Adrian's Study Kit

Year 11 Chemistry

buzz buzz… i'm an electron

Syllabus:
https://educationstandards.nsw.edu.au/wps/portal/nesa/11-12/stage-6-learning-areas/stage-6-science/chemistry-2017

Contents

Properties and Structure of Matter
Module 1

matter matters!

Properties of Matter

Purity of substances

Pure substances - Elements - Compounds	- **Fixed composition** (*i.e* the proportions of atoms remain the same throughout) (*e.g* `NaCl` will always have a ratio of `1Na:1Cl`) - **Fixed properties** (*e.g* H_2O will always have a boiling point of 100°C) - **Chemically bonded** (*e.g* `NaCl` is an ionically-bonded molecule) - **Can only be separated through chemical reactions** (*e.g* combustion, electrolysis)
Impure substances - Mixtures	- **Variable composition** (*e.g* salt water has H_2O and `NaCl` mixed randomly) You can vary the amount of each substance in a mixture - **Variable properties** (*i.e* each substance in the mixture keeps its own properties) - **Not chemically joined** - **Can be separated by physical separation technique** (*e.g* filtration, boiling)

Elements, compounds & mixtures

	Description	**Example**
Elements	- Pure substance - Only one type of atom - Cannot be broken down into a simpler substance	- H_2 - Fe - Ar
Compounds	- Pure substance - Has multiple types of atoms - Can only be separated through chemical equations	- NaCl - H_2SO_4 - CCl_4
Mixtures	- Impure substance - Variable composition and properties - Not chemically joined	- Salt water - Air - Milk

Homogeneous and heterogeneous mixtures

Homogenous mixture	Uniform in composition	*e.g* salt water is homogeneous since the salt is uniformly mixed within the water particles.
Heterogeneous mixture	Variable composition	*e.g* muddy water is heterogeneous since heavier sediment sinks to the bottom, while the more watery mud sits at the top.

Note: Milk is heterogeneous since there are fat lumps that are distributed unevenly… yuck.

Solutions

- A homogenous mixture in which dispersed particles are so small they never settle out
- Cannot be seen by a microscope

Aqueous solutions

- Formed when a substance is dissolved in water
- The substance that is dissolved is called the **solute0**
- The liquid that dissolves the solute is called the **solvent**

Example: When dissolving salt in water, salt is the **solute** and water is the **solvent.**

Polyatomic ions

Name	**Formula**	**Valency**
Hydroxide	OH^-	1–
Nitrate	NO_3^-	1–
Hydrogencarbonate	HCO_3^-	1-
Carbonate	CO_3^{2-}	2-
Sulfate	SO_4^{2-}	2-
Phosphate	PO_4^{3-}	3-

Naming compounds

Polyatomic ions

nitr–**ate**	Largest number of oxygen atoms	NO_3
nitr–**ite**	Least number of oxygen atoms	NO_2
nitr–**ide**	Negatively charged ion	N_2

Ionic compounds (metal/non-metal)

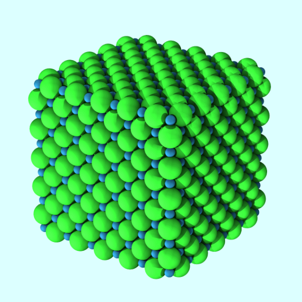
- **Magnesium chloride** $MgCl_2$
- **Calcium hydroxide** $Ca(OH)_2$

Covalent compounds (non-metal/non-metal)

- **Dinitrogen pentoxide** N_2O_5
- **(Di)phosphorus trioxide** P_2O_3

1	2	3	4	5	6	7	8	9	10
mono	di	tri	tetra	pent	hex	hepta	octa	nona	deca

Empirical formula



This is our friend, **NaCl**.

NaCl is in a 1:1 ratio. It is the simplest form in which it exists. Hence, this is known as the **empirical formula**.

Therefore, the empirical formula is the simplest ratio in which atoms in a compound can exist.

Ionic equations

Given an element, an electron can be added to either side of an ionic equation. This will cause it to either gain/lose electrons, giving it a full valence shell.

$$Na \rightarrow Na^{+} + e^{-}$$

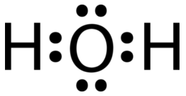
$$(2,\ 8,\ 1) \rightarrow (2,\ 8) + e^{-}$$

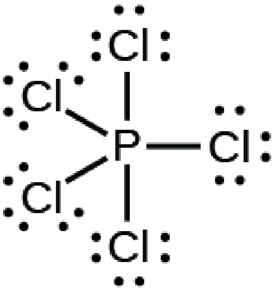
$$F + e^{-} \rightarrow F^{-}$$

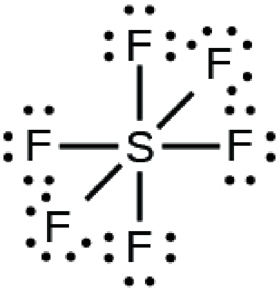
$$(2,\ 7) + e^{-} \rightarrow (2,\ 8)$$

Lewis dot diagrams

Lewis dot diagrams are simplified diagrams used to show the valence electrons and bonding within compounds and substances.







Atomic structure

- The nucleus consists of protons and neutrons
- The outer shells consist of electrons which are of negligible mass

Mass of particles

Subatomic particle	Charge	Mass (amu)
Protons (p^+)	+	1
Neutrons (n^o)	o	1
Electrons (e^-)	–	1/1800

Atomic calculations

Atomic mass

$$atomic\ mass\ = p^{+} + n^{o}$$
$$\therefore A\ =\ mass(p^{+} + n^{o})$$

where **A** is atomic mass.

Atomic number

$$atomic\ number\ = p^{+}$$
$$\therefore Z\ =\ atomic\ number(p^{+})\ = e^{-}$$

where **Z** is atomic number.

Nuclear notation

Atoms and isotopes can be represented in standard nuclear notation.

$${}^{A}_{Z}X$$

For example, oxygen-16 and carbon-12 would be represented as

$${}^{16}_{8}O \qquad {}^{12}_{6}C$$

Radioactivity

Radioactive decay occurs when unstable atoms lose energy via radiation. There are considered to be three types of radiation.

Name	Particle representation	Equivalent to	Has particle mass?
α	${}^{4}_{2}He^{2+}$	**Nucleus**	**Has particle mass**
β	${}^{0}_{-1}e$	**Electron**	
γ			**No particle mass**

Alpha decay

Alpha decay gives off a particle identical to helium-4, gaining protons while losing neutrons.

$${}^{226}_{88}Ra \rightarrow {}^{4}_{2}He + {}^{222}_{86}Rn$$

Beta decay

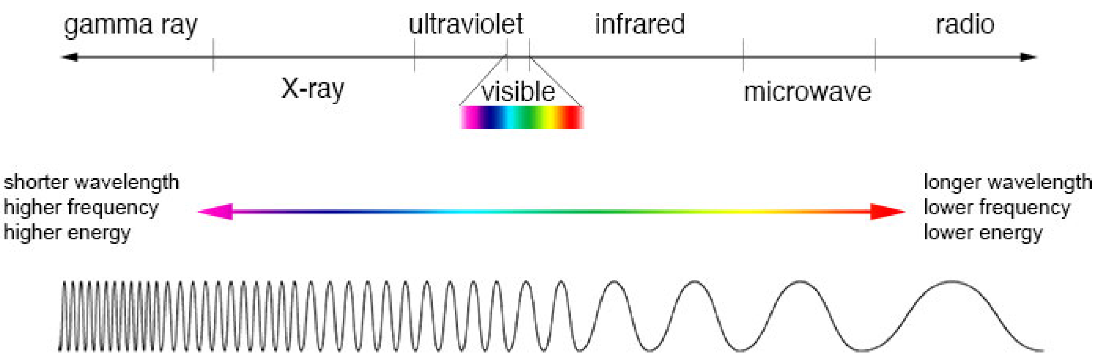
Beta decay gives off an electron as well as a small amount of gamma rays.

$${}^{60}_{27}Co \rightarrow {}^{60}_{28}Ni + {}^{0}_{-1}e \quad (+\ \gamma)$$

Electromagnetic spectrum

The greater the amount of energy to be released, the shorter the wavelength of the radiation emitted. These wavelengths can be characterised into areas such as ultraviolet, visible light and infrared.

Electrons aren't just particles as they also have wave properties. They can be assigned a wavelength or undergo diffraction.



Relative atomic mass

Given two isotopes, with both relative isotopic mass and abundance known, you can find the relative atomic mass.

$$A_r = \frac{(I_{r_1} \times A_1) + (I_{r_2} \times A_2)}{100}$$

where:

- **A_r** is relative atomic mass
- **I_r** is relative isotopic mass
- **A** is abundance

For example, given two isotopes of chlorine exist:

$$^{35}_{17}Cl \qquad ^{37}_{17}Cl$$

with respective abundances 75 and 25, therefore:

$$A_r = \frac{(35 \times 75) + (37 \times 25)}{100} \approx 35.5\ (1dp)$$

Remember: For unknown abundances, let the respective abundance to be x and $100 - x$.

Relative molecular mass

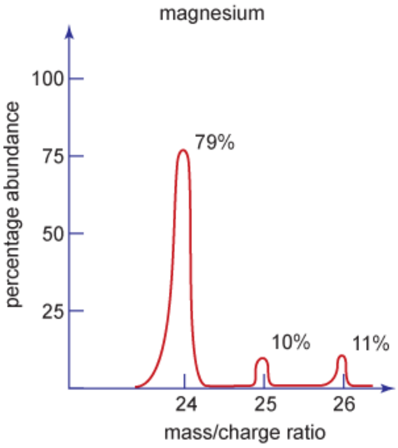
Molecular mass is measured against carbon-12. Hence, to calculate relative molecular (formula) mass, take the sum of all relative atomic mass of all atoms which comprises a molecule.

$$MM = \sum relative\ atomic\ masses$$

Example: For example, in copper nitrate => $Cu(NO_3)_2$

$$Cu + N + O_3$$
$$63.55 + 2[14.01 + 3(16.00)]$$
$$= 187.57\ amu$$

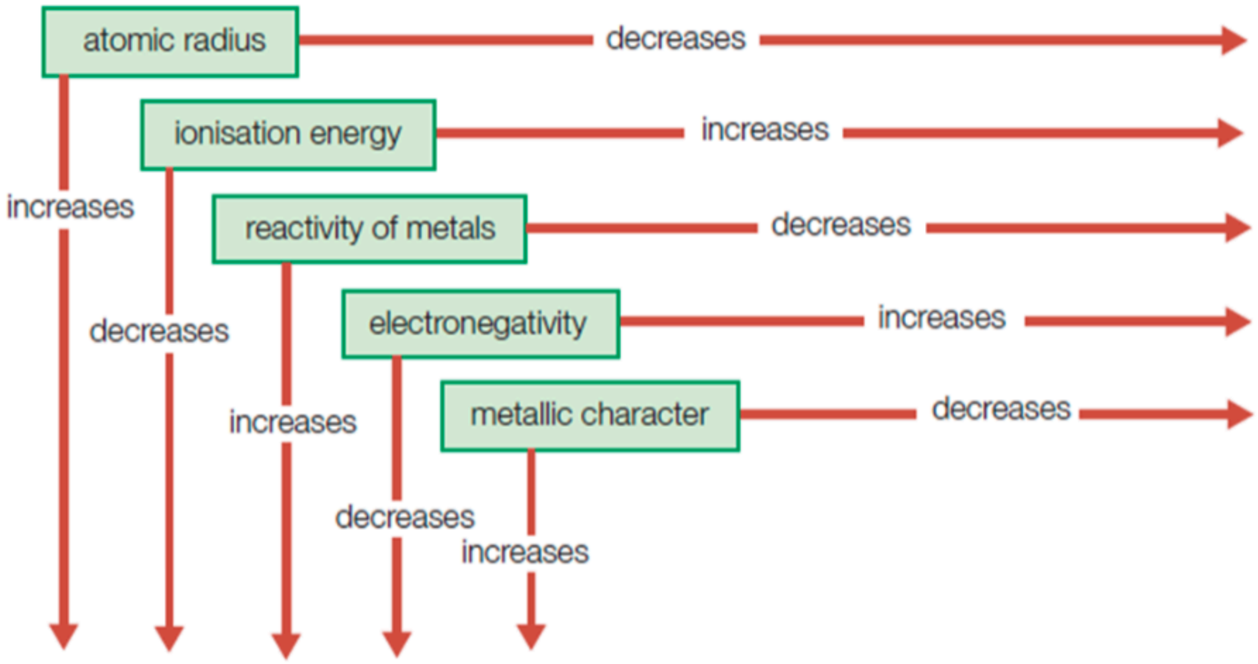
Percentage abundance



Given a simplified mass spectrum such as the one provided for magnesium, you can find the percentage of a particular isotope occuring in nature (i.e percentage abundance). They can be graphed on a mass spectrum which uses data obtained from a **mass spectrometer.**

Periodicity

There are many other trends that we can identify throughout the Periodic Table. It can be summarised in the following diagram.



let’s try and do our best to understand each one!

Atomic radius

Atomic radius is how large an atom is. Yep, that’s about it.

Trends in atomic radius

Excluding the noble gases, atomic radius:

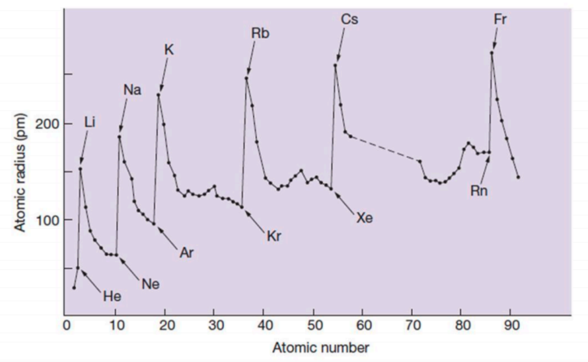
- **decreases** across a period from left to right
- **increases** down a group from top to bottom

Note: Note that along a period, you gain more protons, yet the atomic radius decreases. Therefore, it can be said that **as atomic radius decreases, atomic density increases**.

Atomic radius will suddenly increase at the beginning of every period. This occurs as at the beginning of each period, it gains a new electron shell.

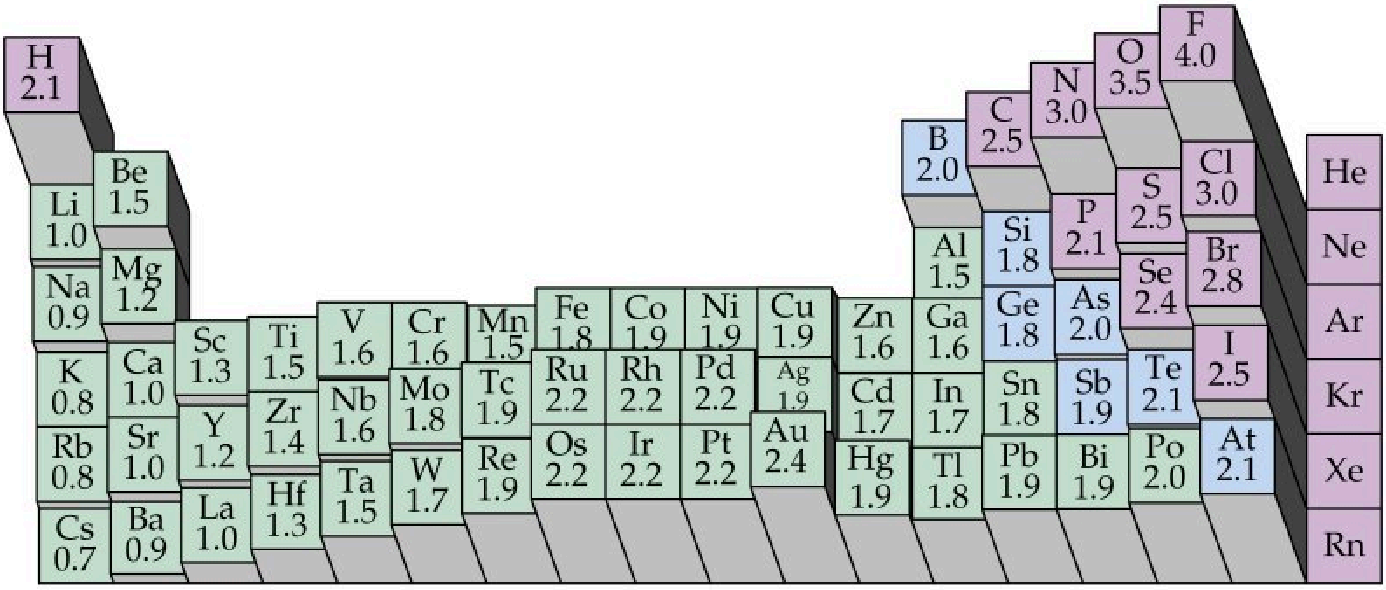
Why does atomic radius decrease across a period and increase down a group? See Explaining core charge.

Atomic radius increases in going down any group of the table.



Electronegativity

Electronegativity is the measure of an atom's ability to attract shared electrons towards itself, when forming compounds.



Explaining core charge

Electronegativity and atomic radius is best explained with the concept of **core charge:**

1. At the beginning of each period, the **core charge** of an atom suddenly increases, given the addition of one electron shell.
2. Down a group, the core charge remains **constant**, however valence electrons have weaker attractions as they're further away from the nucleus.
3. Across a period, core charge increases ∴ the valence electrons are attracted more towards the nucleus.

Core charge:

increases	across a period from left to right
is constant	within a group

The **core charge** of an atom is a measure of the attraction of valence electrons toward the nucleus.

- **Down a group,** atoms will contain the same core charge. But, since they're further away, the valence electrons are held less strongly as the force of attraction is not as strong.
- **Across a period**, this core charge increases. Therefore, valence electrons are attracted more strongly towards the nucleus.

Therefore, core charge can be calculated using the following formula:

$$core\ charge\ =\ number\ of\ p^{+}\ -\ number\ of\ inner\ shell\ e^{-}$$

Example:

Take two Group 7 non-metals.

- **Fluorine** has electron configuration (2, 7) = $9p^+$
  ∴ inner-shell e^- = 2
  ∴ core charge = 7+
- **Chlorine** has electron configuration (2, 8, 7) = $17p^+$
  ∴ inner-shell e^- = 10
  ∴ core charge = 7+

Fluorine and chlorine carry the same core charge of 7+. However, fluorine's valence electrons are closer to the nucleus, therefore it will carry a stronger force of attraction.

Trends in electronegativity

Excluding the noble gases, electronegativity and ionic character:

- **increases** across a period from left to right
- **decreases** down a group from top to bottom

Since ionic character increases across a period, the type of compound will demonstrate trends as well. For instance, take a compound with oxygen, using various elements across period 3.

Across a period

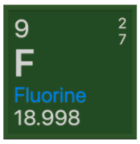
Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Ionic			Covalent		Covalent/ionic	

Down a group

CH_4	Covalent
$SiCl_4$	
$SnCl_4$	Ionic
PbO	

0.7

1.83



3.8

Note: Percent ionic character is the amount of electron sharing between two atoms. A percent ionic character is proportional to Δ electronegativity.

If $\Delta > 2.1$ ∴ ionic compound, otherwise covalent.

Note:

- Electro**negativity** = tendency to **gain** electrons.
- Electro**positivity** = tendency to **lose** electrons

Example:

- Fluorine has the highest electronegativity of any element on the Periodic Table. This means that it attracts electrons better than any other element. It also has the highest ionic character.
- Caesium has a very low electronegativity. It does not attract electrons very well.

Bonding

Revision: Ionic and covalent bonds

- **Ionic bonding:** transfer of electrons between ions
- **Covalent bonding:** sharing of electrons

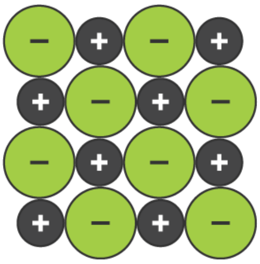
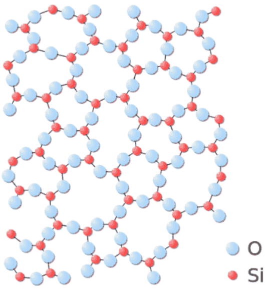
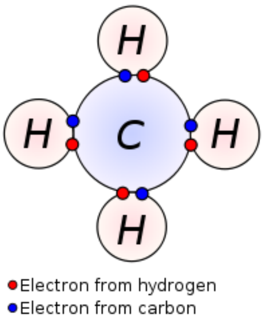
What is intramolecular bonding?

Intramolecular forces are forces between molecules, such as ionic bonds or covalent bonds.

Chemical structures

Summary

	Lattice			
	Ionic networks	**Covalent networks**	**Metallic structure**	**Covalent molecular**
Description	Ionic compounds which extend indefinitely in a regular, repeating arrangement.	Solids where covalent bonding extends indefinitely in a regular, repeating arrangement.	Metals where positive metal ions are arranged in a lattice structure, amongst a sea of delocalised electrons.	Sharing of electrons.
Melting point/boiling point	High (strong ionic bonds)	Very high (strong covalent bonds)	High (strong metallic bonds)	Low
Solubility	Soluble	Insoluble	Insoluble	Insoluble=
Electrical conductivity	Solid: no Molten/liquid: yes as there are free moving electrons	No	Yes, as there are free moving electrons in the sea of delocalised electrons	No
Hardness	Hard and brittle	Hard and brittle	Malleable	Soft
Forces	Electrostatic forces of attraction between oppositely charged ions	Covalent bonding throughout the crystal	Electrostatic forces of attraction between electrons and positive ions (metallic bonding)	Intermolecular forces
Examples	Sodium chloride	Diamond Silicon dioxide (sand)	Copper	Bromine gas Oxygen gas Water

Did you know? Silicon dioxide was the **molecule of the month** for November 2013, as designated by Guillermo Godino Sedano of King's College, Madrid, Spain.

Melting point/boiling point

Melting point and boiling point (mp/bp) is dependent on how much energy is required to break bonds within substances.

- **Ionic networks** have strong electrostatic forces of attraction between the positive and negative ions, and therefore will have a **high mp/bp.**
- **Covalent networks** have strong covalent bonding within the crystal, and therefore will have a **very high mp/bp.**
- **Metallic structures** have strong electrostatic forces of attractions between positive ions and the sea of delocalised electrons, and therefore will have a **high mp/bp.**
- **Covalent molecular solids** have weak covalent forces of attraction, and therefore will have a **low mp/bp.**

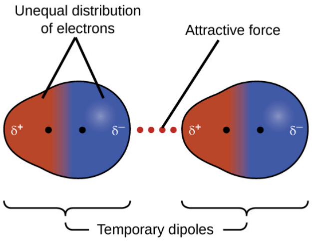
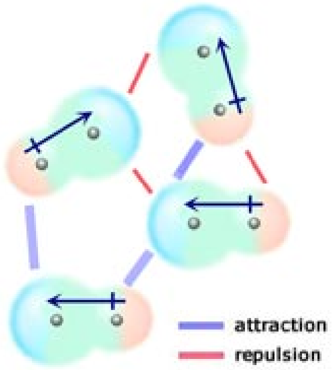
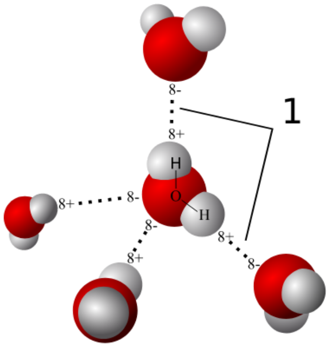
Electrical conductivity

Electrical conductivity is dependent on the number of free flowing electrons.

- **Ionic networks** do not conduct electricity as a solid as there are no free electrons. But in a molten/liquid state, it can, since the ions are free to move around.
- **Covalent networks** can't conduct electricity as there are no free electrons.
- **Metallic structures** can conduct electricity through a sea of delocalised electrons.
- **Covalent molecules** can't conduct electricity as there are no free ions or electrons.

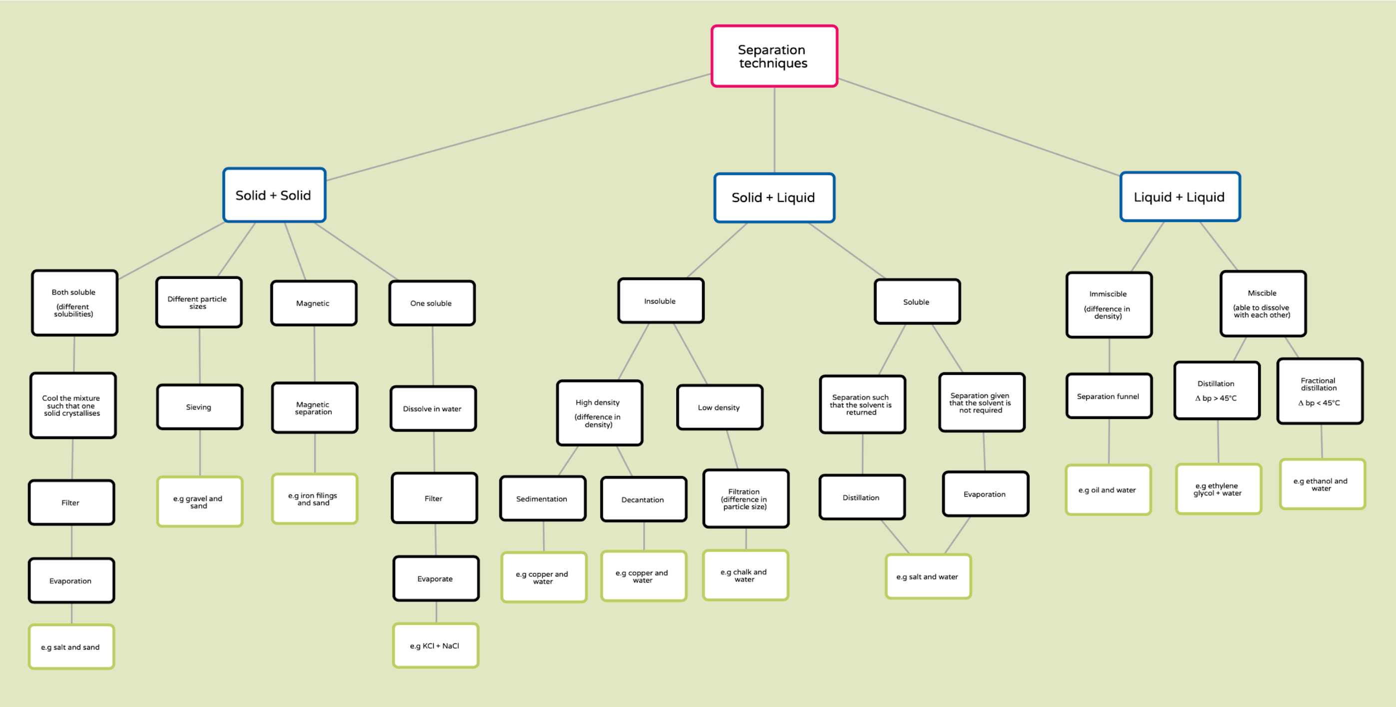
Intermolecular forces

Intermolecular forces are forces between molecules that attract them together.

	Dispersion forces	Dipole-dipole forces	Hydrogen bonds
Between which molecules?	All atoms and molecules	Polar molecules only	Hydrogen attached to fluorine, oxygen or nitrogen.
Explanation	Constant and fluctuating movement of electrons and the nucleus, causes constant repulsion and attraction.	Permanent dipoles between a slightly positive side and slightly negative side of each respective molecule.	A form of dipole-dipole force between bonds with very high electronegativity (H-F, H-O, H-N).
Diagram			
Increases with what?	Increases with higher molecular mass.	Increases with difference in electronegativity (aka dipole moment) which is representative of polarity.	
Relative strength	Weakest	Middle	Strongest
Example	H_2O, CO, Cl_2	CCl_4, HCl, H_2O	HF, H_2O

Remember: Hydrogen bonds only have H attaching to **F, O, N.**

Separation techniques

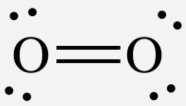
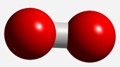
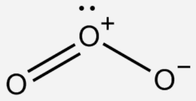
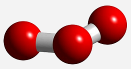


Source link: https://app.popplet.com/#/p/5509089

Allotropes

Allotropes are different structural forms of the same element, each with different physical properties.

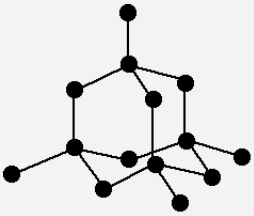
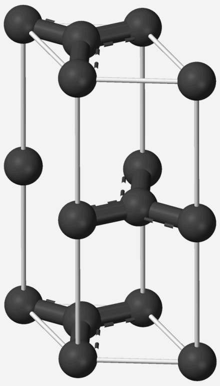
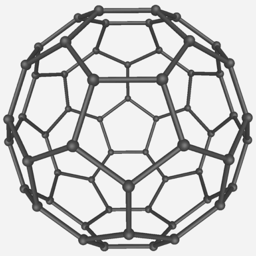
Oxygen

Properties of the two main oxygen allotropes			
Property	**Molecular oxygen**	**Ozone**	**Description**
Chemical formula	O_2	O_3	
Physical structure	Linear  	Bent  	
Melting point/ boiling point	**mp:** -219°C **bp:** -183°C	**mp:** -193°C **bp:** -111°C	O_3 = higher MM = higher mp/bp
Density	1.429 g/L (STP)	2.144 g/L (STP)	Oxygen gas' density is around the same as

			air, while ozone is 1.5x that.
Stability	Stable	Easy to decompose	Ozone has one single covalent bond, while oxygen gas has a double covalent bond, affecting both stability and oxidising strength.
Oxidising strength	Moderately strong	Very strong	
Solubility	Not very soluble	Much more solar	Ozone is slightly polar. O_2 is nonpolar.
Colour	White	Pale blue	

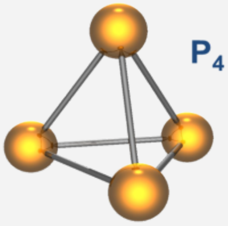
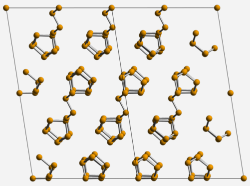
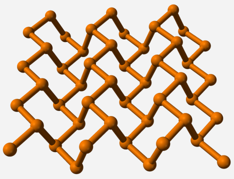
Carbon

There are three main allotropes of carbon: diamond, graphite and fullerene.

<table>
<tr><th colspan="5">Properties of the three main carbon allotropes.</th></tr>
<tr><th>Description</th><th>Diamond</th><th>Graphite</th><th>Fullerene</th><th>Description</th></tr>
<tr><td rowspan="4">Structure</td><td colspan="3">• All covalent bonds
• All lattice structures</td><td></td></tr>
<tr><td colspan="2">• Six-membered carbon rings</td><td></td><td></td></tr>
<tr><td></td><td colspan="2">• Free valence electrons = delocalised electrons.</td><td></td></tr>
<tr><td>Each carbon atom is bonded to **4** other carbon atoms.
No valence electrons.
Tetrahedral structure.
</td><td>Each carbon atom is bonded to **3** other carbon atoms.
One free valence electron in each carbon atom.
Weak IMF between slides, they can slide across each other.
</td><td>Each carbon atom is bonded to three other carbon atoms.
Some delocalised electrons.
Spherical cage structure — buckyballs (non-planar bonds)
</td><td></td></tr>
</table>

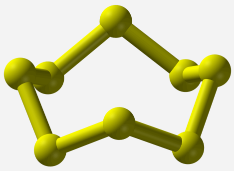
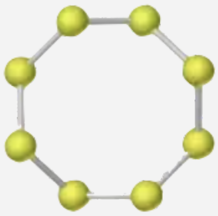
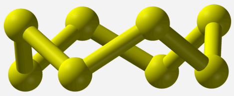
Conductivity	Does not conduct electricity	Conducts electricity		Graphite and fullerene both have free electrons in the delocalised electron cloud, hence they have high conductivity.
Physical properties	• Very hard • Transparent • Brilliant	• Slippery • Lubricating		
Uses	• Drills and saws • Jewellery	• Electrodes • Batteries • Pencils • Lubricants	Biomedical uses: • Gene/drug carrier Technological uses: • Superconductors in flat-panel displays • Polymers	Fullerene is still not used in practical application so much, but is subject to ongoing research.

Phosphorous

Properties of the three main phosphorus allotropes				
Property	**White (yellow)**	**Red**	**Black**	**Description**
Structure	• P_4 molecules • Covalent bonding			
	Each P atom bonded to **3** other P atoms.  	Chain of P atoms with P-P bonds between neighbouring P_4 units. 	Each P atom bonded to **3** other P atoms. Buckled layers/sheets. Strong bonds in P atom pairs within the sheet; weak bonds between sheets. 	
Stability and reactivity	———————————— Stability increases ————————————> ——————————— Reactivity decreases ———————————>			
	Spontaneously	Reacts with oxygen	Not too reactive.	Red phosphorus is

	combusts in air.	with enough energy.		used in matches to ignite a match.
Other physical properties	• Insoluble in water • Soft and waxy		• Flaky	
Uses	• Explosives and munitions • Rockets • Cleaning compounds	• Match heads • Smoke bombs • Fireworks	• Sensors • LEDs • Batteries	
	• Fertiliser and pesticides			

Sulfur

Properties of the three main sulfur allotropes				
Property	**Rhombic**	**Monoclinic**	**Plastic**	**Description**
Bonding	• Covalent bonding			
	8 sulfur atoms covalently bonded in a ring shape		Long chains of sulfur atoms	
Structure	Crystalline. **Octahedral.** 	Needle-like crystals. 	Non-crystalline. 	
Other physical properties			• Rubbery and non-crystalline texture • Dark brown/black • No definite melting point • Changes back to S_8 rings on standing	

Note: Allotropes of sulfur are interconvertible. For example, at 369 °K/96 °C, rhombic sulfur gives out monoclinic sulfur.

Uses

- Vulcanisation (hardening) of rubber
- Many insecticides contain sulfur in a compound

- Manufacturing of bleaching events
- As part of carbon disulfide, which is used in skin ointments and skincare products
- Fertilisers
- Gunpowder, matches and fireworks

Note: In most cases, sulfur is used as part of a compound which is then applied to practical usage.

Introduction to Quantitative Chemistry

Module 2

The mole

- One mole is defined as the amount of substance that contains the same number of 'specified' particles as there are in 12.0g of carbon-12.
- The number of particles in one mole is represented by the symbol **N_A** known as **Avogadro's number** which has the value 6.022×10^{23}.

Important formulae

$n = \frac{N}{N_A}$	● **n** is number of moles ● **N** is number of particles ● **N_A** is Avogadro's Number = 6.022×10^{23}
$n = \frac{m}{MM}$	● **n** is number of moles ● **m** is mass in grams ● **MM** is molar mass
$c = \frac{n}{v}$	● **c** is the concentration of the solution (mol L-1) ● **n** is the number of moles of the solute (mol) ● **v** is the volume of the solution (L)
$c_1 v_1 = c_2 v_2$	● **c** is the concentration of the respective solution (mol L-1) ● **v** is the volume of the respective solution (L)

Limiting reagents

In quantitative calculations it is important to consider that some reactants may not be used fully. To find the **limiting reagent**, or the reactant which is used up fully in a reaction:

1. Find number of moles present
2. Compare against the molar ratio and the number of moles required for a full reaction - which moles are in excess?
3. The limiting reagent is the reactant that is fully used up

Concentration

Given:

$$solute + solvent = solution$$

% w/w	$$c = \frac{weight\ of\ solution}{weight\ of\ solution} \times \frac{100}{1}\ \%\ w/w$$	**mg/kg**
% v/v	$$c = \frac{volume\ of\ solution}{volume\ of\ solution} \times \frac{100}{1}\ \%\ v/v$$	**mL/kL**
% w/v	$$c = \frac{weight\ of\ solution}{volume\ of\ solution} \times \frac{100}{1}\ \%\ w/v$$	**mg/L**

Parts per million

$$\begin{aligned} 1\ ppm &= \frac{1g}{10^6 g} \\ &= \frac{1g}{1000kg} \\ &= \frac{1mg}{1000g} \\ &= \frac{1mg}{1kg} \end{aligned}$$

Molarity

$$c = \frac{n}{v}$$

where:

- **c** is the concentration of the solution (mol L^{-1})
- **n** is the number of moles of the solute (mol)
- **v** is the volume of the solution (L)

Dilution

$$c_1 v_1 = c_2 v_2$$

where:

- **c** is the concentration of the respective solution (mol L^{-1})
- **v** is the volume of the respective solution (L)

Standard conditions

Name	Temperature	Pressure	Molar volume (V_M)
SLC — Standard Laboratory Conditions	25°C	100 kPa	24.79 L
STP — Standard Temperature and Pressure	0°C	100 kPa	22.71 L

Gas laws

Properties of ideal gases

- Gas particles are in **constant, random, straight-line** motion
- Collision with other particles are **perfectly elastic** – no energy is lost
- No attraction or repulsion between particles
- Volumes of particles are **insignificant** compared to the gas' total volume
- Kinetic energy of particles is a measure of the particles' absolute temperature

Unit conversions

Pressure

1 bar = 100kPa
1 atm = 760mmHg = 101.3kPa

Temperature

Let T_C be the temperature in °C and let T_K be the temperature in K (Kelvin).

$$T_C + 273.15 = T_K$$

Laws

Name	Proof	Law	Constant
Boyle's Law	Pressure is inversely proportional to volume: $P \propto \frac{k}{V}$	$P_1V_1 = P_2V_2$	**n**
Charles' Law	Volume is inversely proportional to temperature: $V_1 = kT_1$ and $V_2 = kT_2$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	**P**
Gay-Lussac's Law	Pressure is inversely proportional to temperature: $P_1 = kT_1$ and $P = kT_2$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	**V**
Combined Gas Law	By combining Charles' Law and Gay-Lussac's Law, we obtain the Combined Gas Law.	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	**n**
Avogadro's Law		$\frac{V_1}{n_1} = \frac{V_2}{n_2}$	

Ideal Gas Law	By combining all other gas laws, we obtain the Ideal Gas Law, which works at non-standard conditions.	$PV = nRT$	

Molar volume

At standard conditions, the relationship between volume and molar volume is given as

$$n = \frac{V}{V_M}$$

Reactive Chemistry

Module 3

Redox reactions

Remember **OILRIG**:

- Oxidation is the **loss** of electrons
- Reduction is the **gain** of electrons

Note:

- The **oxidising agent** causes **reduction**
- The **reducing agent** causes **oxidation**

Ionic equations

Complete ionic equation	$Fe_{(s)} + Cu^{2+}_{(aq)} + SO^{2-}_{4(aq)} \rightarrow Fe^{2+}_{(aq)} + SO^{2-}_{4(aq)} + Cu_{(s)}$
Net ionic equation	$Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$
Half equation	**Oxidation:** $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^{-}$ **Reduction:** $Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)}$

Experiments & investigations

Accuracy

- How close is the experimental value close to the theoretical value?

Validity

- Does the experiment achieve the aim?
- Is it maintained at standard condition?

Reliability

- Repetition
- Consistency

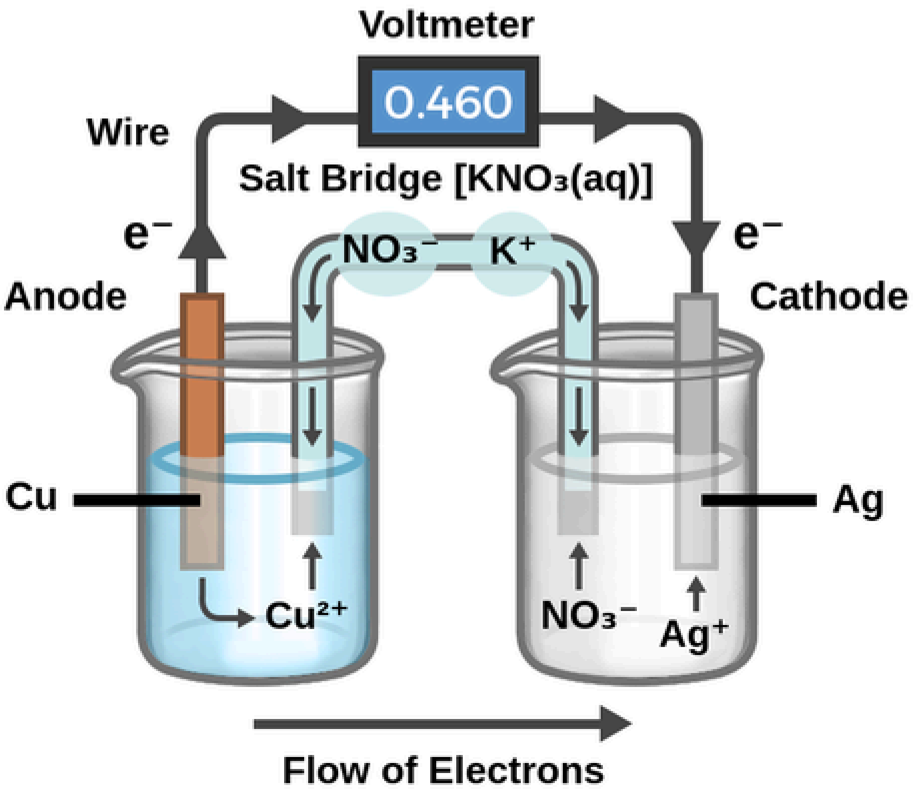
Reactivity

More reactive metals will oxidise more easily, while a stronger oxidant will have a lower reduction potential.

Example: Magnesium is highly reactive and will both rust (oxidise) and combust well, therefore it will oxidise better. Gold isn't highly reactive and therefore won't oxidise easily, which is why it is used to make electrical components.

Galvanic cells

Galvanic cells are able to store energy and form an electrical circuit. The anode loses electrons, while the cathode gains electrons and forms a precipitate.



- Electrons leave the anode and flow to the cathode.
- In the salt bridge, anions leave the cathode and cations leave the anode.
- When drawing a galvanic cell, it is usually safe to use nitrates like KNO_3 as it is soluble with everything.
- Occasionally, you may have an inert platinum electrode which means that the metal that's part of the redox reaction isn't solid (i.e aqueous or liquid).
- At other times, your galvanic cell might be a U-tube which doesn't need a salt bridge for an electrical circuit to form.

Precipitation

Solubility rules

Soluble		Insoluble	
Compounds	**Exceptions**	**Compounds**	**Exceptions**
Group 1 and NH4+ compounds		Carbonates	Except Group 1 and NH4+ compounds
Nitrates		Sulfites	
Acetates		Phosphates	
Chlorides, bromides and iodides	Except Ag+ and Pb2+	Hydroxides	Except Group 1, NH4+, Ba2+, Sr2+, Ca2+
Sulfates	Except Ag+, Pb2+, Ba2+, Sr2+, Ca2+	Oxides	
		Sulfides	Except Groups 1, 2, and NH_4^+

Precipitation reaction

A precipitation reaction occurs when a double displacement reaction produces a solid precipitate.

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow 2KNO_3 + PbI_2$$

In order to determine what the precipitate is, we use the **solubility rules.** The precipitate is insoluble in water.

Synthesis and decomposition

Synthesis

Synthesis reactions take the form:

$$A + B \rightarrow AB$$

Example:

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

Decomposition

Decomposition reactions take the form:

$$AB \rightarrow A + B$$

Example:

$$CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$$

Decomposition reactions occur when subject to energy:

- Heat
- Electricity
- Light

Displacement and replacement reactions

Replacement reactions

Replacement or single displacement reactions take the form:

$$A + BC \rightarrow AC + B$$

where A (metal) is more reactive than B (metal).

Example:

$$Mg_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + MgSO_4$$

Displacement reactions

Displacement or double displacement reactions take the form:

$$AB + CD \rightarrow AC + BD$$

Example:

$$Na_2SO_4 + BaCl_2 = BaSO_4 + 2NaCl$$

Combustion

Combustion is a process of **burning**, which occurs as an **exothermic self-sustaining chemical reaction**.

- Light and heat is emitted
- Completely new products are formed
- Processes cannot be easily reversed

Complete combustion

Complete combustion occurs when there is sufficient oxygen and a full reaction occurs, such that it produces carbon dioxide, water and a large amount of energy.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

Incomplete combustion

Complete combustion occurs when there is insufficient oxygen and a partial reaction occurs, such that it produces carbon dioxide, water and a large amount of energy.

$$4CH_{4(g)} + 5O_{2(g)} \rightarrow 2CO_{(g)} + 8H_2O_{(g)} + 2C_{(s)}$$

Example: The burning of fossil fuels may produce carbon monoxide or soot (carbon) from incomplete combustion.

Types of acid reactions

Neutralisation reactions

Acids react with bases in **neutralisation** reactions to form salts. It is a type of double displacement reaction.

$$acid + base \rightarrow salt + water$$

Example:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(s)} + H_2O_{(l)}$$

Toxins in food

Toxins in food can cause food poisoning in the form of vomiting or diarrhea. Indigenous Australians used chemical reactions to remove toxins from staple foods, such as black beans or cycads.

To remove toxins from black beans:

- Boil (breaking down cell membranes)
- Slice (increase surface area)
- Leach

To remove toxins from a cycad fruit:

- Roast (combustion) or boil (breaking down cell membranes)
- Crack open and remove kernel
- Leach (soak in water and dissolve toxins)
- Ground edible kernels
- Wash in water

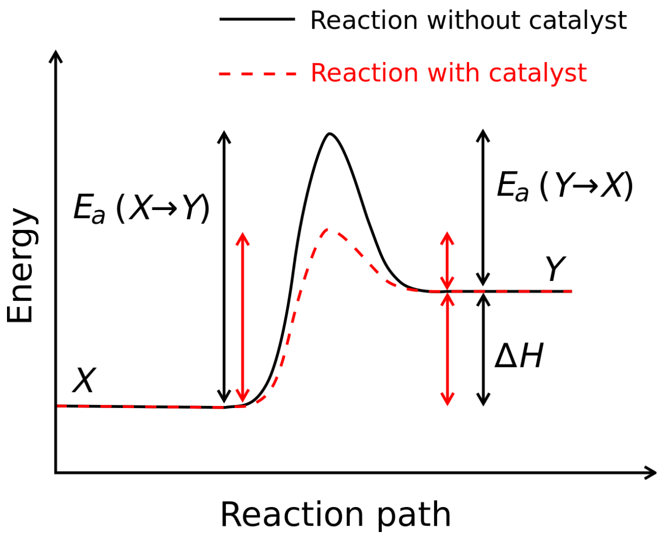
Fermentation can remove toxins by breaking down glucose.

$$C_6H_{12}O_{6(aq)} \rightarrow 2C_2H_5OH_{(aq)} + 2CO_{2(aq)}$$

Rate of reaction

Activation energy

Definition: Activation energy (E_A) is defined as the minimum amount of energy required for a reaction to occur.



Collision theory

For a reaction to occur, the reactant particles must meet all of these requirements:

- Make contact (i.e collide) with each other
- Collide with sufficient energy to break the bonds within the reactants (i.e energy equal to or greater than the **activation energy**)
- Have with the correct orientation to break bonds within the reactants

If a collision does not meet all of these requirements, no reaction occurs.

Reaction kinetics

The rate of reaction is the speed at which a reaction occurs.

- A log when combusting will have a slow rate of reaction
- Chopping up the log increases the surface area and smaller wood pieces will have a faster rate of reaction
- An explosion is a combustion reaction with a high rate of reaction

The rate of reaction is considered to be the gradient of the **change in concentration** of a reaction. Alternatively, it could be considered as a measure of product per unit of time.

$$average\ rate\ = \frac{c_2 - c_1}{t_2 - t_1} = \frac{\Delta c}{t}$$

where:

- **average rate** is the average rate of reaction
- **c** is concentration
- **t** is time

The rate of reaction tends to be logarithmic with t on the x-axis and rate of reaction on the y-axis.

$$as\ t \rightarrow \infty\ and\ c \rightarrow 0 \therefore r_t \rightarrow 0$$

Increasing rate of reaction

To increase the rate of reaction, you can increase the amount of successful collisions:

- Increase frequency of successful collisions
- Increase proportion of collisions which have sufficient energy

Example:

When is the reaction rate highest and lowest in a 1 minute time period?

The reaction rate is fastest when the <u>maximum number</u> of collisions occur. This occurs when the concentration is highest, at t = 10.

The rate is slowest when the minimum number of collisions is occurring. This occurs when the concentration is lowest at t = 60.

Factors affecting rate of reaction

Concentration of reactants	More reactants present = increased frequency of collisions = higher reaction rate
Nature and concentration of any catalyst present	Catalyst = less E_A = greater proportion of molecules which have sufficient energy = more successful collisions = higher reaction rate
Temperature	Higher temperature = more energy = greater proportion of molecules which has sufficient energy = more successful collisions = higher reaction rate
State of division (surface area) of a solid	Increased divisions = increased surface area = increased frequency of collisions = higher reaction rate
Rate of stirring	Increased rate of stirring = increased motion and kinetic energy = increased frequency of collisions = higher reaction rate

Note: The Maxwell-Boltzmann curve illustrates how with greater temperature, the distribution of kinetic energy for particles in a gas will be smoother, and there will be more energy in products past the activation energy.

Catalysts

Catalysts are substances which do not affect the overall reaction, but **reduce the activation energy** through weakening bonds in reactants, thereby increasing the rate of reaction.

Inhibitors

Inhibitors are substances which slow or stop a chemical reaction.

Example: Examples of inhibitors include penicillin and aspirin.

Drivers of Reactions

Module 4

Temperature

- Temperature is a measure of the degree of hotness or coldness of an object or substance
- It is a measure of the average kinetic energy of the particles in a system
- When temperature is uniform throughout two substances, **thermal equilibrium** has been achieved

Heat

- Heat is the total amount of energy of molecular motion in a substance
- **Specific heat capacity** (C) is the amount of heat needed to raise the temperature of a substance by 1°C

Remember: Total amount of heat energy is proportional to the mass of the substance involved.

Example: Water has a high specific heat capacity (4.18 kJ °C) because of the strong intermolecular bonds present in water — it has all three types of bonds up to hydrogen bonds. On a hot day, the temperature of water doesn't change much, compared to a cold day. Hexane on the other hand, would rise in temperature more than water given all other factors remain equal and constant, and has a lower specific heat capacity (2.26 kJ °C).

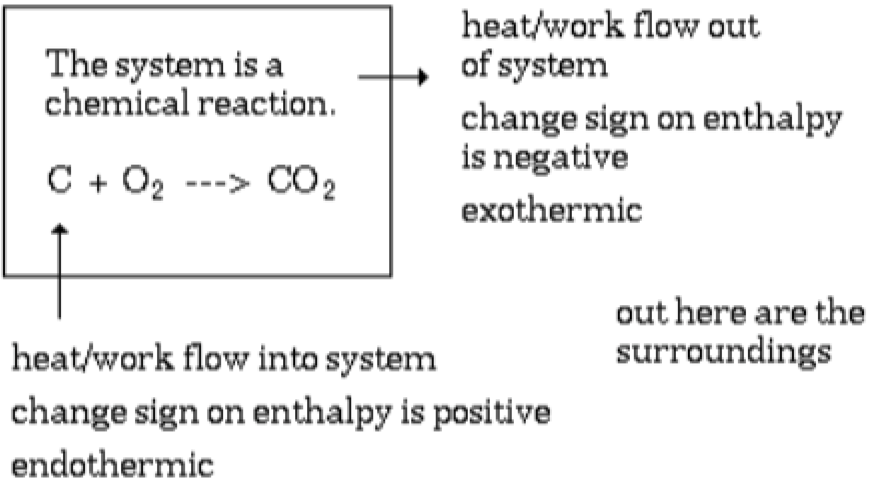
Heat (q)	Temperature (T)
Total energy of molecular motion in a substance — total energy multiplied by total number of atoms	**Average** energy of molecular motion in a substance — how fast atoms are moving Measure of the degree of hotness/coldness in a substance
Measured in Joules.	Measured in degree Celsius, Kelvin or *yuck* Fahrenheit.

Enthalpy

Enthalpy (H) is the heat content of a system. It is stored in the chemical bonds and intermolecular forces as potential energy,

- The **system** is the chemical reaction we focus our attention on
- The **surroundings** are everything else

Hence, the enthalpy change (ΔH) is the difference in enthalpy between reactants and products which can be observed.



+ΔH	**–ΔH**
When heat is added **from surroundings to system** enthalpy (H_{system}) **increases**	When heat is added from **systems to surroundings** enthalpy ($H_{surrounding}$) **decreases**
$H_{products} > H_{reactants}$	$H_{products} < H_{reactants}$
Endothermic	Exothermic

$$q = mC\Delta T$$

where:

- **q** is amount of heat energy (J)
- **m** is mass of substance (kg)
- **C** is specific heat capacity (J kg^{-1} K^{-1})
- **ΔT** is change in temperature

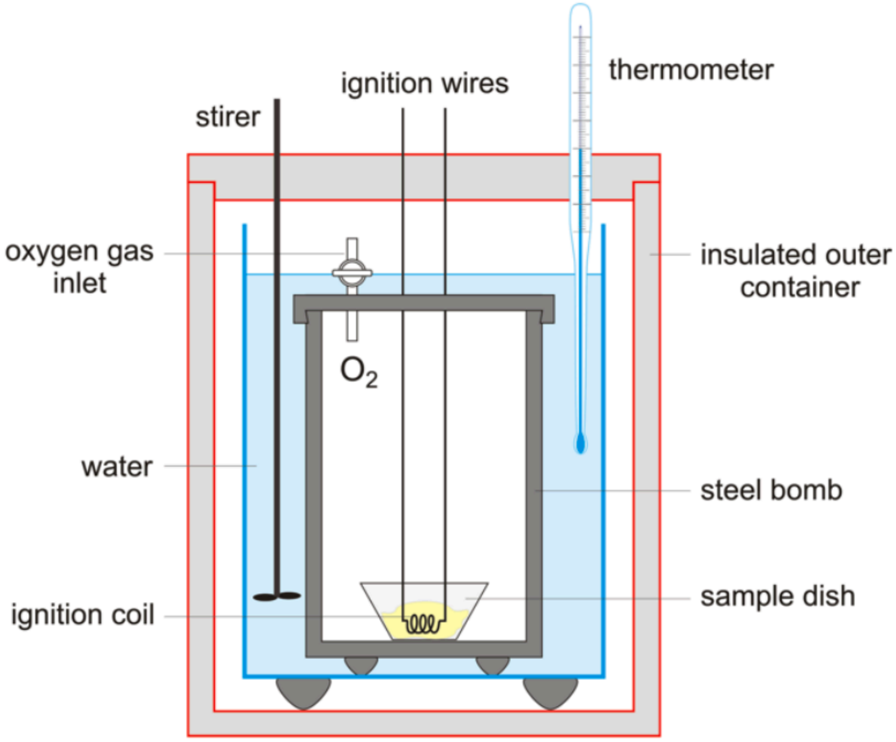
Bonding

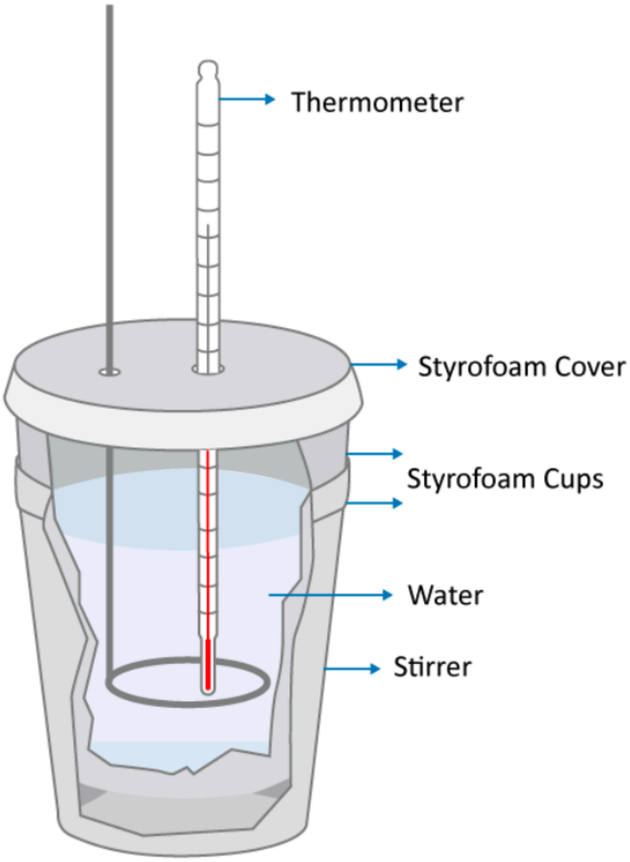
- When bonds **break**, energy is **put in**
- When bonds are **formed**, energy is **released**

Calorimetry

A **calorimeter** is an apparatus used to measure **heat changes** during a chemical reaction or change of state. Hence it can be used to demonstrate endothermic & exothermic reactions, through:

- Combustion
- Dissociation of ionic substances in aqueous solutions





Meme credit idea to Toby :)

The experimental enthalpy (kJ mol^{-1}) may be lower than the theoretical enthalpy value (kJ mol^{-1}) because of:

- Heat loss to the environment
- Imprecise measurement instruments

At a high school setting, the experiment could be improved by:

- Using two cups to increase insulation
- Using a temperature probe instead of a thermometer for greater accuracy

Enthalpy formula

$$\Delta H = \frac{q}{n}$$

where:

- **ΔH** is change in enthalpy (kJ mol^{-1})
- **q** is energy (kJ mol^{-1})
- **n** is number of moles

Remember: The sign of ΔH in an exothermic reaction must be negative.

$$\Delta H = \Sigma H_{products} - \Sigma H_{reactants}$$

Molar heat of solution

The **molar heat of solution** (ΔH_{soln}) of a substance is the heat absorbed when one mole of the substance dissolves in a large excess of water.

- For ΔH_{soln} being positive, the process is **endothermic** — heat absorbed, temperature falls
- For ΔH_{soln} being negative, the process is **exothermic** — heat released, temperature rises

Molar heat of combustion

The **molar heat of combustion** (ΔH_C) of a substance is the heat absorbed when one mole of the substance is completely combusted in sufficient to form water and carbon dioxide.

Since a combustion reaction is always exothermic, ΔH_C is always negative. Enthalpy change in boiling and melting

- Latent heat is energy absorbed or released as a substance undergoes a change of state.
- Latent heat values are a measure of the enthalpy required to melt/boil a substance at its melting/boiling point.

Example: For the vaporisation of water, the latent heat value is +40.7 kJ mol^{-1}. Hence, the amount of heat energy required to boil 5 moles of liquid water to steam is:

$$\Delta H \times n$$
$$= +40.7\ kJ\ mol^{-1} \times 5$$
$$= +203.5\ kJ\ mol^{-1}$$

Standard enthalpy of formations

The **standard enthalpy of formation** is the change in enthalpy when 1 mole of a compound is formed in standard states and conditions. Hence,

$$\Delta H^{\circ}_{reaction} = \Sigma\,\Delta H^{\circ}_{f}(products) - \Sigma\,\Delta H^{\circ}_{f}(reactants)$$

Enthalpy can be calculated given a chemical reaction and the standard enthalpies of formation:

Example: For a reaction between ammonia and hydrogen chloride:

$$NH_{3(g)} + HCl_{(g)} \rightarrow NH_4Cl_{(s)}$$

Given the following standard enthalpies of formation:

$$\tfrac{1}{2}N_{2(g)} + \tfrac{3}{2}H_{2(g)} \rightarrow NH_{3(g)} \therefore \Delta H^{\circ}_{f} = -45.9\ kJ\ mol^{-1}$$
$$\tfrac{1}{2}H_{2(g)} + \tfrac{1}{2}Cl_{2(g)} \rightarrow HCl_{(g)} \therefore \Delta H^{\circ}_{f} = -92.3\ kJ\ mol^{-1}$$
$$\tfrac{1}{2}N_{2(g)} + 2H_{2(g)} + \tfrac{1}{2}Cl_{2(g)} \rightarrow NH_4Cl_{(s)} \therefore \Delta H^{\circ}_{f} = -314.9\ kJ\ mol^{-1}$$

Therefore,

$$\Delta H^{\circ}_{reaction} = (-314.9) - [(-92.3) + (-45.9)]$$
$$\Delta H^{\circ}_{reaction} =- 176.7\ kJ\ mol^{-1}$$

Bond energy

Bond energy is the amount of energy required to break one mole of gaseous molecules into their individual atoms under conditions of 26°C and 1 atm of pressure.

- Bond **breaking** is always an **endothermic (+)** process
- Bond **forming** is always an **exothermic (–)** process

Hence

$$\Delta H = energy\ from\ forming\ bonds + energy\ from\ breaking\ bonds$$

Bond energies differ for different substances, as they may be affected by atomic radius, electronegativity, and the type of bond.

Example: For the enthalpy change per mole of O_2:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O(g)$$

This reaction has:

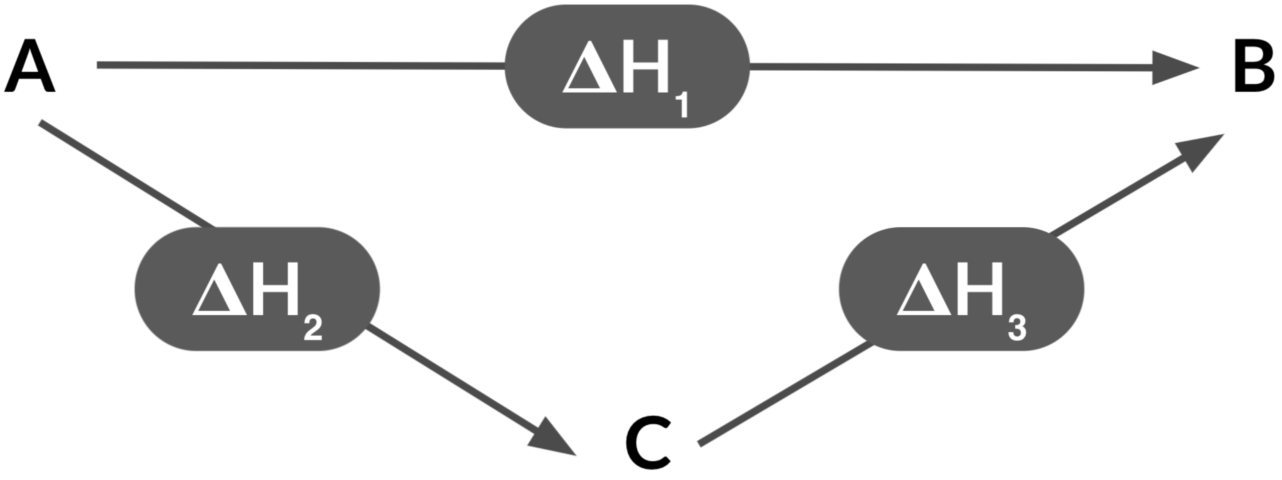
Bonds broken	2x H—H bonds	2(436) = 872
	1x O—O bonds	1(498) = 498
Bonds formed	4x O—H bonds	4(463) = 1852

Therefore,

$$\Delta H = (872 + 498) + [(-1852)]$$
$$\Delta H = -482\ kJ\ mol^{-1}$$

Hess's law

Given the following reaction, given **A—B** and **A—C—B** are two reactions with reactants **A** and products **B**:



$$\Delta H_1 = \Delta H_2 + \Delta H_3$$

Therefore, the amount of heat energy (ΔH) released or absorbed in a chemical reaction remains **constant** irrespective of the number of steps or kind of steps, provided the same products and reactants are involved.

Remember: The number of steps in the reaction pathway doesn't affect the overall enthalpy of the reaction.

Example: Reactions can be added just like algebraic equations. For example:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)} \quad \Delta H = -802\ kJ$$
$$2H_2O_{(g)} \rightarrow 2H_2O(l) \quad \Delta H = -88\ kJ$$

We can cancel out the $2H_2O_{(g)}$ in order to gain the final chemical reaction.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O(l) \quad \Delta H = -890\ kJ$$

Using Hess's law, we therefore can use those reaction enthalpy changes to find the overall enthalpy change for that reaction.

Entropy

Introducing entropy

Entropy (ΔS) is the **measure of the <u>number of possible arrangements</u> of components available within a system.** It is considered to be the measure of order/disorder within a system. It is measured in **$J\ mol^{-1}\ K^{-1}$**

$$\Delta S = \frac{q_{rev}}{T}$$

Note: You do **not** need to know the formula for entropy.

First law of thermodynamics

Let us first consider the first law of thermodynamics: **energy cannot be created or destroyed.**

Hence, there are two types of considerations we should make:

- **System** — object/reaction being studied
- **Surroundings** — everything outside of the system

Although the energy of the universe is constant, there are three types of systems:

- **Open** — can gain/lose mass and energy (e.g human body)
- **Closed** — can absorb/release energy but not mass (e.g coffee cup)
- **Isolated** — cannot exchange matter or energy (i.e adiabatic reaction)

Second law of thermodynamics

The second law of thermodynamics states that: **the entropy of the universe is increasing.**

$$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings} \geq 0$$

where:

- $\Delta S_{universe} = 0$for a **truly reversible** process
- $\Delta S_{universe} > 0$ for a **irreversible** process

Factors affecting entropy

Entropy increases with:

Factor	**Explanation**
Larger system	More possible arrangements with more space available
More moles	With more number of moles, there are more possible arrangements
Higher temperature	More kinetic energy with freer-moving components
Solid to liquid	Particles in a liquid move more freely
Liquid to gas	Gas particles can occupy all available space
Dissolved solid	Solute particles can be arranged in different ways within the solvent

More disorder = increased entropy

Hence

$$\Delta S = \Sigma\Delta S_{products} - \Sigma\Delta S_{reactants}$$

- when **ΔS < 0** ∴ entropy **decreases**
- when **ΔS > 0** ∴ entropy **increases**

Example: For the following reaction:

$$H_{2(g)} + I_{2(s)} \rightarrow 2HI_{(g)}$$

Given the following standard entropies:

$$S°(H_2) = \; + \; 130.6\, J\, mol^{-1}\, K^{-1}$$
$$S°(I_2) = \; + \; 116.7\, J\, mol^{-1}\, K^{-1}$$
$$S°(HI) = \; + \; 206.3\, J\, mol^{-1}\, K^{-1}$$

Hence by applying the formula:

$$\Delta S = \Sigma\Delta S_{products} - \Sigma\Delta S_{reactants}$$

Therefore the entropy for the reaction is given by:

$$\Delta S = 2(+\ 206.3) - [2(+\ 130.6) + 2(116.7)]$$
$$\Delta S = \ +\ 165.3\ J\ mol^{-1}\ K^{-1}$$

Spontaneity

Some spontaneous reactions are endothermic, such as the evaporation of water. Spontaneous reactions occur when:

- **ΔS = positive** such that increasing entropy is a driver of reactions
- **ΔH = negative** such that energy does not need to be constantly supplied

This is because:

- Most spontaneous reactions result in an **increase in entropy**
- Increasing entropy is a **driver of reactions**

Gibbs free energy

What happens when **ΔS** and **ΔH** are the same sign, such that one is favourable towards a spontaneous reaction, and one isn't favourable, as shown below?

- **ΔS = positive** (favourable) and **ΔH = positive** (unfavourable)
- **ΔS = negative** (unfavourable) and **ΔH = negative** (favourable)

In this situation, **temperature** becomes the defining factor. Hence **Gibbs free energy** reflects the *balance* between enthalpy and entropy, and is given by the equation:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

where:

- **ΔG** is Gibbs free energy
- **ΔH** is change in enthalpy
- **ΔS** is change in entropy
- **T** is temperature (K)

	ΔH < 0	**ΔH > 0**
ΔS > 0	Spontaneous at all temperatures ΔG < 0	Spontaneous at high temperatures when TΔS is large
ΔS < 0	Spontaneous at low temperatures when TΔS is small	Not spontaneous at all temperatures ΔG > 0

References and Related Links

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- **Periodicity,** BBC Bitesize (https://www.bbc.co.uk/bitesize/guides/zxc99j6/revision/1)
- **Modelling metals,** BBC Bitesize (https://www.bbc.co.uk/bitesize/guides/z9m6v9q/revision/1)
- **Intermolecular Forces,** UNSW School of Chemistry (https://www.chem.unsw.edu.au/coursenotes/CHEM1/nonunipass/hainesIMF/)
- **Molecule of the Month** (http://www.chm.bris.ac.uk/motm/motm.htm)
- **Year 11 Chemistry Depth Study** (https://docs.google.com/presentation/d/174Nier8TmE6O_hm6bY1_vKWes8I3myPnku74lbiHQck/edit?usp=sharing)

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- Miss D'Souza