Ionic Compound Dissolved In Water)

For electrolysis using an ionic solution…

Solution contains:
- Ions that make up the ionic compound
- The ions in water (OH^- and H^+)

At the negative cathode:
- Least reactive element forms from the cations present
  - i.e. hydrogen (from water) produced **unless…**
    - The cations from ionic compound are from a metal **less reactive** than hydrogen
      - If metal is less reactive the hydrogen, the metal will be produced at cathode instead

At the positive anode:
- Oxygen produced **unless…**
  - The ionic compound contains **halide ions** (e.g. Cl^-, Br^-, I^-)
    - If halide ions present, halogen produced instead (e.g. Cl_2, Br_2, I_2)

Worked Example 15: Electrolysis Of An Ionic Solution (Aqueous Sodium Chloride)

Predict the products at the cathode and anode of the electrolysis of aqueous sodium chloride

Answer: At the cathode: $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$ (electrons gained = reduction)

At the anode: $2Cl^- (aq) - 2e^- \rightarrow Cl_2 (g)$ (electrons lost = oxidation)

Why does hydrogen gas form at the negative electrode?
- Because hydrogen is less reactive than sodium

Why does chlorine form at the positive electrode?
- Because Cl^- is a halide ion

What is left in solution?
- Sodium ions and hydroxide ions
  - Forms sodium hydroxide

SwH LEARNING

Worked Example 16: Electrolysis Of An Ionic Solution (Dilute Sulfuric Acid)

Give the half equations for the electrolysis of dilute sulfuric acid (H_2SO_4) using platinum electrodes

Answer: At the cathode: $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$ (electrons gained = reduction)

At the anode: $4OH^- (aq) - 4e^- \rightarrow 2H_2O (l) + O_2 (g)$ (electrons lost = oxidation)

What will happen at the cathode?
- Hydrogen gas produced (bubbles form)
  - No other positive ions present

What will happen at the anode?
- Oxygen gas produced (bubbles form)
  - No halide ions present

Worked Example 17: Electrolysis Of An Ionic Solution (Aqueous Copper (II) Sulfate)

Give the half equations for the electrolysis of aqueous copper (II) sulfate using an impure copper anode and pure copper cathode

Answer: *At the cathode:* $Cu^{2+} (aq) + 2e^- \rightarrow Cu (s)$ *(electrons gained = reduction)*

At the anode: $Cu (s) - 2e^- \rightarrow Cu^{2+} (aq)$ *(electrons lost = oxidation)*

What will happen to the cathode?
- *Gains pure copper*
- *Increases in mass*

What will happen to the anode?
- *Loses copper ions (and impurities)*
- *Decreases in mass*

What will be formed at the positive and negative electrodes if inert electrodes are used instead?
- *Positive: Oxygen gas*
- *Negative: Copper*

How is charge transferred during electrolysis?
- *Negatively charged electrons move from negative terminal towards positive*
- *New electrons added to circuit at the cathode*
- *Electrons removed at anode*
- *Ions transfer charge in electrolyte to electrode of opposite charge*
  - *Positive ions move to cathode*
  - *Negative ions move to anode*

Why are metal objects electroplated?
- To improve appearance
- To increase resistance to corrosion
  - e.g. coating steel cans with tin
  - e.g. coating jewellery with silver

Describe the process of electroplating a metal
- Use the electroplating metal as anode
- Item to be plated used as cathode
- Electrolyte is solution of soluble compound of the electroplating metal
  - Metal ions reduced at cathode
  - Cathode is plated with metal

4.2 Hydrogen-oxygen fuel cells

Describe production of electrical energy in a simple cell
- Simple cell is two metals and electrolyte
- The redox reactions form a current
- More reactive metal is negative pole
  - Negative pole produces electrons
- Electrons flow along wire forming current
- If ions in solution are less reactive than the anode, they accept electrons

What is a hydrogen-oxygen fuel cell?
- Electrochemical cell
- Uses hydrogen and oxygen to produce electricity
- Water is only chemical product
  - $2H_2 (g) + O_2 (g) \rightarrow 2H_2O (l)$

Outline the advantages and disadvantages of using hydrogen-oxygen fuel cells over petrol engines

Advantage	Disadvantage
Do not produce polluting gases (e.g. CO_2, a greenhouse gas)	Hydrogen is expensive to produce and store
Produce more energy per kg	Hydrogen and oxygen gases are explosive - require high pressure tanks for storage
Less energy lost as heat (no moving parts compared to combustion engine)	Hydrogen-oxygen fuel cells less efficient at low temperatures

5. CHEMICAL ENERGETICS

5.1 Exothermic and endothermic reactions

Define exothermic reaction
- Thermal energy is transferred to surroundings
- Causes increase in temperature

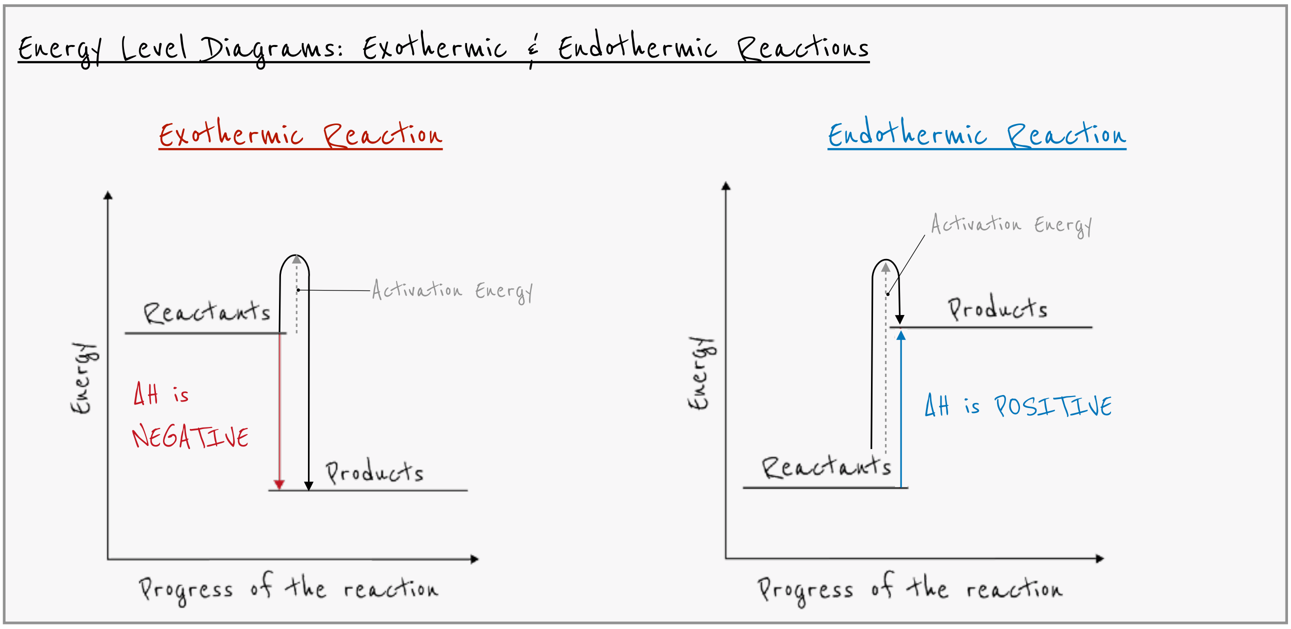
Define endothermic reaction
- Thermal energy is taken in from surroundings
- Causes decrease in temperature

Define ΔH
- *Enthalpy (energy) change of a reaction*
  - *Transfer of thermal energy during a reaction*

What is the value of ΔH for an exothermic reaction?
- *ΔH is negative*
  - *More energy needed to make bonds than break bonds*
  - *Bond making is an exothermic process*
  - *Bond breaking is an endothermic process*

What is the value of ΔH for an endothermic reaction?
- *ΔH is positive*
  - *Less energy needed to make bonds than break bonds*

Define activation energy
- *Minimum amount of energy that colliding particles must have for a reaction to occur*



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Explain the shape of the energy level diagram for an exothermic reaction
- *More energy needed to make bonds than break bonds*
- *Reactants have more energy than products*

Explain the shape of the energy level diagram for an endothermic reaction
- *More energy needed to break bonds than make bonds*
- *Reactants have less energy than products*

Define bond energy

- *Energy needed to break the bond between two atoms*

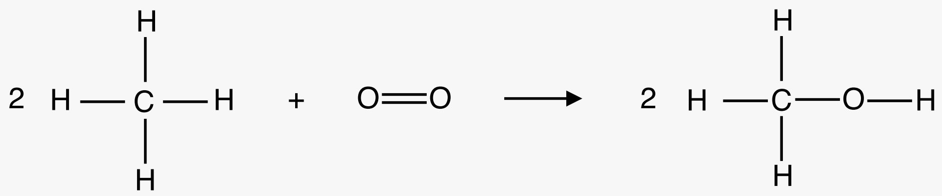
How are bond energies used to calculate the enthalpy change during a chemical reaction?

- *ΔH = bonds broken - bonds formed*

Worked Example 18: Calculating Enthalpy Change Using Bond Energies

Methane (CH_4) and oxygen (O_2) are reacted together to make ethanol. The reaction is exothermic.

Use the bond energies below to calculate the enthalpy change for this reaction.



Bond	Bond energy (kJ)
C — H	435
O = O	497
C — O	336
O — H	464

Step 1: Calculate the bond energy of bonds broken (reactants). Remember the stoichiometry.

Methane: 2 x (4 x C — H) = 2 x 4 x 435
= 3480 kJ

Oxygen: 1 x O = O = 497 kJ

Total bonds broken = 3480 + 497
= 3977 kJ

Step 2: Calculate the bond energy of bonds formed (products). Remember the stoichiometry.

Methanol: 2x ((3 x C — H) + (1 x C — O) + (1 x O — H))

Total bonds formed = 2 x ((3 x 435) + (336) + (464))
= 2 x 2105
= 4210 kJ

Step 3: Calculate the enthalpy change of the reaction

ΔH = bonds broken - bonds formed
ΔH = 3977 - 4210
ΔH = -233 kJ

*ΔH is negative so this is an **exothermic** reaction*

Remember:

*Bond **breaking** = **endothermic** process*

*Bond **making** = **exothermic** process*

6. CHEMICAL REACTIONS

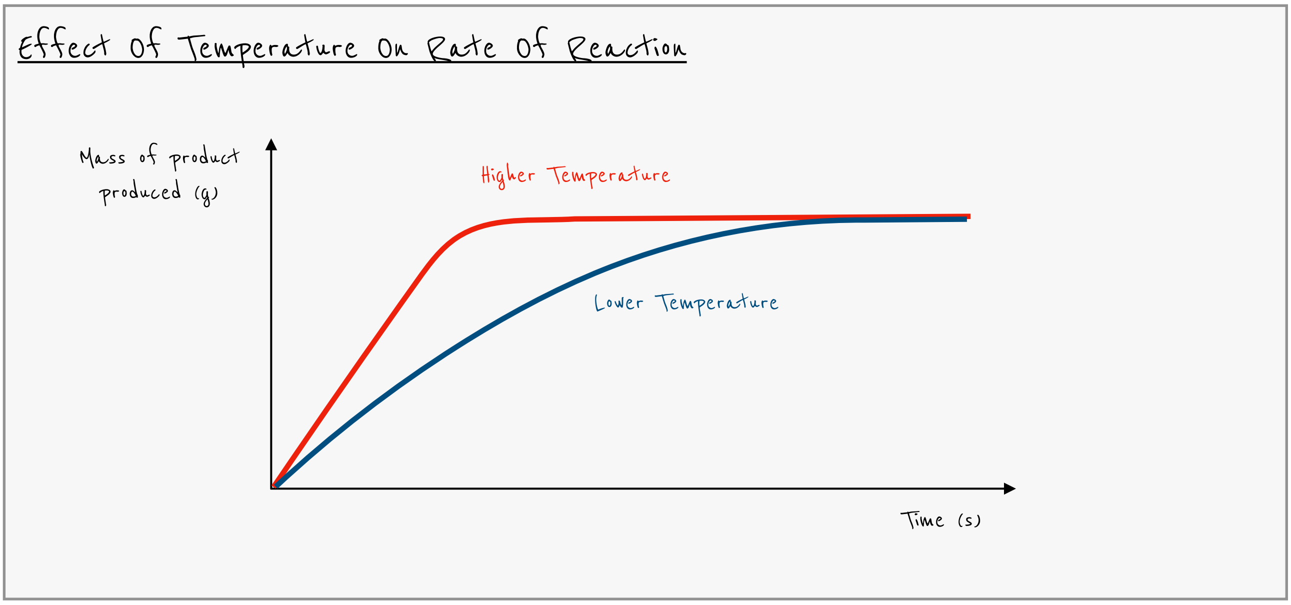
6.1 Physical and chemical changes

What is the difference between a chemical and a physical change?
- Physical changes don't make new chemical substances
- Easy to reverse
  - e.g. mixing, dissolving
- Chemical changes (reactions) produce new chemical substances
- Cause a change in energy (exothermic or endothermic)
- Chemical changes are difficult to reverse
  - e.g. combustion reactions, neutralisation reactions

6.2 Rate of reaction

Use collision theory to explain how chemical reactions occur
- *Chemical reactions occur when particles collide with sufficient energy*
  - *Minimum amount of energy required = activation energy*
- *Rate of reaction is dependent on the energy and number of collisions*
- *Greater kinetic energy and increased frequency of collisions = faster rate of reaction*

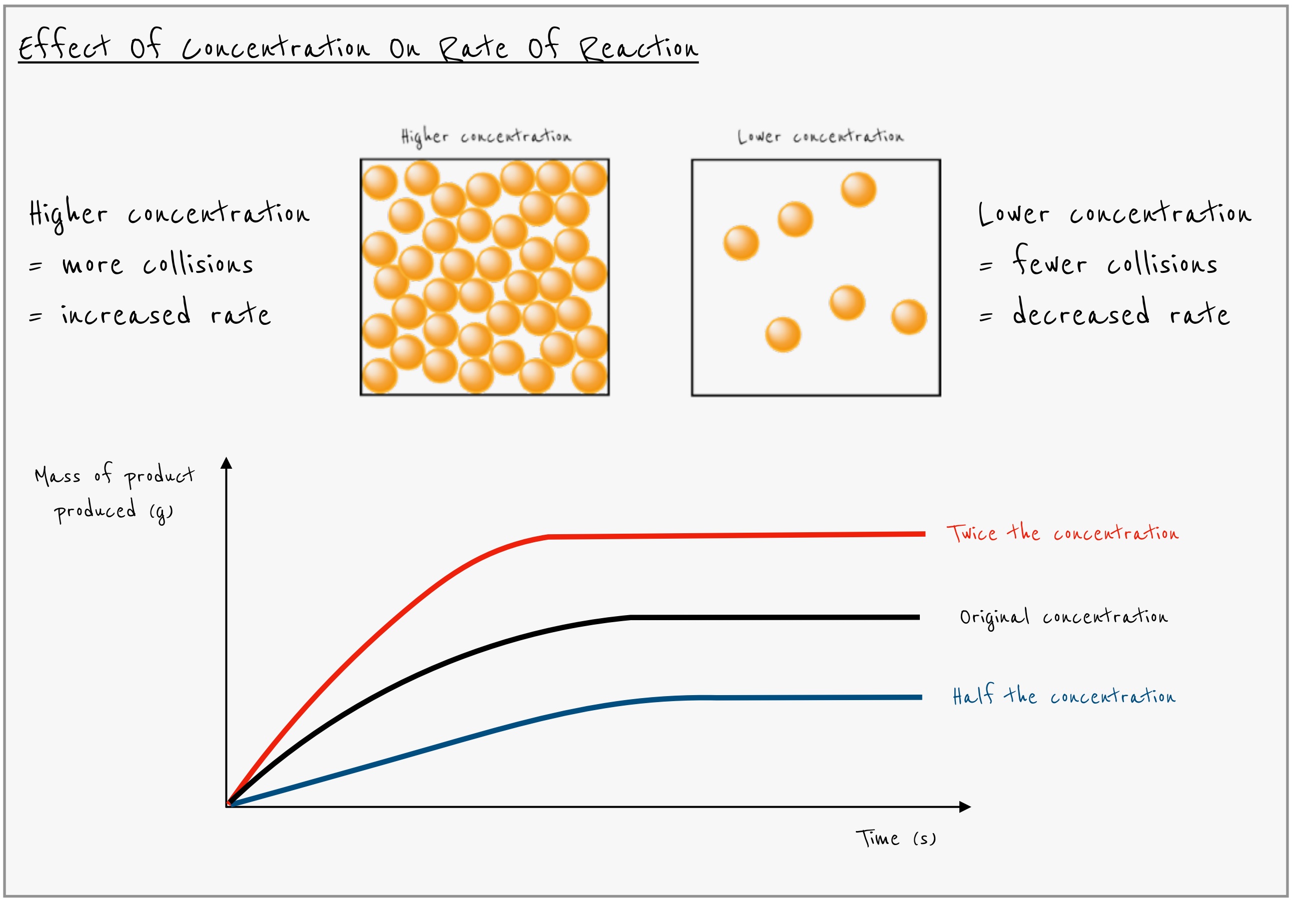
What effect does a higher temperature have on the rate of reaction?
- Rate increases
  - *Particles have greater kinetic energy*
  - *More collisions have energy greater than activation energy*
  - *More successful collisions occur per second*



Credit: Martin Bailey for SwH Learning

What effect does increasing the pressure of a gas have on the rate of reaction?
- Rate increases
  - *More particles in the same volume*
  - *Collisions occur more frequently*

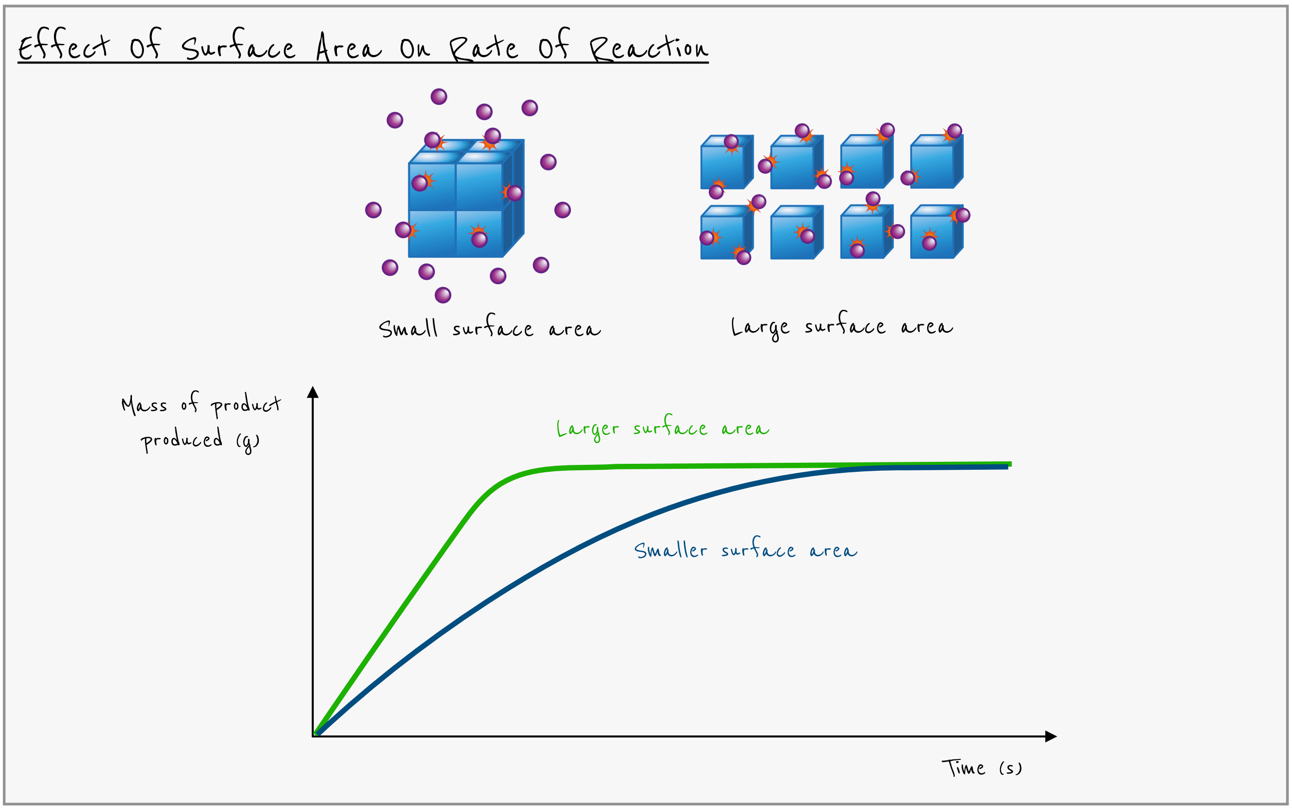
What effect does increasing the concentration have on the rate of reaction?
- Rate increases
  - *More particles in the same volume*
  - *Collisions occur more frequently*



Credit: Tom Morley for SwH Learning

What effect does increasing the surface area have on the rate of reaction?

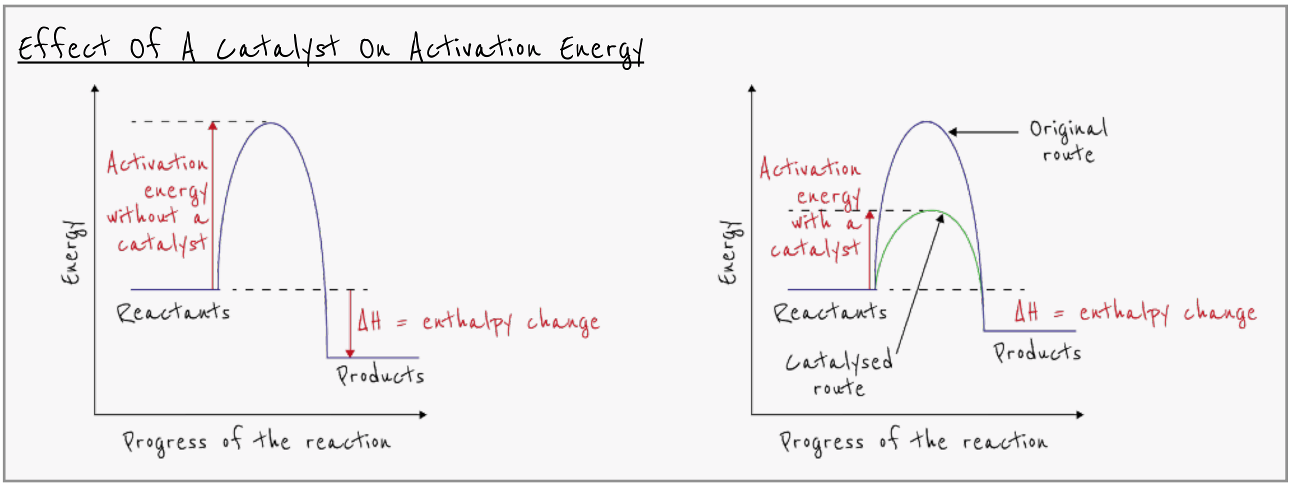
- Rate increases
    - *Collisions occur more frequently*



Credit: Tom Morley for SwH Learning

Define catalyst
- A substance that increases the rate of reaction
- Chemically unchanged at the end of the reaction

How does a catalyst work?
- Speeds up the rate of reaction without being used up
- *Provides alternative reaction pathway with lower activation energy*



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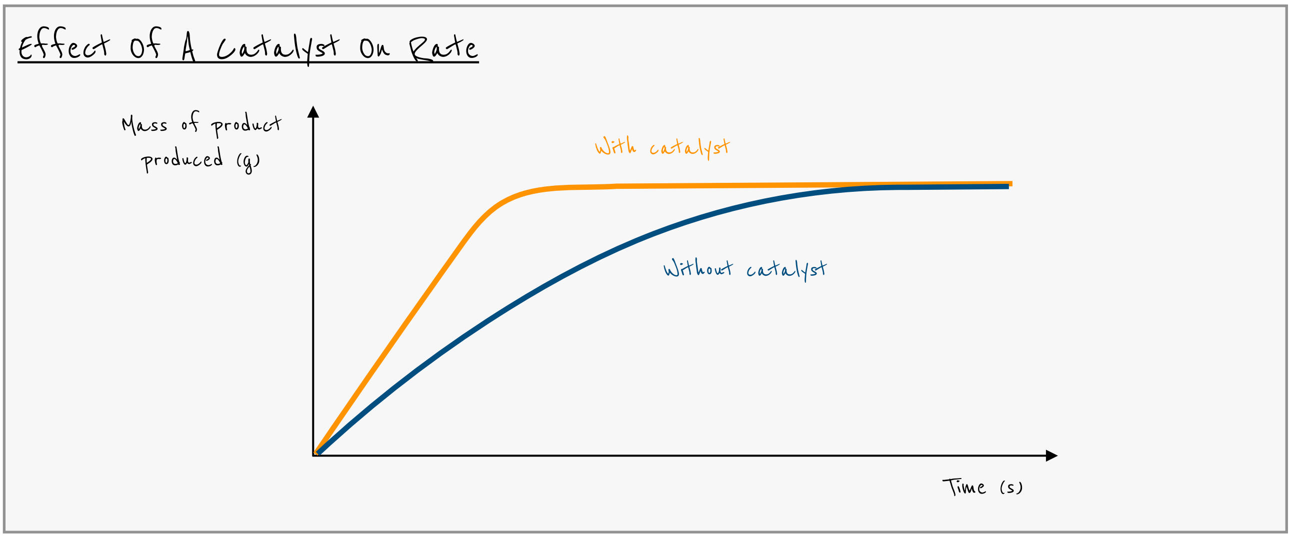

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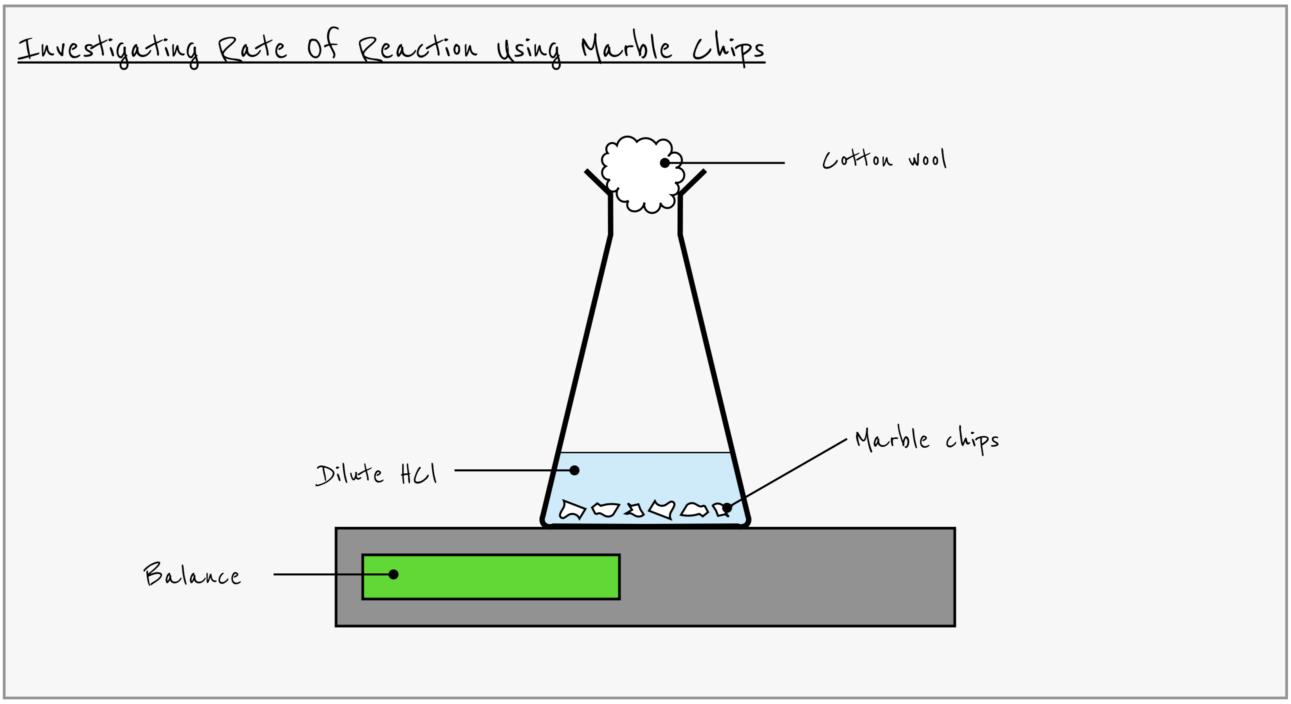
How is the rate of a reaction calculated?
- Either: Rate = $\frac{\text{amount of reactant used}}{\text{time}}$
- Or: Rate = $\frac{\text{amount of product formed}}{\text{time}}$

Describe how to carry out an experiment to measure the rate of reaction when a precipitate is formed
- Example: Sodium thiosulfate solution and dilute hydrochloric acid react together to form a precipitate of sulfur
  - Place 20 cm^3 of sodium thiosulfate solution and 20 cm^3 of water in a conical flask
  - Add 10 cm^3 of dilute hydrochloric acid to the flask
  - Place the flask on a piece of paper marked with a black X
  - Time how long it takes before the X can no longer be seen

Describe how to carry out an experiment to measure the rate of reaction when a gas is produced
- Example: The decomposition of hydrogen peroxide (H_2O_2) into H_2O and O_2
  - Place a known mass of H_2O_2 in a conical flask
  - Seal flask with a bung attached to a gas syringe
  - Time how long it takes for 100cm^3 of oxygen to be collected

Describe how to carry out an experiment to measure the rate of reaction from the change in mass of a solid
- Example: Reaction of marble chips ($CaCO_3$) with dilute hydrochloric acid
  - Place 50cm^3 HCl into a conical flask. Place the flask on a balance
  - Add marble chips and note the starting mass
    - Place cotton wool in the neck of the flask to stop acid spraying out
  - Measure the mass of carbon dioxide lost at intervals
  - Plot a graph of the results
  - Repeat changing one variable only
    - e.g. temperature, surface area, concentration of HCl



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Which variables should be controlled when investigating the effect of temperature on the rate of a reaction?
- Concentration of reactants
- Volume of reactants
- Surface area (if using a solid)

Which variables should be controlled when investigating the effect of concentration on the rate of a reaction?
- Concentration of all reactants other than the one being investigated
- Volume of reactants
- Surface area (if using a solid)
- Temperature

Which variables should be controlled when investigating the effect of surface area on the rate of a reaction?
- Concentration of reactants
- Volume of reactants
- Temperature

Which variables should be controlled when investigating the effect of a catalyst on the rate of a reaction?
- Concentration of reactants
- Volume of reactants
- Temperature
- Surface area

6.3 Reversible reactions and equilibrium

What does this $\rightleftharpoons$ arrow represent?
- Reaction is reversible

Describe the dehydration of copper (II) sulfate crystals
- Heat blue (hydrated) copper (II) sulfate crystals
- Blue crystals turn to white powder
    - Because water is lost i.e. water of crystallisation is lost
- White anhydrous copper (II) sulfate is formed
- Add water and blue (hydrated) copper (II) sulfate crystals form
- Reaction is reversible

Describe the dehydration of cobalt (II) chloride
- Heat pink (hydrated) cobalt (II) chloride paper
- Pink paper turns to blue
    - Because water is lost i.e. water of crystallisation is lost
- Blue anhydrous cobalt (II) chloride is formed
- Add water and pink (hydrated) cobalt (II) chloride forms
- Reaction is reversible

What is a dynamic equilibrium?
- *Forward and reverse reactions occur at the same rate*
- *Concentration of reactants and products remains constant*
- *Requires sealed container (closed system) so substances do not escape*

REMEMBER: When a change is made to a system in equilibrium, the system moves to oppose that change
- *e.g. if you heat up a system, the system will try and make it cooler (carry out more of the endothermic reaction)*

Describe and explain the effect of increasing the temperature on a reversible reaction
- *Increasing temperature favours the endothermic reaction*
    - *Position of equilibrium shifts in favour of the endothermic reaction*

Describe and explain the effect of decreasing the temperature on a reversible reaction
- *Decreasing temperature favours the exothermic reaction*
    - *Position of equilibrium shifts in favour of the exothermic reaction*

Describe and explain the effect of increasing the pressure on a reversible reaction
- *Increasing pressure favours the side with fewer moles of gas*
    - *Position of equilibrium shifts to the side with fewer moles of gas*

Describe and explain the effect of decreasing the pressure on a reversible reaction
- *Decreasing pressure favours the side with more moles of gas*
    - *Position of equilibrium shifts to the side with more moles of gas*

Describe and explain the effect of increasing the concentration of reactants on a reversible reaction
- *Increasing concentration of reactants favours the formation of more products*
  - *Position of equilibrium shifts to the right to reduce the concentration of reactants*

Describe and explain the effect of increasing the concentration of products on a reversible reaction
- *Increasing concentration of products favours the formation of more reactants*
  - *Position of equilibrium shifts to the left to reduce the concentration of products*

What effect does the addition of a catalyst have on the position of equilibrium?
- *Catalyst has no effect on the position of equilibrium*
  - *Because rate of the forward and reverse reactions increase equally*
- *Reaction reaches equilibrium faster*

Worked Example 19: Dynamic Equilibrium

$2NO_2\,(g) \rightleftharpoons N_2O_4\,(g) \qquad \Delta H = -57$ KJ/mol

brown — colourless

Describe and explain the effect of increasing the temperature on the colour of the reaction vessel
- *Forwards reaction is exothermic (produces heat energy)*
- *Increasing temperature favours the endothermic (reverse) reaction*
- *Position of equilibrium shifts to the left*
- *Mixture turns brown*

Describe and explain the effect of increasing the pressure on the colour of the reaction vessel
- *There are 2 moles of gas on the left hand side and 1 mole of gas on the right hand side*
- *Increasing the pressure favours the side with fewer moles of gas (the right hand side)*
- *Position of equilibrium shifts to the right*
- *Mixture turns colourless*

What are the conditions used for the manufacture of ammonia (the Haber process)?
- $N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g) \qquad \Delta H = -92$ KJ/mol
  - *Temperature = 450°C*
  - *Pressure = 200 atmospheres (20,000 kPa)*
  - *Catalyst = Iron*

Why are the conditions used in the Haber process described as 'compromised'?
- *Forward reaction is exothermic - favoured by low temperatures*
  - *But rate of reaction is too slow at low temperature so 450°C temperature is used*
- *Forward reaction results in fewer moles of gas so is favoured by high pressures*
  - *But high pressures are dangerous and expensive so 200 atmospheres is used*

What effect does the iron catalyst have on the yield of ammonia in the Haber process?
- *No effect on yield*
- *Increases the rate of the forward and reverse reactions equally*

Give the sources of hydrogen and nitrogen for the Haber process
- *Hydrogen from methane*
- *Nitrogen from air*

What is the equation for the conversion of sulfur dioxide to sulfur trioxide (the Contact process)?
- *Sulfur dioxide + oxygen ⇌ sulfur trioxide*
- $2SO_2\,(g) + O_2\,(g) \rightleftharpoons 2SO_3\,(g)$

Give the conditions used for the Contact process
- *Temperature = 450°C*
- *Pressure = 1-2 atmospheres (100-200 kPa)*
- *Catalyst = Vanadium(V) oxide, V_2O_5*

Give the sources of sulfur dioxide and oxygen for the Contact process
- *Sulfur dioxide from burning sulfur or roasting sulfide ores*
  - *$S + O_2 \rightarrow SO_2$*
- *Oxygen from air*

Why are the conditions used in the Contact process described as 'compromised'?
- *Forward reaction is exothermic - favoured by low temperatures*
  - *But rate of reaction is too slow at low temperature so 450°C temperature is used*
- *Forward reaction results in fewer moles of gas so is favoured by high pressures*
  - *But high pressures are dangerous and expensive so 2 atmospheres is used*
  - *Higher pressures also cause SO_2 to liquify, removing it from the reaction and moving the position of the equilibrium to the left*

6.4 Redox

What does a Roman numeral indicate in the name of a compound?
- Oxidation number of an element in the compound
  - e.g. copper(II) sulfate means copper has an oxidation number of +2

Define oxidation
- Gain of oxygen
- *Loss of electrons*

Define reduction
- Loss of oxygen
- *Gain of electrons*

> ***Remember: Oil Rig***
>
> ***O**xidation **i**s **l**oss (of electrons)*
> ***R**eduction **i**s **g**ain (of electrons)*

Define redox
- A reaction where both reduction and oxidation take place at the same time

What do oxidation numbers signify?
- *The oxidation state of an atom or ion*
- *The number of electrons that an atom has lost, gained or shared to form a compound*

How To Use Oxidation Numbers

- *An element in its uncombined state has an oxidation number of 0*
- *A positive number means an element has lost electrons (it has been oxidised)*
  - *e.g. Cu^{2+} has an oxidation number of +2*
- *A negative number means that an element has gained electrons (it has been reduced)*
  - *e.g. O^{2-} has an oxidation number of –2*
- *The more positive the number, the more the element has been oxidised. The more negative the number, the more it has been reduced*
- *The oxidation numbers always have a + or – sign (unless they are zero)*

Oxidation Number Rules

Use these rules when assigning oxidation numbers

1. *Uncombined elements have an oxidation number of 0 (e.g O_2, H_2, Mg)*

2. *Most elements always have the same oxidation number when in a compounds*

Element	Usual oxidation state in compound	Example	Exceptions
Hydrogen	+1	HCl	Metal hydrides, e.g. NaH, where it is –1
Group 1 metals	Always +1	NaCl	
Group 2 metals	Always +2	$CaCl_2$	
Aluminium	Always +3	$AlCl_3$	
Oxygen	–2	Na_2O	Peroxides where it is –1 OF_2 where it is +2
Fluorine	Always –1	NaF	
Chlorine	–1	NaCl	Compounds with F and O, where it is positive

3. *The sum of all the oxidation numbers in a balanced compound (e.g. NH_3, CH_4) = 0*

4. *The sum of the oxidation numbers of a complex ion (e.g. NH^{4+} or SO_4^{2-}) equals the charge on the ion*

Worked Example 20: Oxidation Number Questions

What is the oxidation number of sulfur in H_2SO_4?
- *Hydrogen has an oxidation number of +1. There are 2 hydrogens → 2 x 1 = +2*
- *Oxygen has an oxidation number of –2. There are 4 oxygens → 4 x –2 = –8*
- *The overall charge of H_2SO_4 = 0*
  - *2 – 8 + S = 0*
  - *S = +6*

What is the oxidation number of zinc in the compound $Zn(OH)_4{}^{2-}$
- *Hydrogen has an oxidation number of +1. There are 4 hydrogens → 4 x 1 = +4*
- *Oxygen has an oxidation number of –2. There are 4 oxygens → 4 x –2 = –8*
- *The overall charge of the compound = –2*
  - *4 – 8 + Zn = –2*
  - *Zn = +2*

Which elements/ions have been oxidised and which have been reduced in the reaction below?

$$2KBr + Cl_2 \rightarrow 2KCl + Br_2$$

+1 –1 0 → +1 –1 0

- *Bromine has increased from –1 to 0 → it has been oxidised*
- *Chlorine has decreased from 0 to –1 → it has been reduced*

Define oxidising agent

- *Substance that causes another substance to be oxidised*
- *An oxidising agent is itself reduced (it gains electrons)*

Define reducing agent

- *Substance that causes another substance to be reduced*
- *A reducing agent is itself oxidised (it loses electrons)*

Redox Reactions

Magnesium reacts with an aqueous solution of silver nitrate. The ionic equation is as follows:

Reduction

$$Mg(s) + 2Ag^{+}(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$$

Oxidation

Why is magnesium being oxidised?

- *Mg is losing electrons to form* Mg^{2+}

Why are the silver ions being reduced?

- Ag^{+} *is gaining electrons to form Ag*

Why is this a redox reaction?

- *Both oxidation and reduction are occurring*

What is the oxidising agent?

- Ag^{+}
    - *It is gaining electrons (reduced)*

What is the reducing agent?

- *Mg*
    - *It is losing electrons (oxidised)*

7. ACIDS, BASES AND SALTS

7.1 The characteristic properties of acids and bases

Describe the reactivity of metals with acids
- Metals below hydrogen in the reactivity series do not react with acids
- Metals above hydrogen in the reactivity series react with acids, producing hydrogen
  - **Note:** very reactive metals react very explosively with acids e.g. Potassium

Typical Reactions Involving Acids

Lean these different reactions off by heart:

Metal + acid → salt + hydrogen

- e.g. magnesium + hydrochloric acid → magnesium chloride + hydrogen
  - $Mg + 2HCl \rightarrow MgCl_2 + H_2$

Metal hydroxide + acid → salt + water

- e.g. Sodium hydroxide + sulfuric acid → sodium sulfate + water
  - $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Metal oxide + acid → salt + water

- e.g. magnesium oxide + hydrochloric acid → magnesium chloride + water
  - $MgO + 2HCl \rightarrow MgCl_2 + H_2O$

Metal carbonate + acid → salt + water + carbon dioxide

- e.g. potassium carbonate + hydrochloric acid → potassium chloride + water + carbon dioxide
  - $K_2CO_3 + 2HCl \rightarrow 2KCl + H_2O + CO_2$

What ion is responsible for making something acidic?
- H^+
  - Produced by acids in aqueous solutions

Which compounds are bases?
- Metal oxides
- Metal hydroxides

Define alkali
- Soluble base

What ion is responsible for making something alkaline?
- OH^-
  - Produced by alkalis in aqueous solutions

What is a neutralisation reaction?
- Addition of a base to an acid
- Produces a neutral solution

Give the ionic equation for a neutralisation reaction
- $H^+ (aq) + OH^- (aq) \rightarrow H_2O (l)$

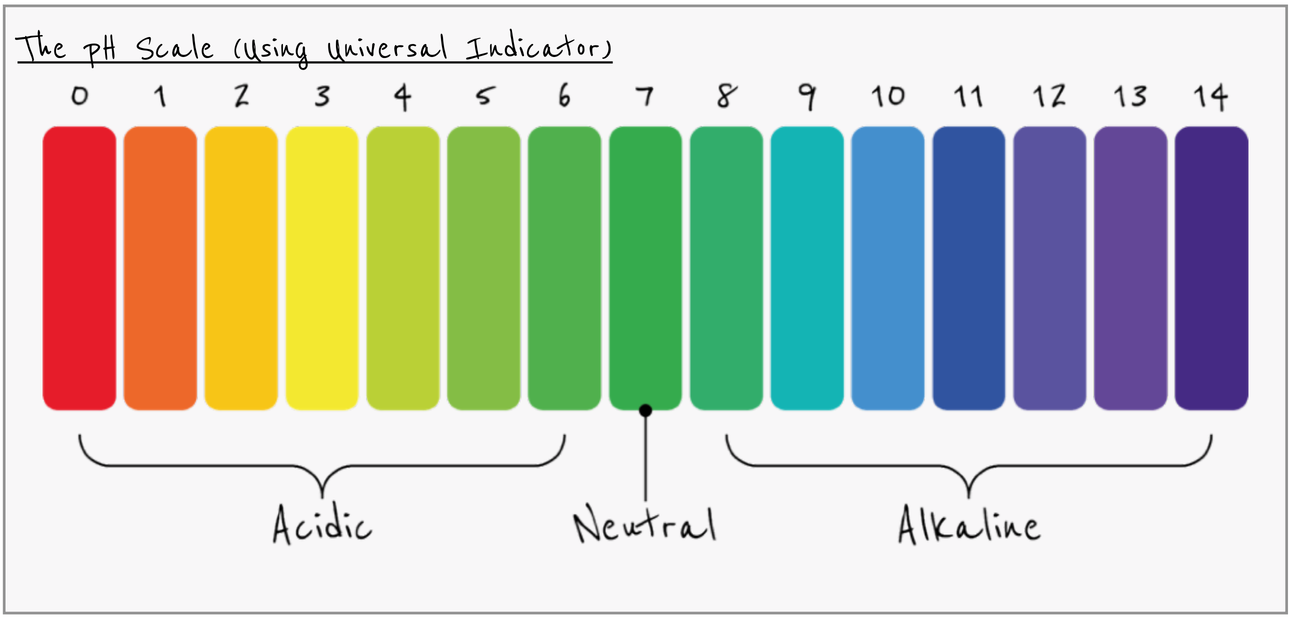
Describe the reaction of a base with an ammonium salt
- Decomposition reaction
- Ammonium ion displaced by another base
  - e.g. ammonium chloride + potassium hydroxide → ammonia + water + potassium chloride
  - $NH_4Cl + NaOH \rightarrow NH_3 + H_2O + NaCl$

Give the Brønsted-Lowry definition of an acid
- *H^+ (proton) donor*

Give the Brønsted-Lowry definition of a base
- *H^+ (proton) acceptor*
  - *e.g. metal carbonate, hydroxide, oxides, ammonia*

Define strong acid
- *An acid that dissociates completely in aqueous solution*
  - *e.g. $HCl (aq) \rightarrow H^+ (aq) + Cl^- (aq)$*

Define weak acid
- *An acid that partially dissociates in aqueous solution*
  - *e.g. $CH_3COOH (aq) \rightleftharpoons H^+ (aq) + CH_3COO^- (aq)$*
  - *Equilibrium established between acid (CH_3COOH) and ion (CH_3COO^-)*



Credit: Tom Morley for SwH Learning

Describe the pH scale
- Ranges from 0 - 14
  - 0-3 = strongly acidic (red / orange)
  - 4-6 = weakly acidic (yellow / light green)
  - 7 = neutral (green)
  - 8-10 = weakly alkaline (blue)
  - 11-14 = strongly alkaline (purple)

What is the pH scale a measure of?
- Concentration H^+ ions
- More acidic solution = greater concentration of H^+ ions

7.2 Oxides

What is an acidic oxide?
- Oxide formed from the reaction of oxygen with a non-metal
  - e.g. sulfur dioxide (SO_2), carbon dioxide (CO_2)

What is a basic oxide?
- Oxide formed from the reaction of oxygen with a metal
  - e.g. copper (II) oxide (CuO), calcium oxide (CaO)

What is an amphoteric oxide?
- *Oxide that can react with both acids and bases*
  - *e.g. aluminium oxide (Al_2O_3), zinc oxide (ZnO)*
- *Reactions produce a salt and water*
  - *e.g. Al_2O_3 (s) + 6HCl (aq) → $2AlCl_3$ (aq) + $3H_2O$ (l) (aluminium oxide reacts with acid)*
  - *Al_2O_3 (s) + 2NaOH (aq) → $2NaAlO_2$ (aq) + H_2O (l) (aluminium oxide reacts with base)*
  - *e.g. Zn_3 (s) + 2HCl (aq) → $ZnCl_2$ (aq) + H_2O (l) (zinc oxide reacts with acid)*
  - *ZnO (s) + 2NaOH (aq) → Na_2ZnO_2 (aq) + H_2O (l) (zinc oxide reacts with base)*

7.3 Preparation of salts

What is a salt?
- Ionic compound formed by the neutralisation of an acid by a base

Useful Information About Making Salts

- All acids contain hydrogen e.g. HCl, H_2SO_4 and HNO_3
- When hydrogen in an acid is displaced with a metal or ammonium ion, a salt is formed
  - e.g. magnesium sulfate, zinc chloride, ammonium chloride and potassium nitrate

Examples of acids and the salts they form:

Acid	Formula	Example of salt	Name of salts
Hydrochloric acid	HCl	KCl	Chlorides
Nitric acid	HNO_3	$NaNO_3$	Nitrates
Sulfuric acid	H_2SO_4	$MgSO_4$	Sulfates

Remembering which ionic compounds (salts) are soluble/insoluble:
- All sodium, potassium and ammonium compounds are soluble
- All nitrates are soluble
- All sulfates are soluble except lead (II) sulfate, barium sulfate, and calcium sulfate
- All chlorides are soluble except lead (II) chloride and silver chloride
- All carbonates are insoluble except ammonium, potassium and sodium salts
- All hydroxides are insoluble except ammonium, potassium and sodium salts
  - Calcium hydroxide is slightly soluble

Ionic Compound Solubility Table

Salts	Soluble	Insoluble
Sodium, Potassium & Ammonium	All	None
Nitrates	All	None
Sulfates	Most	Sulfates of lead(II), barium and calcium
Chlorides	Most	Chlorides of lead(II) and silver
Carbonates	Carbonates of ammonium, potassium and sodium	Most
Hydroxides	Hydroxides of ammonium, potassium and sodium. Calcium hydroxide is slightly soluble	Most

How do you make soluble salts (except ammonium, potassium and sodium salts)?
- Use the crystallisation method
  - REACT
  - FILTER
  - EVAPORATE: heat to evaporate some water
  - COOL: collect crystals that form (crystallisation)
  - DRY: allow the crystals to dry in a warm place or on filter paper
- For the reactants, you can use:
  - Acid + insoluble metal/metal oxide/metal hydroxide/metal carbonate

How do you make soluble salts - that do contain sodium, potassium or ammonium?
- Use the titration method:
  - REACT: an acid (from a burette) with an alkali (in a conical flask)
  - INDICATOR: requires an indicator to show when the alkali has been neutralised (all alkali has been reacted)
  - REPEAT: once amounts required have been worked out, add required volume of acid to alkali without indicator
  - EVAPORATE: heat to evaporate some water, this concentrates the solution
  - COOL: collect crystals that form
  - DRY: allow the crystals to dry in a warm place or on filter paper
- For the reactants, you can use:
  - Acid + soluble metal oxide / hydroxide (acid + alkali)
  - Acid + ammonia solution

Why can't you use the crystallisation method when reacting acid & an alkali to make a soluble salt?
- Because sodium, ammonium and potassium compounds are soluble in water
  - No insoluble excess reactant present that could be removed by filtration

Worked Example 21: Making A Soluble Salt 1

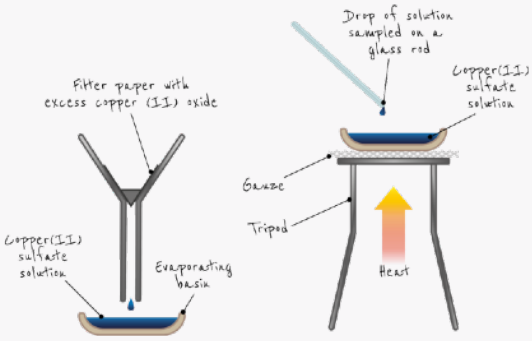
Describe how to prepare a sample of pure, dry hydrated copper (II) sulfate crystals starting from copper (II) oxide

Answer: Copper (II) sulfate is soluble, and does not contain sodium, potassium or ammonia. We therefore need to use the crystallisation method.

To make copper (II) sulfate, the copper (II) oxide needs to react with sulfuric acid

$$CuO\ (s) + H_2SO_4\ (aq) \rightarrow CuSO_4\ (aq) + H_2O\ (l)$$

- REACT: the sulfuric acid and copper oxide in a beaker
  - Keep adding CuO until no more will dissolve
- FILTER: using filter paper
- EVAPORATE: heat gently to evaporate some water
- COOL: collect crystals that form using filter paper
- DRY: allow the crystals to dry in a warm place or on filter paper



Credit: Tom Morley for SwH Learning

Worked Example 22: Making A Soluble Salt 2

Describe how to prepare a sample of pure, dry sodium chloride crystals from sodium hydroxide and hydrochloric acid

$$NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$$

Answer: Sodium chloride is soluble, and contains one of sodium, potassium or ammonia. We therefore need to use the titration method.

- *REACT: an acid (from a burette) with an alkali (in a conical flask)*
- *INDICATOR: requires an indicator to show when the alkali has been neutralised (all alkali has been reacted)*
- *REPEAT: once exact amounts required have been worked out, add required volume of acid to alkali without indicator*
- *EVAPORATE: heat gently to evaporate some water, this concentrates the solution*
- *COOL: collect crystals that form*
- *DRY: allow the crystals to dry in a warm place or on filter paper*

How do you make insoluble salts from two soluble reactants?

- *Use the precipitation method*
  - *REACT*
  - *FILTER*
  - *WASH*
  - *DRY: allow the precipitate to dry in a warm place or on filter paper*
- *For the reactants, you can use:*
  - *Two soluble salts mixed together to form an insoluble salt and a soluble one*
  - *e.g. to make insoluble silver chloride, mix together silver nitrate and sodium chloride*

Explain what happens when silver nitrate and sodium chloride react

- *In silver nitrate solution the Ag^+ and NO_3^- are attracted weakly*
- *In the sodium chloride solution the Na^+ and Cl^- are attracted weakly*
- *When the two solutions are mixed, all the ions mix together*
  - *Ag^+ and Cl^- attract strongly forming solid AgCl precipitate*
  - *Na^+ and NO_3^- attract weakly and remain in solution*

Worked Example 23: Making An Insoluble Salt

The equation for the formation of lead(II) sulfate from lead(II) nitrate and sodium sulfate is:

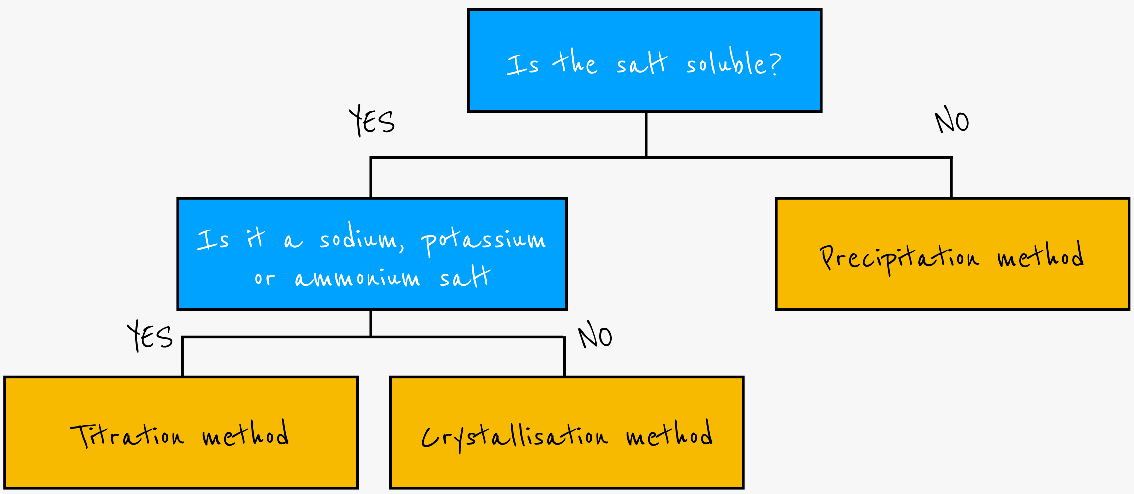
$$Pb(NO_3)_2\ (aq) + Na_2SO_4\ (aq) \rightarrow PbSO_4\ (s) + 2NaNO_3\ (aq)$$

Describe how a pure, dry sample of lead(II) sulfate could be obtained

Answer: The product is an insoluble salt ($PbSO_4$ is a solid). This therefore requires the precipitation method…

- *REACT: in a conical flask and stir with glass rod*
- *FILTER: using filter paper*
- *WASH*
- *DRY: allow the precipitate to dry in a warm place or on filter paper*

How To Approach Questions On Making Salts



Define hydrated substance
- A substance that is chemically combined with water

Define anhydrous substance
- A substance containing no water

Define water of crystallisation
- *The water molecules present in hydrated crystals*
  - *e.g.* $CuSO_4.5H_2O$
  - *e.g.* $CoCl_2.6H_2O$

Worked Example 24: Water of Crystallisation

11.25 g of hydrated copper sulfate, $CuSO_4.xH_2O$*, is heated until it loses all of its water. Its new mass is found to be 7.19 g. What is the value of x?*

Step 1: Calculate the mass of H_2O *in hydrated copper sulfate*

11.25 - 7.19 = 4.06 g

Step 2: Calculate the moles of $CuSO_4$ *and* H_2O

	$CuSO_4$	**$x.H_20$**
Mass	7.19	4.06
M_r	159.5	18
Moles	0.04508	0.2256

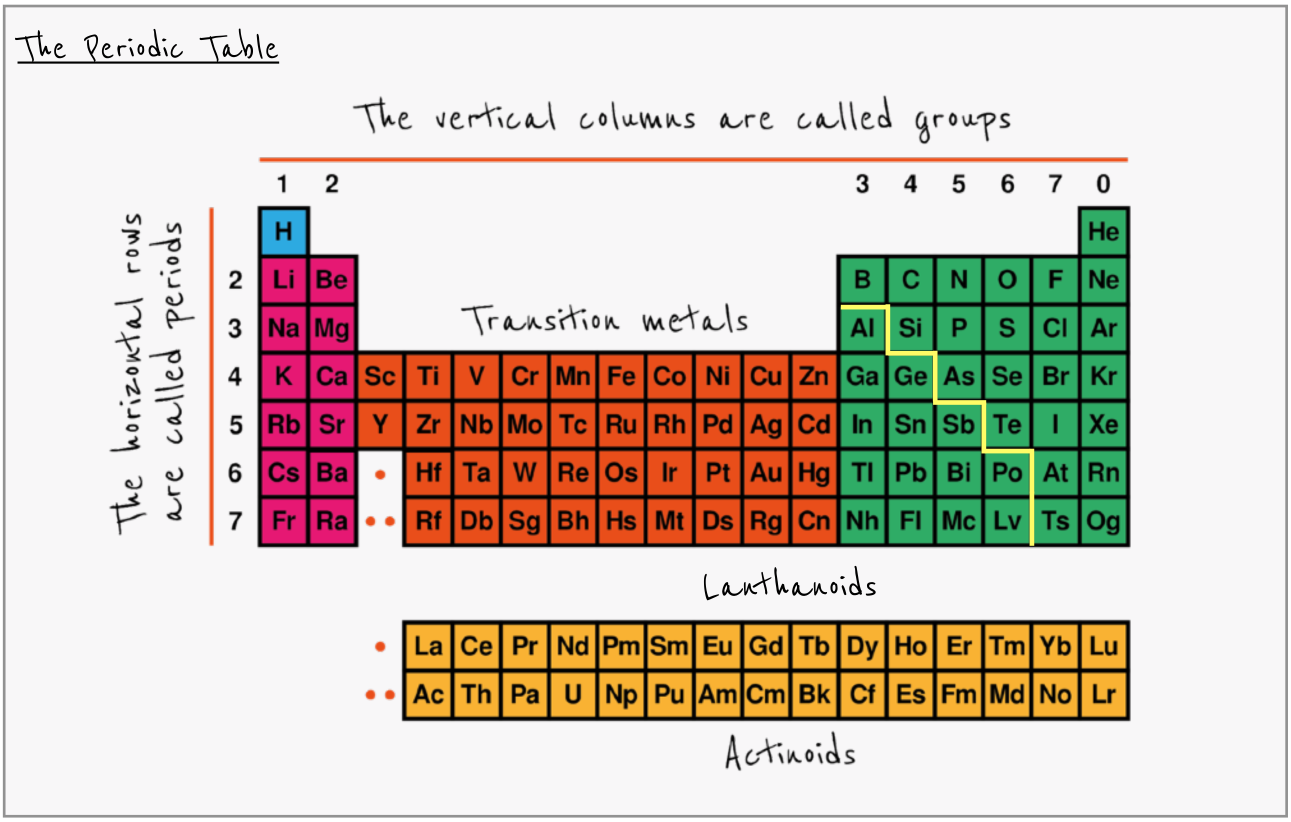
Step 3: Divide by the smallest value (0.04508 in this example)

	$CuSO_4$	**$x.H_20$**
Divide by smallest number	0.04508 / 0.04508 = 1	0.2256 / 0.04508 = 5
Ratio	1	5

Answer: $x = 5$ ($CuSO_4.5H_2O$)

8. THE PERIODIC TABLE

8.1 Arrangement of elements



Credit: Tom Morley for SwH Learning

Describe the layout of the periodic table
- Elements arranged in order of atomic number
- Hydrogen by itself
- Metals found to left of stepped line
- Non-metals found to right of stepped line
- Transition elements are in the middle

Define valency
- Number of electrons in outer shell

What is the group number of the periodic table?
- Tells you the number of electrons in the outer shell (valence electrons)
  - e.g. F has 7 electrons in its outer shell and is therefore in group 7

Why do elements in the same group have similar chemical properties?
- Same number of electrons in outer shell

When working out the charge on an ion, remember:
- For groups 1 to 3, the charge on the ion is the same as the group number
  - e.g. Mg is in group 2 and therefore forms Mg^{2+}
- For groups 5,6 and 7, the charge on the ion is (8 – group number)
  - e.g. N is in group 5. 8 – 5 = 3, therefore N^{3-}

What is the period number of the periodic table?
- Tells you the number of shells of electrons
  - e.g. Ca 2.8.8.2 has four shells of electrons and is therefore in period 4

8.2 Group I properties

What is the name given to Group 1 elements?
- Alkali metals

Describe the physical properties of Group 1 elements
- Soft (can be cut with a knife)
- Low melting and boiling points
- Low density (float on water)
- Shiny (tarnish when exposed to air)

Describe the chemical properties of Group 1 elements
- React vigorously with water to form metal hydroxides e.g. NaOH
- React with oxygen to form metal oxides e.g. Li_2O
  - e.g. $4Li + O_2 \rightarrow 2Li_2O$
- Form ionic compounds e.g. NaCl
- React with halogens e.g. KCl

Why do Group 1 elements have similar chemical properties?
- All have 1 electron in outer shell

How does the melting point of Group 1 elements change down the group?
- Melting point decreases

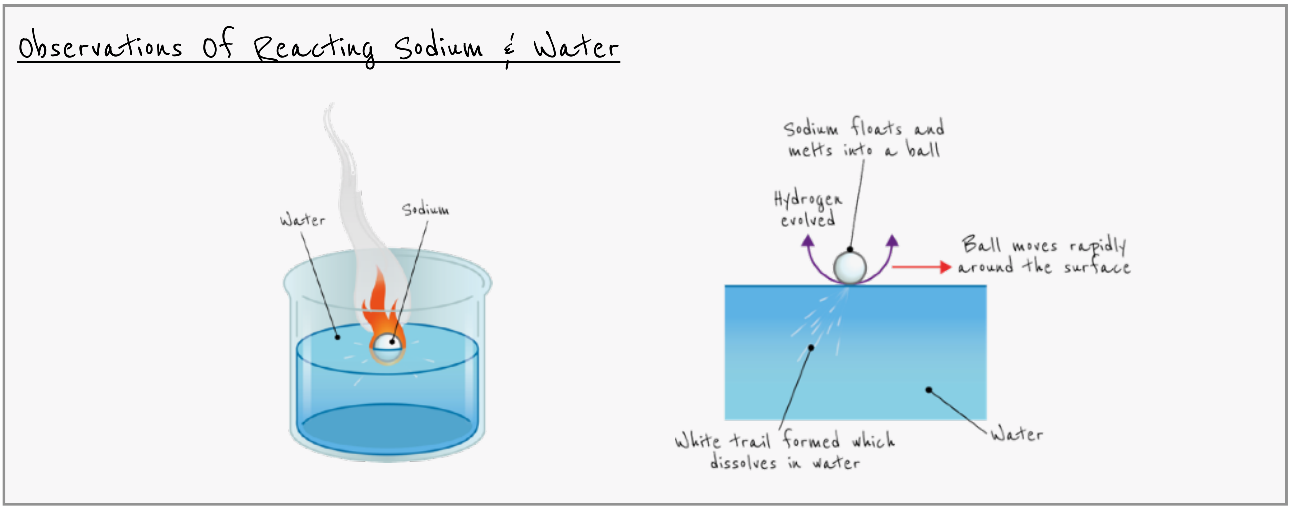
How does the density of Group 1 elements change down the group?
- Density increases

How does the reactivity of Group 1 elements change down the group?
- Reactivity increases

How should Group 1 elements be stored and why?
- In oil
- Very reactive with water

Describe the observations when Group 1 metals are added to water

	Observations when react with water
Lithium	Fizzes gently Moves across surface of water Disappears
Sodium	Melts and forms ball Fizzes Moves across surface of water Gets smaller and disappears Leaves a white trail
Potassium	Melts and forms ball Fizzes vigorously / explodes Sparks Burns with lilac flame Leaves a white trail



Credit: Tom Morley for SwH Learning

Compare the reactions of lithium and potassium with water

	Lithium	Potassium
Similarities	Float	
	Move around on surface of the water	
	Fizz (effervesce)	
	Disappears	
Differences	Fizzes gently	Explodes
	Does not burn	Lilac flame
	Does not melt	Melts into a ball

Why do Group 1 elements become more reactive down the group?
- Elements are larger
- Outer shell electron further from nucleus
  - Electron is more shielded
- Electron more easily lost

Predict the properties of francium
- Group 1 element
  - Soft, low melting and boiling point, low density, shiny (tarnishes when exposed to air)
- Near bottom of Group 1
  - Reactions with air and water will be more violent than other Group 1 metals

8.3 Group VII properties

Describe the trends in the physical properties of the halogens (Group VII elements)
- Colour gets darker down the group
- Boiling point and melting point increase down the group
- Reactivity decreases down the group
- All form diatomic molecules
  - e.g. F_2, Cl_2, I_2

Give the states and colours of Group VII elements at room temperature:

	Physical state at room temperature	Colour at room temperature
Fluorine	Gas	Yellow
Chlorine	Gas	Green
Bromine	Liquid	Red-brown
Iodine	Solid	Grey (forms purple vapour when heated)

Describe the reactions of halogens with hydrogen
- Hydrogen halides formed
  - e.g. H_2 (g) + Br_2 (g) → 2HBr (g)
    - Hydrogen halides are acidic and highly poisonous
    - Very soluble in water
      - e.g. hydrogen chloride gas dissolves to form hydrochloric acid
      - HCl (g) → HCl (aq)

What is a displacement reaction?
- When a more reactive halogen displaces a less reactive halide from its compound

Halogen displacement reactions:

	KCl	KBr	KI
Cl_2	X	$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$	$Cl_2 + 2KI \rightarrow 2KCl + I_2$
Br_2	No reaction	X	$Br_2 + 2KI \rightarrow 2KBr + I_2$
I_2	No reaction	No reaction	X

Give the ionic equation of the displacement reaction between Cl_2 and KBr
- Overall reaction: Cl_2 (aq) + 2KBr (aq) → 2KCl (aq) + Br_2 (aq)
- Separate out ionic compounds
  - Cl_2 (aq) + $2K^+$ (aq) + $2Br^-$ (aq) → $2K^+$ (aq) + $2Cl^-$ (aq) + Br_2 (aq)
- Remove spectator ions that appear on both sides (i.e. $2K^+$)
  - Ionic equation: Cl_2 (aq) + $2Br^-$ (aq) → Br_2 (aq) + $2Cl^-$ (aq)

Why are three of the reactions not carried out in the above table?
- A halogen cannot displace itself
- No reaction would occur

Why do three of the experiments not produce a reaction?
- A halogen cannot displace a more reactive halogen

Describe two experiments to show the order of reactivity of bromine, chlorine and iodine
- Add chlorine (Cl_2) to KBr
  - Solution turns orange
  - $Cl_2 + 2KBr \rightarrow Br_2 + 2KCl$
- Add bromine (Br_2) to KI
  - Solution turns brown
  - $Br_2 + 2KI \rightarrow I_2 + 2KBr$

Why do Group VII elements become less reactive down the group?
- Elements are larger
- Outer shell electron is further from the nucleus and more shielded
- Harder to gain an electron

Predict the state and colour and properties of astatine
- Below iodine in periodic table
- Solid
- Very dense
- Dark grey / black

8.4 Transition elements

What are the transition elements?
- Metals with high densities, high melting points
- Form coloured compounds
- Often act as catalysts
- *Have ions with variable oxidation numbers*
  - *e.g.* Fe^{2+} *= Fe(II) and* Fe^{3+} *= Fe(III)*
  - *e.g.* Cu^{+} *= Cu(I) and* Cu^{2+} *= Cu(II)*

8.5 Noble gases

What are the noble gases?
- Group VIII (also known as group 0) elements
- Unreactive
- Monoatomic
  - e.g. Ar, Xe, Ne

Why are noble gases unreactive?
- They have a full outer shell of electrons
- Stable

9. METALS

9.1 Properties of metals

Define ductile
- May be drawn into a wire

Define malleable
- May be hammered into shape

Describe the general physical properties of metals
- Good conductors of heat and electricity
- Ductile
- Malleable
- Shiny (lustrous)
- High melting and boiling points
- High density

Describe the general physical properties of non-metals
- Poor conductors of heat and electricity
- Brittle
- Dull (non-reflective)
- Low melting and boiling points
- Low density

Outline the general chemical reactivity of metals with dilute acids
- Most metals react with dilute acids
- Produces salt and hydrogen gas
  - e.g. $Zn\ (s) + 2HCl\ (aq) \rightarrow ZnCl_2\ (aq) + H_2\ (g)$

Outline the general chemical reactivity of metals with cold water
- Reactive metals react with cold water
- Produces metal hydroxide and hydrogen gas
  - e.g. $2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$

Outline the general chemical reactivity of metals with steam
- Less reactive metals react with steam
- Produces metal oxide and hydrogen gas
  - e.g. $Mg\ (s) + H_2O\ (g) \rightarrow MgO\ (s) + H_2\ (g)$

Outline the general chemical reactivity of metals with oxygen
- Reactive metals (e.g. Group I) react readily with oxygen
- Copper and iron react slowly with oxygen
- Produces metal oxide
  - e.g. $2Li\ (s) + O_2\ (g) \rightarrow Li_2O\ (s)$
  - e.g. $2Cu\ (s) + O_2\ (g) \rightarrow 2CuO\ (s)$
- Unreactive metals (e.g. gold, platinum) do not react with oxygen

9.2 Uses of metals

Give some uses of aluminium and link them with an appropriate property of aluminium
- Aeroplanes (low density)
- Electricity cables (good conductor of electricity, ductile, low density)
- Food containers (resistant to corrosion)

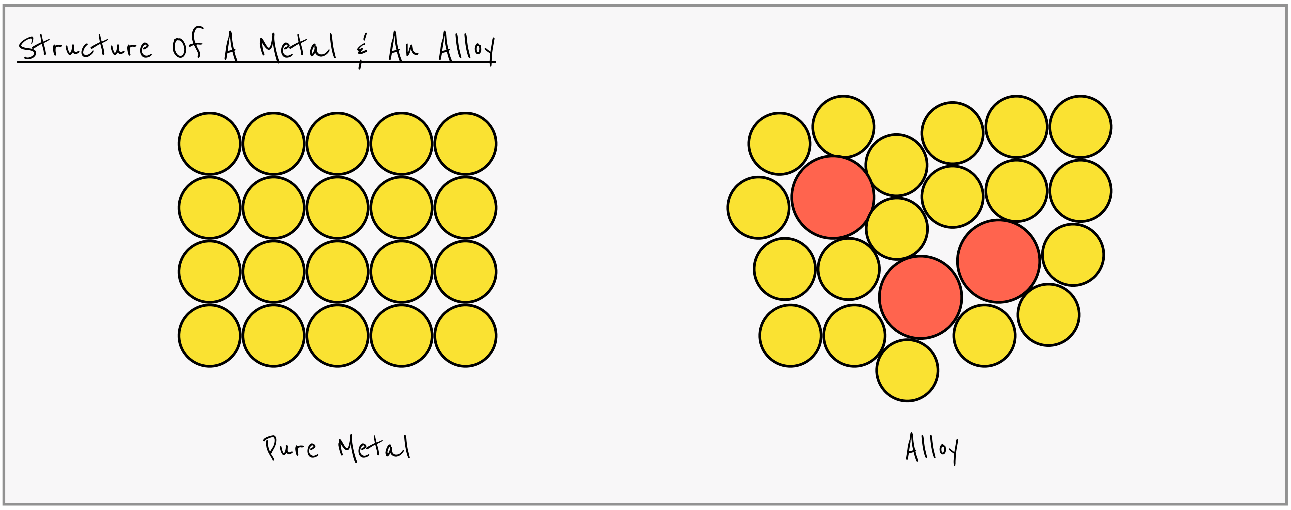
Describe the properties and uses of copper
- Malleable, ductile, good conductor of heat and electricity
- Uses: water pipes, wires

9.3 Alloys and their properties

What is an alloy?
- Mixture of a metal with other elements
  - e.g. brass: mixture of copper and zinc
  - e.g. stainless steel: mixture of iron, chromium, nickel and carbon

Why are alloys often more useful than pure metals?
- Alloys can be made harder and stronger than pure metals

Why do alloys tend to be harder than individual metals?
- *Contain atoms of different sizes*
- *More difficult for layers to slide over each other*



Credit: Martin Bailey for SwH Learning

Describe the properties and uses of stainless steel
- Alloy
- Contains iron, chromium and nickel
- Hard
- Oxide layer prevents corrosion/rusting
- Uses: sinks, saucepans, cutlery, gardening tools

9.4 Reactivity series

What is the reactivity series?
- List of metals in order of their reactivity
  - Most reactive metals at the top
  - Least reactive at the bottom

How can you determine the reactivity of a metal?
- Place in cold water (most will not react) - those which react are the most reactive metals
- If no reaction, test with steam
- If no reaction, test with acid (**note:** only metals above hydrogen in the reactivity series will react with acid)
  - Potassium, sodium, lithium and calcium are too reactive to react with acids (dangerous)

The Reactivity Series (Learn This Off By Heart)

Most reactive ↑

- Potassium
- Sodium
- Calcium
- Magnesium
- Aluminium
- *Carbon*
- Zinc
- Iron
- *Hydrogen*
- Copper
- Silver
- Gold

Least reactive

React with cold water (Potassium, Sodium, Calcium)

React vigorously with steam (Potassium to Magnesium)

React with dilute acid (Potassium to Iron)

(Carbon and hydrogen are included as reference points)

Describe the reactions of the metals in the reactivity series with cold water, steam and dilute acid

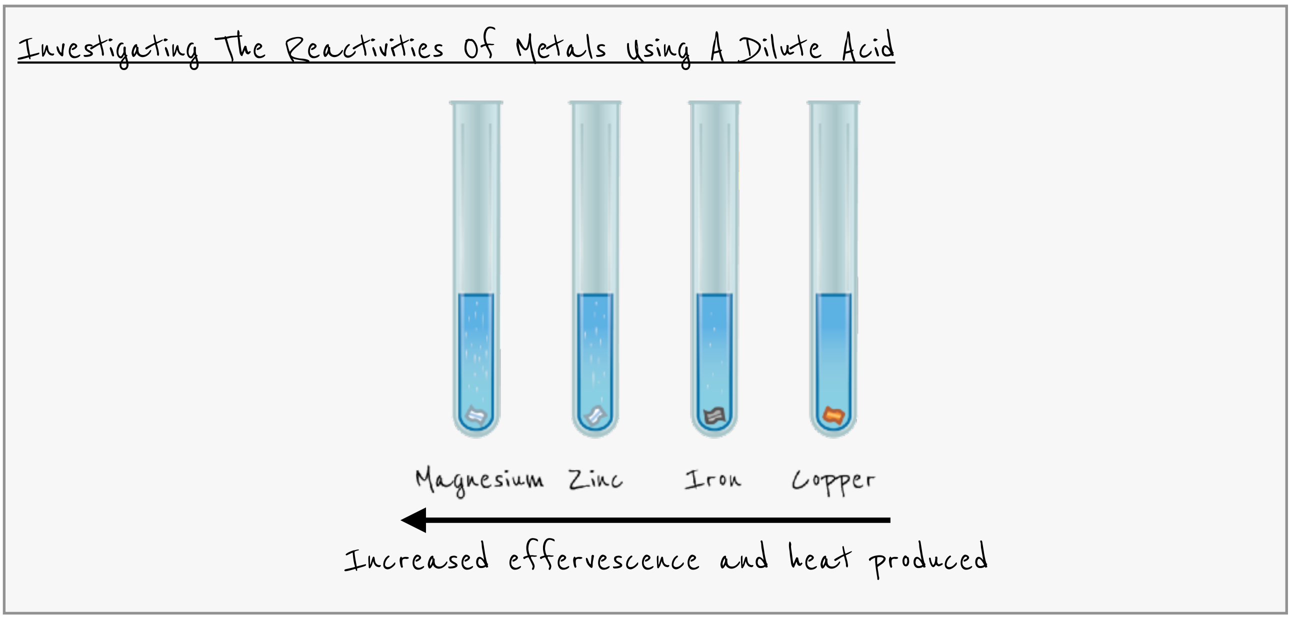
Metal	Reaction with cold water / steam	Reaction with HCl
Potassium	Very violent with cold water	Explosively fast
Sodium	Violent with cold water	Explosively fast
Calcium	Less violent with cold water	Very vigorous
Magnesium	Very slow with cold water, vigorous with steam	Vigorous
Zinc	Quite slow with steam	Quite slow
Iron	Slow with steam	Slow
Copper	No reaction	No reaction
Silver	No reaction	No reaction
Gold	No reaction	No reaction

Explain how displacement reactions can be used to arrange metals in order of reactivity

- *Occur when a less reactive element is pushed out of its compound by a more reactive element*
- *The more reactive the metal, the greater its tendency to lose electrons and form positive ions*
  - *e.g. with metals and metal oxides:*
    - *Magnesium + copper (II) oxide → magnesium oxide + copper*
    - *Mg + CuO → MgO + Cu*
      - *Displacement reaction occurred, therefore magnesium more reactive than copper*
  - *e.g. with metals and aqueous solutions of metal salts:*
    - *Zinc + copper (II) sulfate → zinc sulfate + copper*
    - *Zn +* $CuSO_4$ *→* $ZnSO_4$ *+ Cu*
      - *Displacement reaction occurred, therefore zinc more reactive than copper*

Describe how to investigate reactions between dilute acids and metals
- Pour the same volume of a dilute acid into four boiling tubes
- Place a small piece of different metals (e.g. magnesium, zinc, iron, copper) in each tube
- Observe the changes that occur:
  - A rapid fizzing and a colourless gas (hydrogen) produced
  - Reaction mixture becomes warm as heat is produced (exothermic)
  - More reactive metal = more fizzing and more heat

Explain why hydrogen gas is produced when a metal reacts with a dilute acid
- Metals above hydrogen in reactivity series will react with dilute acids
  - Produce a salt and hydrogen in a displacement reaction
  - e.g. magnesium + hydrochloric acid → magnesium chloride + hydrogen
  - $Mg (s) + 2HCl (aq) \rightarrow MgCl_2 (aq) + H_2 (g)$



Credit: Tom Morley for SwH Learning

Why is aluminium apparently unreactive?
- *Reacts rapidly with oxygen in the air*
- *Thin coat of aluminium oxide formed*
  - $4Al (s) + 3O_2 (g) \rightarrow 2Al_2O_3$
- *Aluminium oxide forms a barrier against further corrosion from water or dilute acids*

9.5 Corrosion of metals

What conditions are needed for rusting?
- Water
- Oxygen

Give the word and symbol equation for when iron rusts
- Iron + oxygen + water → hydrated iron (III) oxide
- $4Fe + 3O_2 + nH_2O \rightarrow 2Fe_2O_3.nH_2O$

What methods are used to prevent iron from rusting?
- Barrier methods
- Galvanising
- Sacrificial protection

How do barrier methods prevent iron from rusting?

- Provide protective layer
- Prevent water / oxygen reaching iron
  - e.g. paint (bridges), grease (car engines), oil (bike chains), coating with plastic
- If the layer is damaged the iron will rust

Describe the process of galvanising

- *Coating iron in a more reactive metal e.g. zinc*
  - *Zinc reacts with oxygen first, forming Zn^{2+}*
  - *Zn loses electrons instead of iron*
    - *Used for car bodies, buckets*

What is sacrificial protection?

- *Method used to stop iron from rusting*
- *Iron coated in a more reactive metal / blocks of more reactive metal attached to iron*
  - *More reactive metal undergoes oxidation in preference to iron*
  - *Reactive metal loses electrons to form positive ions*

9.6 Extraction of metals

How are different metals extracted from their ores?

- Most reactive metals extracted by electrolysis
- Least reactive metals occur naturally as pure element
- Metals less reactive than carbon can be extracted by heating with carbon

How are unreactive metals obtained?

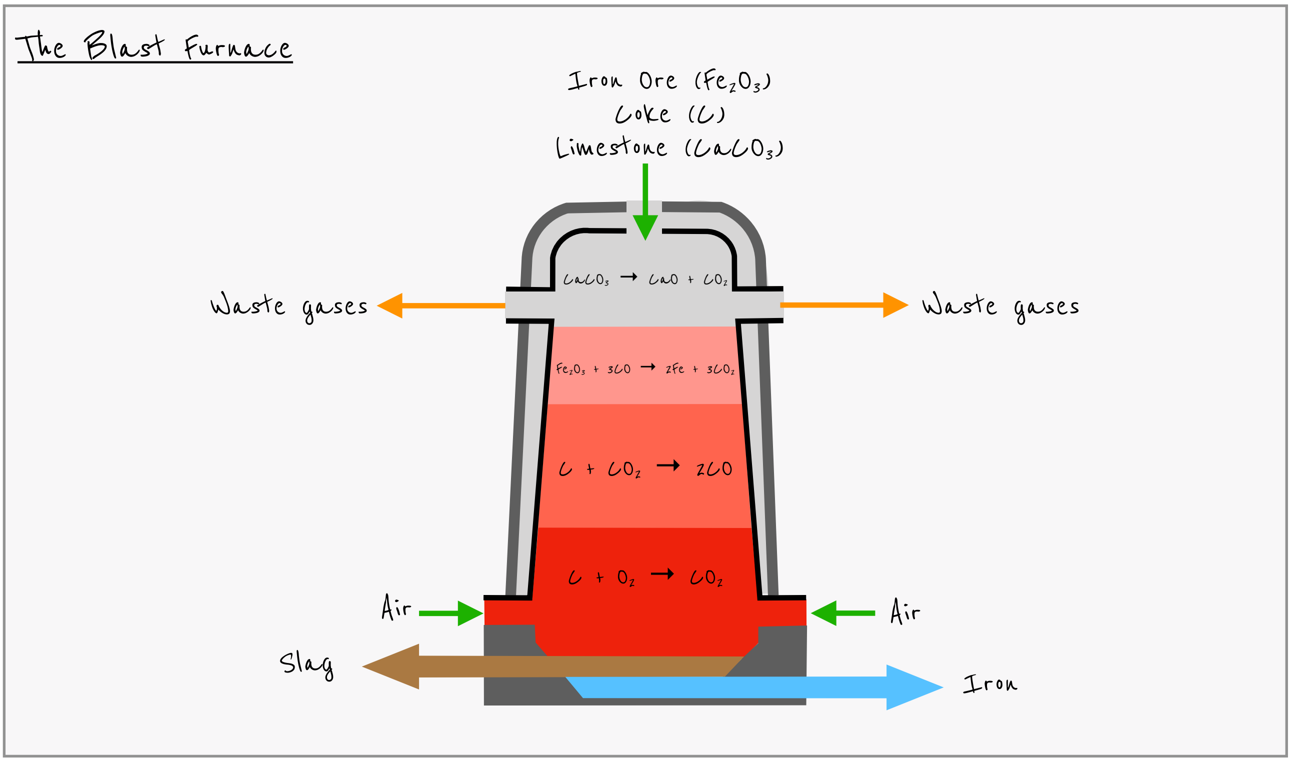
- Found as the uncombined element (i.e. exist naturally in Earth's crust)
- Easy to extract in uncombined state

Why is iron obtained by reduction with coke (carbon) in the blast furnace?

- Iron is less reactive than carbon (below carbon in the reactivity series)

Describe how iron is extracted from hematite in the blast furnace

- Carbon (coke) burnt in oxygen
  - Provides heat, produces carbon dioxide
    - carbon + oxygen → carbon dioxide
    - $C + O_2 \rightarrow CO_2$
- Carbon dioxide reduced to carbon monoxide
  - Requires high temperature
    - carbon + carbon dioxide → carbon monoxide
    - $C + CO_2 \rightarrow 2CO$
- Carbon monoxide reduces iron (III) oxide in iron ore
  - Forms iron
    - iron (III) oxide + carbon monoxide → iron + carbon dioxide
    - $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
- Calcium carbonate (limestone) added to remove impurities
  - Calcium carbonate undergoes thermal decomposition
    - calcium carbonate → calcium oxide + carbon dioxide
    - $CaCO_3 \rightarrow CaO + CO_2$
- Calcium oxide reacts with silicon dioxide to form calcium silicate
  - This melts and collects as slag
    - calcium oxide + silicon dioxide → calcium silicate
    - $CaO + SiO_2 \rightarrow CaSiO_3$
      - *This is a neutralisation reaction*



Credit: Martin Bailey for SwH Learning

Why is electrolysis not used to extract iron?
- *Electrolysis is very expensive*
- *High electricity demands*

How is aluminium extracted?
- *From bauxite ore*
  - *Aluminium oxide is the main constituent of bauxite*
- *Using electrolysis*

Why can't aluminium be extracted using the blast furnace?
- *Aluminium is more reactive than carbon (above carbon in the reactivity series)*
- *Electrolysis used instead*

Why is aluminium oxide dissolved in molten cryolite before extraction?
- *Molten cryolite decreases melting point of aluminium oxide*
- *Less energy required to melt aluminium oxide for electrolysis → less expensive to carry out*

Outline the extraction of aluminium from aluminium ore (bauxite) by electrolysis
- *Purified aluminium oxide dissolved in molten cryolite*
- *Mixture placed in electrolysis cell with graphite (carbon) electrodes*
- *At the negative cathode: $Al^{3+} + 3e^- \rightarrow Al$ (reduction)*
- *At the positive anode: $2O^{2-} - 4e^- \rightarrow O_2$ (oxidation)*
- *Overall equation: $4Al_2O_3 \rightarrow 4Al + 3O_2$*

Why do the carbon anodes need to be regularly replaced?
- *Oxygen produced reacts with the carbon to produce carbon dioxide*

Why is the manufacture of aluminium such an expensive process?
- *High cost of electricity*
- *Regular replacement of carbon anodes required*

10. CHEMISTRY OF THE ENVIRONMENT

10.1 Water

What are the chemical tests for water?
- Turns white anhydrous copper (II) sulphate blue
- Turns blue cobalt (II) chloride paper pink
  - Both changes can be reversed by heating

How do you do a physical test for water?
- Check boiling or melting point
- Water boils at 100°C and melts at 0°C

How do you show that water is pure?
- Check boiling or melting point
- Pure water has single boiling point at 100°C and single melting point at 0°C

Why is distilled water used in chemistry rather than tap water?
- Distilled water contains fewer chemical impurities

Which substances may be found in water from natural sources?

Substance	Beneficial or harmful?	Explanation
Dissolved oxygen	Beneficial	Aquatic life requires oxygen for respiration
Metal compounds	Both	**Beneficial:** Provide essential minerals for life (e.g. iron for haemoglobin, magnesium for chlorophyll) **Harmful**: Some metal compounds are toxic
Plastics	Harmful	Harm aquatic life
Sewage	Harmful	Contains harmful microbes which cause disease
Pathogens/microbes	Harmful	Cause disease
Nitrates (from fertilisers)	Harmful	Cause algal bloom, leading to deoxygenation of water and killing aquatic life
Phosphates (from fertilisers and detergents)	Harmful	Cause algal bloom, leading to deoxygenation of water and killing aquatic life

What is the difference between potable water and pure water?
- Potable: water that is safe for human consumption
- Pure: contains H_2O molecules only

How is domestic water treated?
- Solids removed by sedimentation and filtration
- Water filtered through carbon to remove tastes and odours
- Chlorine added to kill microbes

Describe the process of sedimentation
- Water stored in a reservoir
- Allows large, insoluble particles to sink to bottom
- Clean water at top is extracted

Describe the process of filtration
- Water sprayed over layers of sand and gravel
- Small, insoluble particles removed as water passes through

10.2 Fertilisers

Which chemicals are commonly used as fertilisers?
- Ammonium salts
- Nitrates

Which elements do NPK fertilisers contain?
- Nitrogen (N)
- Phosphorous (P)
- Potassium (K)

Why is nitrogen needed in fertilisers?
- Plants require it to produce chlorophyll and proteins

Why is phosphorous needed in fertilisers?
- Root growth
- Ripening of crops

Why is potassium needed in fertilisers?
- Produce proteins
- Disease resistance

10.3 Air quality and climate

What is the approximate composition of clean, dry air?
- 78% nitrogen
- 21% oxygen
- Remainder is mixture of noble gases and carbon dioxide

Outline the sources and effects of common air pollutants

Pollutant	Source	Effect
Carbon dioxide	Complete combustion of carbon-containing fuels	Increased global warming leading to climate change
Carbon monoxide	Incomplete combustion of carbon-containing fuels	Toxic – death from oxygen depletion
Methane	Decomposition of vegetation, waste gases from digestion in animals	Increased global warming leading to climate change
Oxides of nitrogen	Reaction of nitrogen with oxygen in car engines	Acid rain, photochemical smog and respiratory problems
Sulfur dioxide	Combustion of fossil fuels containing sulfur compounds	Acid rain

What are the adverse effects of increasing volumes of particulates in air?
- Increase risk of respiratory problems
- Cancer

Why is acid rain a concern?
- Destruction of limestone buildings
- Erodes stone
- Damages forests
- Makes lakes and streams too acidic

What are greenhouse gases?
- Carbon dioxide and methane
- Can contribute to global warming

How do greenhouse gases cause global warming?
- *The Sun emits radiation that enters Earth's atmosphere*
  - *The Earth's surface reflects some of this energy back into space*
- *Most energy is absorbed by the Earth and re-emitted back at a longer wavelength (thermal energy)*
  - *Some of this thermal energy is emitted into space*
- *Greenhouse gases also absorb this thermal energy and re-emit it in all directions*
- *Thermal energy becomes trapped within the Earth's atmosphere*
- *Earth's average temperature increases (global warming)*

Explain the effect on the environment of increasing atmospheric levels of carbon dioxide
- *Greenhouse gas → contributes to global warming*
  - *Polar ice caps melt*
  - *Sea levels rise*
  - *Floods low lying land*
  - *Loss of habitats and biodiversity*

What strategies can be used to reduce the effects of climate change?

Action	Explanantion
Planting trees	Increases carbon dioxide uptake from the atmosphere by photosynthesis
Reduction in livestock	Decreases volume of methane emitted into the atmosphere through waste gases from digestion
Decreasing use of fossil fuels	Decreases volume of carbon dioxide released into atmosphere through complete combustion of hydrocarbons
Increasing use of hydrogen and renewable energy (e.g. wind, solar)	Hydrogen and renewable energy sources do not produce carbon dioxide, decreasing the levels of carbon dioxide in the atmosphere

What is photosynthesis?
- Process plants use to produce carbohydrates
- Uses energy from light

What is the word equation for photosynthesis?
- Carbon dioxide + water → glucose + oxygen
  - Requires light and chlorophyll

What is the symbol equation for photosynthesis?

- *$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$*

How are oxides of nitrogen formed?

- *High temperatures in car engines*
- *Cause nitrogen to react with oxygen*
  - *Forms nitrogen oxides*
    - *e.g. nitrogen oxide (NO), nitrogen dioxide (NO_2)*
- *Nitrogen oxides cause acid rain*

What strategies can be used to reduce the effects of acid rain?

Action	Explanantion
Use catalytic converters in vehicles	Reduces emissions of nitrogen oxides (NO and NO_2) into the atmosphere
Using low-sulfur fuels	Decreases volume of sulfur dioxide emitted into the atmosphere through combustion of fuels
Use flue gas desulfurisation with calcium oxide	Calcium oxide reacts with sulfur dioxide to produce solid calcium sulfite, which can be removed before it enters the atmosphere

How do catalytic converters work?

- *Remove oxides of nitrogen in car engines*
- *By causing them to react with other gases in the engine*
- *Producing less harmful gases*
  - *e.g. $2CO + 2NO \rightarrow 2CO_2 + N_2$*
  - *e.g. $C_8H_{18} + 25NO \rightarrow 12.5N_2 + 8CO_2 + 9H_2O$*

What are catalytic converters made from?

- *Platinum and rhodium metals*
- *Honeycomb shaped to provide large surface area for reaction to take place*

11. ORGANIC CHEMISTRY

11.1 Formulae, functional groups and terminology

What does the molecular formula show?
- Exact number of atoms of each element present in a compound

What does the empirical formula show?
- Simplest ratio of atoms of each element present in a compound

What does the general formula show?
- Relationship between the number of atoms of each element within a molecule

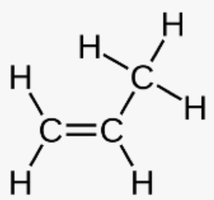
What does the structural formula show?
- *How the atoms in a molecule are arranged*

What does the displayed formula show?
- A drawing of the bonds within a molecule

Types of Chemical Formula

Example: Propene

- **Molecular formula:** C_3H_6
- **Empirical formula:** CH_2
- **General formula:** C_nH_{2n} (alkenes)
- **Structural formula:** $CH_2 = CH - CH_3$
- **Displayed formula:**

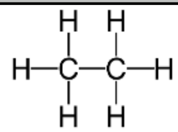
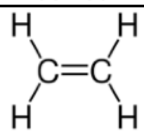
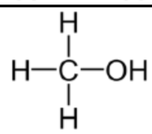
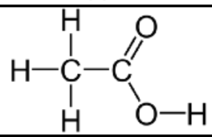


What is a functional group?
- An atom or group of atoms which determine the chemical properties of a homologous series

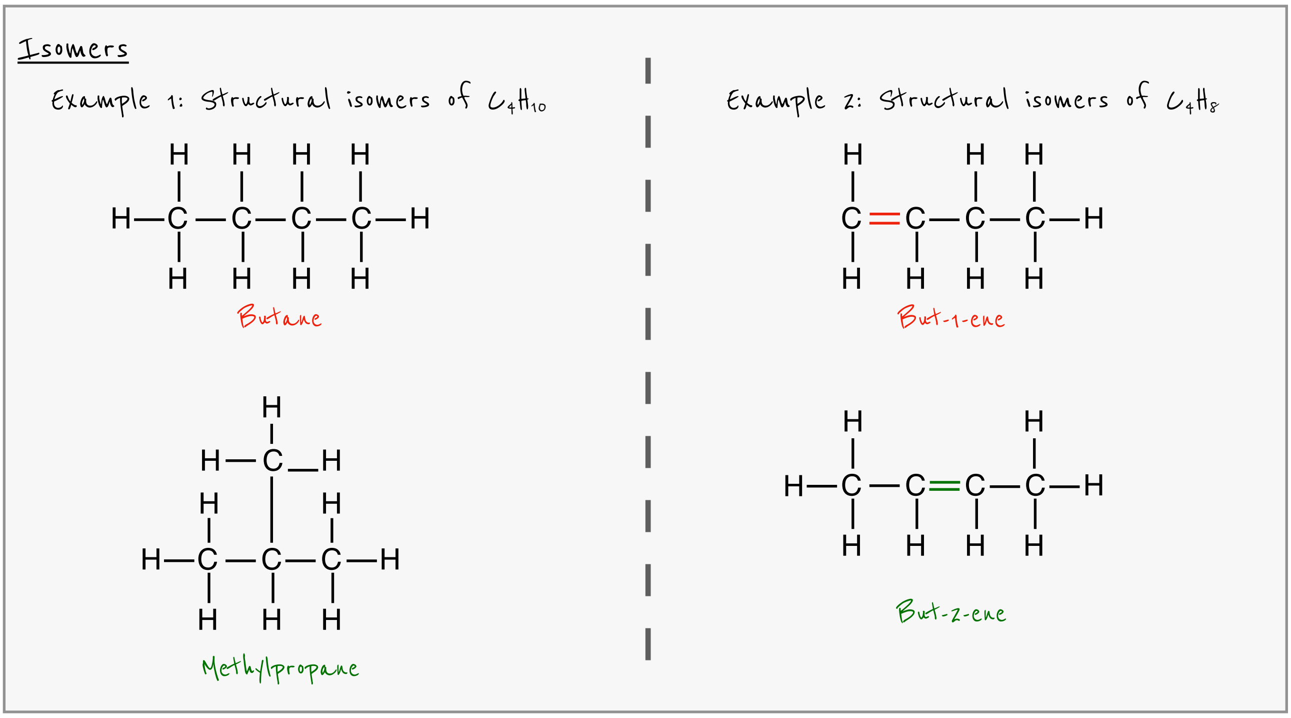
What is a homologous series?
- Group of compounds with same chemical properties because they have same functional group
  - e.g. alcohols (–OH), alkenes (C=C)

What do members of the same homologous series have in common?
- *Similar chemical properties*
- *Trend in physical properties*
- *Same functional group*
- *Same general formula*
- *Differ from one member to the next by a – CH_2 – unit*

Summary of common homologous series:

Homologous Series	General Formula	Example Displayed Formula
Alkanes	C_nH_{2n+2}	
Alkenes	C_nH_{2n}	
Alcohols	$C_nH_{2n+1}OH$	
Carboxylic acids	$C_nH_{2n+1}COOH$	

What does structural isomerism mean?
- *Compounds with same molecular formula but different displayed/structural formula*



Credit: Martin Bailey for SwH Learning

What are ‘saturated’ compounds?
- Molecules in which all carbon-carbon bonds are single bonds
- No C=C double bonds

What are ‘unsaturated’ compounds?
- Molecules that contain one or more C=C double bonds

11.2 Naming organic compounds

Name beginning	Number of carbons
meth-	1
eth-	2
prop-	3
but-	4

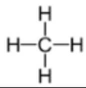
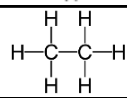
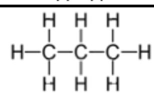
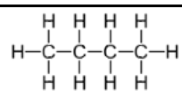
Remember:

A mnemonic for the first 4 carbon chain names:

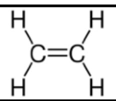
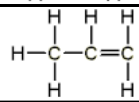
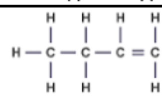
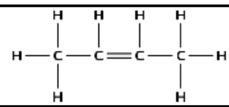
M**onkeys = **meth-
E**at = **eth-
P**eanut = **prop-
B**utter = **but-

Name ending	Type of molecule
-ane	Alkane
-ene	Alkene
-ol	Alcohol
-oic acid	Carboxylic acid

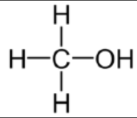
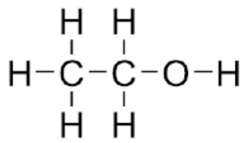
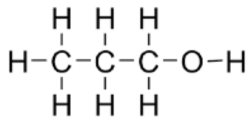
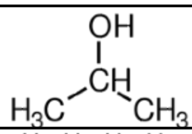
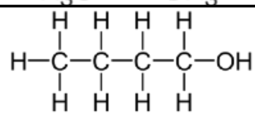
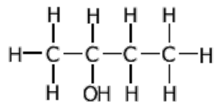
Summary of the first 4 straight-chained alkanes

Number of Carbon Atoms	Name	Structural Formula	Displayed Formula
1	Methane	CH_4	
2	Ethane	CH_3CH_3	
3	*Propane*	$CH_3CH_2CH_3$	
4	*Butane*	$CH_3CH_2CH_2CH_3$	

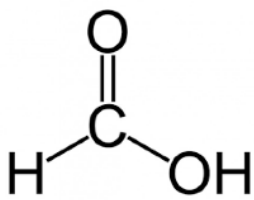
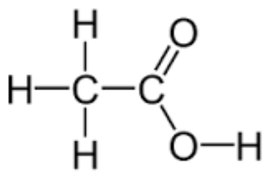
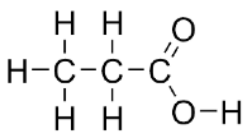
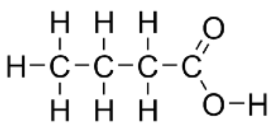
Summary of the first 4 straight-chained alkenes (and isomers)

Number of Carbon Atoms	Name	Structural Formula	Displayed Formula
1	n/a		
2	Ethene	$CH_2{=}CH_2$	
3	*Propene*	$CH_3CH{=}CH_2$	
4	*But-1-ene*	$CH_3CH_2CH{=}CH_2$	
4	*But-2-ene*	$CH_3CH{=}CHCH_3$	

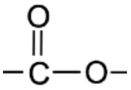
Summary of the first 4 straight-chained alcohols (and isomers)

Number of Carbon Atoms	Name	Structural Formula	Displayed Formula
1	Methanol	CH_3OH	
2	Ethanol	CH_3CH_2OH	
3	*Propan-1-ol*	$CH_3CH_2CH_2OH$	
3	*Propan-2-ol*	$CH_3CH(OH)CH_3$	
4	*Butan-1-ol*	$CH_3CH_2CH_2CH_2OH$	
4	*Butan-2-ol*	$CH_3CH(OH)CH_2CH_3$	

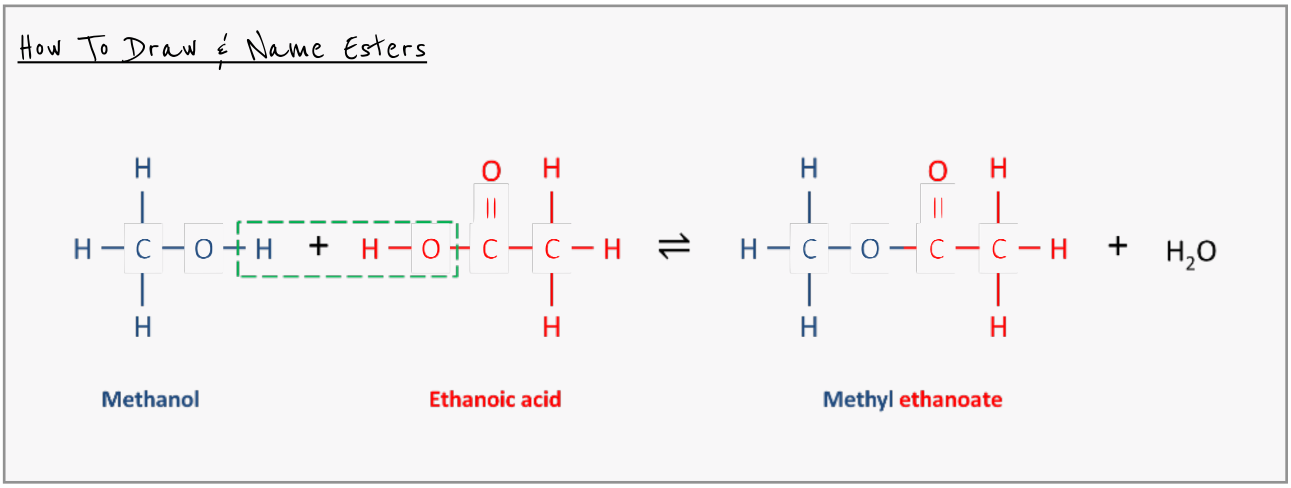
Summary of the first 4 straight-chained carboxylic acids

Number of Carbon Atoms	Name	Structural Formula	Displayed Formula
1	*Methanoic acid*	$HCOOH$	
2	Ethanoic acid	CH_3COOH	
3	*Propanoic acid*	CH_3CH_2COOH	
4	*Butanoic acid*	$CH_3CH_2CH_2COOH$	

What is the functional group of the esters?
- *-COO-*



What is the general equation for the formation of an ester (esterification)?
- *Carboxylic acid + alcohol ⇌ ester + water*



11.3 Fuels

What is a fuel?
- Substance which releases energy when burnt

Name the fossil fuels
- Coal
- Natural gas
- Petroleum

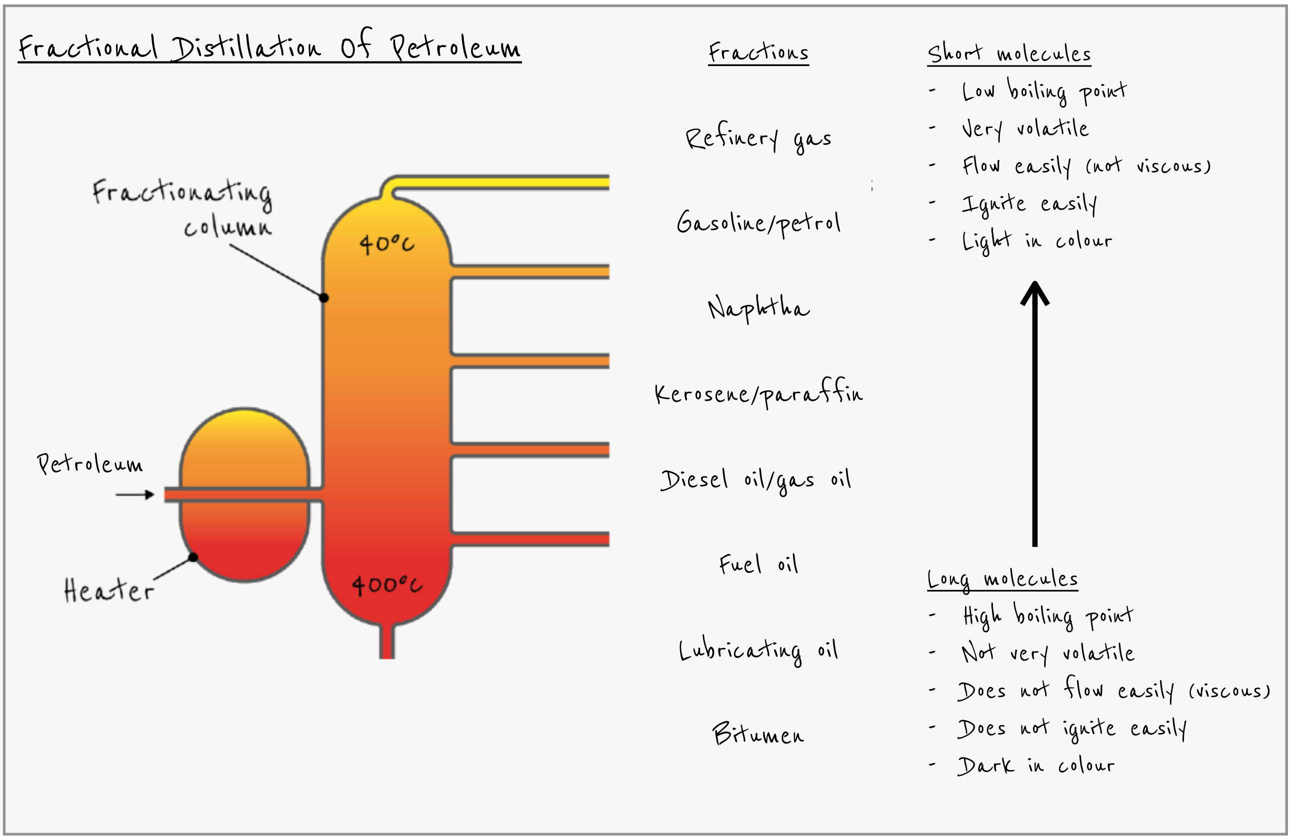
Which gas is the main constituent of natural gas?
- Methane

What is a hydrocarbon?
- A compound which contains hydrogen and carbon atoms ONLY

What is petroleum?
- Mixture of hydrocarbons

What is a fraction?
- Group of substances with similar boiling points

How is petroleum separated into its various fractions?
- Fractional distillation
  - Petroleum is heated
  - Petroleum boils and vaporises
  - Vapour passed into bottom of fractionating column
  - Column hottest at the bottom - longest chain fractions condense here e.g. bitumen
  - Column coolest at the top - shortest chain fractions condense here e.g. refinery gases



Credit: Tom Morley for SwH Learning

What are the main uses of the following fractions:

- Refinery gases (mixture of methane, ethane, propane) - bottled gas for heating and cooking
- Gasoline/petrol - fuel for cars
- Naphtha - chemical feedstock
- Kerosene/paraffin - fuel for planes (jet fuel)
- Diesel/gas oil - fuel for diesel engines (e.g. for buses, lorries)
- Fuel oil - fuel for ships and home heating systems
- Lubricating oil - lubricants, waxes and polishes
- Bitumen - road surfacing

What does viscosity mean?

- How readily a liquid flows
  - e.g. honey - very viscous
  - e.g. water - not very viscous

What does volatile mean?

- Evaporates readily

Compare the colour, viscosity and boiling point of bitumen and refinery gases

	Bitumen	**Refinery gases**
Colour	Dark	Light
Viscosity	Very viscous	Not viscous
Boiling point	High	Low
Chain length	Long	Short

11.4 Alkanes

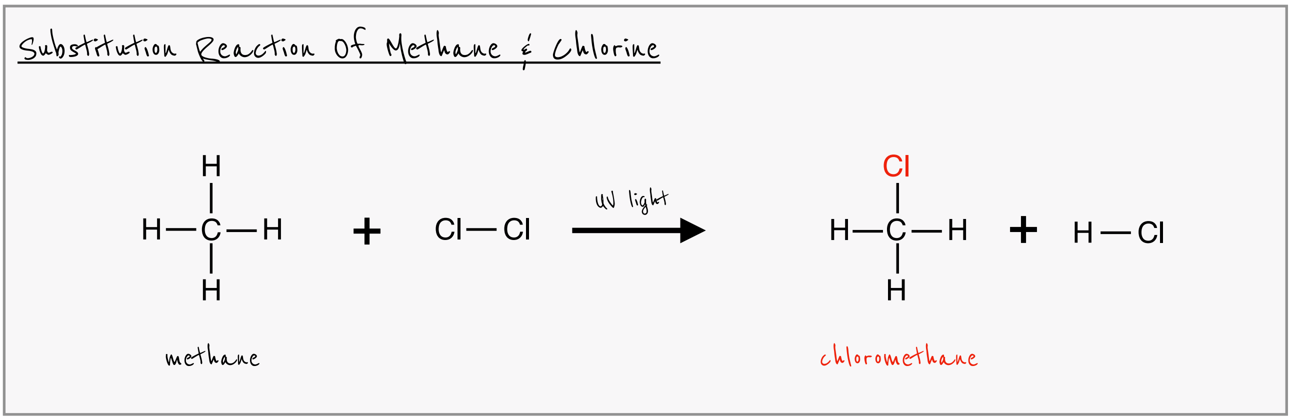
Why are alkanes classified as 'saturated' hydrocarbons?
- All carbons form 4 single, covalent bonds
- No C=C double bonds

Describe the properties of alkanes
- Generally unreactive
  - Except for combustion and substitution by chlorine

Define substitution reaction
- *Reaction in which one atom (or group of atoms) is replaced by another atom (or group of atoms)*

Describe the reaction of alkanes with chlorine
- *Substitution reaction*
- *Photochemical reaction*
  - *UV radiation required to provide activation energy*

Give the equation for the reaction of methane with chlorine
- *Methane + chlorine → chloromethane + hydrogen chloride*
- $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$



Credit: Martin Bailey for SwH Learning

11.5 Alkenes

What is the functional group of the alkenes?
- C=C

What does the '1' in but-1-ene mean?
- Double bond attached to the first carbon in the chain

Why are alkenes classified as 'unsaturated' hydrocarbons?
- Contain C=C double bonds

What is cracking?
- The breaking down of long alkane chains
- Into smaller chain alkanes and alkenes

What reaction conditions are needed for cracking?
- 600-700°C
- Alumina or silica catalyst

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Worked Example 25: Cracking

The chemical equation for a cracking reaction is $C_{16}H_{34} \rightarrow C_8H_{18} + 2C_3H_6$ + Compound X. Deduce the molecular formula of X.

Step 1: Count how many carbons and hydrogens are on each side of the equation

Left hand side:	C = 16 H = 34
Right hand side:	C = 8 + (2 x 3) = 14 H = 18 + (2 x 6) = 30

Step 2: Calculate how many carbons and hydrogens need to be added to the right hand side to make the equation balance

C = 16 – 14 = 2
H = 34 – 30 = 4

Answer: Compound X is C_2H_4 (ethene)

Explain why cracking is important
- Produces shorter chain molecules
- Shorter molecules more useful as fuels
  - Used to make petrol/diesel for vehicles
- Petroleum richer in long chain molecules
- Alkenes also produced by cracking
  - Used to make alcohols / polymers / plastics

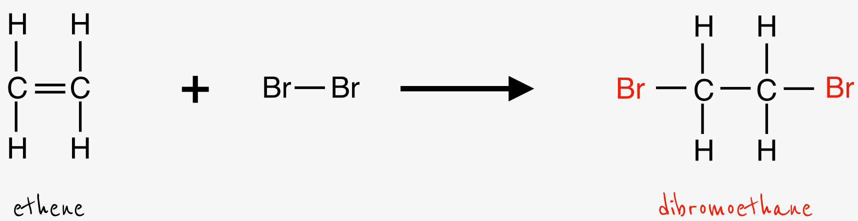
What is the test for an unsaturated compound (an alkene)?
- Add compound to bromine water
- Orange colour turns colourless
- This is an addition reaction
  - Dibromoalkane formed as product

Describe how to use bromine water to distinguish between an alkane and an alkene
- Add bromine water to each compound
- If bromine water turns colourless → compound is an alkene

Why don't alkanes turn bromine water colourless?
- Alkanes have no double bond (they're saturated)

Give the equation for the reaction of ethene with bromine water
- *Ethene + bromine → dibromoethane*
- $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$

Addition Reaction of Ethene & Bromine



Credit: Martin Bailey for SwH Learning

Why is the reaction of ethene with bromine water an addition reaction?
- *Br_2 added to the ethene molecule*
- *Only one product formed*

Describe the reactivity of alkenes
- *More reactive than alkanes due to double bond*
- *Undergo addition reactions*
- *Double bond breaks and atoms can be added*
- *Addition reaction turns unsaturated alkene in to saturated compound*

Give the equation for the addition reaction of ethene with hydrogen
- *Ethene + hydrogen → ethane*
- *$C_2H_4 + H_2 \rightarrow C_2H_6$*
- *Requires nickel catalyst*

Give the equation for the addition reaction of ethene with steam
- *Ethene + water → ethanol*
- *$C_2H_4 + H_2O \rightarrow C_2H_5OH$*
- *Requires acid catalyst*

11.6 Alcohols

What is the functional group of the alcohols?
- OH

Outline the process of manufacturing ethanol by the fermentation of aqueous glucose
- Yeast added to glucose solution
- Left in anaerobic conditions at 30°C
  - Optimum temperature for enzymes
- Yeast respires anaerobically
- Enzymes convert glucose to ethanol and carbon dioxide
  - $C_6H_{12}O_6\ (aq) \rightarrow 2C_2H_5OH\ (aq) + 2CO_2\ (g)$

Outline the process of manufacturing ethanol by the hydration of ethene
- React ethene with steam
- Use phosphoric acid catalyst
- Conditions: 300°C, 60 atmospheres
  - $C_2H_4\ (g) + H_2O\ (g) \rightarrow C_2H_5OH\ (g)$

Compare the processes of manufacturing ethanol

Comparison	Hydration of ethene	Fermentation of sugar
Rate of reaction	Fast	Slow
Type of process	Continuous	Batch
Renewable?	Non-renewable	Renewable
Temperature/Pressure	High temperature (300°C) High pressure	Low temperature (30°C) Low pressure
Purity of Alcohol	Pure	Impure
Equation	$CH_2CH_2 + H_2O \rightarrow CH_3CH_2OH$	$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ (this is anaerobic respiration of yeast)

Why is the product of fermentation not pure ethanol?
- *Yeast killed when ethanol content > 15%*

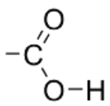
Why is 30°C the optimum temperature for fermentation?
- *Too high → enzymes denature*
- *Too low → kinetic energy too low → rate too slow*

Give the equation for the complete combustion of ethanol
- Ethanol + oxygen → carbon dioxide + water
  - C_2H_5OH (l) + $3O_2$ (g) → $2CO_2$ (g) + $3H_2O$ (l)

Give some uses of ethanol
- Solvent
- Fuel
- Alcoholic drinks

11.7 Carboxylic acids

What is the functional group of the carboxylic acids?
- COOH



Describe the reaction of carboxylic acids with metals
- Metal + acid → salt + hydrogen
  - e.g. magnesium + ethanoic acid → magnesium ethanoate + hydrogen
  - Mg (s) + $2CH_3COOH$ (aq) → $(CH_3COO)_2Mg$ (aq) + H_2 (g)
- Rate of reaction is slower than using HCl

Describe the reaction of carboxylic acids with bases
- Base + acid → salt + water
  - e.g. sodium hydroxide + ethanoic acid → sodium ethanoate + water
  - NaOH (aq) + CH_3COOH (aq) → CH_3COONa (aq) + H_2O (l)
- Rate of reaction is slower than using HCl
  - Because ethanoic acid is a weak acid

Describe the reaction of carboxylic acids with metal carbonates
- Metal carbonate + acid → salt + carbon dioxide + water
  - e.g. sodium carbonate + ethanoic acid → sodium ethanoate + carbon dioxide + water
  - Na_2CO_3 (s) + $2CH_3COOH$ (aq) → $2CH_3COONa$ (aq) + CO_2 (g) + H_2O (l)
- Rate of reaction is slower than using HCl
- Reaction fizzes as CO_2 produced

What acid is found in vinegar?
- Ethanoic acid

How is ethanoic acid produced?
- *Oxidation of ethanol*
- *Method 1: Heating ethanol with acidified aqueous potassium manganate(VII)*
  - *Solution heated under reflux*
  - *Solution turns from purple → colourless*
- *Method 2: Bacterial oxidation of ethanol*
  - *e.g. ethanol oxidises to form ethanoic acid (vinegar)*

What reaction conditions are needed in the formation of esters?
- *Strong acid catalyst (sulfuric acid)*

What is the general equation for the formation of an ester (esterification)?
- *Carboxylic acid + alcohol $\rightleftharpoons$ ester + water*
  - *e.g. ethanoic acid + methanol $\rightleftharpoons$ methyl ethanoate + water*

Why are esters used in perfumes and food flavourings?
- *Highly volatile*
- *Distinctive smells*

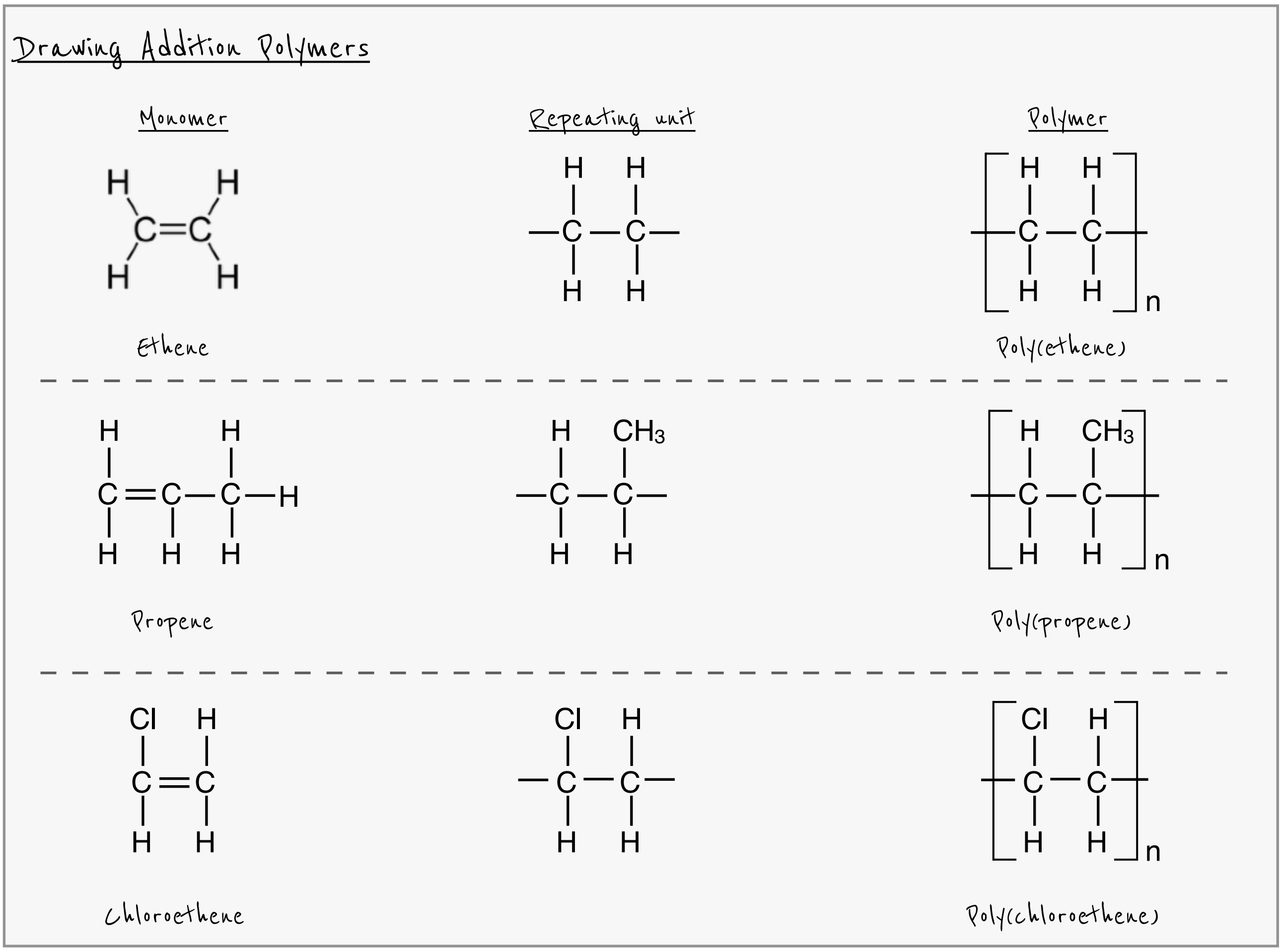
11.8 Polymers

Define monomer
- A small molecule that joins together to form a polymer

Define polymer
- A long chain formed from many small molecules joined together

What is an addition polymer?
- A polymer made from the joining up of many monomers

How are addition polymers formed?
- One bond in the double bond breaks
- Monomers join together
- Form a long chain containing only single bonds

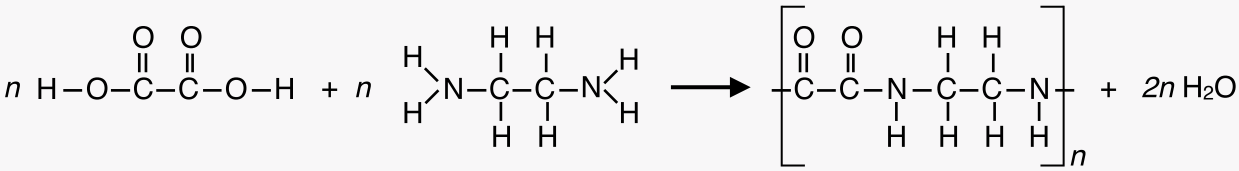


Give a use for common addition polymers
- Poly(ethene) - plastic bags
- Poly(propene) - plastic water pipes
- Poly(chloroethene) - window frames (PVC)

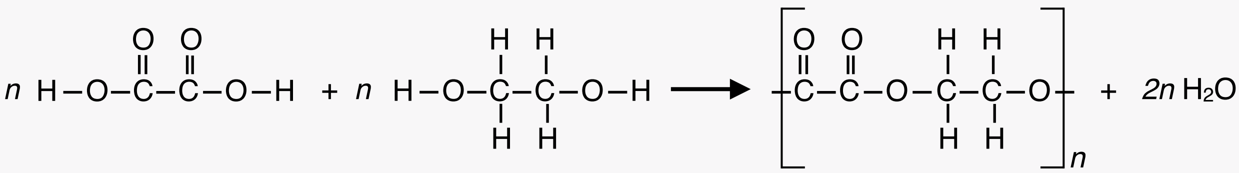
What is condensation polymerisation?
- *Formation of a polymer from monomers with different functional groups*
- *Produces polymer and water*

Drawing Condensation Polymers

dicarboxylic acid + diamine → polyamide + water



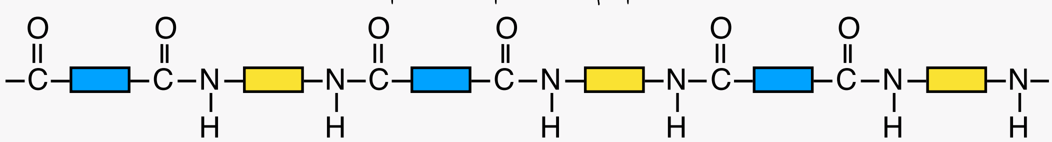
dicarboxylic acid + dialcohol → polyester + water



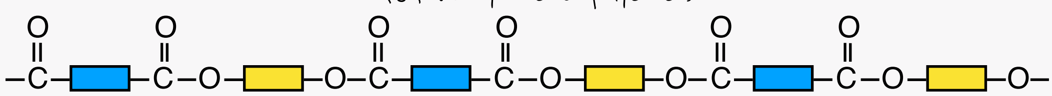
How does condensation polymerisation differ from addition polymerisation?
- *In condensation polymerisation a small molecule is also formed*
- *In condensation polymerisation two different monomers are required*

Drawing The General Structures Of Nylon & PET

Nylon (a synthetic polyamide)



PET (a synthetic polyester)



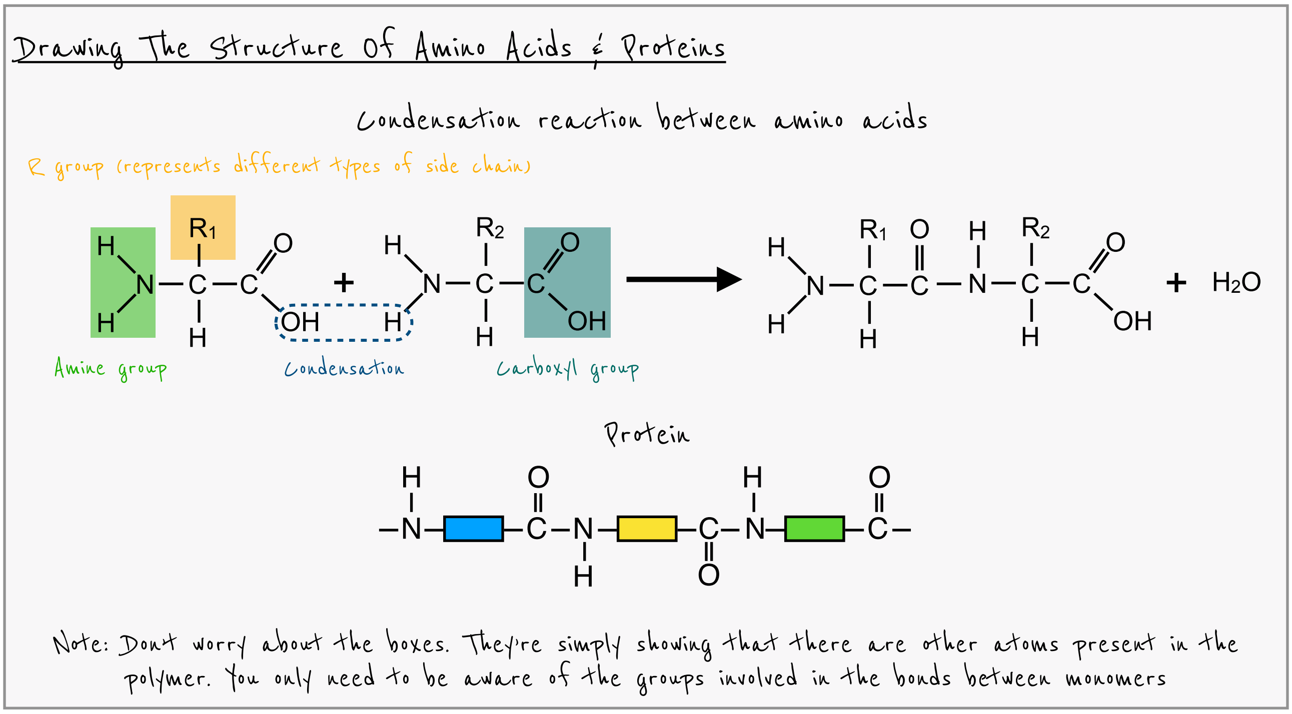
Note: Don't worry about the boxes. They're simply showing that there are other atoms present in the polymer. You only need to be aware of the groups involved in the bonds between monomers

What are the benefits of using PET as a plastic?
- *PET can be converted back into monomers and re-polymerised*
- *Reduces waste and and energy usages*
- *Reduces carbon footprint of production process*

What are the difficulties with the disposal of addition polymers?
- Inert (unreactive)
  - Means they do not-biodegrade
- When burned give off toxic gases

How do non-biodegradable plastics cause pollution problems?
- Fill up landfill
- Birds, fish and mammals choke if try to eat them
  - Cannot be digested so sit in stomach of animals - leads to starvation
- Clog up drains and sewers, causing flooding
- Collect in oceans rivers and harm aquatic life
- Produces toxic gases when burnt, harming surrounding wildlife

What are proteins?
- *Natural polyamides*
- *Formed from amino acid monomers*



Credit: Martin Bailey for SwH Learning

12. EXPERIMENTAL TECHNIQUES AND CHEMICAL ANALYSIS

12.1 Experimental design

Common measuring apparatus

Apparatus	Measure
Stopwatch	Time
Thermometer	Temperature
Balance	Mass
Burette	Volume (titration)
Volumetric pipette	Volume (transferring precise quantities)
Measuring cylinder	Volume
Gas syringe	Gas volume

Define solvent
- A substance that dissolves a solute

Define solute
- A substance that is dissolved in a solvent

Define solution
- A mixture of one or more solutes dissolved in a solvent

Define saturated solution
- A solution containing the maximum concentration of a solute dissolved in the solvent at a specific temperature

Define residue
- A substance that remains after evaporation, distillation or filtration

Define filtrate
- A liquid or solution that has passed through a filter

12.2 Acid-base titrations

How do you carry out an acid-alkali titration?
- Use volumetric pipette to add acid to conical flask
- Add indicator to flask e.g. methyl orange, phenolphthalein
- Place on white tile
- Use burette to add alkali to the conical flask
- Add alkali to conical flask drop-wise towards the end
- Swill contents of conical flask to mix
- Record volume of alkali that caused colour change
  - First attempt gives approximation of alkali volume required to neutralise acid
  - Subsequent repeats give more exact volumes

How is the end-point of a titration identified?
- Indicator changes colour

Indicator summary:

Indicator	Colour in acid	Colour in alkaline
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink
Litmus paper	Red	Blue
Universal indicator	Red	Purple

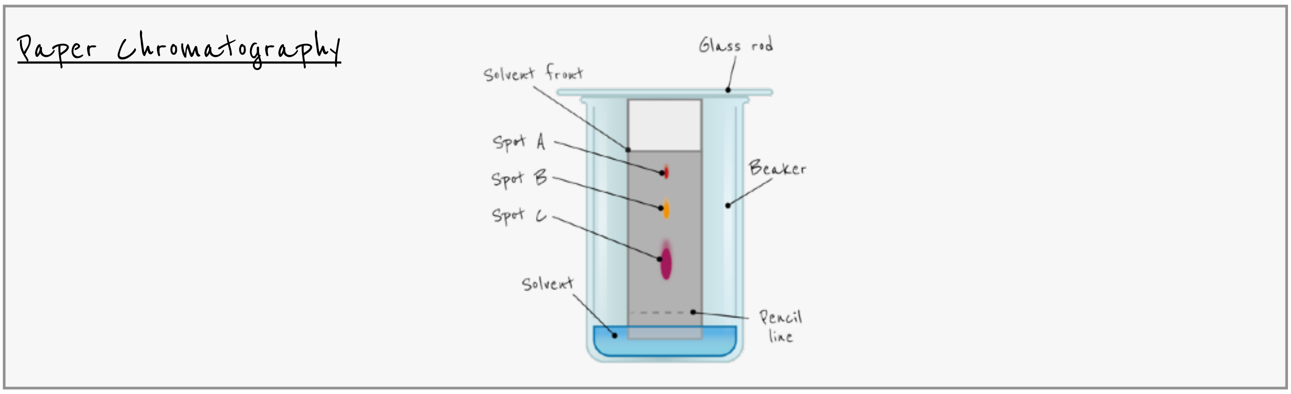
What colour is universal indicator when added to a neutral solution?
- Green

What colour is methyl orange when added to a neutral solution?
- Orange

12.3 Chromatography

What is paper chromatography used to separate?
- Dyes/inks i.e. coloured liquids with different solubilities

How to carry out paper chromatography:
- Use a pencil line as reference line
- Add spots of ink/dye to pencil line
- Place filter paper in suitable solvent
  - Make sure solvent level is below pencil line
- Leave until solvent nearly reaches top of paper
- Allow to dry
- Furthest dot has the greatest solubility



Credit: Tom Morley for SwH Learning

Why should the line be drawn in pencil?
- Pencil does not contain ink so does not run

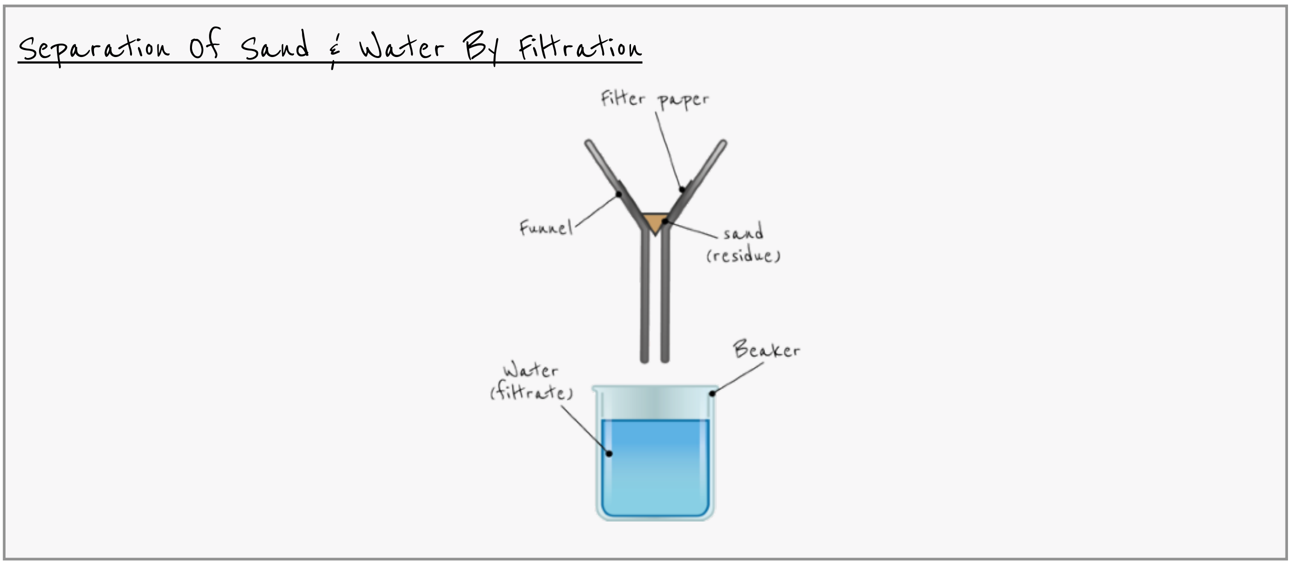
Explain why it is important for the solvent level to be below the spots of dye
- To prevent spots dissolving in the solvent

What is a locating agent used for?
- *Analysing colourless substances*
- *Locating agent reacts to form coloured products (or glow under UV light)*

How do you calculate the R_f value?
- $R_f = \frac{\text{distance travelled by component}}{\text{distance travelled by solvent}}$

12.4 Separation and purification

What method is used to separate an insoluble solute from a solvent?
- Filtration



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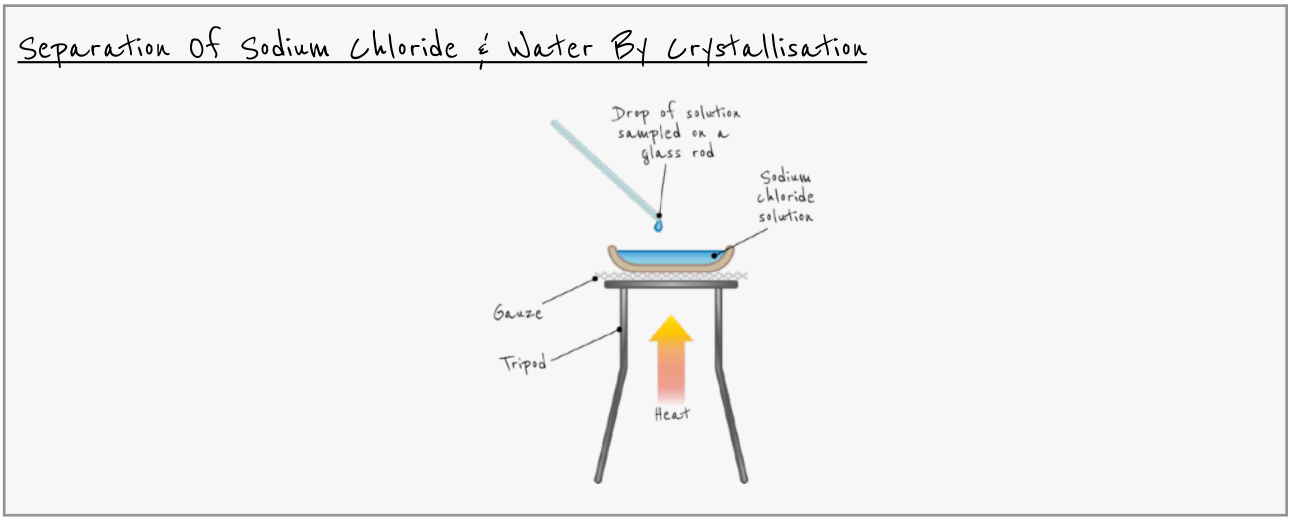
What method is used to separate a soluble solute from a solvent?
- Crystallisation



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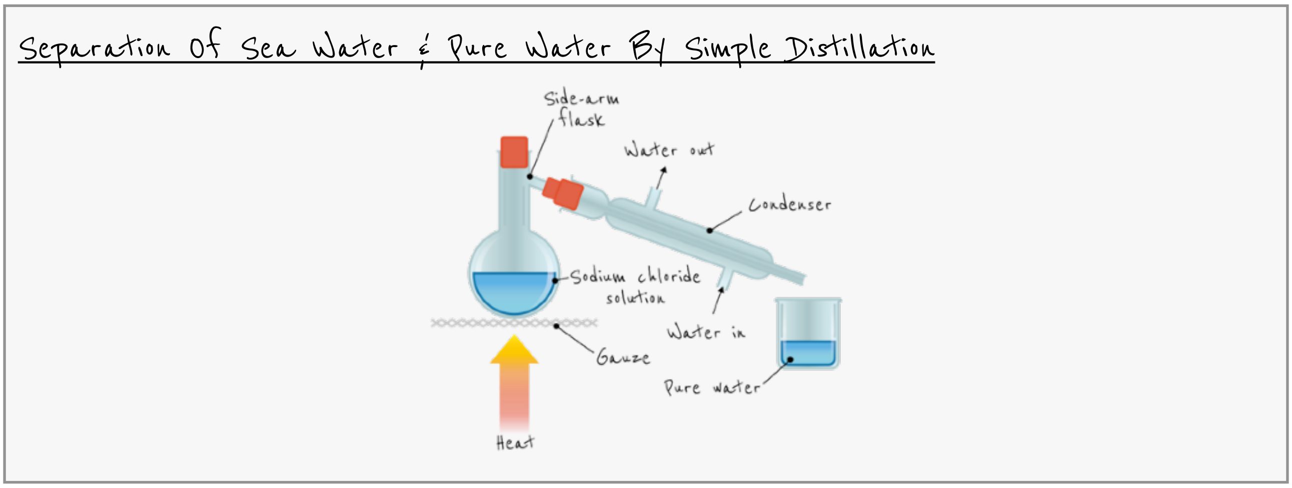
What method is used to separate pure water from sea water?
- Simple distillation



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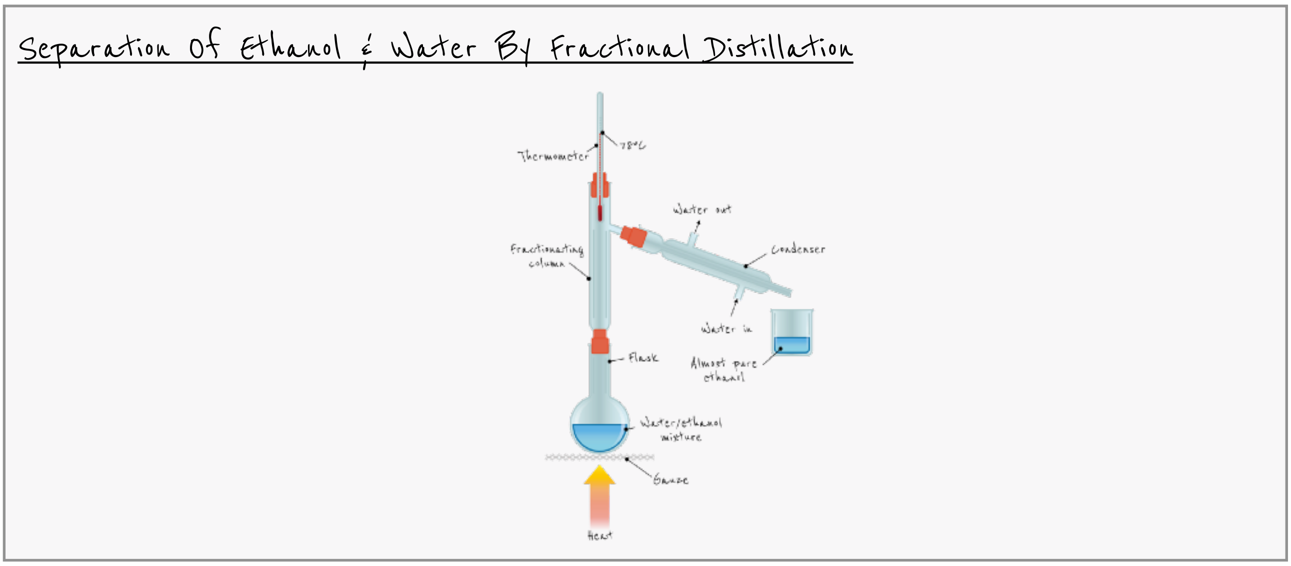
What method is used to separate liquids of different boiling points?
- Fractional distillation



Credit: Tom Morley for SwH Learning

What method is used to separate petrol and water and why is this used?
- Separating funnel
- Petrol and water are immiscible (don't mix)

What is a pure substance?
- Contains one type of material only
    - e.g. one type of element or molecule

Describe the melting and boiling points of pure substances
- Fixed
    - e.g. boiling point of pure water is exactly 100°C
    - e.g. melting point of ice to pure water is exactly 0°C

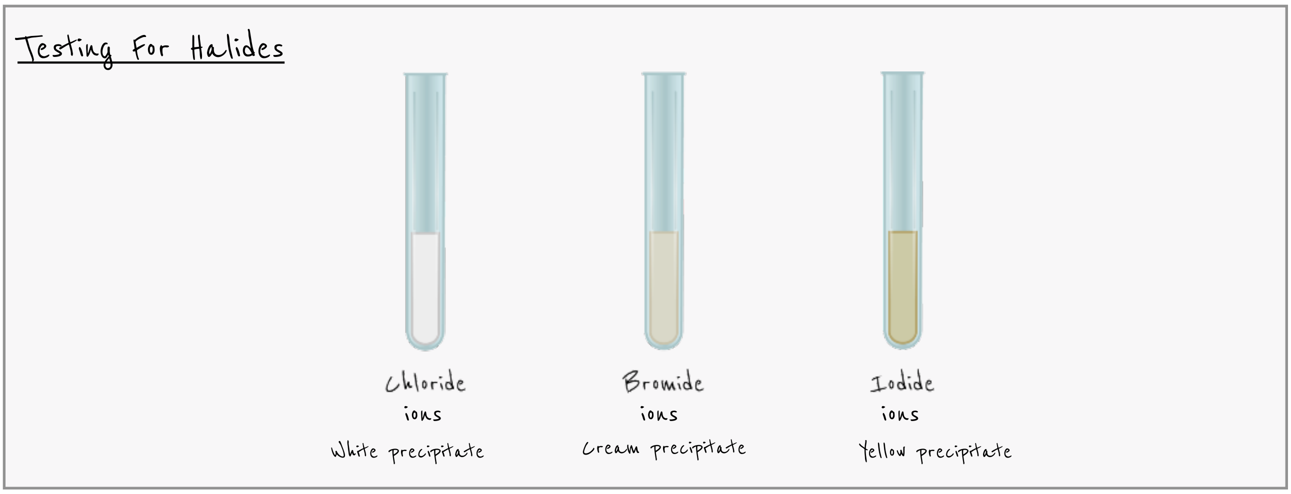
Describe the melting and boiling points of mixtures
- Melts and boils over a range of temperatures

12.5 Identification of ions and gases

How do you test for carbonates (CO_3^{2-})?
- Add dilute hydrochloric acid
    - Fizzing indicates carbon dioxide being produced
- Test for carbon dioxide using lime water
    - Turns milky/cloudy
- $2H^+ (aq) + CO_3^{2-} (aq) \rightarrow CO_2 (g) + H_2O (l)$

How do you test for halides?
- Add dilute nitric acid
    - Removes carbonate ions
- Add silver nitrate
    - Silver halide precipitate forms

Results of adding silver nitrate to halides:
- Silver chloride = white precipitate
- Silver bromide = cream precipitate
- Silver iodide = yellow precipitate
    - e.g. $Ag^+ (aq) + I^- (aq) \rightarrow AgI (s)$



Credit: Tom Morley for SwH Learning

How do you test for nitrates (NO_3^-)?
- Add aqueous sodium hydroxide and aluminium foil
  - Nitrates reduced to ammonia (NH_3)
- Ammonia turns damp red litmus paper blue

How do you test for sulfates (SO_4^{2-})?
- Add dilute nitric acid (removes carbonate ions)
- Add aqueous barium nitrate
- Result: barium sulfate is a white precipitate
  - Ba^{2+} (aq) + SO_4^{2-} (aq) → $BaSO_4$ (s)

How do you test for sulfites (SO_3^{2-})?
- React with acidified aqueous potassium manganate(VII)
  - Acid causes sulfur dioxide (SO_2) gas to be produced
  - SO_2 turns aqueous potassium manganate(VII) from purple to colourless

Summary of tests for aqueous cations:

Cation	Aqueous sodium hydroxide	Aqueous ammonia
Aluminium (Al^{3+})	White precipitate (soluble in excess)	White precipitate (insoluble in excess)
Ammonium (NH_4^+)	Ammonia gas produced (turns damp red litmus paper blue)	N/A
Calcium (Ca^{2+})	White precipitate (insoluble in excess)	No precipitate
Chromium(III) (Cr^{3+})	Green precipitate (soluble in excess)	Grey-green precipitate (insoluble in excess)
Copper(II) (Cu^{2+})	Light blue precipitate (insoluble in excess)	Light blue precipitate (dissolves to give dark blue solution)
Iron(II) (Fe^{2+})	Green precipitate (insoluble in excess)	Green precipitate (insoluble in excess)
Iron(III) (Fe^{3+})	Red-brown precipitate (insoluble in excess)	Red-brown precipitate (insoluble in excess)
Zinc (Zn^{2+})	White precipitate (soluble in excess)	White precipitate (soluble in excess)

What is the test for ammonia?
- Damp red litmus paper turns blue

What is the test for carbon dioxide?
- Limewater turns cloudy when gas bubbled through

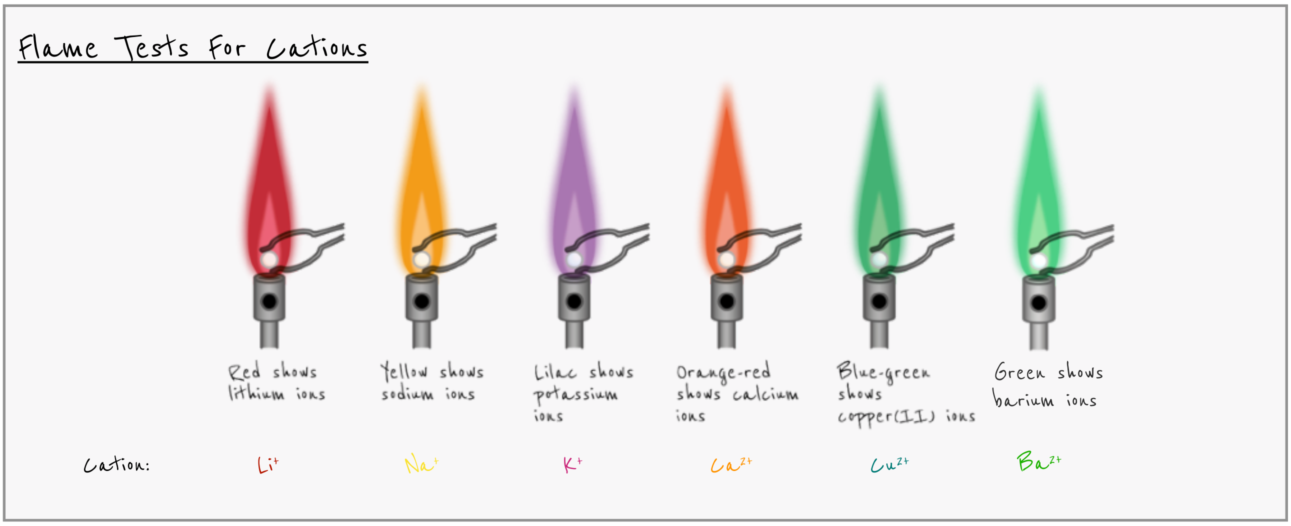
What is the test for chlorine?
- Bleaches damp blue litmus paper

What is the test for hydrogen?
- Lit splint burns with squeaky pop

What is the test for oxygen?
- Glowing splint relights

What is the test for sulfur dioxide?
- Acidified aqueous potassium manganate(VII) turns from purple to colourless

How do you carry out a flame test?
- Dip nichrome wire in hydrochloric acid to clean
- Dip in sample
- Hold in roaring blue Bunsen flame



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PRACTICAL SKILLS ASSESSED IN A WRITTEN EXAMINATION

Note: These skills will be examined in papers 3 & 4, as well as the 'Alternative To Practical' paper 6

When answering experimental design questions, you should always include…
- Independent variable
  - The variable that is being changed
- Dependent variable
  - The variable that is being measured
    - Include how the dependent variable would be measured
- Sensible time frame for taking measurements
- Control variables (minimum 5)
  - The variables that are being kept constant
- Methods of ensuring reliability

What is the purpose of a control?
- Allows a comparison to be made
- Shows what would normally happen so comparison can be made when independent variable is changed

Common ways of improving an investigation:
- Ensure equal sizes/volumes of samples are used
- Repeat the experiment at least 3 times and calculate the mean
- Test a wider range of values for the independent variable

How to make an investigation more <u>reliable</u>:
- Repeat experiment at least three times to increase number of observations
- Identify anomalous results
- Calculate mean

When have a suitable number of results have been obtained for a titration?
- When at least two results are within 0.2 cm^3 or less

How to make an investigation more <u>accurate</u>:
- Carry out more tests within existing range
- Introduce method to ensure no double counting occurs
- Use a narrower range (if appropriate)

How to increase the <u>validity</u> of an investigation:
- Make sure that all control variables are the same for each repeat/investigation
- Collect a wide range of measurements/results

When answering 'describe' questions:
- Write what the data is showing e.g. trends, changes in rate, increases and decreases etc
- If describing a graph, break the graph down into sections
  - Each section should be a describable feature
  - e.g. constant rate from A to B, increasing rate from B to C…
- Use data points provided in the question to illustrate description

When answering 'explain' questions:
- Say why the results have come about
- Use scientific knowledge to explain any patterns and trends
- Make sure explanation is specific to the question