

Year 11 Chemistry Notes - 2018

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- Module 1: Properties and Structure of Matter
- Module 2: Introduction to Quantitative Chemistry
- Module 3: Reactive Chemistry
- Module 4: Drivers of Reactions

Module 1: Properties and Structure of Matter

Mixtures

A mixture is a combination of two or more pure substances (containing one type of molecule) in which each pure substance retains its individual chemical properties - the mixture itself is impure. There are two kinds of mixtures: heterogeneous and homogeneous:

Heterogeneous Mixtures

Two or more substances intermingle, but **remain physically separate**. Often it is possible to separate the original ingredients by **simple physical means**, such as filtering, centrifuge, decanting or sedimentation.

- Examples include: Dirt+Sand, Oil+Water, Salt+Baking Soda
- A Suspension is a specific type of heterogeneous mixture where particles settle at the bottom

Homogeneous Mixtures

Two or more substances have merged into a **uniform phase**. There are no borders between the substances, but they are not chemically bonded. The **physical properties** of each ingredient can be exploited to separate them.

- Examples include: Saltwater, Copper Sulfate solution
- Saltwater can be distilled (boiled) to separate the water

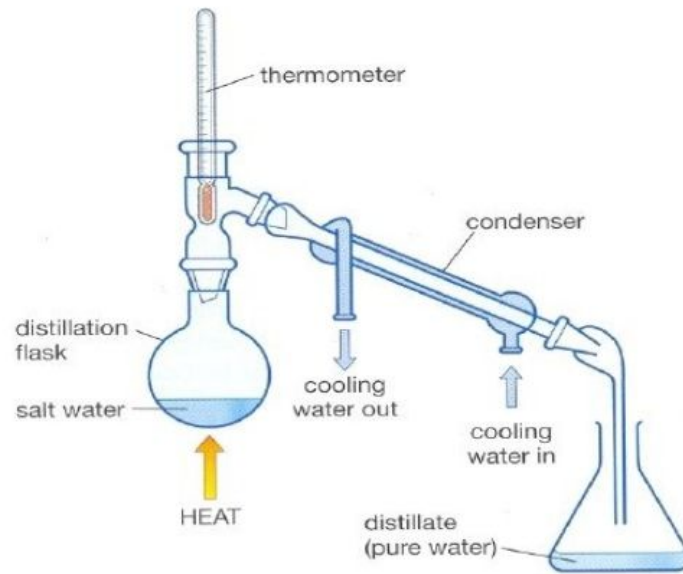
- The two types include solutions and colloids (particles are present, but are very small and do not settle)

Physical properties include magnetism, solubility, density, boiling point, melting point, particle size.

Methods of Separation

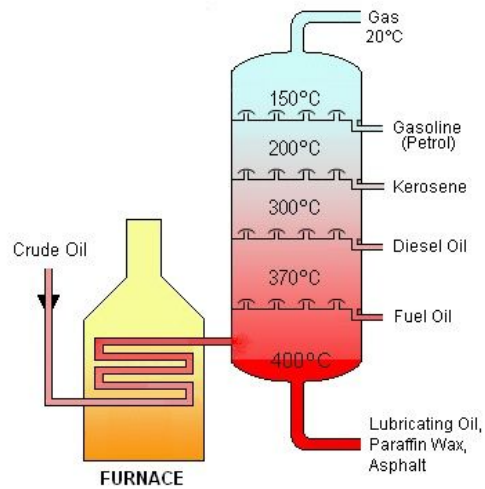
Separation Method	Property used	Example
Filtration	Solid vs Liquid (or much smaller particle sizes in solution)	Solid Impurities in a solution can be separated through filtration and will be left as residue . The liquid that has been filtered is called the filtrate
Sedimentation/ Decantation	Different Densities of solid vs liquid (gravity brings the solid to the bottom)	Grains of sand in water can undergo sedimentation and be decanted out of a beaker
Distillation	Different boiling/condensation points (Separates miscible liquids or ions in a solution by boiling, condensing and collecting)	Distilling saltwater to get salt and water
Fractional Distillation	Very small range of boiling/condensation points A tower is used to distill and separate different oils	Fractionally distilling crude oil
Evaporation + Crystallisation	Different Boiling points and solubility (Ions in solution precipitate out)	Evaporating saltwater without keeping the water
Centrifuging	Different Densities (Centripetal force brings densest component to the bottom)	Centrifuging blood to separate the red blood cells
Sieving	Different Particle sizes of solids (Passed through a material with many holes)	Sieving pebbles and sand
Magnetic Separation	Magnetic properties (Magnet pulls out the magnetic substances)	Separating iron filings out of dirt
Chromatography	Different solubilities of solute	
Froth Flotation		
Separating Funnel	Used for separating immiscible liquids (liquids that don't form a homogenous mixture)	Separating Water and Oil

Distillation Diagram



Fractional Distillation

The crude oil is placed at the bottom of a tall tower. As different hydrocarbons are evaporated, they rise. The higher up they go, the cooler they become. Thus, the one that is vaporized first condenses at the top, the one that is vaporized next condenses at the next level, and so forth.



Percentage Composition

The proportions of each component in a mixture are represented as percentages.

$$\text{Percent by mass} = \frac{\text{mass of component}}{\text{total mass}} \times 100$$

E.g. In a mixture of 12g CaCO_3 and 3g NaCl

$$\text{CaCO}_3 \% = \frac{12}{15} \times 100 = 80\%$$

$$\text{NaCl \%} = \frac{3}{15} \times 100 = 20\%$$

The Periodic Table of Elements

<https://ptable.com>

molybdenum	← element name
42	← atomic number number of protons (Z)
Mo	← atomic symbol
95.94	← atomic mass A (this is an average mass)

The **periodic table** is an ordered compilation of all known elements.

Elements vs Compounds

- Elements are pure substances that cannot be chemically or physically decomposed.
- Compounds are pure substances that are chemical combinations of two or more different elements - they can be decomposed.

Periods

The rows of the periodic table. They increase in atomic number from left-right, and **each period corresponds to the number of electron shells** of the elements in that period.

Groups

The columns of the periodic table. Elements in the **same group** share similar chemical properties, as they have the **same number of valence electrons**.

- For example, Group 1 or 7 elements have only one valence electron, so are highly reactive. Group 8 elements have a full shell already, so are highly unreactive as they are already stable.

Metals, Metalloids and Nonmetals

Uneven chunks of the periodic table that share similar physical properties:

- **Metals** are good conductors of heat and electricity, are malleable and ductile, usually have a silvery shine and are usually solid at room temperature.
- **Nonmetals** are (usually) good insulators of heat and electricity, are brittle; usually dull. Many of the elemental nonmetals are gases at room temperature, while others are liquids and others are solids.
- **Metalloids** have properties of both metals and nonmetals, and can be made to conduct electricity in some circumstances.

Families

Families are named columns (groups) that share even more specific chemical properties.

- **Alkali Metals** - Group 1, with all elements having one valence electron.
 - Physical properties: soft (can be cut with knife), lustrous metallic solids, low densities, high thermal and electrical conductivity, relatively low melting point
 - Chemical properties: highly reactive, vigorous exothermic reaction with water and oxygen, present naturally as salts
- **Alkaline Earth Metals** - Group 2, with all elements having two valence electrons.
 - Physical properties: lustrous metallic solids, high thermal and electrical conductivity, more dense, higher melting points and harder than alkali metals
 - Chemical properties: reactive, oxidise easily, exothermic reaction with water
- **Transition Metals** - Groups 3-12, with elements having varying valencies.
 - Physical properties: white, hard, lustrous, dense metallic solids, high thermal and electrical conductivity, high melting points
 - Chemical properties: less reactive than alkali metals, but chemical properties otherwise vary
- **Halogens** - Group 17, with all elements having 7 valence electrons
 - Physical properties: nonmetals, melting and boiling points increase going down the column, halogens change state going down the column (i.e. Fluorine/Chlorine are gas, Bromine is a liquid, Iodine is a solid), poor thermal and electrical conductivity, unpleasant odours, very toxic
 - Chemical properties: highly reactive, form ions with -1 charge, form diatomic molecules
- **Noble Gases** - Group 18, with all elements having full valencies (8 valence electrons)
 - Physical properties: gases, low boiling points, low densities

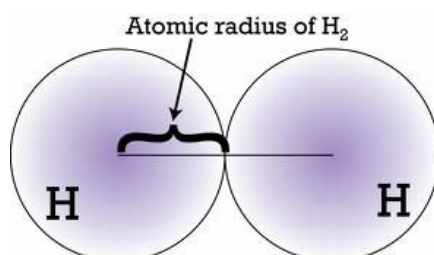
- Chemical properties: highly unreactive, mostly present as monatomic gases, very rarely (usually never) form compounds

Periodicity (Periodic Table Trends)

The three main periodic properties are: Atomic Radius, Ionisation Energy and Electronegativity

Atomic Radius - Half the distance between the centers of two atoms of an element that are touching

- Going **left → right** across a period, atoms have more protons but the same amount of electron shells. Thus, Electrons are attracted to the nucleus more strongly, and the **atomic radius decreases**
- Going **up → down** the group, atoms have more electron shells, which not only put the valence electrons further away, but the inner electrons also repel (or shield) the valence electrons from the nucleus's attraction, so the **atomic radius increases**
- Cations generally have a smaller ionic radius than the neutral atom, and Anions have a larger atomic radius. This is because ions have a different ratio of protons to electrons, so the radius gets bigger or smaller depending on the electrostatic attraction



Ionisation Energy - The energy required to remove one valence electron from a gaseous atom.

- The more strongly bound to the nucleus electrons are, the more ionisation energy is required to remove them
- Smaller atomic radii mean stronger bound electrons, so **ionisation energy increases as atomic radius decreases**

- A low first ionisation energy indicates that an element is a metal, while a high first ionisation energy indicates that it is a nonmetal
- 1st ionisation energy is the energy required to remove the first electron, while 2nd ionisation energy is the the energy needed to remove the second one, etc
 - **Subsequent ionisation energies get higher**, because after removing electrons, the ratio of protons to electron becomes skewed to the protons side, and the electrostatic force between them becomes stronger.
 - If there is a full shell after taking out an electron, it requires exponentially more energy to remove the next one from the full shell (because full shells are stable)
- 3s valence orbital has a higher ionisation energy than 3p orbital

Ionisation equations can be represented like so:



Electronegativity - The measure of the ability of an atom to attract electrons for chemical bonding (measured in *Pauling units*)

- When an atom has a smaller atomic radius, it's valence electrons are closer to the nucleus, and the atom can easily pull external electrons into it. Thus, **as atomic radius decreases, electronegativity increases**
- A high electronegativity difference between atoms indicates a more ionic bond, while a low electronegativity difference indicates a more covalent bond.
- Fluorine is the most electronegative element

Metallic Character - How close an element is to typical metallic properties - The metallic character of an element is proportional to its ability to lose electrons (i.e. if an element has 1, 2 or 3 valence electrons, it is more metallic than 4, 5, 6, 7 or 8 valence electrons)

NOTE: Atomic radius affects all the other properties - i.e. it's easier for an atom with a greater atomic radius to let go of an electron, because it's valence shell is further away from the nucleus (so greater ionisation energy)

Periodic Trends

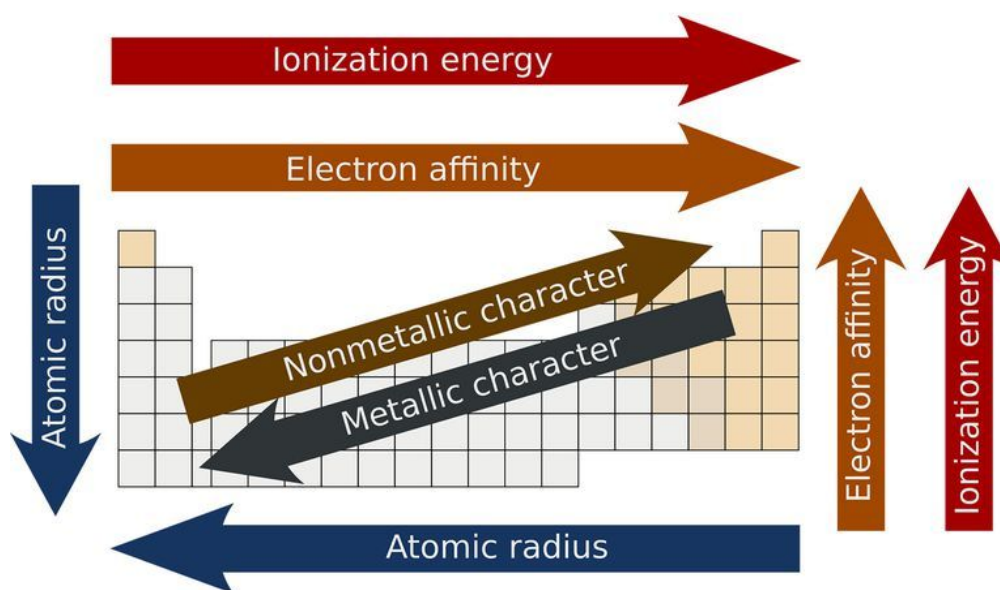
These properties change moving through the periods (left-right) and groups (up-down):

Moving Left → Right (Periods):

- Ionization Energy Increases
- Electronegativity Increases
- Atomic Radius Decreases

Moving Up → Down (Groups):

- Ionization Energy Decreases
- Electronegativity Decreases
- Atomic Radius Increases



Isotopes

While the number of protons defines an element, the number of neutrons indicates the Isotope (different versions) of the element - e.g. a Hydrogen atom can have 0, 1 or 2 neutrons, but it is still hydrogen.

Isotope Stability

The nucleus is held together by a binding energy, and so the **ratio of protons to neutrons affects the stability of an isotope.**

- **Stable Isotopes** have sufficient binding energy to keep the nucleus together. They do not undergo radioactive decay

- **Unstable Isotopes** have an imbalance of neutrons - the binding energy can't hold the nucleus together properly. To become stable, they undergo radioactive decay - and so are also known as radioisotopes

Isotope Trends

- Isotopes with atomic number > 82 are all unstable.
- Isotopes with atomic number < 20 and a 1:1 proton-neutron ratio are much more likely to be stable
- All elements with atomic numbers < 82 have one or more stable isotopes, except for technetium and promethium
- Atoms with odd numbers of protons and neutrons in the nucleus are more likely to be unstable
- Atoms with an even number of protons and neutrons are more likely to be stable

Isotope Notation

When Isotopes are written as words, the name of the element is given, with a hyphen and number indicating the mass number: For example, helium-3 or carbon-14.

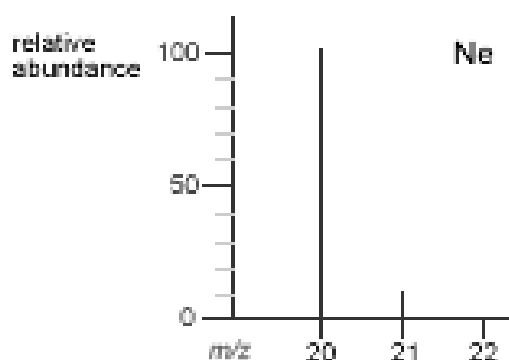
When written as symbols, the chemical symbol is given, with a superscript (mass number) on the upper left, and a subscript (atomic number) on the bottom left. For example, ${}^3_2\text{He}$, or ${}^{14}_6\text{C}$

Relative Atomic Mass

The naturally occurring form of an element is usually a mixture of all its isotopes. The relative atomic masses as given in the periodic table are decimals because they are an average of all the isotopes of that element - dependent on the how common each isotope is.

A **mass spectrometer** is a device that uses electromagnetic fields to sort the isotopes present in a substance by atomic mass, which then allows us to see how abundant each isotope is.

For example, for neon:



The abundance of ^{20}Ne is much greater than ^{21}Ne , in a ratio of 10:1 - so naturally occurring neon is ~90% ^{20}Ne and ~10% ^{21}Ne .

To calculate the relative atomic mass from isotopic composition, multiply the percentages with the atomic mass of each isotope:

$$0.9 \times 20 + 0.1 \times 21 = 20.1 \text{ amu}$$

(amu = atomic mass units)

1 amu = mass of Carbon-12 divided by 12

Radiation

When an atom undergoes radioactive decay, it is basically breaking apart and releasing energy as particles or waves.

- Neutrons prevent the protons in the nucleus from repelling and breaking away. Atoms decay because the forces holding the nucleus together sometimes aren't strong enough to hold together large nuclei - this occurs when the optimal ratio between protons and neutrons deviate.
- Unstable isotopes/ Radioisotopes undergo radioactive decay. All elements greater than atomic number 92 (Uranium) undergo radioactive decay - these are known as transuranium elements.
- Many transuranium elements are artificially synthesised, such as technetium-95 and promethium-146

The **half life** of a substance is a measure of the time it takes for half the atoms in that substance to decay. Half-lives can range from seconds to billions of years, and can be represented as a logarithmic graph.

The three main types of radiation are Alpha, Beta and Gamma radiation:

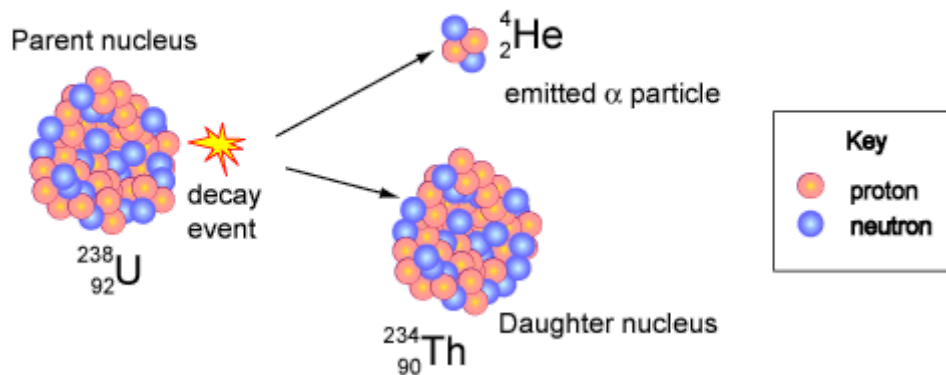
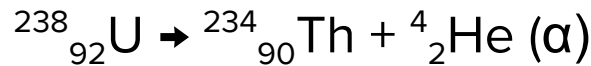
Alpha Decay (Too Much Mass)

Alpha decay (α) occurs when an atom emits an alpha particle - which is made of two protons and two neutrons joined together (a Helium nucleus). This alpha particle is ejected out of the nucleus of the atom.

- Since the number of protons changes, alpha decay causes the atomic mass and element of the atom to change. For example, uranium-238 transforms into thorium-234.

- After Alpha decay, the atomic number decreases by 2, and the mass number decreases by 4

For example, when Uranium-238 undergoes alpha decay:

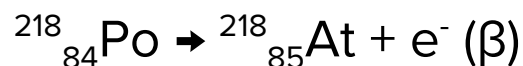


Beta Decay (Too Many Neutrons)

Beta Decay (β) occurs when an atom emits a beta particle - which is either an electron or positron (this is known as positron emission). For normal beta decay, an electron is ejected from the nucleus after a neutron splits up into an extra proton and electron.

- Since the neutron turns into a proton, the mass number stays the same, but the atomic number changes. Therefore, beta decay causes an atom to change into another element with the same atomic mass.
- After Beta decay, the atomic number increases by 1, and the mass number stays the same
- Beta Decay also emits a neutrino, but this is negligible

For example, when Polonium-218 undergoes beta decay:



Positron Emission occurs when a proton splits into a neutron and positron (basically a positive electron). It occurs in isotopes that have **too many protons**.

The atomic number decreases by 1, the mass number stays the same.

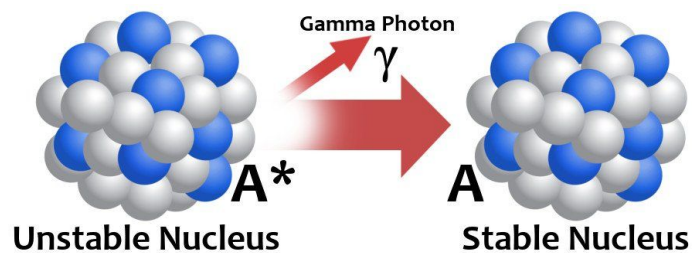
Electron Capture occurs when a proton captures an electron, and becomes a neutron. The atomic number decreases by 1, and the mass number stays the same.

- A neutrino is emitted for both of the above, but this is negligible.

Gamma Radiation

Gamma Radiation (γ) occurs when an atom emits gamma rays. It usually occurs after alpha or beta decay, where the nucleus is still excited after decaying. The excited nucleus then releases gamma ray photons to become more stable.

- Since no protons or neutrons are removed/added, the element, atomic number and atomic mass stay the same
- Gamma Radiation technically isn't a type of decay because only energy is released



Properties of Alpha, Beta and Gamma Radiation

Characteristic	Alpha rays	Beta rays	Gamma rays
Nature	Helium nucleus	Electrons	Photons
Penetrative power	Few centimetres in air	Few millimetres of aluminium	Many centimetres of lead
Charge	+ 2 e	-e	Zero
Mass	6.64×10^{-27} Kg	9.1×10^{-31} Kg	Zero
Detection	Affects photographic plates Affected by electromagnetic fields.	Affects photographic plates. Affected by electromagnetic fields	Affects photographic plates. Not affected by electromagnetic fields.

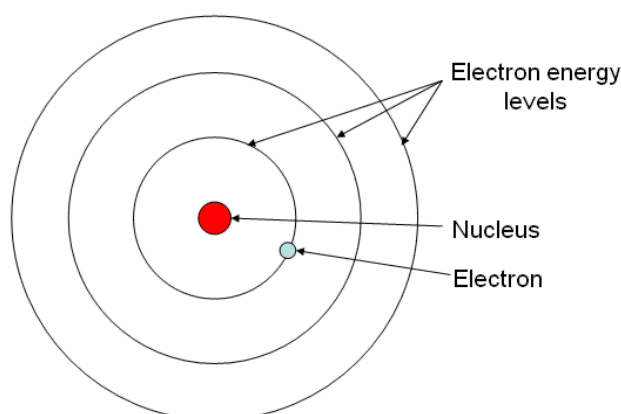
The Atomic Models

Dalton, Thompson and Rutherford Atomic Models

- **Dalton's Model** - atoms were indivisible, solid spheres
- **Thompsons Model** - after performing the Cathode Ray experiment- discovered electrons, and developed the plum-pudding model, where electrons were embedded in a positive solid sphere
- **Rutherford Model** - after the Gold Foil experiment, he determined that atoms were mostly empty space with a positive centre, with electrons floating around the centre
- **Bohr Model** - expanded on Rutherford's, but defined fixed energy levels

Bohr's Model (Planetary Model)

Niels Bohr developed a new model of the atom in 1913 called the “**planetary model**”



His model had four main points (postulates):

1. Electrons are particles that occupy **fixed orbits** around the nucleus - called stationary orbits
2. Each orbit has an **energy level** associated with it
3. Energy is absorbed when an electron jumps from a lower orbit to a higher one and energy is emitted when an electron falls from a higher to a lower orbit
 - The energy and frequency of light emitted can be calculated using the difference between the two orbital energy levels
4. Electrons cannot exist between the energy levels/orbits

The flame test supports Bohr's model - the different lights being emitted from the flame test indicated that energy was emitted when electrons returned to their ground shells, and the single spectral lines indicated that there were set routes/shells for the electrons to follow.

However, Bohr's model had limitations:

- It could only be applied to atoms with one valence electron
- It didn't explain the different intensities of lines/colours in a hydrogen spectrum

- It didn't explain the Zeeman effect (idek)

Schrodinger's Model (Quantum Mechanical Model)

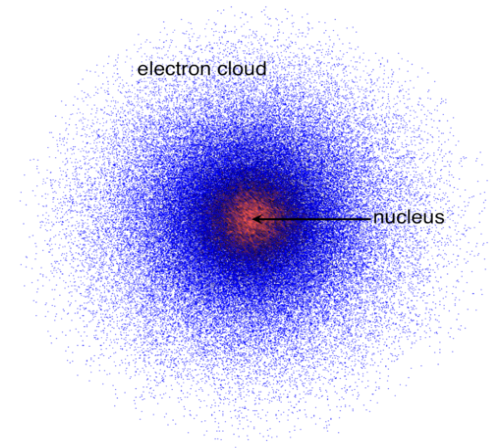
Bohr believed that electrons orbited the nucleus in a **fixed radius** from the nucleus

Schrodinger proposed that they moved around a lot more irregularly, and **didn't completely follow strict paths** - this explained the issues of Bohr's model.

- Schrodinger's model was based on the wave-particle duality of electrons.

In 1926, Schrodinger developed the quantum mechanical model of the atom, which **predicts the probability of an electron being at a certain location around the atom instead of defining a set radius** - basically, proposing that electrons do not travel in fixed radii, but in a particular area/tubes around the atom.

It was modeled as a nucleus surrounded by an "electron cloud". Electrons are most likely to be found where the cloud is most dense.



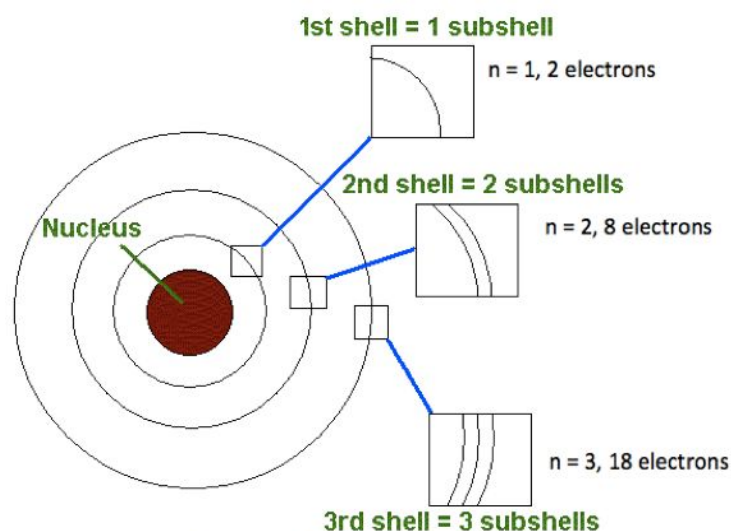
Energy Levels and Electron Configuration

Electrons do not orbit the nucleus in fixed orbits, they move around in energy levels called **orbitals - which are regions around the atom where the electron is very likely to be found at any given time.**

Aufbau Principle

The Aufbau Principle states that in order to configure electrons in an atom, electrons are added to the lowest energy level until it is filled. Then electrons are added to the next energy level until that is filled, etc.

Electron Subshells



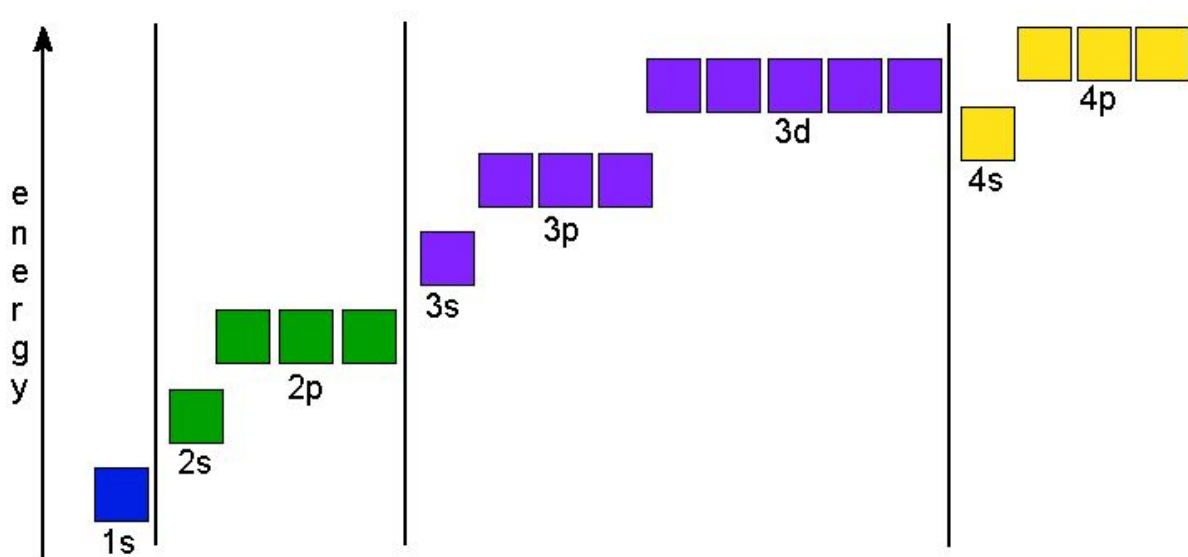
Each electron shell is actually composed of various subshells with different capacities - just like a road is composed of different sized lanes.

- **Each subshell is a group of atomic orbitals** - the orbitals that make up each subshell have distinct shapes, and each orbital can hold 2 electrons

Orbitals come in groups - There is always one s orbital, three p orbitals, five d orbitals and so on.

- **I.e. even though the d subshell in the 3rd principal shell is called 3d, there are Five d orbitals in it (Hella confusing ik).**

Principal Shell → Subshells → Orbitals



The principal shells are named in increasing order of energy as 1,2,3, etc, while the subshells are named s, p, d, f, g. Each subshell can hold a maximum amount of electrons, depending on the orbitals that it contains.

Subshell	No. and type of Orbitals	Electron capacity
s	One s orbital	2
p	Three p orbitals	6
d	Five d orbitals	10
f	Seven f orbitals	14
g	Nine g orbitals	18

The number of subshells in a shell is equivalent to its number - for example, principal shell 1 has one subshell (1s), while principal shell 3 has three subshells (3s, 3p, 3d).

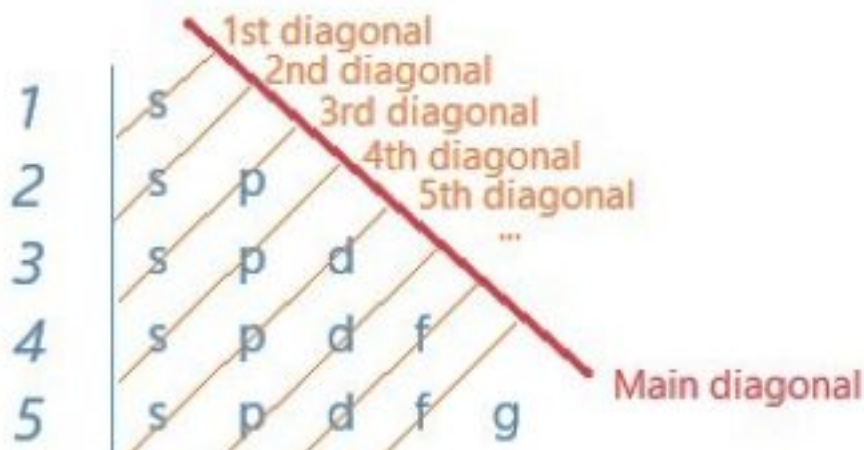
- 3s can hold 2 electrons, 3p can hold 6, and 3d can hold 10, so shell 3 can hold a total of 18 electrons.

To clarify, the 3s subshell has one s orbital, which is why it can only hold 2 electrons, while the 3p subshell has three p orbitals, which is why it can hold 6 electrons.

Each principal shell has a capacity of $2n^2$ electrons.

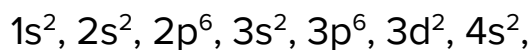
Madelung Rule

Electron configuration follows a specific pattern, known as the Madelung Rule - which states that electrons will always go for the lowest energy subshell, even if it is inside a higher shell:



- 4th diagonal: 3p (6 electrons), 4s (2 electrons) → 2 electrons left
- 5th diagonal: 3d (can hold 10 electrons, but add 2) → 0 electrons left

Once calculated, the configuration should be written as:

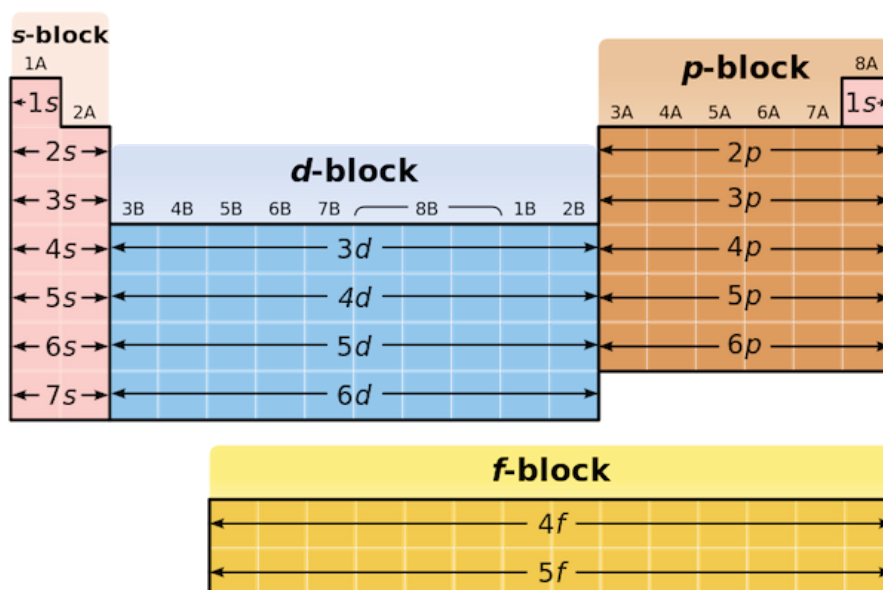


with a superscript to indicate the number of electrons in each subshell.

By adding the electrons in the subshells to give the number of electrons in each main shell, the electron configuration for titanium is: **2, 8, 10, 2**

Blocks of the Periodic Table

The periodic table can also be divided into four blocks - s-block, p-block, d-block and f-block, corresponding to the orbitals. I.e. s-block elements all have their valence electrons in the s-orbital, p-block elements all have their valence electrons in the p-orbital, etc.



Pauli Exclusion Principle

Hund's Rule

Flame Test

The flame test is an analytical experiment used to detect the presence of particular metals in a sample. When heated, the electrons in a metal atom become excited, and move to higher energy levels. **As they return to their normal energy levels, they release the excess energy as photons (light).**

The colour of the light (frequency of photon) depends on the energy gap between the energy levels - thus, the colour of the flame indicates which metal it is, as different metals have different energy gaps between energy levels:

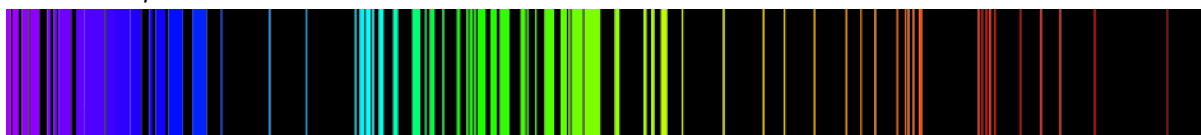
This can be summarised in the formula $E = hf$ (Energy = planck's constant \times frequency of light)

1. Dip a paddle-pop stick into a metal ion and solution
2. Hold the stick at the edge of a bunsen burner flame
3. Observe the colour of the flame, and note what metal it is
4. Repeat steps 1-4 with the other samples

Chlorides have a weaker energy of the lattice structure than Oxides - all the metals are bonded with chlorine because it is easier to break apart the ions and thus see the flame colour more clearly, so hydrochloric acid solution is often used.

A device called a **mass spectrometer** splits the colours into much more specific orders, displaying these colours in an **emission spectrum** - which is unique to each element, like its fingerprint. This enables more precise identification of elements.

Emission Spectrum of Iron:



Bonding

When an atom has too little/too many electrons in the valence shell, it wants to remove/gain electrons to make it stable (full valence shell). Thus, when it makes contact with another element that has the amount of electrons that it needs, they bond together.

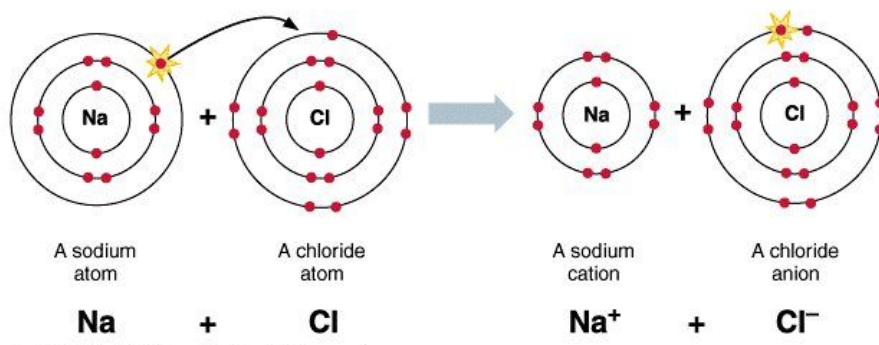
This concept is related to **Electronegativity**, and is the main mechanism behind bonding.

Ionic Bonds (Metal + Nonmetal)

Ionic Bonds involve a Metal donating electrons to a Nonmetal atom

- The **metal** atom becomes **positively charged (cation)** after losing electrons
- The **nonmetal** atom becomes **negatively charged (anion)** after gaining electrons
- The oppositely-charged atoms (cation (+) and anion (-)) attract to create an ionic compound

For example, when Sodium bonds with Chlorine - Sodium donates its one valence electron to Chlorine, making Sodium positive (cation), and Chlorine negative (anion). They attract and bond to form Sodium Chloride.



Polyatomic Ions

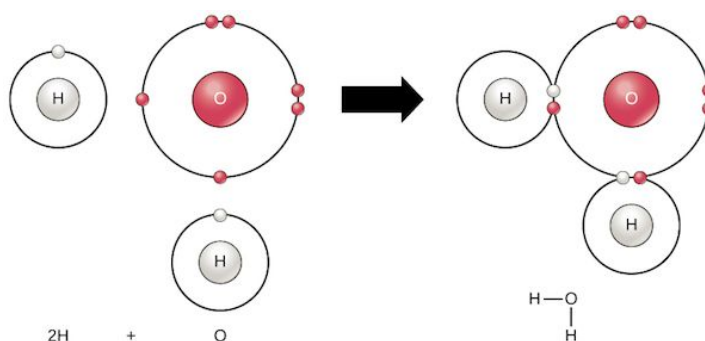
Polyatomic ions are molecules that are composed of bonded atoms, but still have an overall charge. Examples include Nitrate, (NO_3^-), Nitrite (NO_2^-) and Ammonium (NH_4^+).

Covalent Bonds (Non-metal + Non-metal)

Covalent Bonds involve a Nonmetal sharing electrons with another Nonmetal atom

- The atom with excess valence electrons shares them with the atom that needs them
- They bond as these electrons are shared to form a covalent or molecular compound

For example, when Hydrogen bonds with Oxygen, Two hydrogen atoms each share an electron with an Oxygen atom and form Water (H_2O).



Polarity of Covalent Bonds

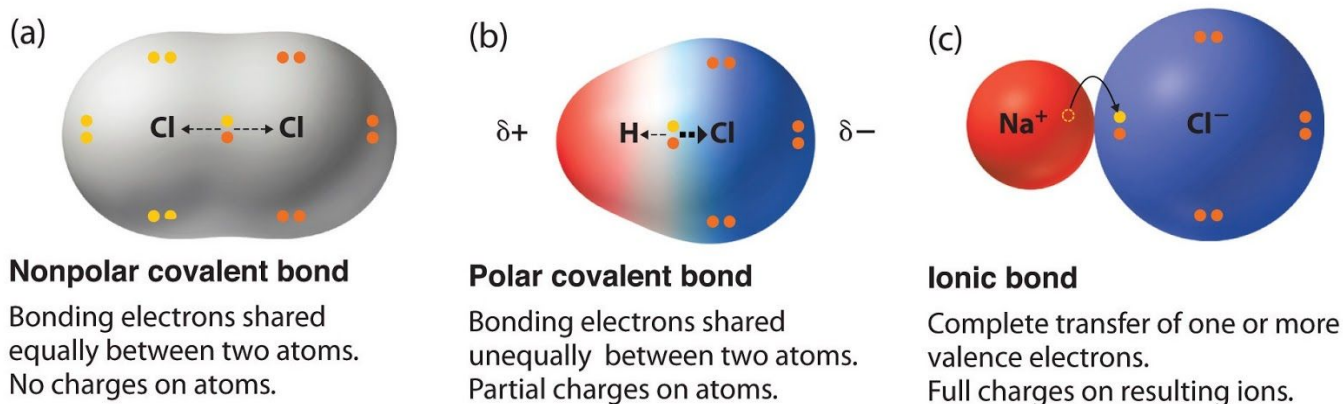
Although electrons are shared in covalent bonds, they are **not always shared equally**. This occurs when one atom has **higher electronegativity (i.e. wants the electrons more)** than the other:

In **Non-Polar Covalent Bonds**, Electrons are **shared equally** between the atoms, because the electronegativity of the atoms is the same or very similar (<0.5 Pauling units).

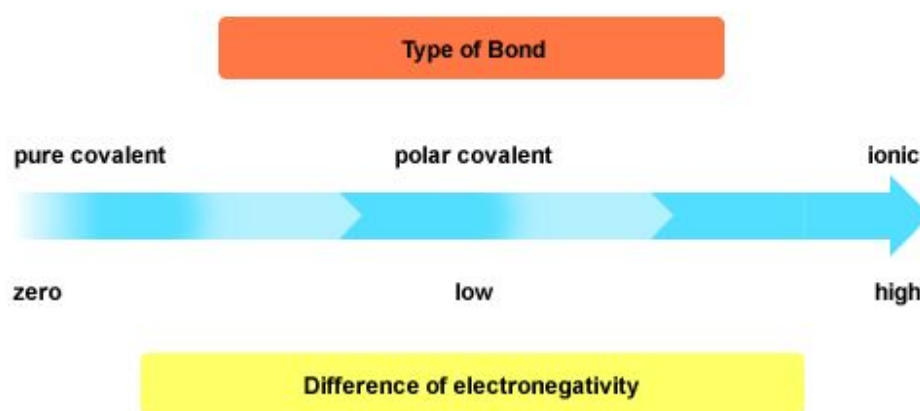
- Nonpolar bonds often occur between atoms that are the same, such as two chlorine atoms (Cl_2)

In **Polar Covalent Bonds**, Electrons are **NOT shared equally** between the atoms, because the electronegativity of the atoms is different.

- The atom with higher electronegativity (that is tugging harder on the electrons) becomes slightly negatively charged (δ^-), while the atom with lower electronegativity becomes slightly positively charged (δ^+). **This makes the bond partially ionic.**
- Polar covalent bonds are the most common, and usually occur between different atoms - such as in H_2O , where the Oxygen atom has higher electronegativity.



This indicates that **covalent and ionic bonds don't have a clear cut difference**. Nonpolar, Polar and Ionic bonds can in fact be represented on a spectrum:



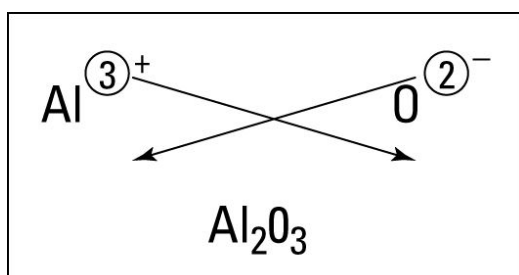
The higher the electronegativity difference, the more ionic the bond is.

- The electronegativity of an element can generally be predicted by its position in the periodic table (periodicity) - so two elements such as Hydrogen and Bromine which are located far from each other would make a very polar covalent bond.

Criss-Cross Method

To easily determine the final product of an ionic bond (or some polar covalent bonds), you can find the two reactants and their oxidation numbers, then swap the numbers and make them the subscripts for the product.

E.g. for Aluminium and Oxygen, the chemical formula for aluminium oxide is given by:



This ensures that the ratios of ions are correct - you need two Aluminium atoms for every three Oxygen atoms to make aluminium oxide (total charges of +6 and -6 cancel out)

If you have something like Sn_2O_4 after crisscrossing, you can simplify that to most basic ratio (1 Sn: 2 O) using it's highest common factor and write it as SnO_2

However, you can criss cross but not simplify polyatomic ions - e.g. Na_2SO_4 can't be simplified

Oxidation Numbers

The **oxidation number** is the charge an atom would have if it were an ion (charged atom). To clarify:

Atoms want to remove electrons when they have 1, 2, or 3 valence electrons, and want to gain electrons when they have 4, 5, 6 or 7 valence electrons.

- When an atom wants to get rid of electrons, it has a positive oxidation number. For example, Sodium has one valence electron, so it has an oxidation number of +1. Aluminium has 3 valence electrons, so has an oxidation number of +3.
- When an atom wants to gain electrons, it has a negative oxidation number. For example, Fluorine wants to gain one electron (it has 7 valence electrons), so has an oxidation number of -1.

The **amount of valence electrons (oxidation number) dictates the ratio of atoms in a bond**. For example - with Potassium +1 and Oxygen -2, two atoms of Potassium are needed for every one oxygen atom. Thus, the chemical formula for Potassium Oxide is K_2O

Lewis Dot Diagrams

Lewis Dot Diagrams are a form of shorthand notation for covalent compounds, showing their valence electrons and bonds.

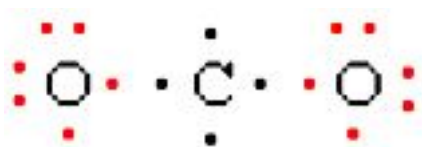
- Atoms are represented by the element symbol
- Valence Electrons are represented as dots and are placed in pairs around the atom
- Covalent bonds with other atoms are represented as lines

For example, To draw a Lewis dot diagram for CO_2 :

1. Find the number of valence electrons of each element in the compound.
 - In this case, Carbon has 4 valence electrons, and Oxygen has 6
2. Place the atom with the lowest electronegativity in the centre, then place the other atoms around it.
 - In this case, carbon is in the centre as it is the least electronegative

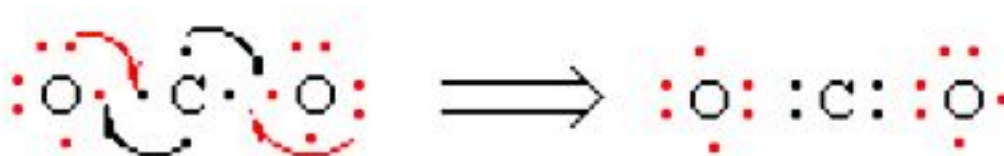


3. Arrange the electrons so that each atom contributes one electron to a bond (a single bond is made of two electrons from separate atoms)



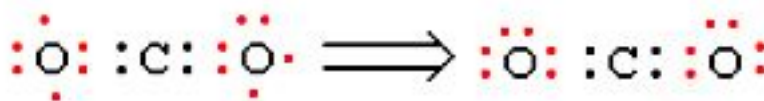
If all the valence shells are now stable (i.e. have 8 electrons), then the Lewis diagram is done.

- However, in this case, Oxygen has 7 valence electrons (six + the one carbon shared) each while carbon has 6 (4 + the two both oxygens shared)
4. If the valence shells are incomplete, then add another electron to make a double bond



Step 4 should be repeated until the valence shells are stable. In this case, the valence shells of oxygen and carbon both have 8 electrons, so it is stable.

5. Arrange the electrons into pairs



There are exceptions to Lewis diagrams, E.g., some stable compounds may have elements without full valence shells, and elements that can hold 18/32 electrons in the valence shell do not work with dot diagrams.

The pairs of electrons that are not bonded are called - non-bonding electron pairs.

Lewis Structures for Ionic compounds

In an ionic bond, one atom becomes a cation, and the other becomes an anion. This is represented in Lewis diagrams by the cation having a positive superscript, and the anion having a negative superscript, with square brackets around it. For example, for Magnesium Oxide:



Intermolecular bonds and Intramolecular Forces

Intramolecular bonds are the bonds that hold atoms together *within* a molecule.

Intermolecular forces are the forces that exist *between* molecules.

Intramolecular bonds are generally much stronger than Intermolecular forces - breaking apart the atoms inside a molecule requires a lot more energy than breaking molecules away from each other.

While intramolecular forces always arise from chemical bonds, intermolecular forces can also arise from electromagnetic attractions holding molecules together.

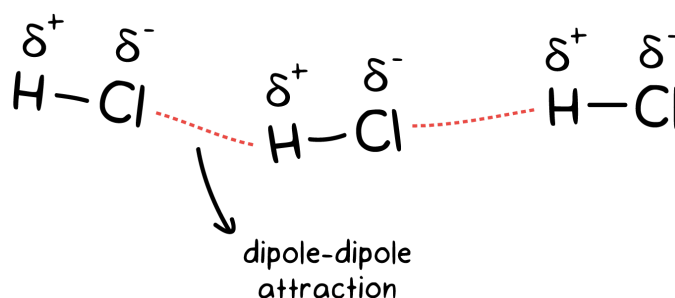
Intramolecular bonds

Types of intramolecular bonds include ionic, covalent and metallic bonds.

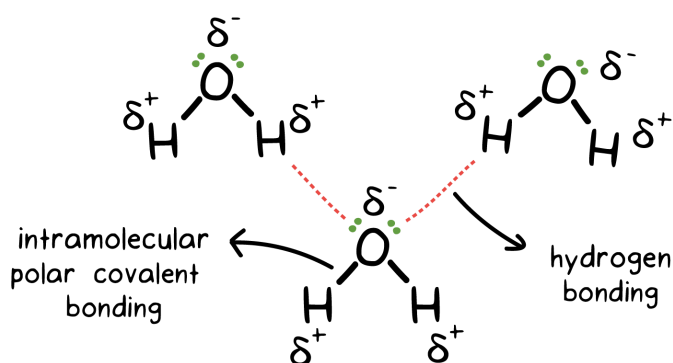
Intermolecular forces

The more electrons an atom has (i.e. larger atomic radius), the stronger it's intermolecular forces. The state of matter at room temperature can be used to compare the strength of the intermolecular forces between two elements/compounds - if it is a solid, it has stronger intermolecular forces than a substance that is a gas at room temp.

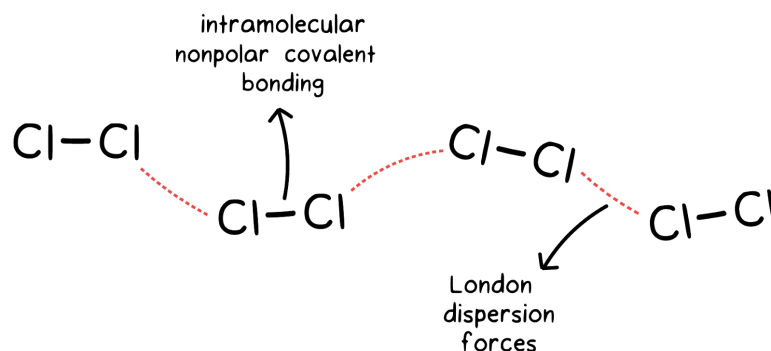
- **Dipole-Dipole Interactions** - When multiple molecules are in polar covalent or ionic bonds, one side of the molecule is slightly positive, and one side is slightly negative. These sides are known as dipoles. When the negative dipole of one molecule attracts to the positive dipole of another, they bond. This is the strongest intermolecular force.



- **Hydrogen Bonds** - a special kind of dipole-dipole interaction that occurs between hydrogen atoms bonded to either oxygen, nitrogen, or fluorine atoms. The positive dipole of Hydrogen is attracted to the negative dipole of oxygen/nitrogen/fluorine. Due to the large electronegativity difference, this is the strongest version of dipole-dipole interactions.



- **London Dispersion Forces (Van der Waal forces)** - are the weakest intermolecular force. They intrinsically exist between all molecules (doesn't have to be ionic/polar), and occur due to the movement of electrons forming temporary dipoles. The more electrons a molecule has, the stronger the Van der Waal forces.



Physical Properties of Elements/Compounds due to Intra/Intermolecular forces

Polar covalent/Ionic compounds have strong dipole-dipole interactions between them, while Nonpolar compounds have only weak London dispersion forces. The stronger a bond is, the more energy is required to break it, so ionic and polar compounds tend to have higher melting and boiling points than nonpolar compounds.

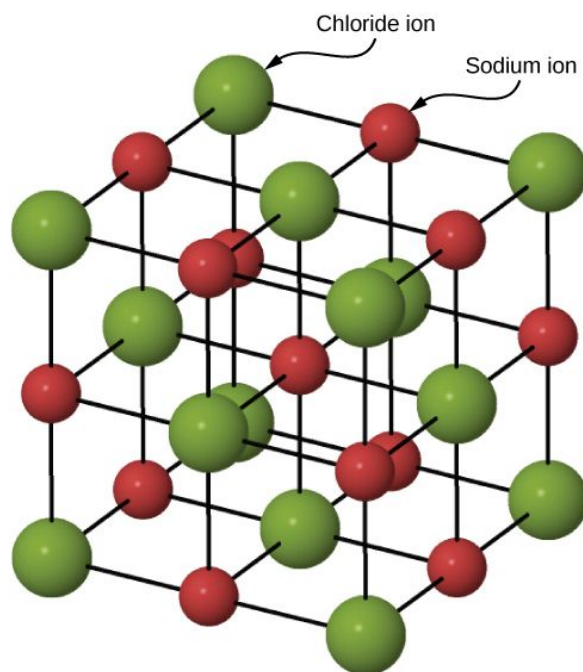
For example, methane boils at -161.5°C , while water, a polar covalent compound, boils at 100°C . An ionic compound such as sodium chloride has a much higher boiling point, at 1413°C .

Chemical Structures

The four main types of chemical structures are: Ionic network, Covalent network, Covalent Molecular and Metallic.

Ionic Network

Ionic Compounds form lattice structures, where each cation is surrounded by 6 anions, and each anion is surrounded by 6 cations - forming something like a cube like NaCl does:

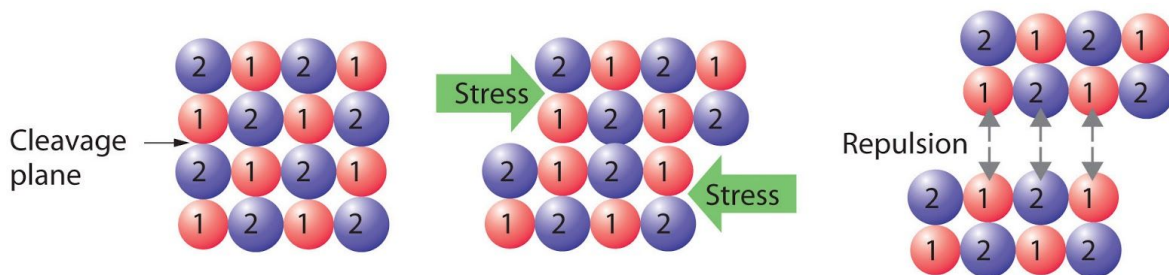


- Although one Sodium atom only gives an electron to one Chlorine, the attraction is equal in all directions, so the bond is **non-directional**.
 - This makes the bonds very strong, so ionic compounds have high melting and boiling points
- A single molecule of an ionic compound cannot exist, they always exist in lattices - so they are referred to as **Formula Units** instead of molecules, and are always in the simplest ratio

There are two methods of electrical conductivity - movement of electrons, or the movement of ions.

- Ionic solids have poor electrical conductivity - they have no free electrons, because they are locked into the oppositely charged atoms, and the ions are stuck in a lattice structure.
- Molten ionic substances have good electrical conductivity - the application of heat energy breaks the lattice, so the ions can move more and conduct electricity
- Aqueous ionic substances have extremely good electrical conductivity, as the solid dissociates into its ions, they have a large and easy range of movement to conduct electricity

Ionic solids are **brittle** (they break upon impact).



When an impact displaces some of the lattice layers, two or more of the same ion may align. Since they have the same charge, the ions that are in line repel each other, and the solid breaks.

Covalent Molecular

Sometimes, Covalent compounds exist as **discrete molecules**, with **weak intermolecular forces between each molecule**.

Depending on the strength of the intermolecular forces, covalent molecular compounds can be liquids, solids or gases - thus, they have a wide range of melting and boiling points

Covalent Molecular substances do not conduct electricity because there are no ions or free floating electrons

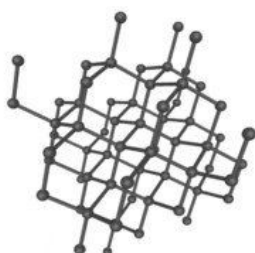
Covalent Network

Can come as 2D or 3D networks.

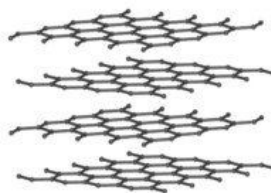
The most prominent example is Graphite (2D Network) vs Diamond (3D Network). They are both carbon allotropes (different structures of the same element), and have very different properties because of their different structures.

Graphite occurs as stacked layers of carbon in a 2D hexagonal lattice structure. Within the layers, there are strong covalent bonds, but between the layers, there are weak intermolecular forces. Thus, the layers slide over each other easily.

- This is why graphite is soft enough to rub off on paper as pencil, and is also used as a solid lubricant



Diamond

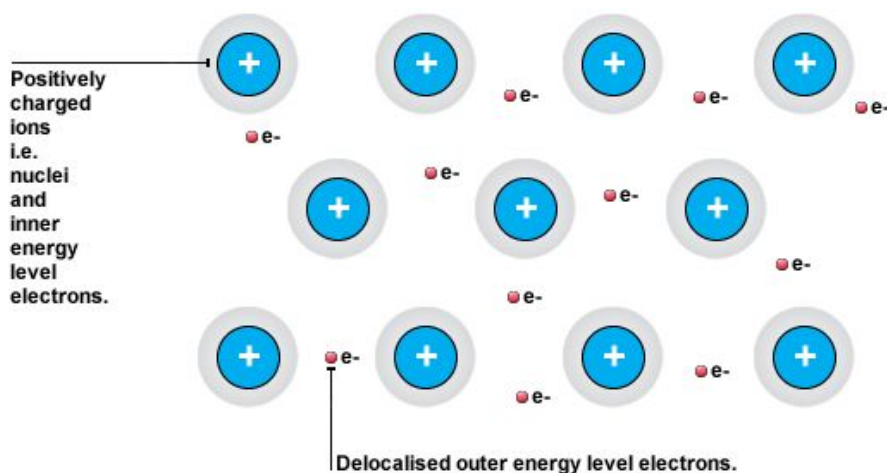


Graphite

Diamond occurs as a tetrahedral structure, with strong covalent bonds everywhere. Because of this, it is extremely hard and not malleable/conductive at all, and has extremely high melting and boiling points.

Metallic

A lattice of positive metal ions are held together by delocalised electrons (electron cloud/ electron sea) in metallic lattices.



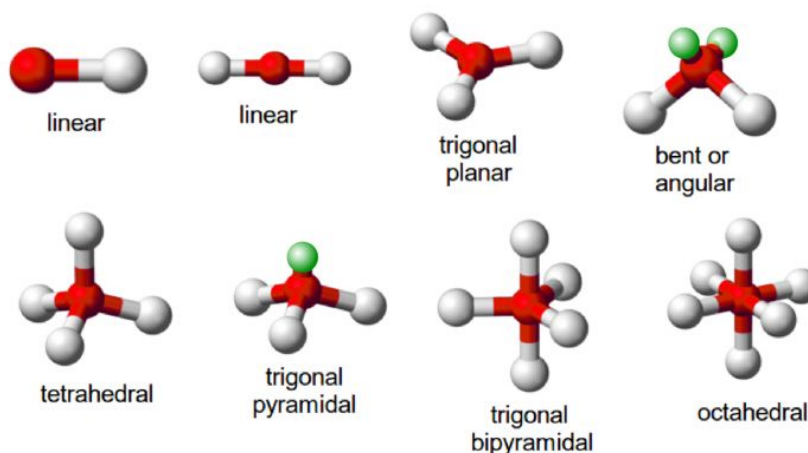
- The delocalised electrons carry electricity very well, making metals great conductors of electricity.
- They also carry heat energy in the form of kinetic energy by moving around and hitting the colder parts of the lattice, making metals great conductors of heat
- Delocalised electrons are also responsible for the malleability/ductility of metals - when metal ions are displaced, the electron cloud acts as a 'glue' and stops the ions from repelling and breaking apart

Allotropes

Allotropes are different structures of the same element - due to their different structures, they have different physical and chemical properties - i.e. diamond vs carbon as above. Sometimes, an allotrope of an element may not have a structure - this is known as amorphous. An example is carbon as soot.

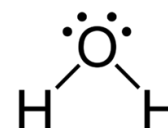
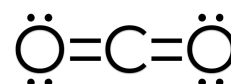
Chemical Shapes (Molecular Geometry)

Different molecules form different shapes depending on their composition. The main types of molecular shapes are:



The shape of a molecule depends on the electron configurations of its atoms - known as Valence shell electron pair repulsion (VSEPR) theory.

- For example, CO_2 has its nonbonding electrons distributed equally on opposite sides (it is nonpolar), so has a linear structure, since the repulsion is equal on either side.
- H_2O has its nonbonding electrons on top of oxygen, so they repel the hydrogen atoms downwards. As the hydrogen atoms come closer together, they also repel each other, creating a bent molecular shape



Naming Compounds

Naming Ionic Compounds

- Name the cation (metal) first, then the anion (nonmetal) second
- Add the appropriate suffix to the anion (-ide, -ate, -ite)
- For example. NaCl is Sodium Chloride, AlO_2 is Aluminium Oxide, etc

Naming Covalent (Molecular) Compounds

- Name the element with lowest electronegativity first
- Add the appropriate suffix to the second element (-ide)
- Add a prefix to both the first and second element to indicate the number of atoms
 - i.e. mono, di, tri, tetra, penta, hexa, hepta, octa, nona, deca
 - Don't add the mono prefix to the first element

- For example, Carbon Dioxide or Dichlorine Monoxide

IUPAC Nomenclature

IUPAC (The International Union of Pure and Applied Chemistry) has more rules (including the ones above) for naming inorganic compounds:

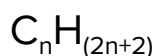
1. Single atom anions are named with an -ide suffix.
 - a. For example, a H^- ion is Hydride, O^{2-} ion is Oxide, etc
2. Cations with more than one possible charge must have the charge indicated after it in roman numerals.
 - a. For example, copper (i) oxide and copper (ii) oxide
3. Oxyanions (polyatomic anions containing oxygen) are named with -ite or -ate, for a lesser or greater quantity of oxygen, respectively.
 - a. For example, NO_2^- is nitrite, while NO_3^- is nitrate.
4. If four oxyanions are possible, the prefixes hypo- and per- are used:
 - a. For example, hypochlorite is ClO^- , perchlorate is ClO_4^-
5. The prefix bi- is a way of indicating the presence of a single hydrogen ion in the compound.
 - a. For example, sodium bicarbonate (NaHCO_3)

Naming Alkanes

No. of Carbons	Name	No of Carbons	Name
1	methane	7	heptane
2	ethane	8	octane
3	propane	9	nonane
4	butane	10	decane
5	pentane	11	undecane
6	hexane	12	dodecane

Ane or ene mean the carbon must have a single or double bond respectively.

The general formula for an alkane is:



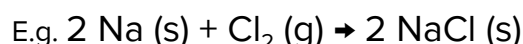
I.e. Hexane (6 Carbon) has $(2 \times 6 + 2)$ Hydrogen, so it's formula is C_6H_{14}

Module 2: Introduction to Quantitative Chemistry

Stoichiometry

Since the Law of Conservation of Mass states that matter cannot be created or destroyed, the atoms before and after the reaction must be the same.

Thus, Chemical Equations can be balanced by adding **stoichiometric coefficients** in front of the molecules.



This means that for every two separate Sodium atoms (total mass: 45.98 amu) and diatomic Chlorine molecule (70.90 amu) that react, 2 molecules worth (275.9 amu) of Sodium Chloride is produced.

Moles and Molar Mass

A mole is a unit of measurement, defined as **6.022×10^{23}** (Avogadro's Constant) particles of a compound/element, with the symbol **n**.

- E.g. 1 Mole of Sodium is equal to 6.022×10^{23} Sodium atoms

This system of measurement is relative to the Carbon-12 atom (1 mole of carbon 12 weighs exactly 12 grams)

The mass in grams of one mole of a molecule is known as the **Molar Mass** of the molecule, measured in grams per mole (g/mol). The Molar Mass of a molecule is equal to its atomic mass but in grams.

- E.g. Sodium (atomic mass: 22.99 amu) has a molar mass of 22.99 g/mol

The relationship between Moles, Actual Mass and Molar mass is given by the formula:

$$n = m/MM$$

$$\text{Number of Moles} = \text{Mass(g)} \div \text{Molar Mass(g/mol)}$$

Other Mole Formulae include:

$$n = N/N_A$$

(Number of Moles = Number of Particles / Avogadros Number)

$$n = CV$$

**(Number of Moles = Concentration of Solution (moles/litre) *
Volume (litres))**

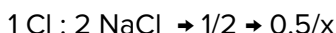
$$n = \text{Volume of Gas} / 22.71 \text{ OR } 24.79 \text{ (STP or RTP)}$$

Molar Stoichiometry

To describe the Sodium + Chlorine \rightarrow Sodium Chloride reaction in molar terms, it would be 2 moles of Na reacting with 1 Mole of Cl_2 to produce 2 Moles of NaCl.

These coefficients may not always be integers - you may need to calculate How much product is produced if 0.5 Moles of Chlorine is reacted.

To do this, find the ratio between the reactant and product in the balanced equation, and use your maths skills:



$$x = 1 \text{ (1 mole of NaCl is produced)}$$

This concept can also be applied to reactant vs reactant. The ratio between two reactants/ reactant and product is known as the **stoichiometric ratio**.

ALWAYS CONVERT WEIGHT ETC TO MOLES BEFORE DOING CALCULATIONS

Limiting Reagents

The limiting reagent in a chemical reaction is the substance that is totally consumed when the chemical reaction is complete. The amount of product formed is limited by this reagent, since the reaction cannot continue without it, and will leave an excess of another substance.

The maximum amount of product that can be produced is called the **theoretical yield**, and requires knowing the limiting reagent to be calculated.

Concentration and Molarity

Concentration refers to the amount of solute divided by the volume of solvent. It can be measured in g/L, n/L, g/kg, mL/L, ppm (parts per million).

- Colour is a good indicator of how concentrated a solution is, since the more solute is added, the darker the colour gets.
- Molarity is a unit of concentration, defined as moles of solute per litre of solvent (n/L).

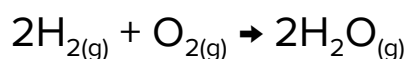
$$n = C \times V$$
$$C_1V_1 = C_2V_2$$

*Convert concentrations to mol/L if required in calculations

Gay-Lussac's Law of Combining Volumes

When measured at constant temperature and pressure, the volume of gases taking part in a chemical reaction will combine in simple whole number ratios

E.g.



2 volumes of Hydrogen gas reacts with 1 volume of Oxygen gas to produce 2 volumes of Water Vapour.

Avogadro's Gas Law

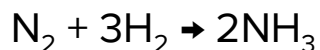
When measured at the same temperature and pressure, equal volumes of different gases contain the same number of molecules.

$$V \propto n$$

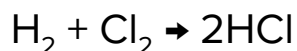
(Volume is proportional to number of Moles)

$$V_1/n_1 = V_2/n_2$$

E.g.



If Nitrogen was 100ml, then 300ml of H₂ is needed for 200ml of Ammonia



100 mL Hydrogen gas reacts with 100 mL Chlorine gas to produce 200ml of Hydrogen Chloride gas

Boyle's Law

The **pressure** exerted by a gas is **inversely proportional to** its **volume** when measured at a constant temperature

$$P \propto 1/V$$

$$P_1V_1 = P_2V_2$$

E.g. Calculate the volume of a gas at 200 kPa, if volume of the same gas at 100 kPa is 2.5L (at 25°C in both cases)

$$100 \times 2.5 = 200 \times V_2$$

$$V_2 = 1.25 \text{ L}$$

Charles' Law

The volume of a gas is directly proportional to its temperature when measured at a constant pressure.

$$V \propto T$$

$$V_1/T_1 = V_2/T_2$$

E.g. Calculate the volume of a gas at 20°C, if the volume of the same gas at 40°C is 3.0L (at 100 kPa in both cases)

$$3/40 = V_2/20$$

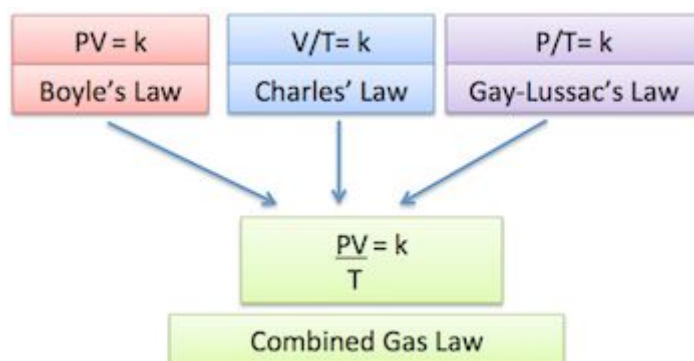
$$V_2 = 1.5\text{L}$$

Gay-Lussac's Pressure and Temperature Law

The **pressure** exerted by a gas is **directly proportional to its temperature** (in Kelvin), when Volume is constant.

$$P \propto T$$
$$P_1/T_1 = P_2/T_2$$

Combined Gas Law



Ideal Gas Law

An Ideal Gas is a theoretical model of a gas that obeys the gas laws exactly. It follows certain assumptions:

- The gas molecules move randomly in straight lines
- Pressure is due to collisions between the walls of the container and the molecules
- All collisions are perfectly elastic
- There are no intermolecular forces between gas molecules
- The gas molecules themselves take up no volume

Real Gases don't follow the gas laws perfectly. The last two assumptions are the main things that differentiate real gas from the ideal gas.

The pressure, volume and temperature are related through a formula called the ideal gas law:

$$PV = nRT$$

NOTE:

R is the universal gas constant.

When $R = 0.082 \text{ K}\cdot\text{mol/L}\cdot\text{atm}$:

- Pressure is measured in *atm*, Volume in L, Temperature in Kelvin (K)

When $R = 8.31 \text{ J/K}\cdot\text{mol}$:

- Pressure is measured in Pascals (Pa), Volume in m^3 , Temperature in Kelvin (K)

Standard Solutions

A standard solution is a solution of known and fixed concentration. A known weight of solute is dissolved to make a specific volume.

Primary standards are used to prep them - a primary standard is a reagent which can be weighed easily, and which is so pure that its weight is truly representative of the number of moles of substance contained. Not all substances can be used as primary standards.

- Sodium Hydroxide absorbs water rapidly, so it's difficult to know the actual mass

The properties of a good primary standard include: high purity, low reactivity, non-toxic, not likely to absorb moisture, high equivalent weight

How to make a Standard Solution:

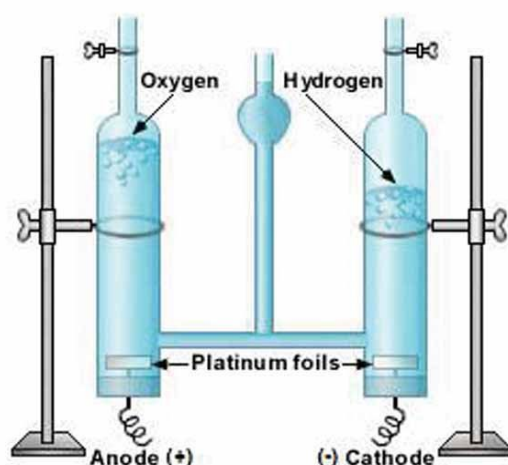
1. Weigh the primary standard - the (pure) solid dissolved to form the standard solution.
2. Dissolve the primary standard in a small amount of distilled water in a beaker.
3. Pour the contents of the beaker into a volumetric flask, rinsing the beaker with slightly more distilled water and pouring in to ensure all contents are transferred.
4. Add distilled water to the flask up to the calibration point. Use a graduated pipette to deliver the final amount.

Each step in the procedure is designed to increase accuracy and precision as much as possible, to ensure that concentration of the solution is accurate and precise.

Electrolysis

Application of an electric current to decompose a compound

Consider: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$



In an acid environment, an electric current is applied to water. Gases are collected over the anode and cathode respectively. On testing, the gases are shown to be O_2 at the anode and H_2 at the cathode (using glowing splint and pop test). The ratio of gases is 1:2 respectively, which validates Gay-Lussac's Law of combining gases.

Electrolysis requires much more energy than boiling water, since it is actually breaking the intramolecular covalent bonds instead of only the intermolecular bonds. It is also a chemical reaction (bonds are being broken and reformed), while boiling water is a physical change.

Dilutions

Module 3: Reactive Chemistry

Chemical Reactions

Chemical reaction occur when Chemical Bonds are broken, rearranged and established to form new substances. They involve reactants turning into products.

Indicators of a chemical reaction include:

1. Bubbles
2. Colour change
3. Change in Energy (change in temperature)
4. Appearance of a solid (due to precipitation)
5. Disappearance of a solid (not due to dissolving)
6. New Substance ← this is the only thing that guarantees a chemical reaction occurring, the rest are *indicators*

Exothermic vs Endothermic

Exothermic reaction: the reactants have more energy than the products

- This **energy is released** from bonds being formed and goes into the surroundings, usually causing the **temperature to rise**
- Examples include most Combustion reactions

Endothermic reaction: energy is taken from the surroundings and the reactants have less energy than the products

- Because **energy is being used** to break bonds, endothermic reactions usually cause the **temperature to drop**
- Examples include most Decomposition reactions

However, even physical changes can be endothermic or exothermic.

- For example, ice melting is endothermic due to the energy required for the bonds being broken.
- Sodium Hydroxide dissolving in water is exothermic, because even though the bonds between the Na and OH require energy to break, the negative poles of the water molecules attach to the Na ions (it is **hydrated**), and the formation of these “attachments” (intermolecular forces, not new bonds) releases energy.

Types of Reactions

Synthesis Reactions ($A + B \rightarrow AB$)

Synthesis is the formation of compounds by combining simpler substances such as elements

- Synthesis reactions are exothermic (energy is given off when bonds are broken)
- Examples of synthesis reactions include:
 - $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$

Displacement Reactions ($A + BC \rightarrow AC + B$ or $AB + CD \rightarrow AD + CB$)

When salts dissolve in water, the bonds between the cation and anion separate inside the solution

- When two of these solutions containing soluble salts (that have unpaired cations and anions) are mixed, the cation-anion pairings sometimes swap creating new salts
 - **This swapping of ions when two salt solutions are mixed is called displacement**
 - Both reactants have to be soluble
 - An example displacement reaction is: $\text{Na}_2\text{S} + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{S}$

Precipitation Reactions (Soluble Salt A (aq) + Soluble Salt B (aq) → Precipitate (s) + Soluble Salt C (aq))

After displacement, sometimes one of the new salts is insoluble and falls to the bottom if heavy enough. This settling of the heavy insoluble salt is called **precipitation**, and the solid particles formed are called the **precipitate**

- An example of this is: $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$
- **AgCl** (silver chloride) is insoluble, so forms a white precipitate

Whether or not such a reaction occurs can be determined using the **solubility rules**

Decomposition Reactions (AB → A + B)

- In **decomposition**, the atoms of a compound are separated to form two or more products
- Decomposition reactions are used in smelting metal ores to separate the pure metal from the ores as well as in electrolysis of water (creating hydrogen and oxygen)
- **Rules for decomposition:**
 1. All binary compounds (compounds with only two different elements) will break down into their elements.
 2. All carbonates break down to the oxide and carbon dioxide.
 3. Chlorates will break down to the binary salt and oxygen.

Types of Decomposition Reactions

- **Thermal**
 - Decomposed through the application of heat, making it an endothermic reaction
 - *E.g. $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$*
 - *Calcium carbonate → Calcium Oxide + Carbon Dioxide gas*
- **Electrolytic**
 - Decomposition due to an electric current being passed through an aqueous solution of a compound
 - Electrolytic cells are used for electrolytic decomposition
 - *E.g. $\text{NaCl} \rightarrow \text{Na} + \text{Cl}$*
 - *Sodium Chloride → Sodium + Chloride*
 - *E.g. $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$*
 - *Water → Hydrogen gas + Oxygen gas*
- **Photo Decomposition**
 - Decomposition due to use of light energy/photons
 - *E.g. $2\text{AgCl} \rightarrow 2\text{Ag} + \text{Cl}_2$*
 - *E.g. $2\text{AgBr} \rightarrow 2\text{Ag} + \text{Br}_2$*

Decomposition reactions are **usually endothermic** as they require energy to break the bonds

Reactions involving Acids

Acid + Base → Salt + Water

When an Acid and Base react, the (OH-) and (H+) create water and neutralise each other, forming a salt

- This is called a **neutralisation** reaction, because the H⁺ in acids and OH⁻ in bases react and produce water, and this brings the pH closer to 7 (since pH is based on concentrations of H⁺ or OH⁻ ions)
- Different acids will produce different salts

E.g. $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

Hydrochloric Acid + Sodium Hydroxide → Salt + Water

Acid + Metal → Salt + Hydrogen gas

When acid reacts with a metal, hydrogen gas and a salt is produced

- Some metals react more rapidly than others i.e. magnesium
- Some metals require heat to begin the reaction
- Other metals such as gold will not react with weaker acids

E.g. $\text{Mg} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2$

E.g. $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$

E.g. $\text{Pb} + \text{H}_2\text{SO}_4 \rightarrow \text{PbSO}_4 + \text{H}_2$

Metal Carbonate + Acid → Salt + Carbon Dioxide + Water

When an acid reacts with a metal carbonate or bicarbonate, a salt, carbon dioxide and water is produced

- Citric acid reacting with sodium bicarbonate produces the fizzy sensation of sherbet

Carbonate vs Acid	Hydrochloric Acid	Nitric Acid	Sulfuric Acid
Copper Carbonate	Copper chloride	Copper nitrate	Copper sulfate
Zinc Carbonate	Zinc Chloride	Zinc nitrate	Zinc sulfate
Iron Carbonate	Iron Chloride	Iron Nitrate	Iron sulfate

E.g. $2\text{HCl} + \text{Na}_2\text{CO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + 2\text{NaCl}$

E.g. $\text{H}_2\text{SO}_4 + \text{CuCO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2 + \text{CuSO}_4$

Metal Oxide + Acid \rightarrow Salt + Water

When a Metal Oxide and an Acid react, salt and water is produced

E.g. $\text{CuO} + 2\text{HCl} \rightarrow \text{CuCl}_2 + \text{H}_2\text{O}$

Combustion reactions

Things burning are indicative of reactions with oxygen, where flames (light, heat and sound energy) can be observed - these are known as **combustion reactions**, and are exothermic

- Other reactions with oxygen are corrosion reactions - these are much slower

Incomplete combustion (not enough oxygen)

Hydrocarbon + Oxygen \rightarrow Carbon Monoxide + Carbon + Water

Complete Combustion

Hydrocarbon + Oxygen \rightarrow Carbon Dioxide + Water

E.g. Complete combustion of methane

$\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

E.g. Incomplete combustion of ethane

$\text{C}_2\text{H}_6 + 2\text{O}_2 \rightarrow \text{CO} + \text{C} + 3\text{H}_2\text{O}$

Solubility Rules

These rules provide guidelines that tell which ions form solids and which remain in their ionic form in aqueous solution.

If all the ions in a reaction are shown to be soluble, then no precipitation reaction occurs.

The rules are to be followed from the top down, meaning that if something is insoluble (or soluble) due to rule 1, it has precedence over a higher-numbered rule.

1. Salts formed with group 1 cations and NH_4^+ cations are **soluble**. There are some exceptions for certain Li^+ salts.
2. Acetates ($\text{C}_2\text{H}_3\text{O}_2^-$), nitrates (NO_3^-), and perchlorates (ClO_4^-) are **soluble**.
3. Bromides, chlorides, and iodides are **soluble**.
4. Sulfates (SO_4^{2-}) are **soluble** with the exception of sulfates formed with Ca^{2+} , Sr^{2+} , and Ba^{2+} .

- Salts containing silver, lead, and mercury (I) are **insoluble**.
- Carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), sulfides, oxides, and hydroxides (OH^-) are **insoluble**. Sulfides formed with group 2 cations and hydroxides formed with calcium, strontium, and barium are exceptions.

Ion	General Solubility Rule
NO_3^-	All nitrates are soluble
$\text{C}_2\text{H}_3\text{O}_2^-$	All acetates are soluble ($\text{AgC}_2\text{H}_3\text{O}_2$ only moderately)
Cl^- , Br^- , I^-	All chlorides, bromides and iodides are soluble except Ag^+ , Pb^+ and Hg_2^{2+} . (PbCl_2 is slightly soluble in cold water and moderately soluble in hot water.)
SO_4^{2-}	All sulfates are soluble except those of Ba^{2+} , Pb^{2+} , Ca^{2+} and Sr^{2+}
CO_3^{2-} and PO_4^{3-}	All carbonates and phosphates are insoluble except those of Na^+ , K^+ and NH_4^+ . (Many acid phosphates are soluble).
OH^-	All hydroxides are insoluble except those of Na^+ and K^+ . Hydroxides of Ba^{2+} and Ca^{2+} are slightly soluble.
S^{2-}	All sulfides are insoluble except those of Na^+ , K^+ , NH_4^+ and those of the alkaline earths: Mg^{2+} , Ca^{2+} , Sr^{2+} and Ba^{2+} . (Sulfides of Al^{3+} and Cr^{3+} hydrolyze and precipitate as the corresponding hydroxides.
Na^+ , K^+ and NH_4^+	All salts of sodium ion, potassium ion and ammonium ion are soluble except several uncommon ones.

Aboriginal Chemistry

The fruit of the cycad *Macrozamia* was exploited as an important food source in spite of its being highly toxic and carcinogenic. Another poisonous food includes the Moreton Bay Chestnut.

The toxicity of cycads is mainly attributed to BMAA, Cycasin and an unidentified toxin.

- Aborigines cooked the seeds in a fire pit to denature the toxic compounds - denaturing refers to using heat to destroy the characteristic properties of a protein or other biological macromolecule, by disrupting its molecular structure.
- High temperatures would break the bonds between the molecules in the structures shown below, removing their ability to disrupt liver and neurological functions.

They then sliced the nuts finely and put them in a basket to soak in running water for a few days, to help the compounds leach out (dissolve) into the water, to help ensure that only a negligible amount of toxic compounds remained. Cycasin especially can be hydrolyzed (it breaks down as it reacts with water). They would often cook it again before eating.

Reactivity Series of Metals

The reactivity series of metals is an arrangement of metals from highest to lowest reactivity (top → bottom).

The official chem datasheet also orders the metals from most → least reactive in the list of half equations.

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

It is made by comparing the reactivity of metals with four criteria:

- How does it react with oxygen?
- How does it react with water?
- How does it react with dilute solutions of acids?
- How does it react with other metal ions in solutions?

Metals with Oxygen - Na, Li, K spontaneously ignite, Mg reacts brightly when heated, Al, Zn, Fe, Nickel burn in pure oxygen if finely divided, Sn, Pb, Cu, Hg oxidise if heated in pure oxygen, Ag, Au, Pt do not react

Metals with Water - K, Na, Ca react with cold water, Mg reacts with hot water, Al, Zn, Fe, Nickel react with steam, Sn, Pb, Cu, Hg, Ag, Au, Pt do not react

Metals with Acids - K, Na, Ca, Mg, Al, Zn, Fe, Nickel react with dilute acid, Pb reacts with warm dilute acid, Sn, Cu react with hot concentrated acids to form a salt and a gas other than hydrogen

Metals above hydrogen in the series can displace hydrogen from solution - the further it is above, the easier it is to displace hydrogen

Metals below hydrogen have difficulty displacing hydrogen from solution, the further below it is, the harder it is to displace.

Metal Reactivity in terms of Periodic Trends

Since metals want to give away electrons, the more reactive metals give away electrons more easily.

When an atom has a large atomic radius, the valence electrons are further away from the nucleus, and it is easier for the atom to give away those electrons.

Thus, metals with larger atomic radius (and therefore lower ionisation energies, and lower electronegativity) are more reactive.

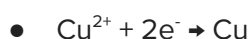
Redox Reactions

An oxidation-reduction reaction is any chemical reaction in which the oxidation number (the charge) of a molecule, atom, or ion changes by gaining or losing an electron.

Oxidation is a **loss** of electrons - thus, an **increase in oxidation state**



Reduction is a **gain** of electrons - thus, a **decrease in the oxidation state**



OILRIG = Oxidation Is Loss, Reduction Is Gain

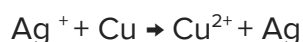
Oxidising and Reducing Agents

The **Oxidising Agent** is the atom that is used when oxidising the other - it gains electrons and **is the one being reduced** in a chemical reaction

The **Reducing Agent** is the atom that is used when reducing the other - it loses electrons and **is the one being oxidised**



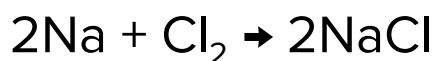
- The oxidising agent here is Hydrogen, as Hydrogen gains the electrons that Zinc loses



- Silver is being reduced, so is the oxidising agent, Copper is being oxidised, so is the reducing agent

Identifying Redox Reactions

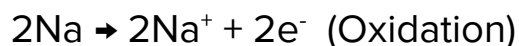
To identify if an equation is a redox reaction, the oxidation numbers of all its atoms must be identified. If there is a change in oxidation number of an atom before and after a reaction, then it is a redox reaction. For example:



Before: Na (charge of 0), Cl (0)

After: Na (+1), Cl (-1)

Therefore, this is a redox reaction. This can then be split into two half-equations:



But for a reaction like the next one, none of the atoms actually gain or lose electrons, since their oxidation numbers are the same before and after:



Balancing Complex Redox Reaction Equations

Balancing Redox equations is similar to balancing normal equations, but with a number of added steps. These steps are dependant on whether the equation is being balanced in a neutral solution, basic solution or acidic solutions.

Balancing involves the use of H^+ ions, water molecules, hydroxide ions and electrons to two half equations.

1. Split the equation into two half equations (oxidation and reduction).
2. Balance elements in the equations other than O and H.
3. Balance the oxygen atoms by adding the appropriate number of water (H_2O) molecules to the opposite side of the equations.
4. Balance the hydrogen atoms (including those added in step 2 to balance the oxygen atom) by adding H^+ ions to the opposite side of the equations.
5. Add up the charges on each side. Make them equal by adding enough electrons (e^-) to the more positive side. (Rule of thumb: e^- and H^+ are almost always on the same side.)
6. The e^- on each side must be made equal; if they are not equal, they must be multiplied by appropriate integers (the lowest common multiple) to be made the same.
7. The half-equations are added together, canceling out the electrons to form one balanced equation. Common terms should also be canceled out.

- (If the equation is being balanced in a basic solution, through the addition of one more step, the appropriate number of OH^- must be added to turn the remaining H^+ into water molecules.)

There are various examples on the following site:

[https://chem.libretexts.org/Textbook_Maps/Analytical_Chemistry/Supplemental_Modules_\(Analytical_Chemistry\)/Electrochemistry/Redox_Chemistry/Balancing_Redox_reactions](https://chem.libretexts.org/Textbook_Maps/Analytical_Chemistry/Supplemental_Modules_(Analytical_Chemistry)/Electrochemistry/Redox_Chemistry/Balancing_Redox_reactions)

Ionic Equations

Net Ionic Equations are another way of describing equations that happen in water. The following provides a really good explanation:

<https://www.khanacademy.org/science/chemistry/chemical-reactions-stoichiome/types-of-chemical-reactions/a/complete-ionic-and-net-ionic-equations>

The ions that do not participate in the reaction (i.e. they get “cancelled out”) are called spectator ions.

Standard Reduction Potentials

The table of standard reduction potentials is a table outlining the electric energy potential (voltage) of various half equations.

- A positive voltage means that is the amount of energy released in the spontaneous reaction
- A negative voltage means that amount of energy must be provided per mole of reactant for the reaction to occur (it is not spontaneous)
- All of these values have been obtained experimentally under standard conditions (25°C , 100 kPa, 1 mol/L concentration of electrolyte)
- The standard oxidation potential for each of these is just the reaction flipped, with the voltage having the opposite sign.
 - E.g. the reduction potential for Iron is 0.77V, but it's oxidation potential would be -0.77V

The values are obtained experimentally, and measured against the hydrogen standard half-cell (which has a potential of 0.00V)

The table is ordered in the reactivity series of metals, from most to least reactive from top to bottom

Some standard potentials

$\text{K}^+ + \text{e}^-$	\rightleftharpoons	K(s)	-2.94 V
$\text{Ba}^{2+} + 2\text{e}^-$	\rightleftharpoons	Ba(s)	-2.91 V
$\text{Ca}^{2+} + 2\text{e}^-$	\rightleftharpoons	Ca(s)	-2.87 V
$\text{Na}^+ + \text{e}^-$	\rightleftharpoons	Na(s)	-2.71 V
$\text{Mg}^{2+} + 2\text{e}^-$	\rightleftharpoons	Mg(s)	-2.36 V
$\text{Al}^{3+} + 3\text{e}^-$	\rightleftharpoons	Al(s)	-1.68 V
$\text{Mn}^{2+} + 2\text{e}^-$	\rightleftharpoons	Mn(s)	-1.18 V
$\text{H}_2\text{O} + \text{e}^-$	\rightleftharpoons	$\frac{1}{2}\text{H}_2(\text{g}) + \text{OH}^-$	-0.83 V
$\text{Zn}^{2+} + 2\text{e}^-$	\rightleftharpoons	Zn(s)	-0.76 V
$\text{Fe}^{2+} + 2\text{e}^-$	\rightleftharpoons	Fe(s)	-0.44 V
$\text{Ni}^{2+} + 2\text{e}^-$	\rightleftharpoons	Ni(s)	-0.24 V
$\text{Sn}^{2+} + 2\text{e}^-$	\rightleftharpoons	Sn(s)	-0.14 V
$\text{Pb}^{2+} + 2\text{e}^-$	\rightleftharpoons	Pb(s)	-0.13 V
$\text{H}^+ + \text{e}^-$	\rightleftharpoons	$\frac{1}{2}\text{H}_2(\text{g})$	0.00 V
$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	$\text{SO}_2(\text{aq}) + 2\text{H}_2\text{O}$	0.16 V
$\text{Cu}^{2+} + 2\text{e}^-$	\rightleftharpoons	Cu(s)	0.34 V
$\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2\text{O} + 2\text{e}^-$	\rightleftharpoons	2OH^-	0.40 V
$\text{Cu}^+ + \text{e}^-$	\rightleftharpoons	Cu(s)	0.52 V
$\frac{1}{2}\text{I}_2(\text{s}) + \text{e}^-$	\rightleftharpoons	I^-	0.54 V
$\frac{1}{2}\text{I}_2(\text{aq}) + \text{e}^-$	\rightleftharpoons	I^-	0.62 V
$\text{Fe}^{3+} + \text{e}^-$	\rightleftharpoons	Fe^{2+}	0.77 V
$\text{Ag}^+ + \text{e}^-$	\rightleftharpoons	Ag(s)	0.80 V
$\frac{1}{2}\text{Br}_2(\text{l}) + \text{e}^-$	\rightleftharpoons	Br^-	1.08 V
$\frac{1}{2}\text{Br}_2(\text{aq}) + \text{e}^-$	\rightleftharpoons	Br^-	1.10 V
$\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+ + 2\text{e}^-$	\rightleftharpoons	H_2O	1.23 V
$\frac{1}{2}\text{Cl}_2(\text{g}) + \text{e}^-$	\rightleftharpoons	Cl^-	1.36 V
$\frac{1}{2}\text{Cr}_2\text{O}_7^{2-} + 7\text{H}^+ + 3\text{e}^-$	\rightleftharpoons	$\text{Cr}^{3+} + \frac{7}{2}\text{H}_2\text{O}$	1.36 V
$\frac{1}{2}\text{Cl}_2(\text{aq}) + \text{e}^-$	\rightleftharpoons	Cl^-	1.40 V
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^-$	\rightleftharpoons	$\text{Mn}^{2+} + 4\text{H}_2\text{O}$	1.51 V
$\frac{1}{2}\text{F}_2(\text{g}) + \text{e}^-$	\rightleftharpoons	F^-	2.89 V

Galvanic Cells

A Galvanic/Voltaic Cell is a device that uses redox reactions to create electricity.

It contains two half cells - made up of an solid metal/carbon electrode submerged in an electrolyte (an ion soup that can conduct electricity) - this electrolyte is often a salt solution, such as zinc sulfate or copper sulfate, but can be a paste sometimes.

A “Salt Bridge” allows ions to pass between the cells - this is often made of a strip of folded filter paper, soaked in 1 mol/L Potassium Nitrate (used because it is relatively inert).

NOTE: It is the Potassium Nitrate Ions that move to the solutions, neutralising them

Anode vs Cathode

The **anode** is the electrode **where oxidation occurs** (loss of electrons), and the **cathode** is the electrode **where reduction occurs** (gain in electrons)

- The anode in a galvanic cell is negative, while the cathode is positive
- Electrons always move from negative to positive, not anode to cathode

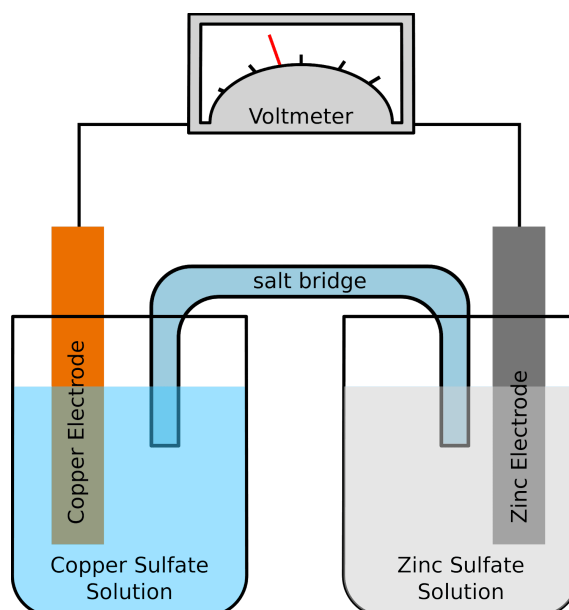
An easy way to remember this is **RED CAT** and **AN OX**

NOTE: The electrolyte must be the same metal solution as the electrode (e.g. Magnesium electrode would be submerged in Magnesium Nitrate solution)

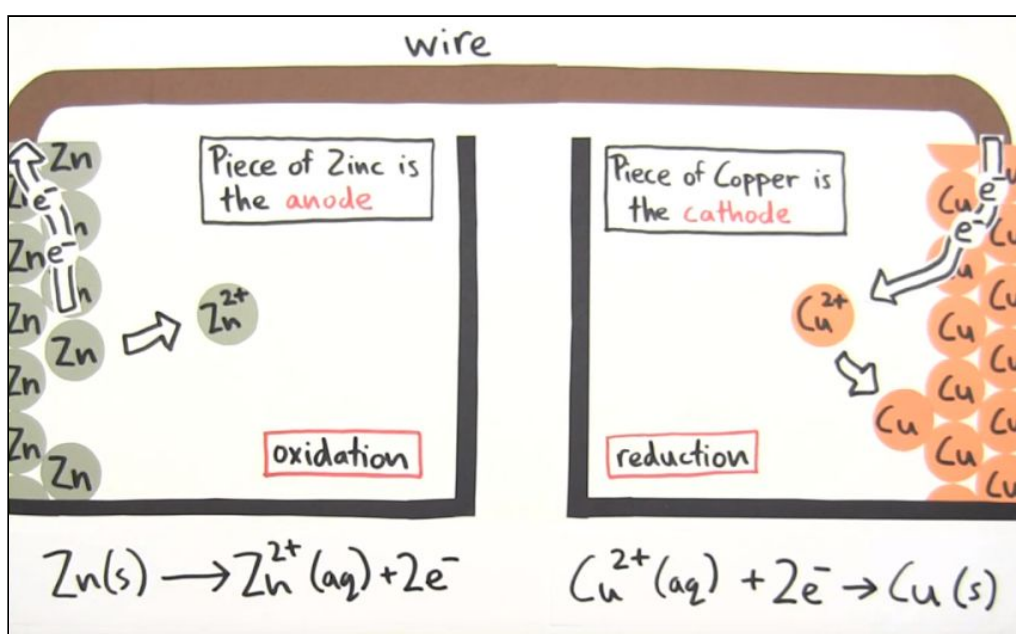
ALSO, the metal with the greatest electronegativity is the positive electrode (cathode), since it pulls the electrons towards it, and undergoes reduction

How they Work

The following is an example, metals other than Zinc and Copper can be used



- Each cell begins as one beaker full of **aqueous Zinc Sulfate with a piece of Zinc in it (Zinc Electrode)**, and another beaker full of **aqueous Copper Sulfate with a piece of copper in it (Copper Electrode)**, and the electrodes are connected by a wire
 - The **copper solution already has copper cations floating around**, and the **zinc solution has zinc cations floating around**, as they are both dissolved in water
- When the Copper and Zinc cations are dissolved in the water, they each want to pull for electrons - but Copper ions have greater electronegativity, so Zinc gives electrons
- Since the Zinc cations don't have many electrons to give, the Zinc metal/electrode gives away electrons through the wire
 - Since zinc is giving away electrons, this is an oxidation reaction, and is occurring at the anode
- When the Zinc metal/electrode gives away electrons, it makes the Zinc atoms into positively charged Zinc ions which then detach from the metal and dissolve into the solution
 - This makes the Zinc metal slowly corrode away as it dissolves
- When these electrons reach the copper solution through the wire, the copper cations grab them and turn neutral
 - Since neutral atoms don't dissolve, the copper cations turn into solid copper and deposit onto the copper electrode
 - Since the copper cations are gaining electrons, this is a reduction reaction, and occurs at the cathode



6. But we must also remember that both solutions have Sulfate anions floating around, to balance the overall charges of the Zinc and Copper cations.
 - a. When Zinc cations start dissolving into the solution, and Copper cations turn into solid copper, one side has too many Sulfate anions, and one side has too little. This would decrease the rate of reaction, since the solutions themselves become charged.
 - b. The **Salt Bridge** is there to balance out the charges in each solution - the Potassium cations move to the excessively negative side, and the Nitrate anions move to the excessively positive side

Half-Equations and Voltage

The two Half-equations in this Voltaic Cell are:

- $\text{Zn}_{(s)} \rightarrow \text{Zn}_{(aq)}^{2+} + 2\text{e}^- (+0.76\text{V})$
- $\text{Cu}_{(aq)}^{2+} + 2\text{e}^- \rightarrow \text{Cu}_{(s)} (+0.34\text{V})$

Therefore, the cell produces 1.10V overall. This can be obtained from the standard potentials table. When choosing electrolytes for a galvanic cell, only those elements that produce a net positive voltage can be used.

The movement of electrons from Zinc to Copper through the wire is what produces electric current.

Rate of Reaction

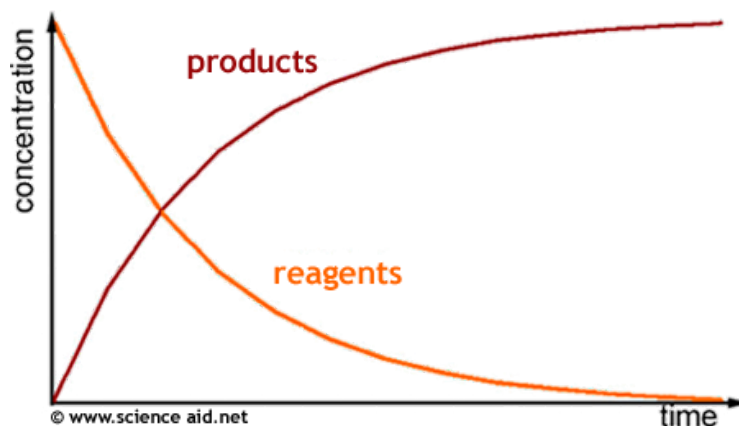
The **Rate of reaction** is a measure of how quickly the reactants are turned into products - it is represented by a change in the concentration of reacts/products with time.

$$R \propto [\text{Concentration of Products or Reactants}] / \text{Time}$$

For chemical reactions to occur, particles must collide with each other at high enough speeds - this is known as **collision theory**.

Anything that increases the chance or speed of collision, consequently increases the rate of reaction.

The reaction rate exponentially slows down over time - as the concentration of reactants decrease and products increase, and there is thus a lower chance of collision of reaction particles:



There are several ways to increase the rate of reaction

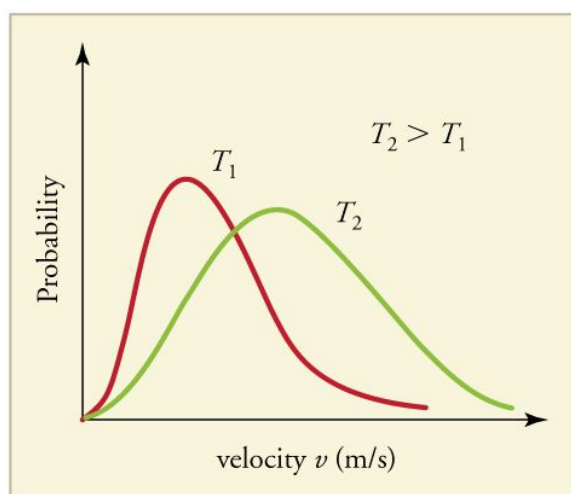
- **Increasing Temperature:** gives the particles kinetic energy for collisions, they move faster and there is a greater chance of successful (fast-enough) collisions
- **Increasing Concentration:** more particles of reactant in the same space, so collisions are more likely
- **Increase the Surface Area:** more of one reactant is exposed to the other, so more collisions
- **Add a Catalyst:** a substance which speeds up the reaction but doesn't get used up
 - Catalysts increase the rate of reaction by decreasing the activation energy required
- **Increasing Pressure:** which pushes particles closer together, more particles of reactant in the same space so collisions are more likely
 - Increasing pressure decreases volume
 - $C = n/V$
 - If volume goes down, concentration goes up
- Removing products as they form, and adding reactants as they are used up
- Mixing reactants and products in **optimal ratios**

Activation Energy

Activation Energy is the minimum amount of energy that is required for a set of reactants to begin reacting. This energy serves two main purposes:

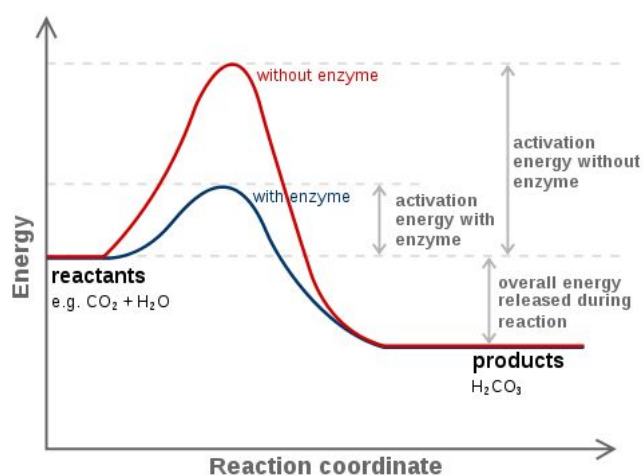
- Existing bonds must be broken in the reaction
- Reactant molecules must collide at high enough speeds

Maxwell-Boltzmann Distribution (Temperature and the Rate of Reaction)



This is a curve showing the likelihood of the particles in a substance having different velocities. At higher temperatures (T_2), more particles are likely to have greater speeds than at a lower temperature (T_1). This increases the probability of successful collisions, increasing the rate of reaction.

Activation Energy Curve



The activation energy is the peak of the curve. A catalyst (such as an enzyme) decreases the activation energy required.

Module 4: Drivers of Reactions

Energy Changes in Chemical Reactions

All chemical reactions have some energy changes occurring, since bonds are being broken and formed.

- The **formation of bonds produces energy**
- The **breaking of bonds consumes energy**

These energy changes follow the law of conservation of energy - the things storing Chemical Potential Energy are:

- Chemical Bonds
- Intermolecular Forces
- Repulsion between Electrons
- Repulsion between Nuclei
- Kinetic Energy of particles

Thus, the energy of the system isn't the same before and after the reaction - some reactions release energy, while some consume energy (i.e. exothermic and endothermic reactions). This is dependent on what and how many bonds are being broken/formed.

Enthalpy

Enthalpy is defined as the energy content of a system.

$$H = U + PV$$

$$\text{Enthalpy(J)} = \text{Internal Energy(J)} + \text{Pressure(atm)} \times \text{Volume(L)}$$

It is very difficult to know the exact enthalpy of a substance, so the enthalpy change is measured/calculated instead.

Change in Enthalpy is measured as Joules per mol, and can be calculated using the following formula:

$$\Delta H = -Q/n$$

$$\text{Change in Enthalpy(J/mol)} = -\text{Heat (J)/moles}$$

The enthalpy change in a reaction is **positive** if heat is absorbed (endothermic), and **negative** if heat is released (exothermic).

Hess' Law

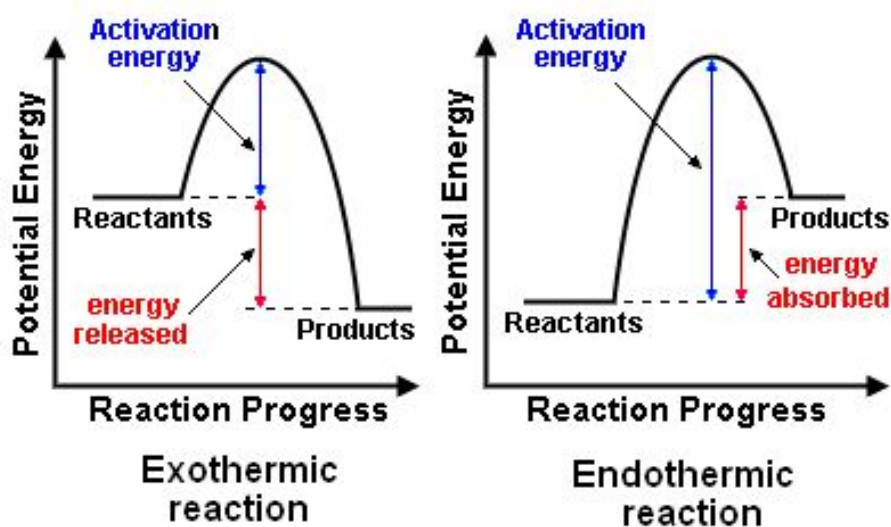
Hess's Law states that the total enthalpy change for a chemical reaction does not depend on the pathway that it takes - it only depends on the initial and final states.

The total enthalpy change for a reaction carried out in multiple steps is also equal to the sum of the enthalpy changes of each individual step.

Energy Profile Diagram

Energy Profile Diagrams are representations of the energy (or enthalpy) level of the reaction as the reaction progresses.

For example, it might start at one point, then gain energy as bonds are broken, then release energy as new bonds are formed, until the final products have an net lower enthalpy than the initial reactants. This shows that the reaction has overall released some energy, making it exothermic



If the **Energy of Products** < **Energy of Reactants**, it is **exothermic** (because that energy is released as heat), and $\Delta H < 0$.

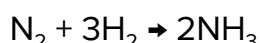
If the **Energy of Products** > **Energy of Reactants**, it is **endothermic** (because additional energy is consumed), and $\Delta H > 0$.

Also note that the Activation Energy in an exothermic reaction is generally lower than an endothermic reaction, because endothermic reactions are actually consuming energy.

Bond Energy

Bond Energy/Enthalpy refers to the amount of energy required to break the bonds of 1 mol of a substance into its constituent atoms under STP.

E.g. Given that an N-N triple bond has a bond energy of 945 kJ/mol, H-H bond has a bond energy of 436 kJ/mol, and N-H bond has a bond energy of 391 kJ/mol, calculate the enthalpy change of the following reaction:



N₂: One N-N triple bond

3H₂: Three H-H bonds

2NH₃: Six N-H bonds

Energy consumed to break bonds = 1 x 945 + 3 x 436 = 2253

Energy released by new bonds forming = 6 x 391 = 2346

$$\Delta H = +2253 - 2346 = -93 \text{ kJ/mol (exothermic)}$$

Therefore,

$$\Delta H = \sum \Delta H_{(\text{bonds broken})} - \sum \Delta H_{(\text{bonds formed})}$$

Standard Enthalpy of Formation

The standard enthalpy of formation of a compound is the **change of enthalpy during the formation of 1 mole of the substance** from its constituent elements, with all substances in their standard states.

Given the standard enthalpies of formation (ΔH_f) of different substances, the change in enthalpy of a reaction can be calculated:

$$\Delta H_{(\text{Reaction})} = \sum \Delta H_{(\text{Products})} - \sum \Delta H_{(\text{Reactants})}$$

Remember to multiply the

The following website has some great examples:

[https://chem.libretexts.org/Textbook_Maps/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_\(Physical_and_Theoretical_Chemistry\)/Thermodynamics/Energies_and_Potentials/Enthalpy/Standard_Enthalpy_Of_Formation](https://chem.libretexts.org/Textbook_Maps/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Thermodynamics/Energies_and_Potentials/Enthalpy/Standard_Enthalpy_Of_Formation)

Stepped Reactions

This video explains it better than words can:

<https://www.youtube.com/watch?v=iETCSFit-zA>

A stepped reaction is one that occurs in multiple steps. When given information about the enthalpy changes for each step, the total enthalpy change of the reaction can be calculated.

E.g. Calculate ΔH for the following reaction: $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$

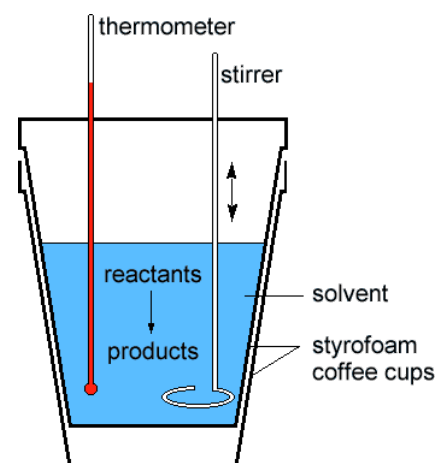


1. Manipulate the given reactions to resemble the actual reaction
 - a. $2\text{CO} \rightarrow 2\text{C} + \text{O}_2$ (equation flipped so flip ΔH , $\Delta H = 221.0 \text{ kJ/mol}$)
 - b. $2\text{C} + 2\text{O}_2 \rightarrow 2\text{CO}_2$ (equation x2 so ΔH x2 as well, $\Delta H = -787 \text{ kJ/mol}$)
2. Add the manipulated reactions
 - a. $2\text{CO} + 2\text{C} + 2\text{O}_2 \rightarrow 2\text{C} + \text{O}_2 + 2\text{CO}_2$
 - b. ΔH is added too, so $221 - 787 = -566 \text{ kJ/mol} = \Delta H$ of this reaction
3. Cancel out the same compounds on both sides, then check if the equation matches the original
 - a. Cancel out 2C and one O_2
 - b. $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$ is left, this matches the initial reaction
 - c. So the answer is $\Delta H = -566 \text{ kJ/mol}$

Calorimetry

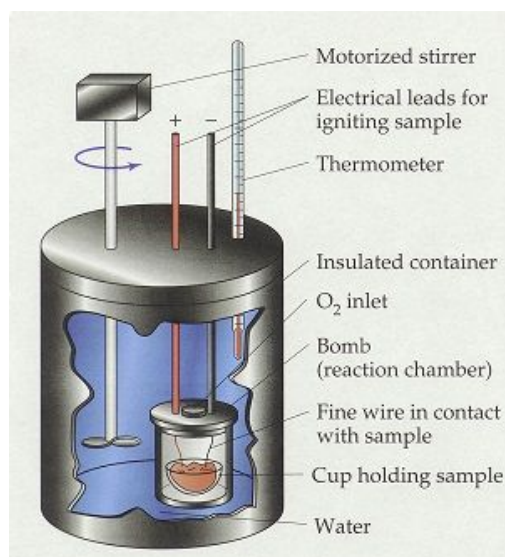
A calorimeter is a device used to measure the quantity of heat flow in a chemical reaction. The type shown in the diagram below is a “coffee cup calorimeter”, used to measure neutralisation and similar reactions

- It consists of a covered container surrounded by insulation (such as styrofoam), to minimise heat getting in from the surroundings, because it only wants to measure the heat flow in the chemical reaction.
- A thermometer measures the temperature of the system as the reactants are stirred.



Bomb calorimeters are a more effective alternative to accurately determine enthalpy changes in combustion reactions. This encompasses a copper ‘bomb’ containing measured amounts of reagents, which is submerged in a calorimeter of water. The chemical is ignited,

and the heat released warms the water, making use of the *specific heat of water* to determine change in energy within the system.



Once the change in temperature is observed, the **specific heat formula** can be used to calculate the heat/enthalpy change:

$$Q = mc\Delta T$$

Q = heat energy (Joules, J)

m = mass of a substance (g - this can also be kg depending on the units used in specific heat capacity)

c = specific heat capacity (units J/g/K)

ΔT = change in temperature (Kelvins, K, or Celsius, C - this would be the same)

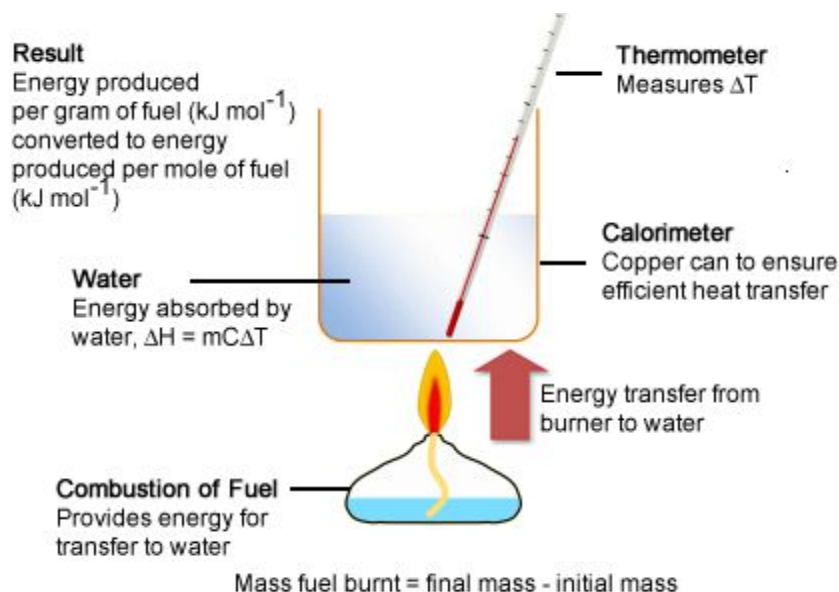
The specific heat capacity of a substance is the amount of energy needed to increase 1 gram of that substance by 1°C.

- Note in bomb calorimeters the mass and specific heat capacity of water should be used as m and c

Molar Heat of Combustion

Molar Heat of Combustion refers to the energy released when 1 mole of a substance undergoes complete combustion under STP - it is measured in kJ/mol. This is always a **positive** value, as combustion is an *exothermic* reaction, and energy is released.

- The **Molar Enthalpy of Combustion** is always a **negative** value, because it is the change in energy of the system per mole of substance that undergoes combustion.



One method used to determine enthalpy of combustion of *liquid fuels* (eg. ethanol, methanol) is to burn the fuel in a spirit burner, and using the heat from the combustion of the ethanol to heat a measured volume of water, held in a double-walled copper calorimeter.

- Copper ensures efficient heat transfer, as it is a conductor, but it is often **double-walled** to prevent the heat exiting the calorimeter
- Make sure the flame is touching the calorimeter
- Leave only a hole at the top for the thermometer

Calculating Molar Heat of Combustion

The total amount of heat released by a substance is the number of moles of that substance x the molar heat of combustion:

$$Q = mc\Delta T$$

$$Q = nH_C \text{ (when there is 100\% heat transfer)}$$

$$\text{Therefore, } H_C = -(m_{\text{water}} c_{\text{water}} \Delta T_{\text{water}}) / n_{\text{fuel}}$$

For Example:

3.15g of CH_3OH heating up 250mL of water with a temperature difference of 21.5°C .

$$c_{\text{water}} = 4.18$$

$$Q = m_w c_w \Delta T$$

$$n_{\text{CH}_3\text{OH}} = 3.15/32 = 0.0984$$

$$nH_C = mc\Delta T$$

$$H_C = 228.328 \text{ kJ/mol}$$

Enthalpy of Combustion: use $\Delta H = -mc\Delta T/n$

NOTE: m = mass of *water* used, not of the fuel

Unfortunately, the results of this technique are not very accurate; a great deal of heat can be lost in the surroundings, and the combustion may be incomplete (indicated by the deposition of black soot on the bottom of the water container).

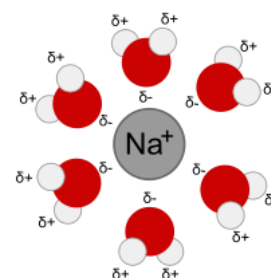
Molar Heat of Dissolution

Molar Heat of Dissolution is the amount of energy consumed/released when 1 mol of a substance dissolves in water under STP.

Dissolving also causes energy changes. When an ionic substance is dissolved in water:

1. Ionic Bonds are broken
2. Intermolecular Hydrogen bonds are broken between water molecules
3. Water molecules surround these ions, and link to them through strong ion-dipole forces. These ions are now called **hydrated ions**.

1 and 2 require energy, while 3 releases energy (formation of new "bonds").



- The dissolution of NaNO_3 is endothermic because the energy required to break bonds is greater than the energy released

A coffee-cup calorimeter can be used to measure the temperature change in the dissolution.

Entropy

Entropy is a measure of the degree of molecular disorder, or randomness, of a system.

- The number of possible arrangements of particles in a system is a good representation of entropy - but is too difficult to measure
 - E.g. Solids have the least entropy (since the particles are ordered), while gases have the most (particles move around randomly, so more possible arrangements)
- Systems tend to favour a more disordered system - reactions tend toward the side with an increase in entropy

The change in entropy can be calculated by:

$$\Delta S = \Delta Q/T$$

$$\Delta \text{Entropy (Joule/Kelvin)} = \Delta \text{Heat(J)}/\text{Temperature(K)}$$

Open vs Closed systems

- **Open system** - where matter and energy can go outside the system
- **Closed system** - matter can't exit the system but energy can
- **Isolated system** - neither matter nor energy can exit the system

Factors that affect Entropy:

- Temperature (more particle motion,)
- Dissolution
- Change in State
- Generally, more enthalpy means more entropy

Entropy Reaction Example (*Add in Barium Hydroxide and Ammonium thiocyanate*)

<https://www.britannica.com/science/entropy-physics>

Spontaneous Reactions and Gibbs Free Energy

<https://www.khanacademy.org/science/chemistry/thermodynamics-chemistry/gibbs-free-energy/a/gibbs-free-energy-and-spontaneity>

Spontaneous Reactions are those that occur without the addition of external energy, once the reaction has started - the two things governing reaction spontaneity are enthalpy and entropy.

Spontaneous Reactions favour the formation of the products under the conditions that the reaction is occurring in - i.e. they naturally want to react.

- Reactions are favorable when they result in a decrease in enthalpy and an increase in entropy of the system. Thus, most, but not all spontaneous reactions are exothermic
- However, this doesn't mean the reaction occurs instantly, spontaneity is not related to the rate of reaction. This also doesn't mean that the reaction does not need activation energy.

Examples include: Baking Soda and Vinegar, Rusting, or Wood Burning

- For the example of wood burning, some activation energy is required to ignite the flame. However, the reaction will continue spontaneously after that.

http://www.softschools.com/notes/ap_chemistry/gibbs_free_energy/