CM1131, CM2131, CM3131 Notes

Full Foundations in Chemistry I - III Series

This is a compilation of questions and notes given in the Foundations in Chemistry Courses.

Made by LQ, Joel and edited by many wonderful people (yey ty all)

Links of All LQ Notes: LQ Notes Links Document

Author's Note and Introduction

These notes cover all important topics in CM1131 and CM2131.
CM3131 notes are in progress.

If you find any mistake, please drop LQ (or Joel, maybe) a note anywhere or just suggest it in this document (everyone has commenting rights).

If a sentence has the tag [verify], it means that it could be inaccurate. If you know how to improve on it, please do help out, any help is appreciated :D

Everything that says "FYI", including all sections labelled as "Extension", are purely for additional reading, and is not tested (at that level).

For **Pause and Ponder**, the answers can be found at the end of the notes, **here**.

Please DO NOT share these notes outside of NUS High,
as NUSH Lesson Materials are used as a reference.

CM1131

Foundations in Chemistry I

Topic 1: The Particulate Nature of Matter

1.1 The Kinetic Particle Theory

The Kinetic Particle Theory says that all matter is made of particles.

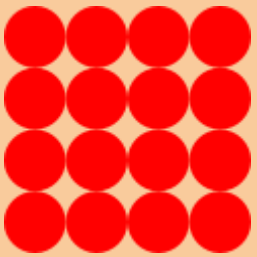
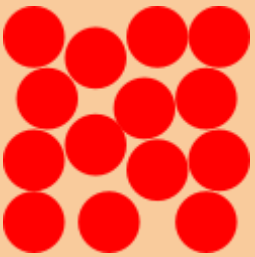
	Solid	Liquid	Gas
Diagram			
Arrangement	Orderly, Closely packed	Disorderly, Less closely packed	Disorderly, Very far apart
Attractive Forces	Strong	Less strong	Very weak
K.E. of particles	Very low	Low	Very high
Movement of particles	Vibrate about fixed positions	Able to slide over one another	Moving about randomly at high speeds

Table 1.1.1 - Properties of Matter as described by Kinetic Particle Theory

Some limitations of the Kinetic Particle Theory includes:

The kinetic theory model assumes that particles collide and bounce off each other. But, in reality, **particles often do not simply bounce off each other because they have forces of attraction between them**. This affects properties such as melting point.

1.2 Changes of State in Matter

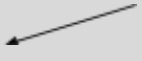
	Solid	Liquid	Gas
Solid		Freezing	Deposition
Liquid	Melting		Condensation
Gas	Sublimation	Evaporation/Boiling	

Table 1.2.1 - The Processes involved in Changes of State

1.3 Evaporation vs Boiling

Evaporation	Boiling
Can occur at temperatures below boiling point	Only occurs at boiling point
Occurs only at the surface of the liquid	Occurs throughout the liquid
Slower	Faster
Produces a cooling effect because heat energy is absorbed from the surrounding	Temperature remains constant (for pure substances)

Table 1.3.1 - Evaporation vs Boiling

1.4 Relative Atomic/Molecular Mass

The symbol of relative atomic mass is $\mathbf{A_r}$.
To find A_r, we look up the element on the periodic table to see its relative atomic mass.

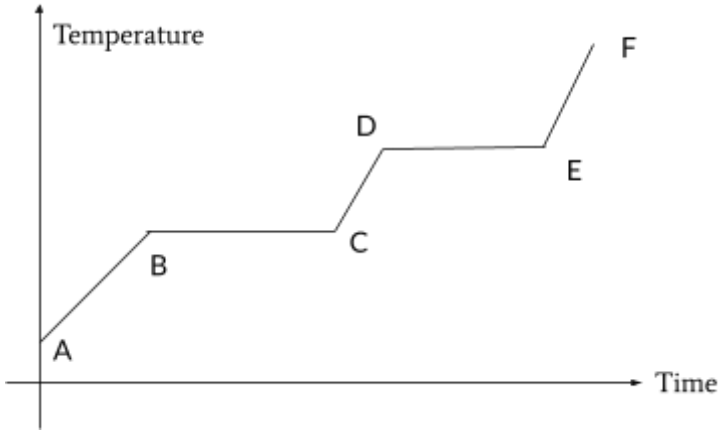
The symbol of relative molecular mass is $\mathbf{M_r}$.
To calculate M_r, we sum up the A_r of all the atoms in the molecule.

E.g. H_2O (Water) has a relative atomic mass of $2 \times 1.008 + 16.00 = 18.02$ (2 dp).

Do take note that A_r and M_r have no units.
As the name suggests, they are relative to a preset value (Specifically, $\frac{1}{12}$ × the mass of one carbon-12 molecule).

1.5 Heating Curve

This is what a regular heating curve looks like:



Graph 1.5.1 - Regular Heating Curve

Refer to *Graph 1.5.1* above.

AB is the substance as a **solid**.
BC is the substance **melting** and is composed of a mix of solid and liquid.
CD is the substance as a **liquid**.
DE is the substance **boiling** and is composed of a mix of liquid and gas.
EF is the substance as a **gas**.

Topic 2: Elements and the Periodic Table

2.1 Elements

An element is a pure substance that cannot be broken down into simpler substances by chemical methods.

2.1.1 The Periodic Table

The Periodic Table of Elements is a table that places the elements based on their atomic number, electron configurations and chemical properties.

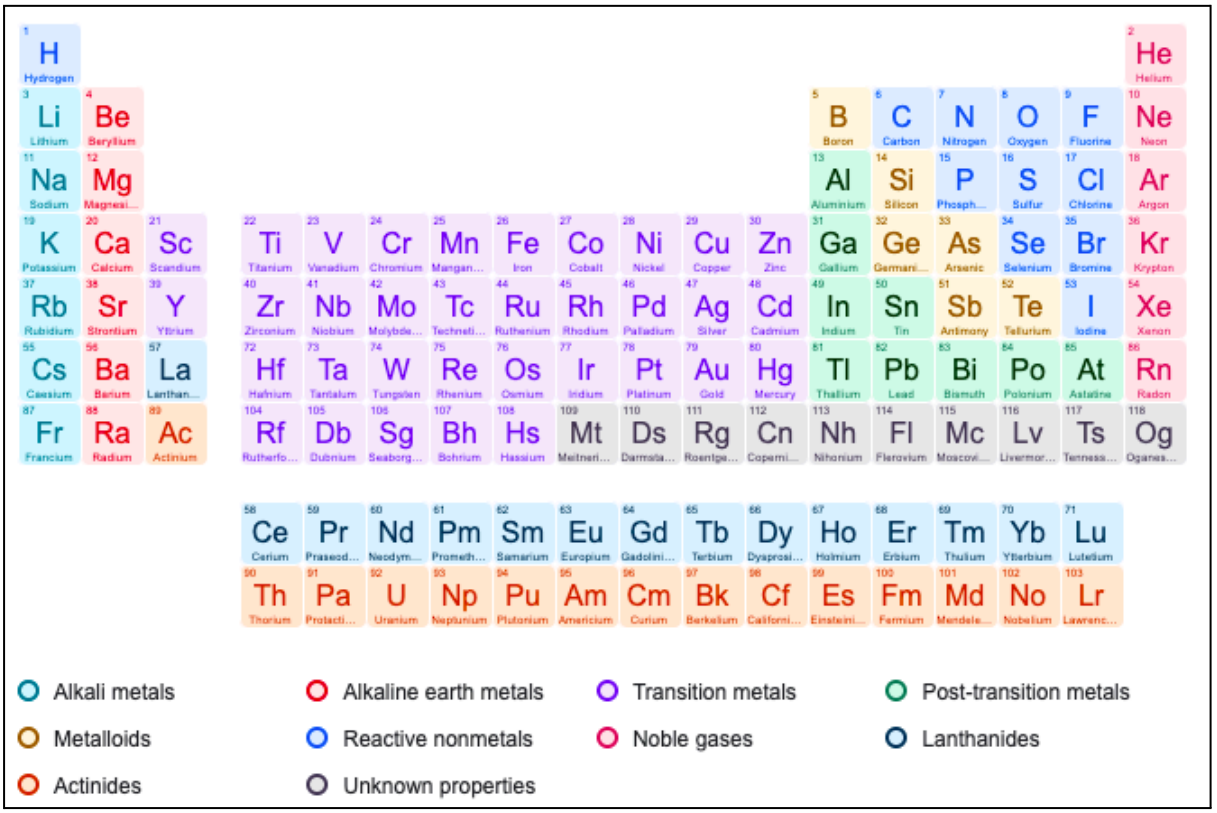


Figure 2.1.1.1 - The Periodic Table

The rows of the Periodic Table are called **periods**.
The columns are called **groups**.

2.1.2 Comparison of Properties of Metals and Non-metals

	Metals	Non-Metals
Appearance	Shiny	Dull
Physical State at Room Temperature	Solids (Except	Usually

	Mercury)	liquids/gases
Density	High	Low
Malleability (Can be hammered into diff. shapes without breaking)	Malleable	Brittle when solid
Ductility (Can be drawn or pulled into wires)	Ductile	Non-ductile
Sonorosity (Makes a ringing sound when struck)	Sonorous	Non-sonorous
Conductivity	Good conductors of heat and electricity	Poor conductors of heat and electricity (Except Carbon in the form of Graphite)

Table 2.1.2.1 - Properties of Metals vs Non-metals

2.1.3 Metalloids

Metalloids, or semi-metals, refer to the elements that lie on a diagonal line as shown below. They are called so because they exhibit properties intermediate between metals and non-metals.

The metalloids are: Boron, Silicon, Germanium, Arsenic, Antimony, and Tellurium.

This is also colloquially referred to as “the staircase”.

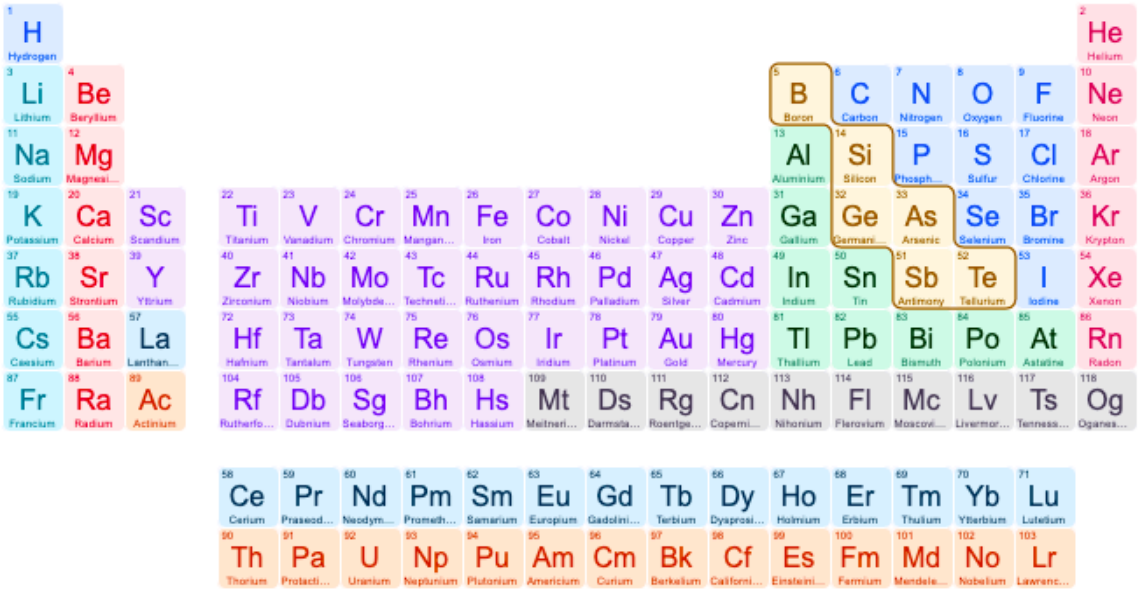


Figure 2.1.3.1 - Metalloids

2.1.4 Groups in the Periodic Table

We need to memorise 4 groups, 1, 2, 17 and 18.

Group	1	2	17	18
Name	Alkali Metals	Alkaline Earth Metals	Halogens	Noble Gases
Reactivity	Very Very Reactive	Very Reactive	Very Very Reactive	Very Unreactive
State of Matter at r.t.p & colour	Silvery Solids	Silvery Solids	Coloured Solids, Liquids or Gases	Colourless Gases
Melting and Boiling Points	Low compared to other metals	High	Low	Low
Other Properties	Densities are low compared to other metals		Reactivity decreases down the group	
	Soft and shiny			

Table 2.1.4.1 - Important Groups in the Periodic Table for CM1131

2.1.5 Valence Electrons

Valence Electrons refer to electrons in the outermost electron shell.

Group	No. of Valence Electrons
1	1
2	2
13	3
14	4
15	5
16	6
17	7
18	8 (Except Helium with 2)

Table 2.1.5.1 - Valence Electrons

Note: Transition metals are not tested in the syllabus for CM1131.

Take note that only **bromine** and **mercury** are liquids at room temperature; the rest are solids or gases.

Topic 3: Elements, Compounds and Mixtures

3.1 Physical and Chemical Changes

3.1.1 Differences Between Physical and Chemical Changes

Physical Changes	Chemical Changes
No new substances are formed	**New substances** are formed
Properties of products are the **same** as those of the reactants	Properties of products are **different** from those of the reactants
Usually **reversible**	Usually **irreversible**

Table 3.1.1.1 - Physical vs Chemical Changes

3.1.2 Chemical Reactions

During a chemical reaction, **heat and light** (or both) **can be taken in, or given out**.
A **word equation** consisting of reactants and products is used to represent a **chemical reaction**.

For example:

Sodium + Chlorine → Sodium Chloride (Common Salt/Table Salt)

There are 3 types of Chemical Reactions.

3.1.2.1 Combination

Combination refers to two or more substances reacting to form a single compound.

For example:

Iron + Sulfur → Iron (II) Sulfide

3.1.2.2 Combustion

Combustion refers to a substance being heated with oxygen present, releasing a lot of heat. When fuels burn in oxygen, large amounts of heat are released with the formation of carbon dioxide and water.

For example:

Methane + Oxygen → Carbon Dioxide + Water

Sulfur + Oxygen → Sulfur Dioxide

3.1.2.3 Decomposition

Decomposition refers to breaking down a single substance into simpler substances.

For example:

$$\text{Sugar} \rightarrow \text{Carbon} + \text{Water Vapour}$$

Decomposition also includes **Electrolysis**, which involves passing an electric current through a liquid or solution.

For example:

$$\text{Water} \rightarrow \text{Hydrogen} + \text{Oxygen}$$

3.1.3 Indicators of a Chemical Change

There are a few main indicators of a chemical change:

- Formation of a gas (bubbling or effervescence)
- Change in colour
- Change in temperature (heating or cooling)
- Precipitation (formation of a solid when two liquid react)

3.1.4 Law of Conservation of Mass

The Law of Conservation of Mass says, when a reaction is carried out in **a closed system**, the **reactants and products both MUST have the same total mass**.

3.2 Molecule

Molecules are made of atoms that are held together by **chemical bonds**, like water, as shown by *Figure 3.2.1* below.

Figure 3.2.1 - A Water Molecule

3.3 Element

A molecule of an element contains the atoms of the same type of element bonded together or just one atom of a certain element.

Hence, **ozone (O_3) is an element but water (H_2O) is not.**

Figure 3.3.1 - Ozone

3.4 Compound

A compound is a **pure substance** that contains **two or more different elements chemically combined**, like water and common salt, shown by *Figure 3.4.1* and *Figure 3.4.2*.

Figure 3.4.1 - Water

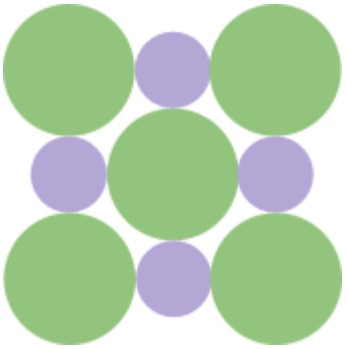


Figure 3.4.2 - Common Salt (Sodium Chloride)

In the case of water, molecules make up the compound.
In the case of sodium chloride, however, ions make up the compound.[verify - See Note Below]

Note: We only call Common Salt a compound in CM1131 Semester 1. As covered in Semester 2, we never call an ionically bonded substance a compound.

3.4.1 Chemical Formula

A compound can be represented with a **chemical formula**.

The chemical formula of a compound tells us **what the compound is made of and the ratio of its components**.

For example, take the chemical formula for Lead (II) Nitrate: **$Pb(NO_3)_2$**

There is(are): 1 Lead (Pb) atom, 1 × 2 = 2 Nitrogen (N) atoms, 3 × 2 = 6 Oxygen (O) atoms.

3.4.2 Composition of Compounds

Every compound has a **fixed composition** of the elements it contains. Hence, every compound has a constant **ratio of atoms present**.

3.4.3 Properties of Compounds

A compound has different properties from the elements it contains.

For example, take the chemical equation:

$$\text{Magnesium} + \text{Oxygen} \rightarrow \text{Magnesium Oxide}$$

Magnesium is a shiny grey solid, Oxygen is a colourless gas, while Magnesium Oxide is a white solid.

3.4.4 Decomposition of Compounds

Compounds can be broken down into simpler substances by **chemical methods**.

3.5 Diatomic and Polyatomic Molecules

A **diatomic** molecule is formed by **two atoms** chemically combined. (E.g. H_2, O_2, NaCl)

Certain elements (Hydrogen, Nitrogen, Oxygen, Fluorine, Chlorine, Bromine, Iodine)

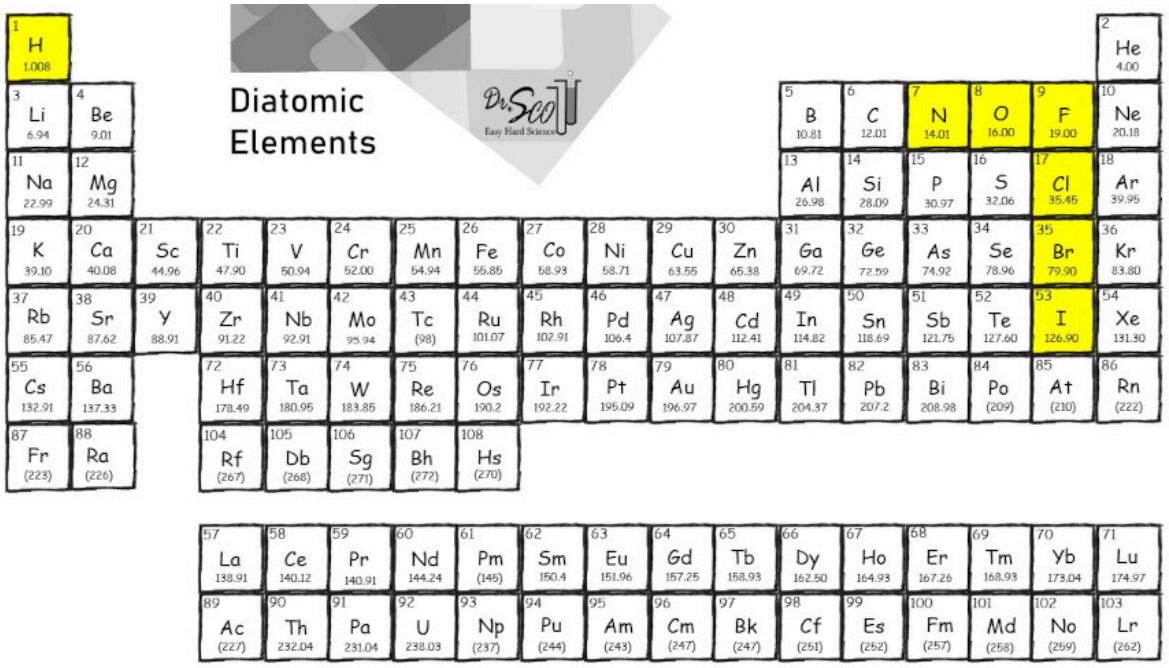


Figure 3.5.1 - Diatomic Molecules

A **polyatomic** molecule is formed by **three or more atoms** chemically combined. (E.g. O_3, H_2O, P_4, S_8)

3.6 Mixtures

Mixtures are made up of **two or more substances** that are **NOT chemically combined**, like air.

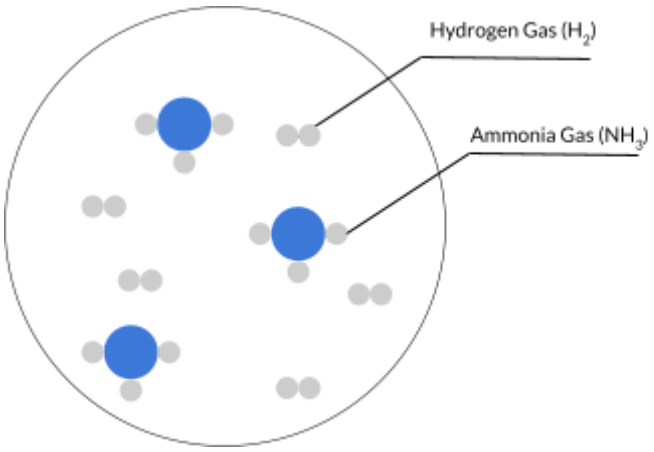


Figure 3.6.1 - A mixture of hydrogen and ammonia gas

As shown, a mixture can be made of an element and a compound.
It can also be made of two elements or two compounds.

3.6.1 Composition of Mixtures

A mixture does **not have a fixed composition** of the substances it contains, so the components can be in **any proportion**.

3.6.2 Properties of Mixtures

A mixture has **similar properties** to its components.

3.6.3 Separation of Mixtures

A mixture can be separated into its components by physical methods. This is covered later in **Topic 4: Separation and Purification of Substances**.

Topic 4: Separation and Purification of Substances

4.1 What is a Pure Substance?

A pure substance is one that consists of **only one type of particle**.
Hence, a mixture is **NOT** a pure substance.

4.2 Homogeneous and Heterogeneous Mixtures

A homogeneous mixture is basically a well-mixed mixture, in which you cannot tell the boundaries of two substances.

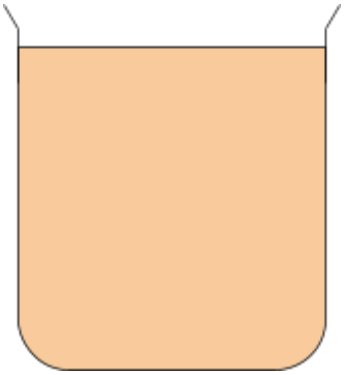


Figure 4.2.1 - A homogeneous mixture

A heterogeneous mixture is a mixture in which there is a **clear boundary** between two substances.

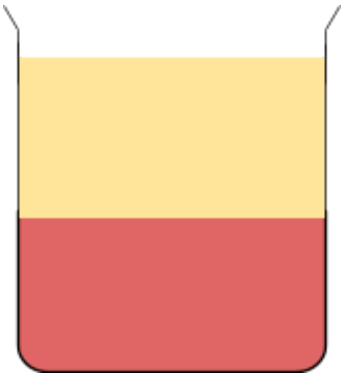


Figure 4.2.2 - A heterogeneous mixture

4.3 How do we check for the Purity of a Substance?

We can check the melting and boiling points of the substance.

Impurities **decrease the melting point and increase the boiling point**.

Pure Substance	
Impure Substance	

Figure 4.3.1 - Pure vs Impure Substance Melting and Boiling Points

Furthermore, we can use Chromatography as covered in **Section 4.10 - Chromatography**.

4.4 Immiscible and Miscible

A mixture is said to be immiscible when **two liquids form two different layers** (is heterogeneous). This means it does not mix.

A mixture is said to be miscible when either of the following is true:

1. A **solid is soluble in a liquid** and is dissolved in it.
2. **Two or more liquids mix** and form a homogeneous mixture.

4.5 Filtration

When you have a solid that is insoluble in a liquid, filtration is the main method to separate the two.

4.5.1 Setup

We fold the filter paper as such:

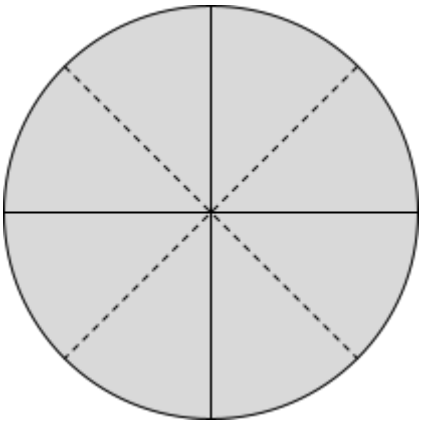


Figure 4.5.1.1 - How to Fold Filter Paper

This is known as a “fluted filter paper”.

Figure 4.5.1.2 below shows the setup for filtration.

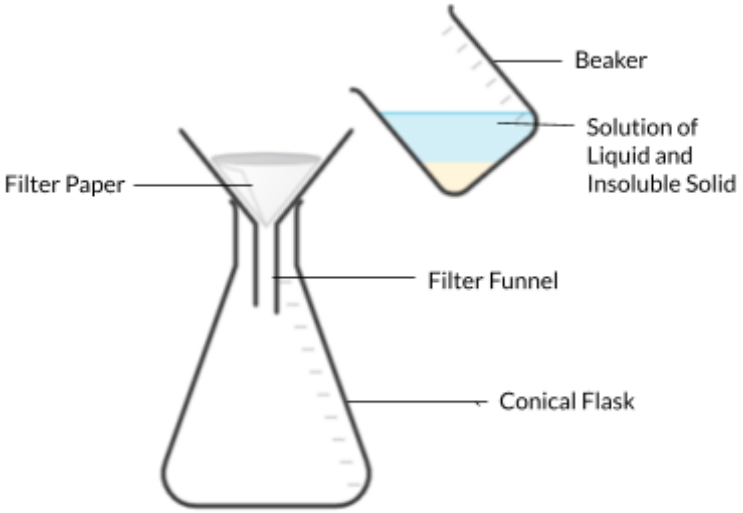


Figure 4.5.1.2 - Filtration Setup

After filtration, the residue will be left on the filter paper and the filtrate will be collected in the conical flask, as seen in *Figure 4.5.1.3* below.

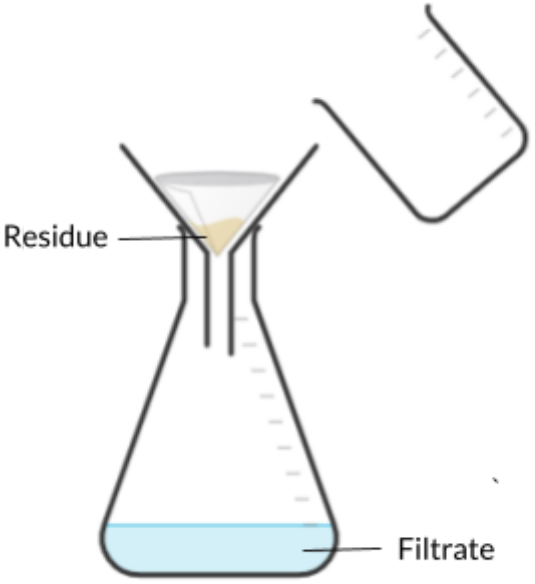


Figure 4.5.1.3 - Filtration Setup after Filtration

Note that residue refers to the **insoluble solid** and filtrate refers to the **liquid**.

4.6 Decanting

When you have a solid that is insoluble in a liquid, decanting can be used. Decanting is simply, in a nutshell, pouring away the liquid without letting the solid spill. This, of course, causes the liquid to not be as pure as other methods such as filtration.

Note: In general, decanting is NEVER used because of the impurity of the substances obtained.

4.6.1 Setup

This is the setup for decanting.

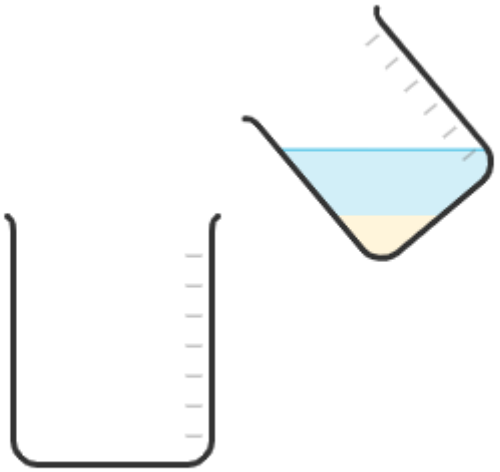


Figure 4.6.1.1 - Setup for Decanting

4.7 Centrifugation

When you have a solid that is insoluble in a liquid, centrifugation can be used.

Centrifugation is used to separate plasma and red blood cells, typically.

4.8 Evaporation to Dryness

When you have a solution of a solid dissolved in a liquid, we can use evaporation to dryness if we want to obtain the solid.

4.8.1 Setup

This is the setup for Evaporation to Dryness.

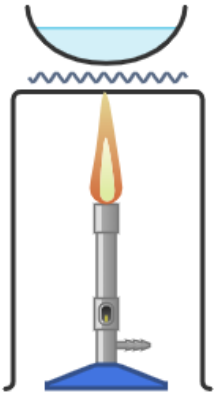


Figure 4.8.1.1 - Setup for Evaporation to Dryness

It is most commonly used with a solution of salt in water to obtain the salt. We heat the solution until all the water has evaporated, leaving just the solid salt crystals.

However, for certain substances that **decompose on heating**, evaporation to dryness is ineffective for these non-heat-stable substances like sugar.

4.9 Crystallisation

When we have a solution of a solid dissolved in a liquid, we can use crystallisation if we want to obtain the solid. Crystallisation works because **the solubility of most solids decreases as temperature decreases**.

Using this, we can heat the solution to evaporate off some of the liquid, and **upon cooling, the excess solid dissolved in the liquid will be precipitated out as crystals**.

Some water should be left in the evaporating mixture when aqueous solutions (aqueous solution: a solution in which the solvent is water) of salts are being crystallised as the salt crystals need water of crystallisation in the solid structure to form a hydrated salt.

4.10 Chromatography

Chromatography is the technique used to separate two or more substances.

Paper chromatography involves two **phases:** the **mobile phase** (i.e. the solvent) and the **stationary phase** (i.e. the paper).

4.10.1 Setup

This is the setup for paper chromatography.

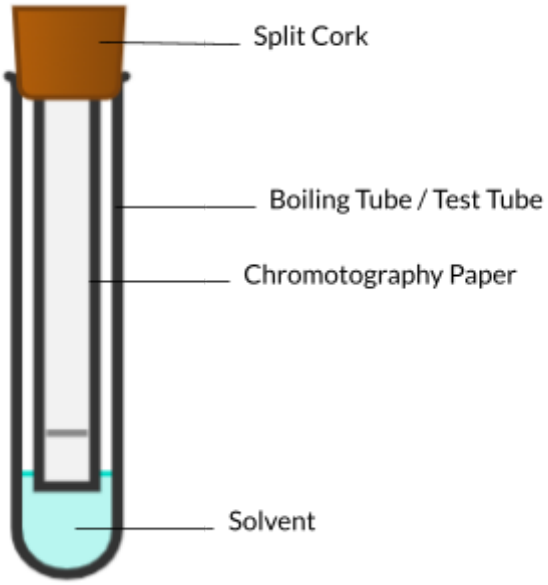


Figure 4.10.1.1 - Setup for Chromatography

4.10.2 Precautions for Chromatography

There are many precautions when doing chromatography.

1. The substances **must be soluble in the same solvent**.

2. The starting line must be **drawn with a pencil**.

   Otherwise, the ink used to draw the starting line will also undergo chromatography and interfere with the results.

3. The solvent front must not **exceed the end of the paper**.

   Otherwise, the R_f values cannot be accurately calculated as we would not be able to accurately estimate the distance travelled by the solvent.

4. **Not too much of the substance** should be added to the paper.

   Otherwise, the substance might interfere with the other substances on the same paper and create a large stain.

4.10.3 R_f Value

The R_f value, or retention factor value of one of the substances is calculated by:

$$R_f = \frac{Distance\ travelled\ by\ the\ substance}{Distance\ travelled\ by\ the\ solvent}$$

The R_f value is greater or equal to 0, but always less than 1.

R_f value **is usually** left to 3 s.f.
However, it can also be left to 2 s.f. when the smallest division given is 1 cm.

In general, R_f should be given to 1 more than the sf of the smallest division of the scale.[verify]

4.10.4 How does it work?

The paper chromatogram works because substances have **different solubilities** in a specific solvent, and a substance with a higher solubility will move up faster and further. This separates the substances.

4.11 Distillation

There are 2 types of distillation - Simple Distillation and Fractional Distillation.

4.11.1 Simple Distillation

When you have a solution of solids dissolved in a liquid, and you want to obtain the liquid, you can use Simple Distillation. The liquid obtained is then the **distillate**.

4.11.1.1 Setup

Figure 4.11.1.1.1 below shows the setup for simple distillation.

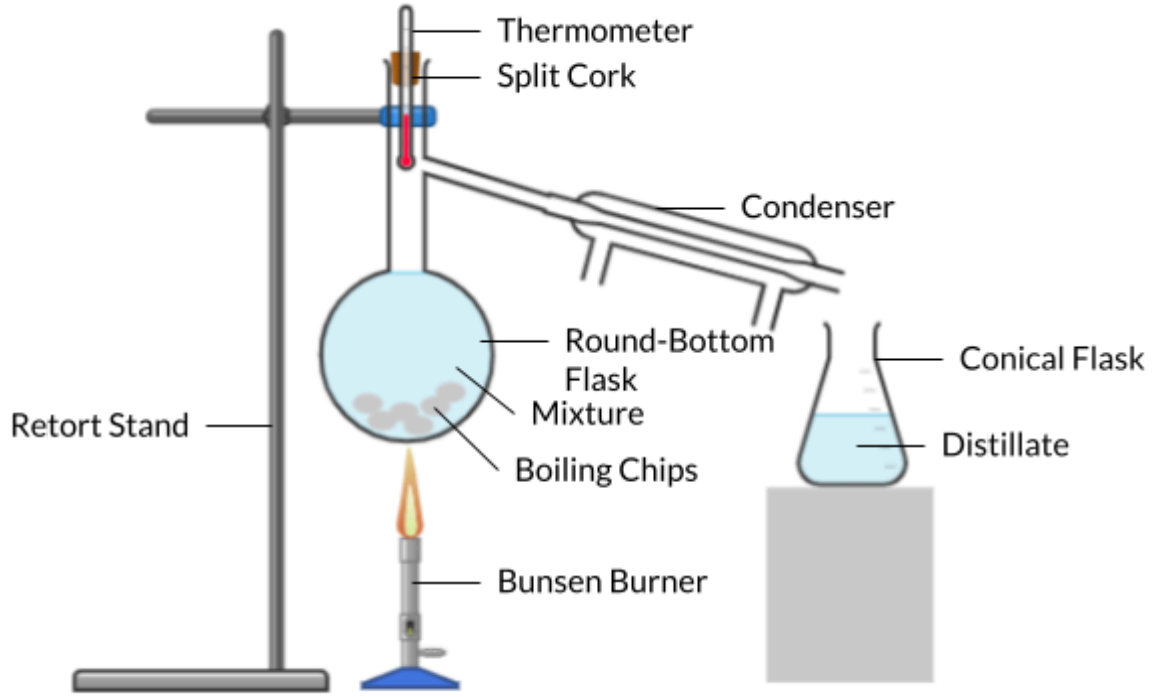


Figure 4.11.1.1.1 - Simple Distillation Setup

4.11.2 Fractional Distillation

When you have a mixture of two miscible liquids, we can use fractional distillation to separate the two.

Some common examples include alcohol and water, and ethanol and water.

4.11.2.1 Setup

The setup of Fractional Distillation is the same as that of Simple Distillation, but with an extra fractionating column, as in *Figure 4.11.2.1.1* below.

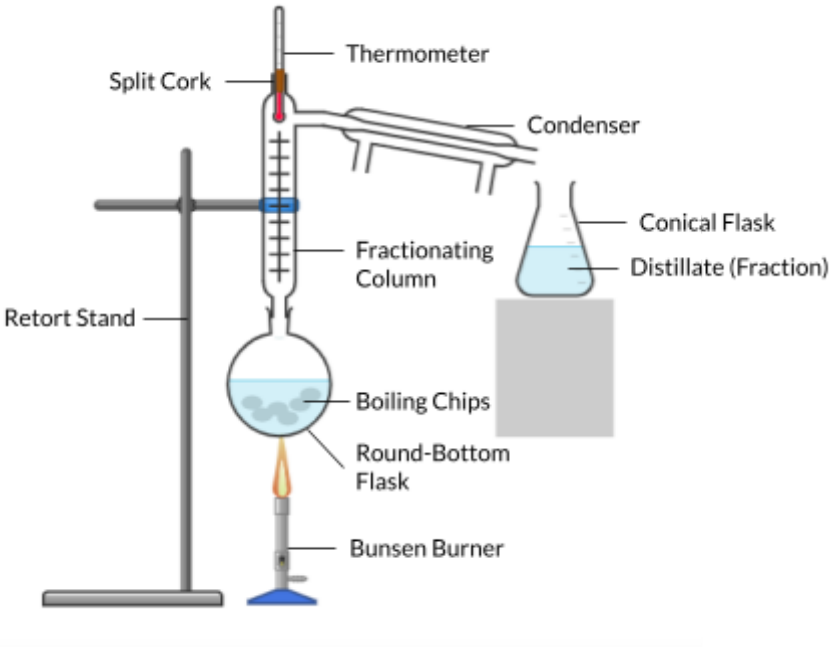


Figure 4.11.2.1.1 - Fractional Distillation Setup

4.12 Separation Using a Separating Funnel

Liquids that are immiscible can be separated using a Separating Funnel.

The less dense liquid will float on top, while the denser liquid will sink to the bottom. This allows the liquids to be separated into different containers by opening a tap attached on the separating funnel.

4.12.1 Setup

This is the setup for the separating funnel.

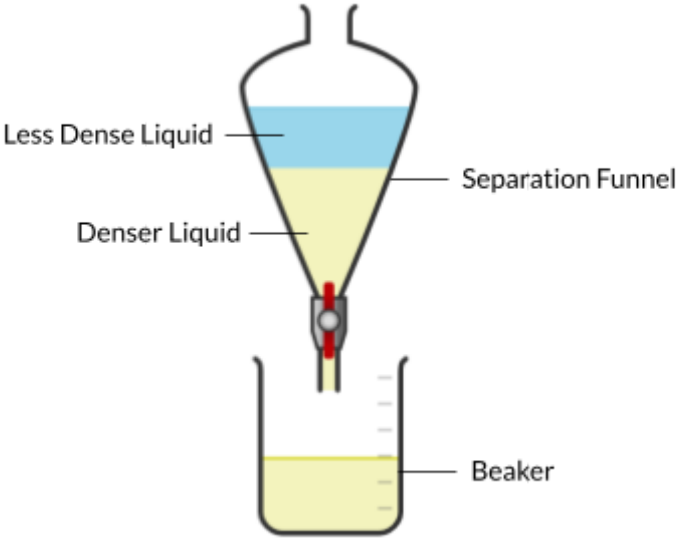


Figure 4.12.1.1 - Setup for Separation Funnel

4.13 Separation Using Magnets

We can separate a magnetic material from a non-magnetic material using a magnet.

4.14 Separation by Sublimation

Certain solids can sublimate on heating. The two solids that we are required to know that sublime are **ammonium chloride (NH_4Cl) and iodine (I_2)**.

When one of these solids is together with a solid that does not sublime, we can separate these solids by sublimation.

By heating the mixture, the solid that sublimes will gain heat and sublime. When it then comes into contact with the cooler inner surface of the filter funnel, it loses heat and deposits on the funnel, returning to its solid state.

The solid deposited is called the **sublimate**.

4.14.1 Setup

This is the setup for Separation by Sublimation.

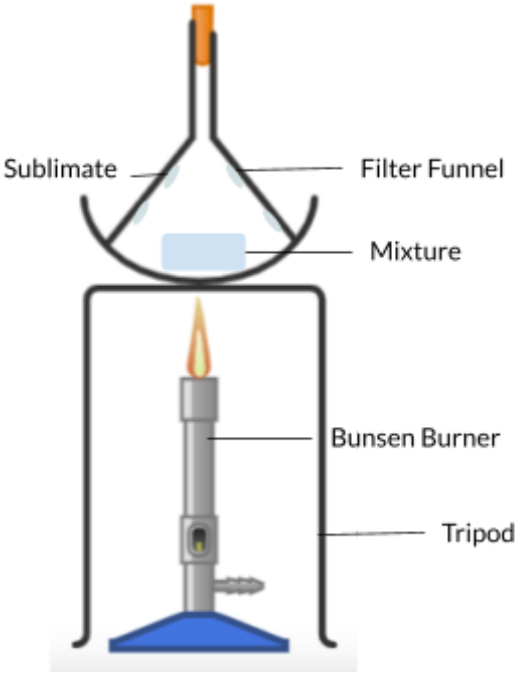


Figure 4.14.1.1 - Sublimation Setup

Topic 5: Atomic Structure

5.1 What is inside of an Atom?

According to the Bohr's model of an atom, inside every atom, we find a nucleus at its centre, with electrons orbiting around in electron shells.

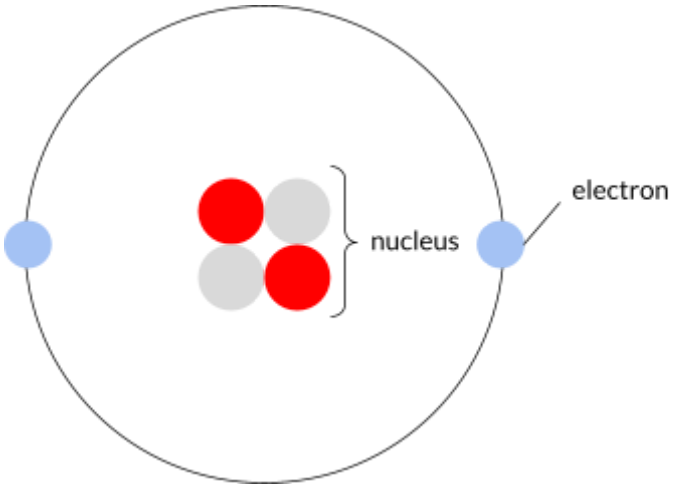


Figure 5.1.1: A helium-4 atom

Note: In an actual atom, the electrons actually orbit extremely far away from the nucleus, and hence the atom is almost fully empty space.

5.2 Protons, Neutrons and Electrons

The atom is made up of 3 types of particles, called **subatomic particles**, namely the proton, neutron and electron. Here are some of their properties:

	Proton (p^+)	Neutron (n^0)	Electron (e^-)
Relative Mass	1	1	$\frac{1}{1840}$
Relative Charge	+1	0	-1

Table 5.2.1 - Properties of Subatomic Particles

Note: The masses and charges of these subatomic particles are so small that we use *Relative* Masses and Charges to describe them, so there are no units. Furthermore, since the relative mass of an electron is so tiny, it is often negligible when performing calculations.

It is also **NECESSARY** to include the signs in front of charges (i.e. +1 for proton instead of 1)

5.2.1 Subatomic Particles under an Electric Field

Since opposite charges attract, different subatomic particles behave differently under an electric field.

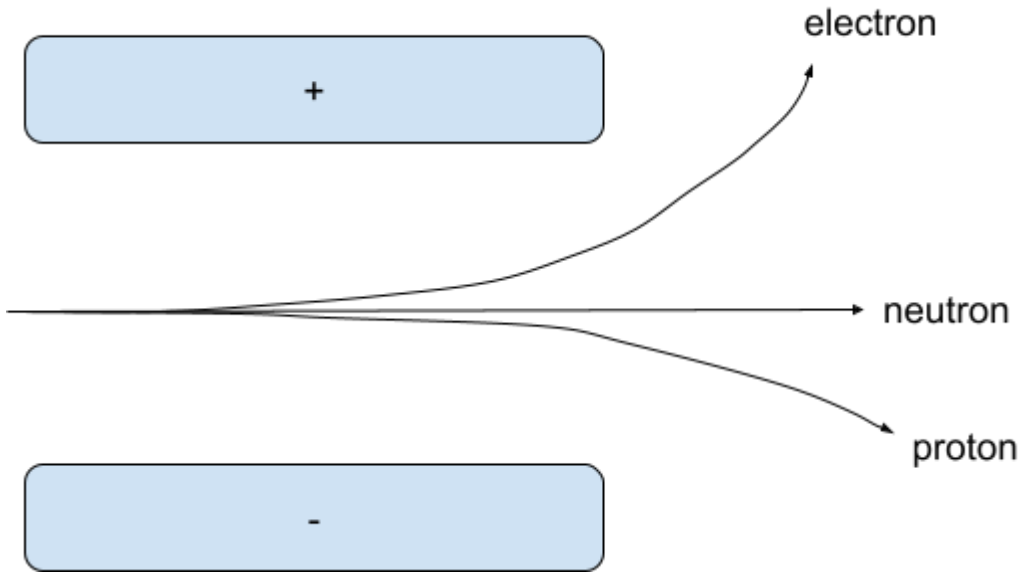


Figure 5.2.1.1: Subatomic Particles under an Electric Field

In *Figure 5.2.1.1*, we can see that an electron is strongly deflected from its original path towards the positively charged region of the electric field, as opposite charges attract. It is deflected the most due to its lower mass and hence lower inertia, so it can be deflected more.

A neutron remains on its original path as it has no charge, so it is unaffected by the electric field.

A proton is also deflected from its original path towards the negatively charged region of the electric field as opposite charges attract. However, it is deflected less than the electron as it has a greater mass and hence higher inertia.

5.3 Proton and Nucleon Number

An atom is identified using its Proton and Nucleon number.

5.3.1 Proton Number

The proton number, or atomic number, refers to the number of protons in an atom's nucleus. This number is what defines the identity of an atom. For example, carbon has 6 protons, so any atom that has 6 protons is, by definition, carbon, and vice versa.

5.3.2 Nucleon Number

The nucleon number, or mass number, refers to the total number of protons and neutrons in an atom's nucleus.

5.3.3 Notation - Nucleon and Proton Number

We denote an element as such:

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

Where X is the element's Chemical Symbol, A is the Nucleon Number, and Z is the Proton Number. For example, we write ${}^{4}_{2}He$.

For convenience, sometimes it is written as ${}^{A}X$, or X-A, for example: ${}^{4}He$ or Helium-4.

5.4 Isotopes of an Element

Isotopes are atoms of the same element with the **same numbers of protons and electrons** but **different numbers of neutrons**.

Pause and Ponder 5.4

Why are isotopes considered the same element?

[Click to Go to Solution]

The word isotopes is derived from the words "iso" (same) and "topos" (place), which means that different isotopes of an element occupy the same place on a periodic table.

For example, Neon (Ne) has three main isotopes, Neon-20, Neon-21 and Neon-22. These isotopes are written as ${}^{20}_{10}Ne$, ${}^{21}_{10}Ne$ and ${}^{22}_{10}Ne$ respectively.

Isotopes have the **same chemical properties** but slightly different **physical properties**.

Since only electrons (or specifically, valence electrons) participate in chemical reactions, isotopes have the same chemical properties due to their common number of electrons.

5.4.1 Natural Abundance and Average Atomic Mass

Natural Abundance refers to how many percent of a random sample in nature of a specific element belongs to a certain isotope.

For example, around 905 neon atoms randomly chosen from 1000 would be Neon-20, or $^{20}_{10}Ne$, so the abundance of Neon-20 is around 90.5%.

The Average Atomic Mass is calculated by taking a **weighted average** of the atomic masses of the isotopes, based on their natural abundance.

In general, for each isotope, we take its Atomic Mass, multiply by its abundance (in percent), then sum up all the masses throughout the atom's isotopes, then divide by 100%.

Let's use Neon as an example.

Pause and Ponder 5.4.1

Given Neon has 3 stable isotopes, Neon-20, Neon-21 and Neon-22, and that they have a natural abundance of 90.48%, 0.27% and 9.25% respectively, calculate the Average Atomic Mass of Neon.

[Click to Go to Solution]

5.5 Electron Arrangement

Electrons are arranged in shells going outwards from the nucleus.

5.5.1 Recap - Valence Electrons

The valence electrons lie in the **valence shell** (or outer shell), which is the outermost shell from the nucleus. Since all the inner shells have to be filled up, **valence (outer) electrons** are more of importance to chemists. These valence electrons are the key to all **chemical reactions.** The number of valence electrons also determines whether they tend to give up or gain electrons in chemical reactions.

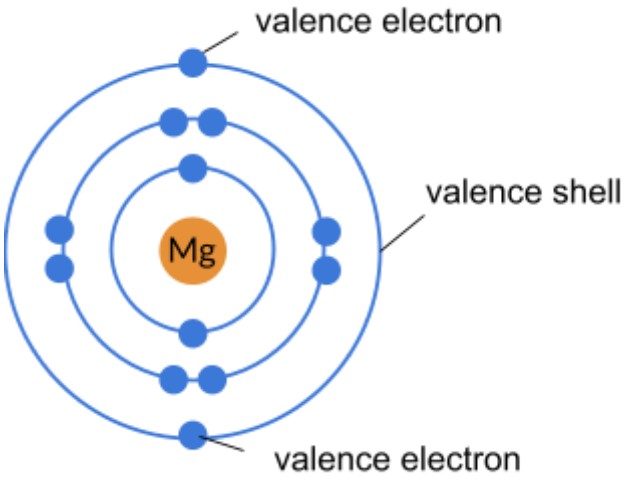


Figure 5.5.1.1 - Valence Electrons in a Magnesium Atom

5.5.2 Notation - Electronic Configuration

The electronic configuration of an atom is written by counting the number of electrons in each shell, starting from the innermost shell and going outwards, with a dot (.) separating each number.

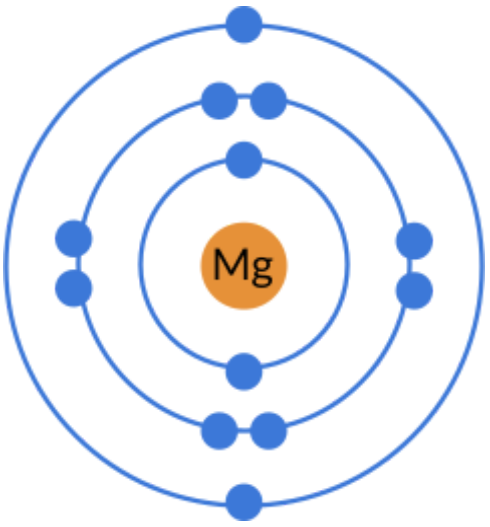


Figure 5.5.2.1 - Electronic Configuration of a Magnesium Atom

For example, a Magnesium Atom has 2 electrons in the first (innermost) shell, 8 in the second, and 2 in the third (valence) shell.

Hence, Magnesium has the electronic configuration: **2.8.2**

5.5.3 Noble Gas Structure

Noble Gases (Elements in Group 18 of the Periodic Table), have a Octet Valence Electron Configuration, which means it has 8 electrons in its outer shell.

There is an exception - Helium (He) has only 2 electrons in its outer shell, and has what is called a Duplet Valence Electron Configuration.

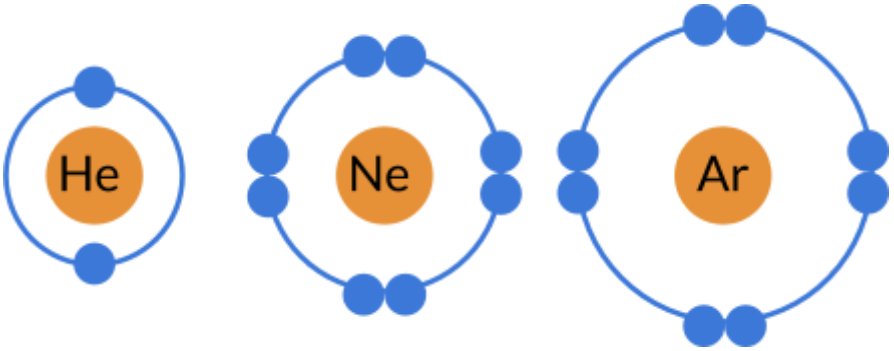


Figure 5.5.3.1 - Examples of Noble Gases and Their Valence Shell Structures

5.5.4 Formation of Cations

Metals (those from Groups 1, 2, 13) usually form Cations, or positive ions.
Take the example of Magnesium, which is from Group 2.

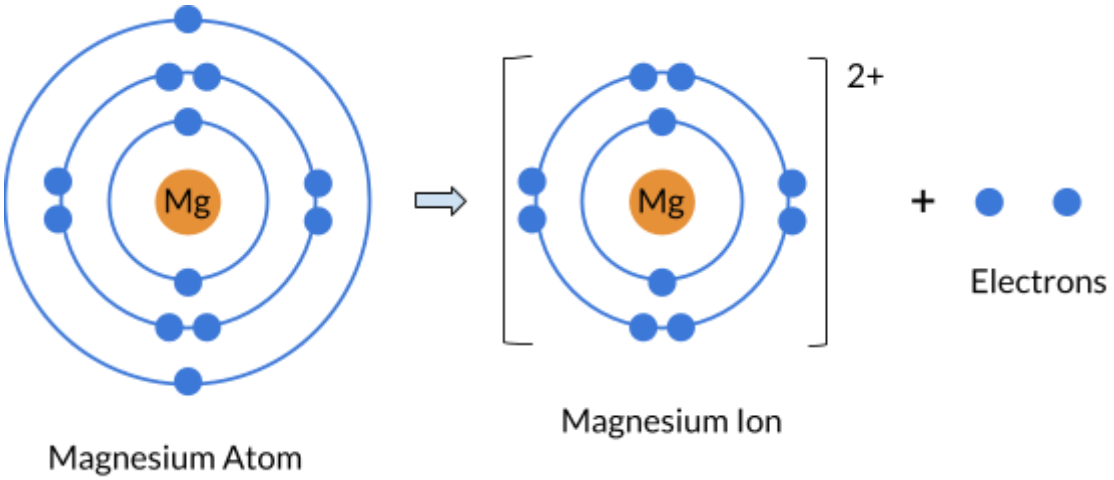


Figure 5.5.4.1 - The formation of Magnesium Ions

The Magnesium Atom loses 2 electrons from its valence shell to obtain a stable noble gas valence shell electronic structure, to form a Magnesium Ion. Since 2 electrons were lost, the magnesium ion lost 2 × (+1) relative charge, which means it now has a charge of +2. These extra electrons can be "given" to other atoms to form ionic compounds (more about this at **Topic 6: Structure and Bonding (Ionic)**)

The Magnesium Ion is then denoted as Mg^{2+}.

Note: When writing the charge, we write +2, but when writing the ion itself, we write 2+ instead of +2.

Pause and Ponder 5.5.4a

Since Aluminium is in group 13, what is the charge on an Aluminium Ion?

[Click to Go to Solution]

5.5.5 Formation of Anions

Non-metals (those from groups 15, 16, 17) usually form anions, or negative ions.

These ions' names are modified to end in **-ide**. (E.g. Oxygen → Ox**ide**, Chlorine → Chlor**ide**)

Take the example of Oxygen, which is from group 16.

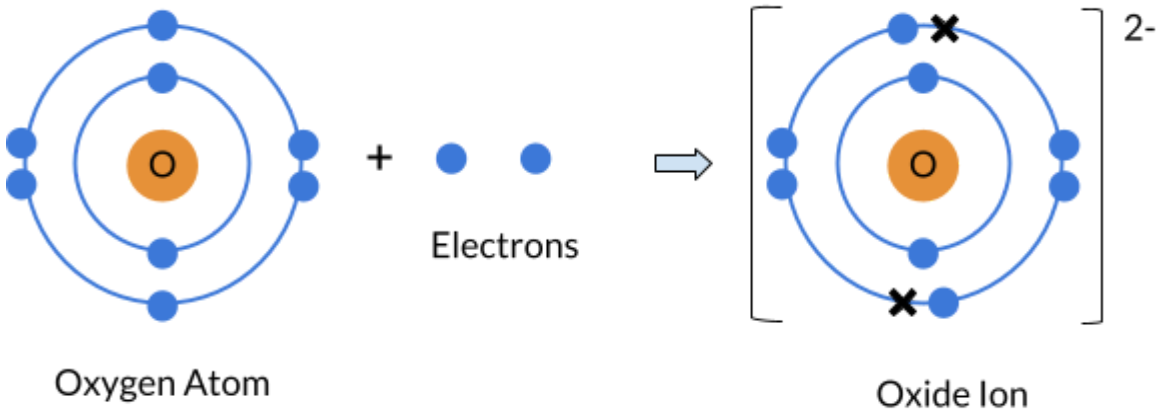


Figure 5.5.5.1: Formation of an Oxide Ion

The Oxygen gains 2 electrons in its valence shell to obtain the stable valence shell configuration of a noble gas, to form an Oxide Ion. Since two electrons were gained, the oxide ion gained 2 × (-1) relative charge, which means now it has a charge of -2.

Since 2 electrons were gained from a different atom, the 2 extra electrons are denoted with ..crosses (×). See **Section 5.6: Order to Fill Electrons** for the order to fill electrons into a shell.

The Oxide Ion is then denoted O^{2-}.

Note: Again, when writing the charge, we write -2, but when writing the ion itself, we write 2- instead of -2.

Pause and Ponder 5.5.4b

Since Chlorine is in group 17, what is the charge on an Chloride Ion?

[Click to Go to Solution]

Here is a table of anion names:

Atom	Anion
Hydrogen (H)	Hydr**ide** (H^-)
Nitrogen (N)	Nitr**ide** (N^{3-})
Oxygen (O)	Ox**ide** (O^{2-})
Fluorine (F)	Fluor**ide** (F^-)
Phosphorus (P)	Phosph**ide** (P^{3-})
Sulfur (S)	Sulf**ide** (S^{2-})
Chlorine (Cl)	Chlor**ide** (Cl^-)

5.6 Order to Fill Electrons Into an Electron Shell

When drawing a shell of electrons, each "side" of the shell will be filled with 1 electron first, before electrons start forming pairs.

Refer to *Figure 7.6.1* below.

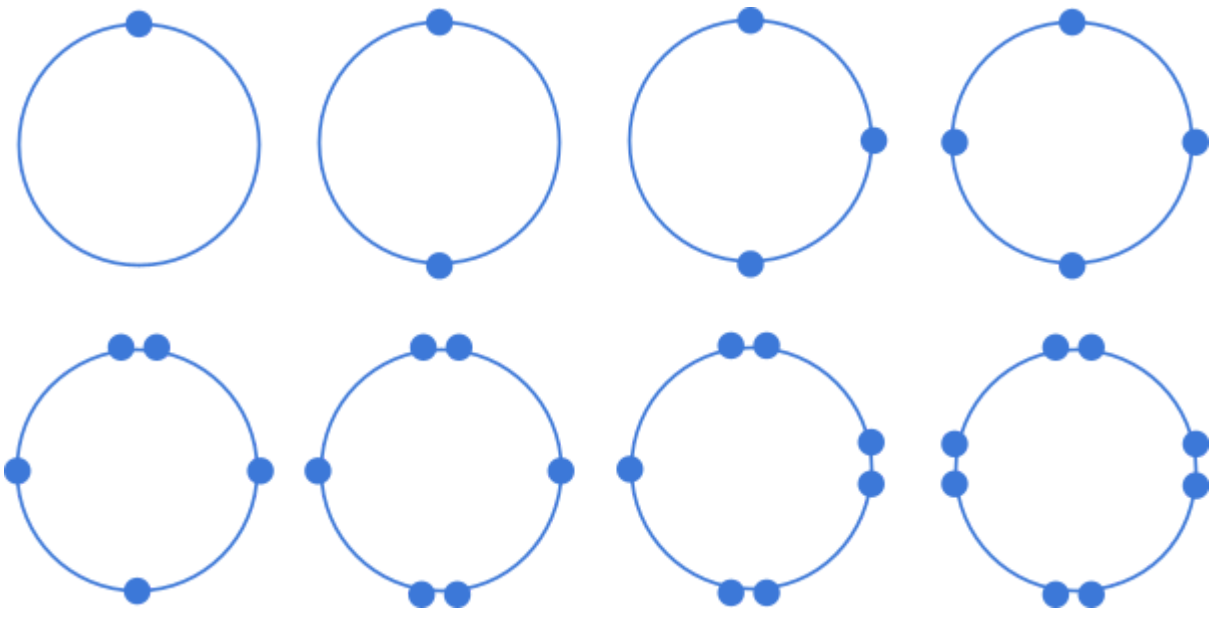


Figure 5.6.1 - Electron shells with 1 to 8 Electrons Respectively

Note that this also applies to dot and cross diagrams (Refer to *Section 6.2*). For now, know that something like *Figure 5.6.2* below is wrong.

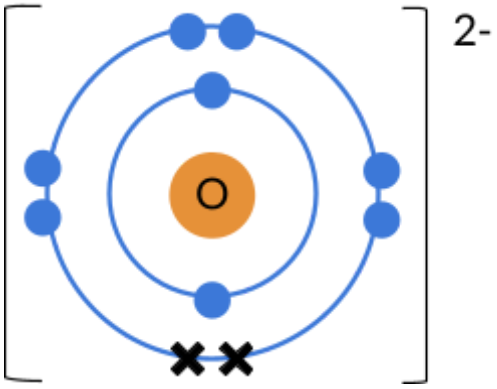


Figure 5.6.2: Wrong Electronic Configuration of Oxide Ion

Instead, before forming pairs of electrons from the original oxygen atom, one electron is assigned to each side first, like in *Figure 5.6.3* below.

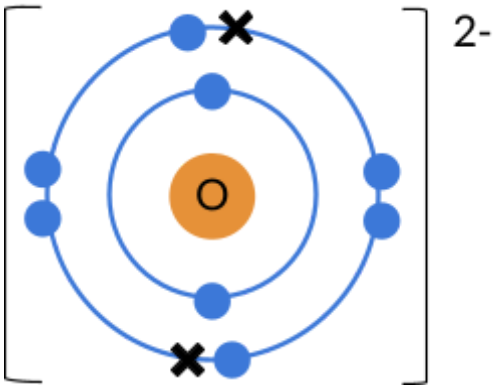


Figure 5.6.3: Correct Electronic Configuration of Oxide Ion

Topic 6: Structure and Bonding (Ionic)

6.1 How is an Ionic Bond formed?

Ions do not usually form in isolation. Instead, electrons are transferred from a metal atom to a non-metal atom, forming a metal ion and a non-metal ion. These two ions are then oppositely charged.

Since oppositely charged things attract, the ions are brought together as a result of strong electrostatic forces of attraction. As a result, an Ionic Bond is formed.

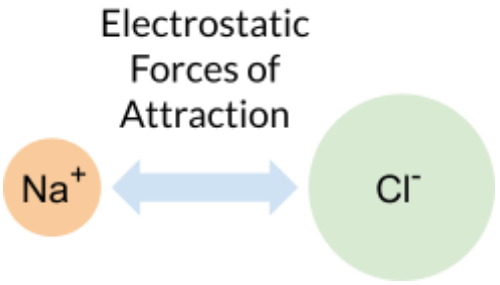


Figure 6.1.1: Electrostatic Forces of Attraction in a Formula Unit of NaCl

However, ionic bonds typically aren't alone. Ionic Compounds have a **giant ionic lattice**,

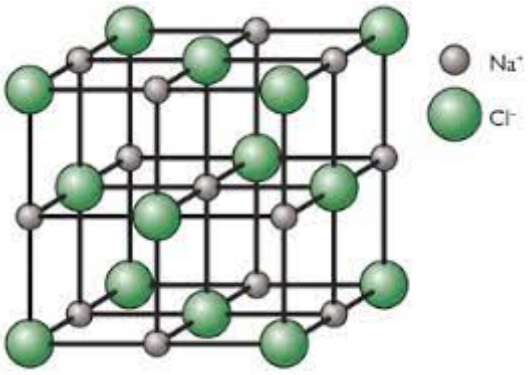


Figure 6.1.2: Giant Ionic Lattice of NaCl

As we can observe from *Figure 6.1.2*, the ratio of Sodium ions to Chlorine ions is 1:1 in sodium chloride, making the formula unit (which can be thought of as the base building block of a lattice) NaCl.

6.2 Dot and Cross Diagrams

We can use dot and cross diagrams to visualise the arrangement of electrons when an ionic bond is formed.

When drawing a dot and cross diagram, only the valence electrons of the non-metal element is drawn, and no electron shells need to be drawn. Also, only one **formula unit** of the compound needs to be drawn.

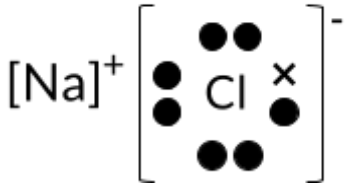


Figure 6.2.1: Dot and Cross Diagram of NaCl

With reference from *Figure 6.2.1*, we can see that the metal should be written before the non-metal, and the charges should be included outside square brackets.

The electron that was transferred from the sodium ion to chlorine ion is labelled with a cross, while all the other electrons are labelled with dots, to show the different ions that the electron originated from.

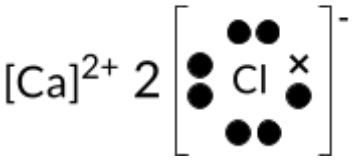


Figure 6.2.2: Dot and Cross Diagram of $CaCl_2$

If more than 1 of an element is present in a compound, we simply add a number before the square brackets, as in *Figure 6.2.2*.

6.3 Deriving Chemical Formula of an Ionic Compound

Write down the ions with their charges.
We then move the numbers diagonally and remove the charge, to find the compound, as shown in *Figure 6.3.1*.

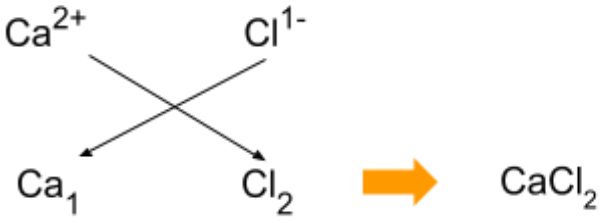


Figure 6.3.1: Finding Chemical Formula of Calcium Chloride

6.4 Important Ions to Remember

Here is a table of Important Ions to remember.

The ions in orange are most important to memorise, while those in grey are likely not tested, so you don't need to memorise them.

Charge	Name	Chemical Formula
+1	Ammonium	NH_4^+
	Silver	Ag^+
+2	Zinc	Zn^{2+}
-1	Hydroxide	OH^-
	Nitrate	NO_3^-
	Nitrite	NO_2^-
	Permanganate	MnO_4^-
-2	Sulfate	SO_4^{2-}
	Carbonate	CO_3^{2-}
	Sulfite	SO_3^{2-}
	Chromate	CrO_4^{2-}
	Dichromate	$Cr_2O_7^{2-}$
-3	Phosphate	PO_4^{3-}

Table 6.4.1 - Important Ions

Note: If an element has a roman numeral after it, like Iron (III), the roman numeral refer to its charge. (***FYI:*** *To be more precise, it refers to the oxidation number, which is covered in* ***CM2131 Section 18.3.2: Oxidation Number***) For example, Iron (III) refers to the ion Fe^{3+}.

6.5 Hydrated Ionic Compounds

Some ionic compounds, referred to as **hydrates**, give a product of water when heated. Each hydrated ionic compound has a certain number of water molecules for every formula unit that are trapped within the ionic lattice.

For example, take the compound Copper (II) Sulfate Pentahydrate.

It's formula is $CuSO_4 \cdot 5H_2O$.

A dot (·) is attached before the water molecules in the formula. Note that the dot is in the middle. (So it's not a fullstop)

To show the number of water molecules, a greek prefix is attached.

This is the table of prefixes used:

Number	Prefix
1	Mono-
2	Di-
3	Tri-
4	Tetra-
5	Penta-
6	Hexa-
7	Hepta-
8	Octa-
9	Nona-
10	Deca-

Table 6.5.1 - Greek Prefixes

6.6 Physical Properties of Ionic Compounds

The strong electrostatic forces of attraction within ionic compounds give it certain physical properties.

6.6.1 Melting Point

Recall from Section 1.1 The Kinetic Particle Theory, we covered that for a compound to melt, the molecules inside the compound will be able to slide over each other.

However, in an ionic lattice, for the ions to be able to slide over each other, **the strong ionic bonds must be broken**. However, since there is a **strong electrostatic force of attraction**

between the **oppositely charged ions**, **a lot of energy is required to overcome** these forces and hence Ionic Compounds have a high melting point.

Main Points:

1. The Ionic Compound has a giant ionic lattice.
2. Strong electrostatic force of attraction between oppositely charged _____ and _____ ions.
3. A lot of energy is required to overcome these forces.

Plus, the **larger the charge**, the greater the electrostatic forces of attraction, the more energy required to overcome those forces, **the higher the melting point**.

For example, MgO (From a 2+ and a 2- ion) has a much higher melting point of 2852 °C as compared to NaCl (From a 1+ and a 1- ion), which has a melting point of 801 °C.

Furthermore, **the smaller the ions involved** in the ionic compound, the more tightly packed the ions are, so the greater the electrostatic forces of attraction, so the more energy required to overcome those forces and **the higher the melting point.**

6.6.2 Electrical Conductivity of Ionic Compounds

Ionic compounds do not conduct electricity in the solid state, as the **ions are in fixed positions in the ionic lattice**, so there are **no mobile charge carriers**, so it **cannot carry a charge**, and hence the compound doesn't conduct electricity.

However, in the liquid and aqueous(when dissolved in water) states, they **can conduct electricity**. In the liquid and aqueous states, the ions are freed from the lattice and **there are mobile charge carriers** to carry a charge, so the compound can conduct electricity.

Pause and Ponder 6.6

Why (and how) are the ions freed from the lattice in the liquid and aqueous states?

[Click to Go to Solution]

6.6.3 Solubility of Ionic Compounds

Why the properties under this section are how they are need not be explained during the exam, but you should still know the properties themselves. In fact, they are covered in **CM3131 Section 26.7: Ion-Dipole Forces**

In solubility, there is a trick called “Like dissolves like”.

Compounds with the same type of bonding are soluble with one another.

In general, polar solvents can dissolve ionic compounds like NaCl.

However, since most organic solvents like ethanol and hexane are covalent, it cannot dissolve ionic compounds.

In Summary, Ionic Compounds are soluble in polar solvents like water but not organic solvents like ethanol or hexane.

Topic 7: Structure and Bonding (Simple Molecular)

7.1 What is a Covalent Bond?

A Covalent Bond is formed when two **elements** or **compounds** share electrons.

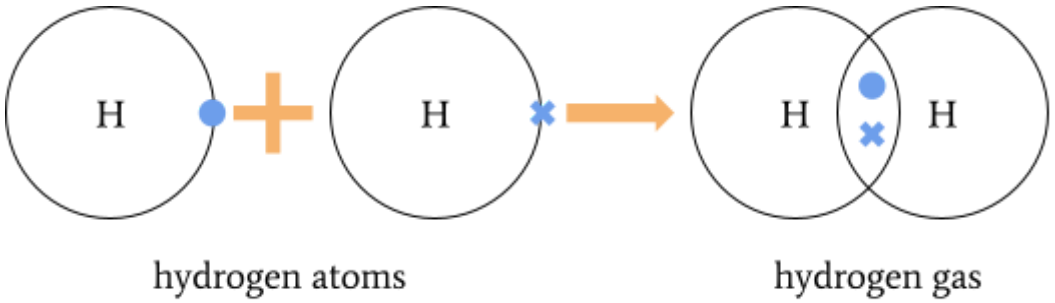


Diagram 7.1.1 - Formation of Hydrogen Gas

Note that the dot and the cross denote electrons from **different atoms**.

The two electrons in between the two hydrogens are the **shared electrons**. Since a pair of electrons are shared, a **single covalent bond** is formed.

We can denote this covalent bond as $H - H$.

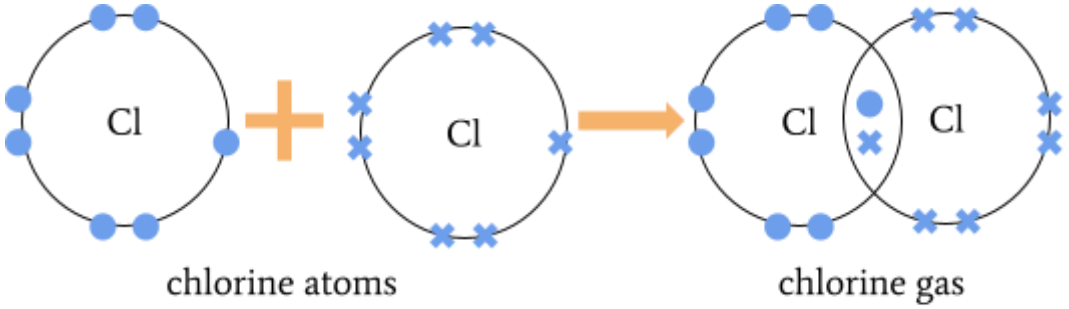


Diagram 7.1.2 - Formation of Chlorine Gas

Again, since one pair of electrons is shared in this case, we denote this bond as $Cl - Cl$.

7.2 Covalent Bond Forming Compounds

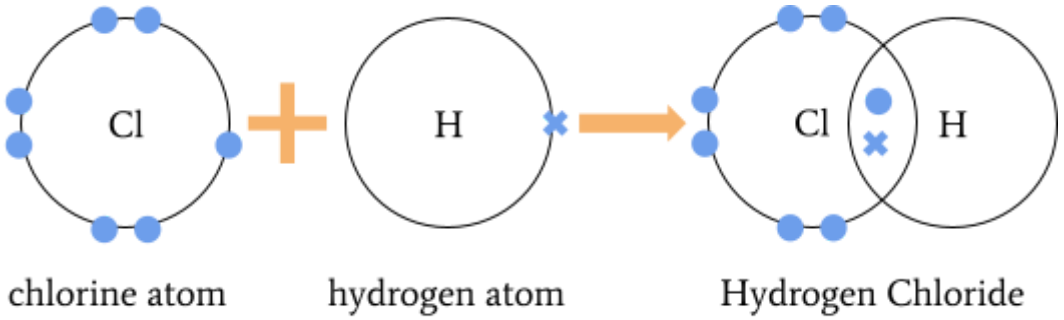


Diagram 7.2.1 - Formation of Hydrogen Chloride

7.3 Structural Formula

The Structural Formula is simply **disregarding all electrons not involved in bonding**, and replacing **pairs of electrons involved with covalent bonds** with a dash.

When two pairs of electrons are shared, we use a double dash, and so on.

Water	$H - O - H$
Ammonia	$\begin{array}{c} H \\ \vert \\ H - N - H \end{array}$
Carbon Dioxide	$O = C = O$
Methane	$\begin{array}{c} H \\ \vert \\ H - C - H \\ \vert \\ H \end{array}$

Table 7.3.1 - Some Common Structural Formulae

Refer to the example of water above. Each dash between O and H actually refers to oxygen and hydrogen each donating one electron to share. How caring!

7.4 Simple Molecular Structure

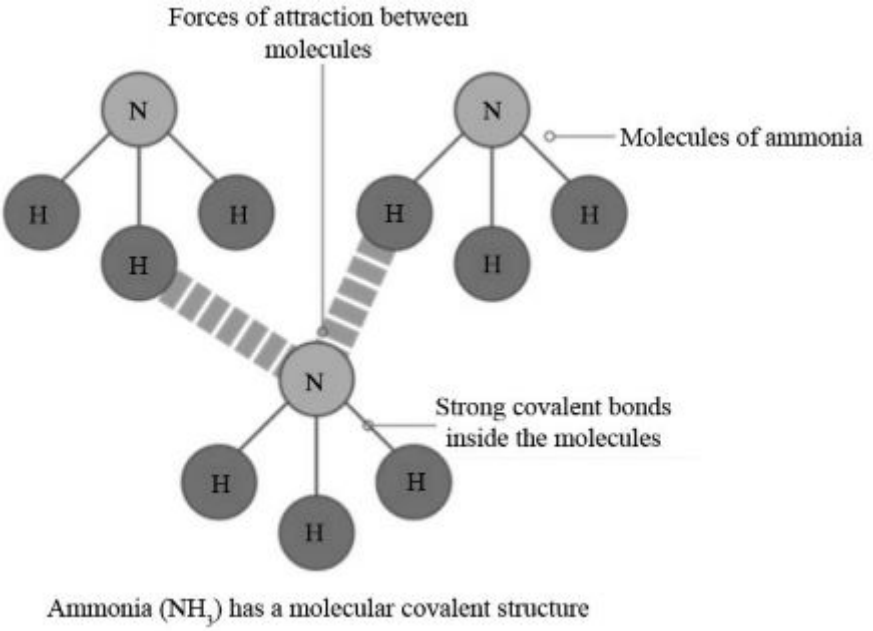


Figure 7.4.1 - Forces in a Simple Molecular Structure

Inside a simple molecular structure, we find **weak intermolecular forces of attraction**, so the ammonia molecules in the diagram above are weakly attracted to each other.

However, the **covalent bonds** $N - H$ are very strong and hence it is very hard to separate **ammonia** into its component **nitrogen** and **hydrogens**.

7.5 Physical Properties of Covalent Compounds

The weak intermolecular forces of attraction give Covalent Compounds certain properties.

7.5.1 Melting and Boiling Point

Since little energy is required to **overcome the weak intermolecular forces of attraction** between covalently bonded molecules, compounds with a simple molecular structure have a **low melting point and boiling point**.

7.5.2 Conductivity of Electricity

Since the molecules in a simple molecular structure are **neutral** (i.e. have no charge) particles, it has **no mobile charge carriers,** so it cannot conduct electricity in any state.

7.5.3 Solubility

Most simple molecular substances are **soluble** in **organic substances like hexane**, while they are **insoluble** in **polar solvents like water**.

Topic 8: Acids and Bases (I)

8.1 Arrhenius Definition of Acids and Bases

Svante Arrhenius, a Swedish chemist, proposed that:

- an **acid** is a substance that **produces hydrogen ions** (H^+) **when dissolved in water;**
- a **base** is a substance that **produces hydroxide ions** (OH^-) when **dissolved in water.**

This **"dissolving"** is known as **dissociation**.

When Hydrogen Chloride is dissociated in water, it forms H^+ and Cl^- ions. Since it forms H^+ ions, it is thus an acid, better known as **hydrochloric acid**.

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

When Sodium Hydroxide is dissociated in water, it forms Na^+ and OH^- ions. Since it forms OH^- ions, it is thus a base.

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

8.2 Neutralisation

When you react an acid with a base, the H^+ and OH^- ions form water, H_2O.

The full reaction is:
$HCl(aq) + NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow NaCl(aq) + H_2O(l)$

Hence, we have a general formula for reactions between acids and bases:
$acid + base \rightarrow salt + water.$

This is called **neutralisation**.

8.3 pH

The pH scale is a scale that shows how **acidic** or how **alkaline** (note: an alkali is defined as a **base soluble in water**) a solution is.

When pH is exactly 7, that means the solution is **neutral**.
When pH drops below 7, that means the solution is **acidic**.
When pH increases above 7, that means the solution is **alkaline**.

The further the pH gets from 7, the more **acidic/alkaline** the solution is.

8.4 Strength of Acids

There are 3 acids which you need to know the formula and name of in CM1131.

Formula	Name
HCl	Hydrochloric Acid
H_2SO_4	Sulfuric Acid
HNO_3	Nitric Acid

The definition of a strong acid is one where **all acid molecules dissociate** (fully dissociate) in water, while the definition of a weak acid is one where **only some acid molecules dissociate** (partially dissociate) in water **to form H^+ions.**

Acids like hydrochloric acid and nitric acid are strong acids, while acids like ethanoic acid (CH_3COOH) or carbonic acid (H_2CO_3) are weaker acids.

8.4.1 Monoprotic, Diprotic and Triprotic Acids

Monoprotic Acids are acids that dissociate to produce **one hydrogen ion** per unit of acid. This includes Hydrochloric acid, HCl and Nitric acid, HNO_3.

Similarly, Diprotic Acids are acids that dissociate to produce **two hydrogen ions**. This includes Sulfuric Acid, H_2SO_4 and carbonic acid, H_2CO_3.

Lastly, Triprotic Acids are acids that dissociate to produce **three hydrogen ions**. This includes phosphoric acid H_3PO_4.

8.4.2 Acidic Properties

Acids have a **sour taste**.

Since acids dissociate in aqueous solution to produce **mobile ions**, **aqueous acidic solutions** can conduct electricity.

Acids can change the colour of certain plant dyes, which are used as **pH indicators**.

8.5 Acid Reaction with Metal

An acid can react with a **reactive metal** to produce a **salt** and **hydrogen gas**.

$$H_2SO_4(aq) + Zn(s) \rightarrow H_2(g) + ZnSO_4(aq)$$

Non-reactive metals include **copper (Cu), gold (Au) and silver (Ag)**.

8.5.1 Test for Hydrogen Gas

To test for hydrogen gas, we can insert a lighted splint into the test tube. If the lighted splint extinguishes with a 'pop' sound, hydrogen gas is present.

8.6 Acid Reaction with Carbonate

Acids can react with carbonate to produce a **salt**, **carbon dioxide** and **water**.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

8.6.1 Test for Carbon Dioxide

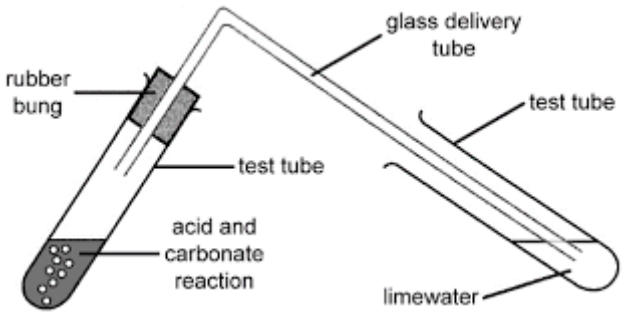


Figure 8.6.1.1 - Test for CO_2 Setup

To test for Carbon Dioxide, we introduce the gas into a test tube with **limewater** *(FYI: $Ca(OH)_2$)* using a **delivery tube** as shown above.

If a **white precipitate** ("cloudy") is formed in the limewater, then the gas is in fact **carbon dioxide**.

8.7 Bases

Bases are either **metal oxides** or **metal hydroxides** (according to Arrhenius' Definition) that react with an acid to form salt and water only.

Ammonia is also an Arrhenius Base, despite not having a hydroxide ion, it actually forms a hydroxide ion when it reacts with water, so we still classify it as an Arrhenius Base.

8.7.1 Alkalis and their Properties

Alkalis have a **bitter taste**.

Alkalis feel **soapy and slippery**.
Alkalis can cause colour change in certain plant dyes, which are used as **pH indicators**.

Alkaline solutions dissociate in water to produce **mobile ions**, so alkaline solutions can conduct electricity.

8.8 Types of Oxides

As we know, an oxide is a compound where an element reacts with oxygen.

There are 4 main types of oxides.

Metal Oxides can be divided into two types - **Amphoteric** and **Basic** oxides.

Non-metal Oxides can also be divided into two types - **Neutral** and **Acidic** oxides.

8.8.1 Amphoteric Oxides

There are 3 amphoteric oxides, ZnO, Al_2O_3 and PbO *(Lead (II) Oxide)*.
These three oxides can be remembered with the acronym **ZAP**.

Amphoteric oxides are oxides that react with **both acids and bases** to neutralise them.

8.8.2 Basic Oxides

Basic Oxides are any metal oxides that are **not amphoteric**. They react with **acids** to neutralise them.

8.8.3 Neutral Oxides

There are 4 main neutral oxides we need to know, **carbon monoxide** (CO), **water** (H_2O), **nitrogen monoxide** (NO) and **dinitrogen monoxide** (N_2O).

Neutral Oxides do not react with **acids**, **bases** or **water**.

8.8.4 Acidic Oxides

Acidic Oxides are any non-metal oxides that are not **neutral**. They react with **bases** to neutralise them.

Topic 9: Solubility Rules

9.1 The Solubility Rules

There are 6 main solubility rules.

All Halides (i.e. Compounds with Group 17 anions, Cl^-, Br^-, F^-, etc.) are **soluble** except compounds formed with **Silver** (Ag) and **Lead (II)** (Pb).

All Sulfates (compounds with SO_4^{2-}) are **soluble** except **Lead (II)** Sulfate, **Barium** Sulfate, and **Calcium** Sulfate.

*(**Note:** $CaSO_4$ is sparingly soluble, but we treat it as **insoluble**)*

All Nitrates (compounds with NO_3^-) are **soluble**.

All Ammonium Compounds (compounds with NH_4^+) and Group 1 Compounds (i.e. Na^+, Li^+, etc.) are **soluble**.

All Carbonates (compounds with CO_3^{2-}) are **insoluble** except **ammonium** carbonate and **Group 1** carbonates.

All Oxides and Hydroxides (compounds with O^{2-}, OH^-) are **insoluble** except **ammonium** oxide/hydroxide, **Group 1** (i.e. Na, Li, etc.) oxides/hydroxides and **Heavier Group 2** (i.e. Sr, Ba) oxides/hydroxides.

There are **AHEM** certain acronyms for this but uhhhhhh… I'm not gonna put it. Refer to here at your own discretion.

CM2131

Foundations in Chemistry II

Topic 10: Acids and Bases (II)

10.1 Arrhenius Definition of Acids (Extended)

As defined in **Section 8.1: Arrhenius Definition of Acids and Bases**, an **acid** is a substance that **produces hydrogen ions** (H^+) **when dissociated in water.**

However, H^+ ions do not persist in water. Instead, it combines with water as such to form hydronium, H_3O^+.

$$H^+ (aq) + H_2O (l) \rightarrow H_3O^+ (aq)$$

Hence the new definition which is more accurate, as introduced in CM2131, is that an acid is a **substance that produces hydronium ions** (H_3O^+) **when dissociated in water**.

When we say a solution is acidic, we say an acidic solution is a solution with a **significant concentration of hydronium ions**.

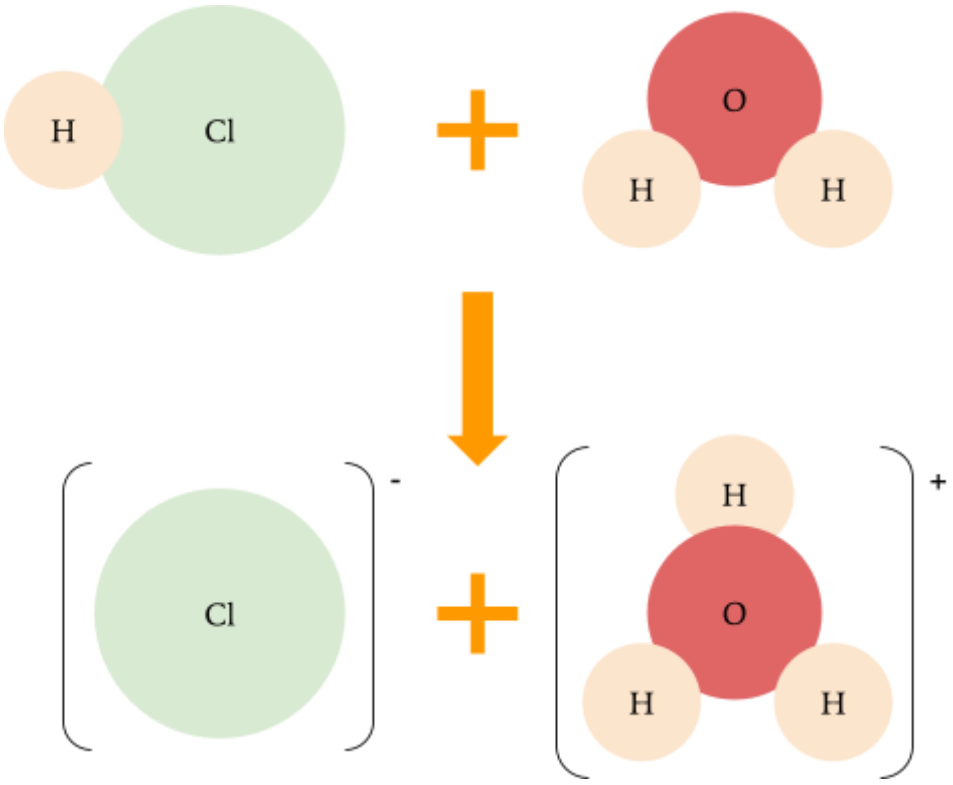


Figure 10.1.1 - Dissociation of Hydrogen Chloride

Refer to *Figure 10.1.1* above. It describes the equation $HCl(g) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$. Since we see HCl giving H_2O a hydrogen ion, which as we know, can be referred to as a **proton**, we call HCl a **proton donor**.

10.2 Types of Arrhenius Acids

In terms of chemical structure, Arrhenius Acids can be divided into several smaller subsets. We will go through three - Binary Acids, Oxyacids and Organic Acids.

10.2.1 Binary Acids

As in CM2131, Binary Acids are defined as acids which have the formula $HX\ (aq)$, where X is one of the **first four halogens** (i.e. Fluorine, Chlorine, Bromine or Iodine).

FYI: More generally, any combination of hydrogen and one other element, H_nE_m, is a binary acid for some integers $n,\ m$ and an element E.

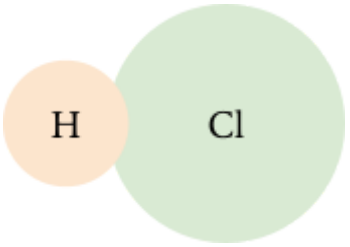


Figure 10.2.1.1 - A Binary Acid

The Binary Acids are $HF\ (aq),\ HCl\ (aq),\ HBr\ (aq)$ and $HI\ (aq)$.

10.2.2 Oxyacids

Oxyacids have the formula $H_aX_bO_c$, where X is an element other than Hydrogen and Oxygen, and where a, b, c are just arbitrary constants.

This means that Oxyacids are made up of **3 elements exactly** - **hydrogen, oxygen** and **one other element**.

The most common oxyacids are Nitric Acid, $HNO_3\ (aq)$ and Sulfuric Acid, $H_2SO_4\ (aq)$.

10.2.3 Organic Acids

Organic Acids are acids that are composed of **hydrogen, oxygen** and **carbon** (which also makes organic acids oxyacids). For example, acetic acid (also known as ethanoic acid), CH_3COOH, is an organic acid.

Acetic Acid can be written as CH_3COOH or CH_3CO_2H.

The reason why the last hydrogen is written separately is because the structure of Acetic Acid only allows that specific hydrogen atom to be transferred when a reaction occurs. This final hydrogen is then called the **acidic hydrogen**.

10.3 Strong vs Weak Acids

A Strong Acid dissociates to form **nearly one hydronium ion** for each molecule of acid dissolved in water, while a weak acid dissociates to form **significantly less than one hydronium ion** for each molecule of acid dissolved.

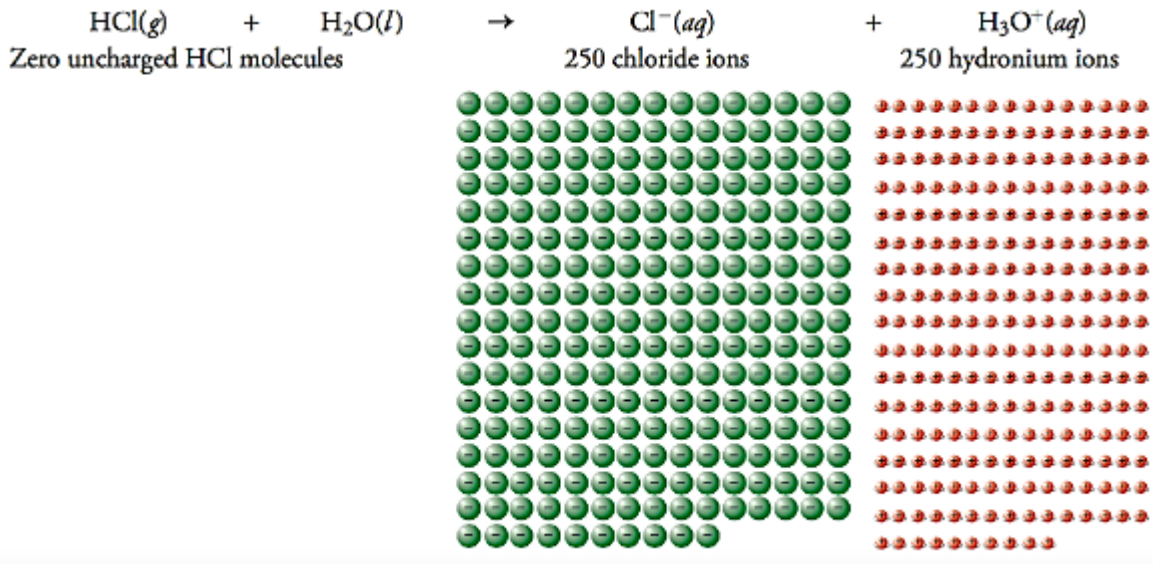


Figure 10.3.1 - Dissociation of a Strong Acid

Refer to *Figure 10.3.1* above. When we dissolve 250 $HCl\ (g)$ molecules in water, 250 hydronium ions are produced, which means for every $HCl\ (g)$ molecule, 1 hydronium ion is released.

We can hence deduce that $HCl\ (aq)$, or hydrochloric acid, is hence a **strong acid**.

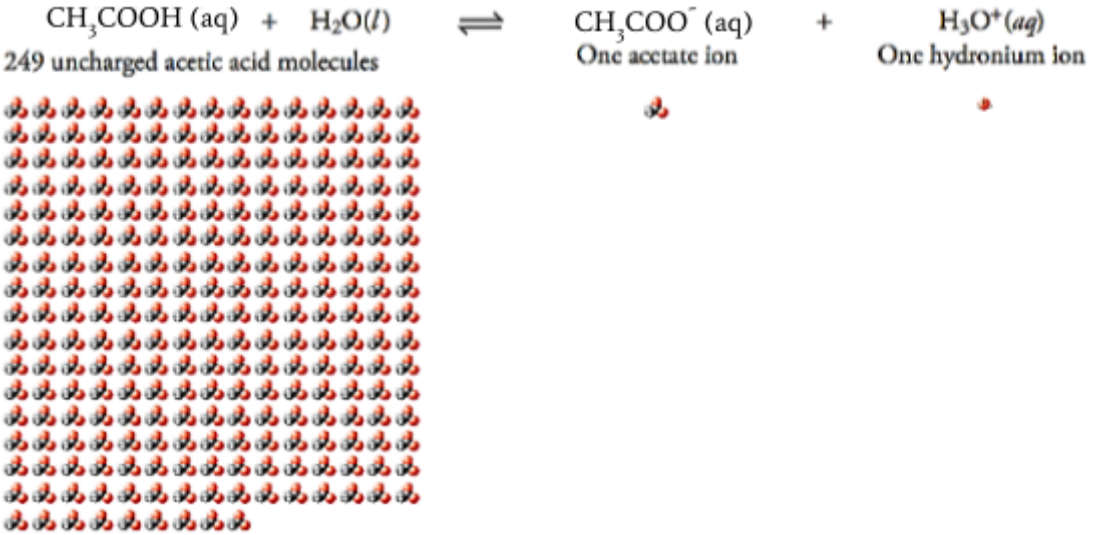


Figure 10.3.2 - Dissociation of a Weak Acid

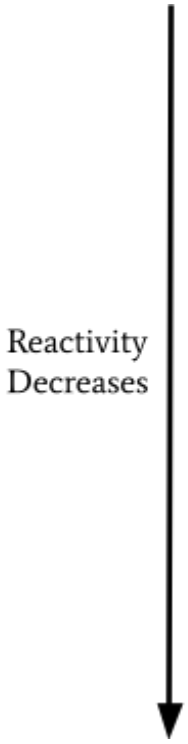
Refer to *Figure 10.3.2* above. When we dissolve 250 CH_3COOH molecules in water, only 1 hydronium ion is produced. This means for every one CH_3COOH molecule dissolved in water, on average, significantly less than 1 hydronium ion is produced.

Also note that in *Figure 10.3.2*, the arrow denoting the equation, ⇌, is double-sided, denoting a reversible reaction. This is true for most weak acid reactions.

Hence, we deduce that $CH_3COOH\ (aq)$, or acetic/ethanoic acid, is a **weak acid**.

10.4 Reactivity Series of Metals

For more information, go to **Topic 19: Reactivity Series of Metals**,



Metal	Reacts with Dilute Acid?
Potassium (K)	Yes (Explosive)
Sodium (Na)	Yes (Explosive)
Calcium (Ca)	Yes (Explosive)
Magnesium (Mg)	Yes
Aluminium (Al)	Yes
Carbon (C)	
Zinc (Zn)	Yes
Iron (Fe)	Yes
Tin (Sn)	Yes
Lead (Pb)	Yes
Hydrogen (H)	
Copper (Cu)	No
Mercury (Hg)	No
Silver (Ag)	No
Gold (Au)	No
Platinum (Pt)	No

Table 10.4.1 - Reactivity Series of Metals

Note: Carbon and Hydrogen, despite not being metals, are often included for comparison purposes.

Dilute Nitric Acid **does not react** with unreactive metals like Copper. However, when a more concentrated Nitric Acid is used, the following reaction occurs:

$$4HNO_3\,(aq)\ +\ Cu\,(s) \rightarrow 2H_2O\,(l)\ +\ 2NO_2\,(g)\ +\ Cu(NO_3)_2(aq)$$

This reaction is not like the other Acid-Metal reactions, as hydrogen gas is **not produced**, instead **nitrogen dioxide** is released.

Lead appears to not react with hydrochloric acid and sulfuric acid. Why is this so? Let's consider the case for hydrochloric acid.

The following reaction occurs when lead is introduced:

$$Pb\,(s)\ +\ 2HCl\,(aq)\ \rightarrow PbCl_2\,(s)\ +\ H_2\,(g)$$

Did you catch it? The state symbol for $PbCl_2$ is written as (s) instead of (aq), and this is because Lead (II) Chloride is **insoluble** in water. Refer to **Topic 9: Solubility Rules**.

This solid Lead(II) Chloride quickly forms an insoluble layer around the lead, causing the reaction to stop quickly and **prohibiting any further reaction**.

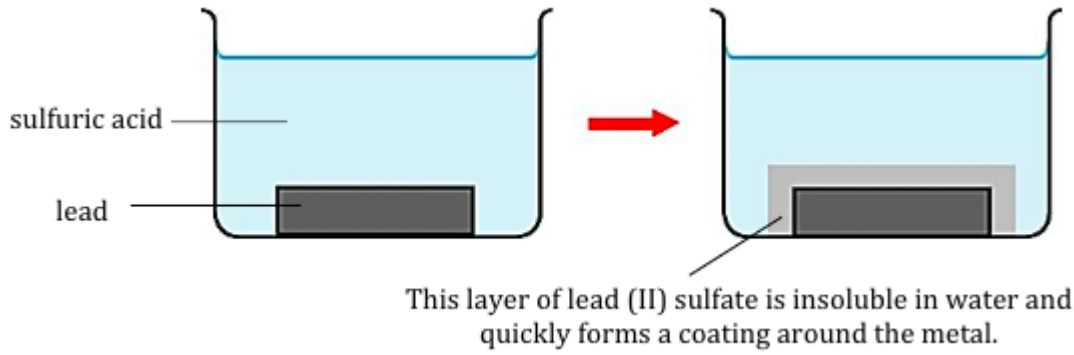


Figure 10.4.2

10.5 Arrhenius Acid-Base Reaction

Let's consider the Acid-Base Reaction between Nitric Acid, HNO_3, and Sodium Hydroxide, $NaOH$.

Let's start with nitric acid. This is what the solution looks like when nitric acid is added, immediately after the nitric acid dissociates into ions.

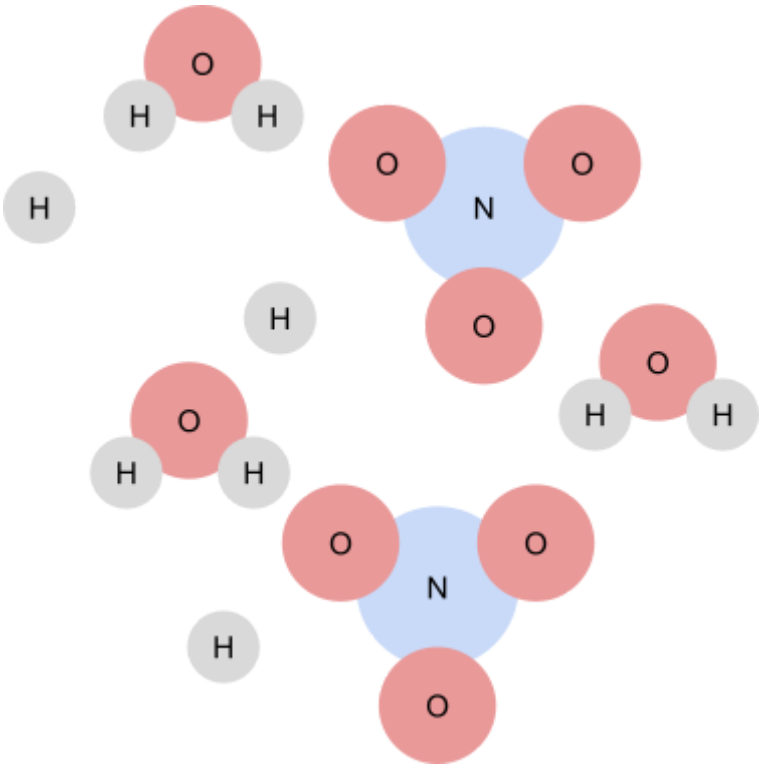


Figure 10.5.1

Then, all of the H^+ ions react with H_2O to form hydronium, as shown in *Figure 10.5.2* below.

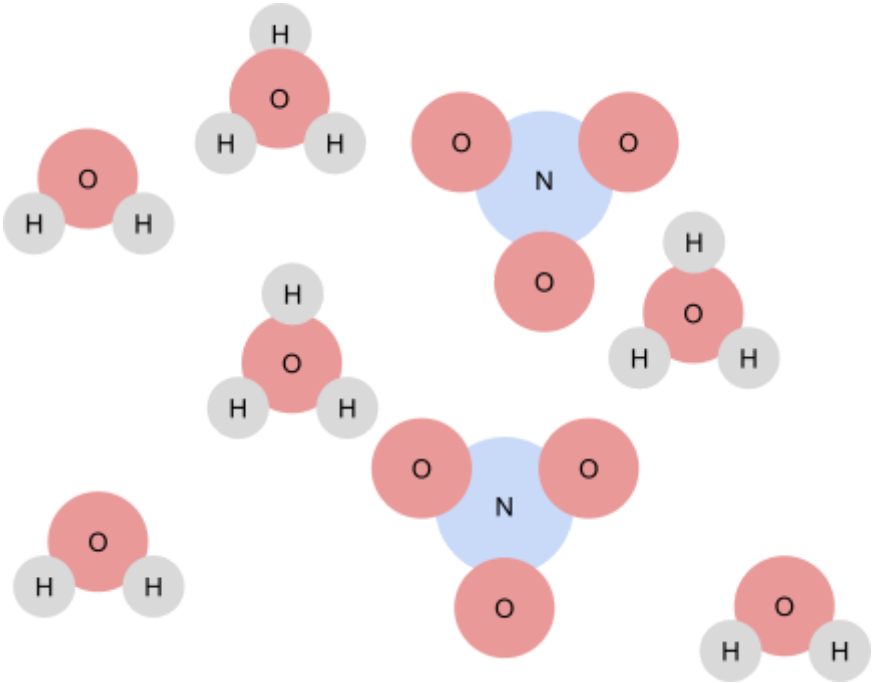


Figure 10.5.2

Hence, we see that in an aqueous solution of HNO_3, the following reaction occurs:

$$HNO_3\,(aq) + H_2O\,(l) \rightarrow NO_3^-\,(aq) + H_3O^+\,(aq)$$

When aqueous $NaOH$ is then added, Na^+ and OH^- ions can also be found in the solution.

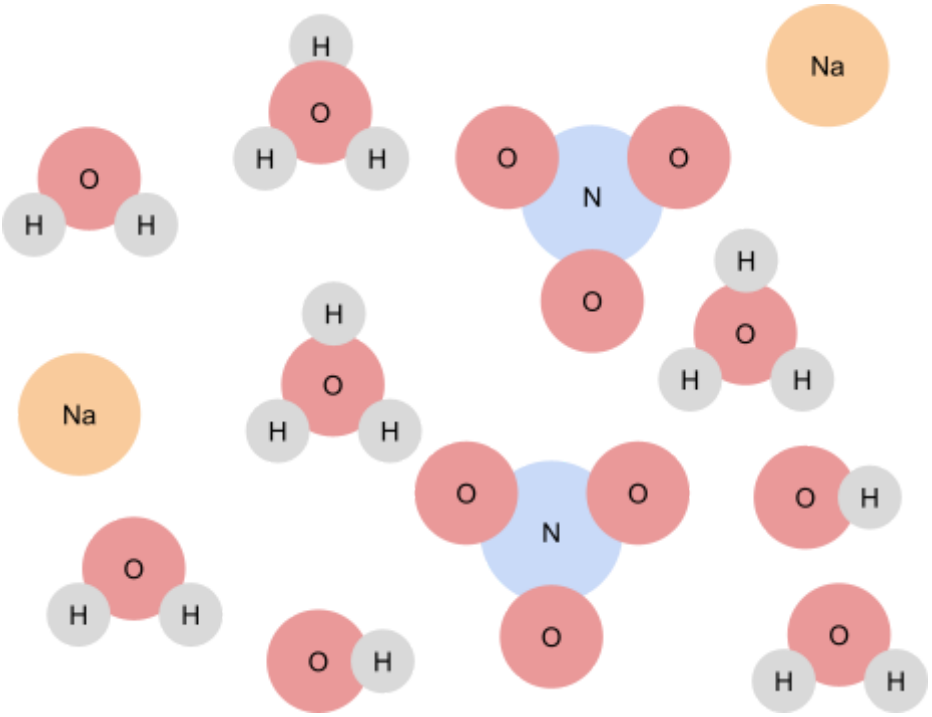


Figure 10.5.3

The OH^- ions then react with the H_3O^+ ions in the following reaction:

$$OH^-\ (aq)\ +\ H_3O^+\ (aq) \rightarrow 2H_2O\ (l)$$

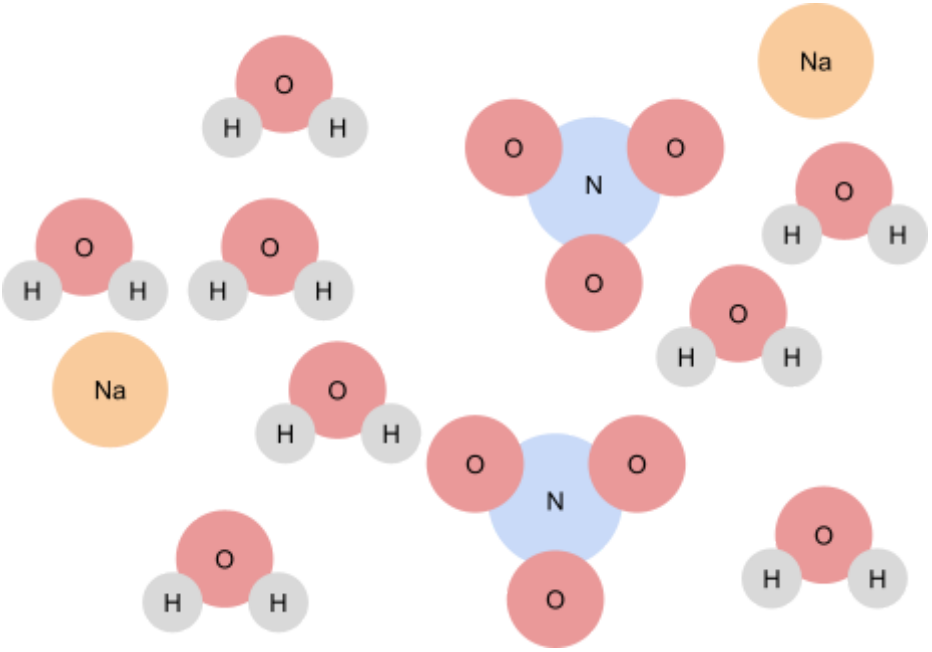


Figure 10.5.4

As you can see from *Figure 10.5.4* above, there are no more hydronium and hydroxide ions.

We can hence express this reaction as such:

$$H_3O^+(aq) + NO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow 2H_2O\,(l) + Na^+(aq) + NO_3^-(aq)$$

Equation 10.5.5

Note: In this case, we don't write $NaOH\,(aq)$ or $NaNO_3\,(aq)$ but instead write it as separate ions because they are soluble in water and will dissociate to form ions anyway.

10.5.1 Net Ionic Equation

Refer to *Equation 10.5.5* above.

Since Na^+ and NO_3^- do not participate in the reaction and remain unchanged, they are called **spectator ions** and can be removed from the equation to obtain the **net ionic equation**.

Hence the net ionic equation of this reaction is $H_3O^+\,(aq) + OH^-\,(aq) \rightarrow 2H_2O\,(l)$.

Pause and Ponder 10.5.1

Consider this reaction. "Aqueous sodium hydroxide reacts with aqueous copper (II) chloride to give a blue precipitate of copper (II) hydroxide."

What is the **net ionic equation** of this reaction?

[Click to Go to Solution]

10.5.2 Precipitation Reaction

A Precipitation Reaction is when two solutions react to form an **insoluble product**, known as a **precipitate**.

For example, $2NaOH\,(aq) + FeSO_4\,(aq) \rightarrow Fe(OH)_2\,(s) + Na_2SO_4\,(aq)$ is a precipitation reaction because it forms an insoluble product, $Fe(OH)_2$.

Tip: Look out for the state symbol (s), as that will tell you that a product is insoluble.

10.6 Arrhenius Definition of Bases (Extended)

As defined in **Section 8.1: Arrhenius Definition of Acids and Bases**, a **base** is a substance that **produces hydroxide ions (OH^-) when dissociated in water.**

Similar to **Section 10.3**, a strong base and a weak base is differentiated by the amount of OH^- ions released when a base is dissociated in water.

A strong base produces **nearly 1 hydroxide ion** for **every molecule of base** dissolved in water, while a weak base produces **significantly less than 1 hydroxide ion** for **every molecule of base** dissolved.

10.6.1 Ammonia - An Arrhenius Base?

Ammonia, NH_3, is considered an **Arrhenius Base**.

Wait, don't Arrhenius bases have OH^- ions? Where's the OH^- in NH_3?

Well, let's consider the reaction that occurs when ammonia reacts with water. Here is the reaction:

$$NH_3\,(g) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

Ah ha! Ammonia **does in fact produce a hydroxide ion** when it is added to water. Additionally, since the reaction is reversible (remember the ⇌ arrow), ammonia is classified as a **weak base**.

10.6.2 Carbonates and Hydrogen Carbonates

Many Carbonates and Hydrogen Carbonates (i.e. compounds with HCO_3^-) are **weak bases**. For example, take the compound Sodium Carbonate, Na_2CO_3.

When we dissociate it in water, we get:

$$Na_2CO_3\,(s) \rightarrow 2Na^+\,(aq) + CO_3^{2-}(aq)$$

Now, when the carbonate ion reacts with water, we have the following:

$$CO_3^{2-}\,(aq) + H_2O\,(l) \rightleftharpoons OH^-(aq) + HCO_3^-(aq)$$

Again, similar to Ammonia, a "hidden" hydroxide ion is produced. Since this reaction is reversible, Na_2CO_3 is a **weak base**.

Pause and Ponder 10.6.2

What are the reactions that take place for:

(a) Lithium Carbonate

(b) Sodium Hydrogen Carbonate ($NaHCO_3$)

...to release a hydroxide ion?

[Click to Go to Solution]

10.7 Test for Ammonia Gas

Ammonia Gas, NH_3, can be identified by its **pungent smell**, and by a **MOIST red litmus paper** (refer to *Figure 10.7.1*).

(rip to all the noses that died during that one practical, you know what i'm talking about if you don't you're lucky)

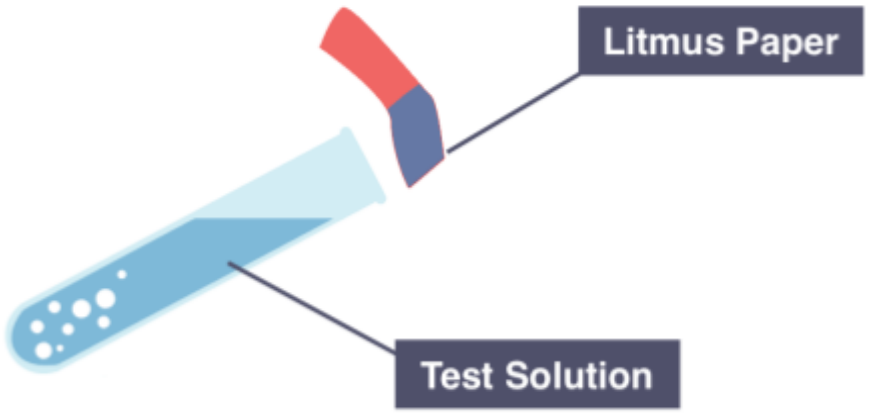


Figure 10.7.1 - Test for Ammonia Gas

If Ammonia Gas is present, the moist red litmus paper will **turn blue**.

Note: Remember to make sure the litmus paper is moist before conducting the test.

Topic 11: The Mole Concept

Credits to Aden Chong (M22101) for contributing to part of Sections 11.1, 11.2 and 11.11.

11.1 Avogadro's Constant

Avocado? aVOcAdOooO?!

The number 6.02×10^{23} is known as **Avogadro's constant** and is denoted N_A.

But where did this number 6.02×10^{23} come from in the first place?

6.02×10^{23} is the number of carbon-12 atoms present in exactly 12 grams of carbon-12, due to the abundance of carbon[verify] and the stability of carbon and carbon compounds.

This is a standard that was agreed upon by chemists in 1961.

Note: Avogadro's Constant is also known as **Avogadro's Number.**

11.2 The Mole

The Mole is a **Base SI Unit** written as **mol**.

A mole of any substance is 6.02×10^{23} particles *(Avogadro's Number!)* of that substance, which could be atoms, molecules, or ions.

11.3 Calculations Involving Moles

This is how the number of particles is related to the number of moles of a substance:

$$No.\ of\ particles = No.\ of\ moles \times N_A$$

For example, let's calculate the number of Zinc Atoms there are in 0.24 mol of Zinc.

$$\begin{aligned} No.\ of\ Zinc\ Atoms &= 0.24 \times (6.02 \times 10^{23}) \\ &= 1.44 \times 10^{23} \\ &= 1.4 \times 10^{23}\ (2sf) \end{aligned}$$

Note: For calculations involving moles, round to the same s.f. as the data provided. If there are multiple numbers provided in the question, take the one with the lowest s.f.

Pause and Ponder 11.3

How many oxygen atoms are in 5 moles of oxygen gas?

[Click to Go to Answer]

11.4 Molar Mass

The Molar Mass of a substance is the **mass per mole** of that substance and has units **g/mol** (or g mol^{-1}).

Tip: If you can't remember what the unit for molar mass is, just remember "mass per mole". The unit for mass is the gram, and the unit for moles is mol, and hence the unit for molar mass is g/mol.

For an element, the relative atomic mass (A_r) *in amu* of that element is the same as its molar mass *in g/mol*, numerically.

For example, the relative atomic mass of oxygen is 16.00 amu, and the molar mass of oxygen is 16.00 g/mol.

11.5 Relative Molecular Mass and Relative Formula Mass

Relative Molecular Mass, denoted M_r, is the average mass of one molecule of an element or compound, compared to $\frac{1}{12}$ of the mass of a carbon-12 atom.

It can also be calculated by summing the relative atomic mass (A_r) of all the atoms in a molecule.

For example, let's calculate the relative molecular mass of water. As we know, it has the formula H_2O.

$$\begin{aligned} M_r &= 2 \times A_r\ of\ Hydrogen + 1 \times A_r\ of\ Oxygen \\ &= 2 \times 1.008 + 1 \times 16.00 \\ &= 18.016 \\ &= 18.02\ (2sf) \end{aligned}$$

For Ionically Bonded substances, it is **INACCURATE** to use the term "Relative **Molecular** Mass", as they do not exist as molecules (cannot be referred to as molecules or compounds).

Hence, we use the term "Relative **Formula** Mass" instead. Relative Formula Mass is also denoted M_r.

11.6 Molecular Mass and Formula Mass

Molecular Mass is the sum of the atomic masses of the component atoms in a molecule.

Pause and Ponder 11.6

Why do Molecular Mass (in amu) and Molar Mass (in g/mol) have the same numerical value?

[Click to Go to Solution]

11.7 Relationship Between Molar Mass, Moles and Mass

The equation relating these three values is:

$$Molar\ Mass \times No.\ of\ Moles = Mass$$

For example, let's find the number of moles of Zinc there are in 23.3g of Zinc.

$$\begin{aligned} No.\ of\ Moles &= Mass \div Molar\ Mass \\ &= 23.3g \div 65.38\ g/mol \\ &= 0.356\ mol\ (3\ sf) \end{aligned}$$

Pause and Ponder 11.7

What is the mass of **one molecule** of oxygen gas?

[Click to Go to Solution]

11.8 Percentage Composition of Compounds by Mass

Let's say we wanted to find the percentage composition of each element in CH_3COOH.

(Do you recognise this yet from all the Chem you've been doing? Yep, it's acetic acid)

Quick Recap:

Name two subgroups of Arrhenius Acids that CH_3COOH belongs to.

If you don't know, time to revisit **Section 10.2: Types of Arrhenius Acids**!

Note: The following Tables (i.e. *Table 11.8.1*, *11.8.2*, *11.8.3*) are not required during exams. They are just for explanation purposes, and you can just use the formula at the end of this Section.

The first step is to find the number of atoms of each element present in acetic acid. Just… count. Basic maths. Yes.

Element	Count
C	2
H	4
O	2

Table 11.8.1

Let's then construct a table for the total mass of each element present.

Element	Count	Atomic Mass (amu)	Total Atomic Mass (amu)
C	2	12.01	24.02
H	4	1.008	4.032
O	2	16.00	32.00
Total	8		58.04 $(4\ sf)$

Table 11.8.2

Now, we can calculate the percentage composition by mass. It is simply the total atomic mass of that element divided by the total atomic mass. (Remember to multiply by 100%!)

Element	Total Atomic Mass (amu)	Percentage Composition
C	24.02	$\frac{24.02}{58.04} \times 100\% = 41.39\%\ (4sf)$
H	2.016	$\frac{2.016}{58.04} \times 100\% = 3.473\%\ (4sf)$
O	32.00	$\frac{32.00}{58.04} \times 100\% = 55.13\%\ (4sf)$
Total	58.04 $(4\ sf)$	

Table 11.8.3

But wait! If we sum up all the percentage compositions, we get 99.99(3)%, and not the expected 100%! What happened?!

Don't worry, nothing is wrong. If this happens during the exam, in general, if the value you get is reasonably close to 100% (*i.e. 99.7% should be okay, but 69% isn't*), there is likely nothing wrong. Instead, it is likely that the error is caused by rounding of Atomic Mass on the Periodic Table, or by rounding for s.f.

Now, we can derive a formula to obtain percentage composition.

$$Percentage\ Composition = \frac{A_r\ of\ element \times n}{M_r\ of\ compound} \times 100\%,$$

where n is the number of that element that occurs in its formula.

11.9 Empirical Formula

The Empirical Formula of a compound is like the **simplest** formula of that compound that shows the ratio of elements in it.

For example, hydrogen peroxide has molecular formula H_2O_2.
However, when talking about its empirical formula, we can "factor out" a two from H and O, obtaining the simplest formula that shows the ratio of elements: HO.

Hence hydrogen peroxide has empirical formula HO.

Let's say we have the following problem.

A compound is made up of 11.19% Hydrogen by mass and 88.81% Oxygen by mass, and we are given that its molecular formula is identical to its empirical formula. So how should we find what compound it is?

Let's consider 100g of this compound. Then the mass of hydrogen and oxygen is just their respective percentage compositions numerically.

Element	Percentage Composition	Mass in 100g of compound (g)
H	11.19%	11.19
O	88.81%	88.81

Table 11.9.1

Now, we can use the formula $No.\ of\ moles = \frac{Mass}{Molar\ Mass}$ to get the number of moles of each element present in 100g of compound. Recall that the Molar Mass of an element is equal to its relative atomic mass, which can be found on the periodic table.

Element	Mass in 100g of compound (g)	Molar Mass (g/mol)	No. of Moles in 100g of compound (mol)
H	11.19	1.008	11.10
O	88.81	16.00	5.551

Table 11.9.2

Let's then compare the ratio of Moles obtained in *Table 11.9.2*.

No. of Moles of Hydrogen (mol)	No. of Moles of Oxygen (mol)	
11.10	5.551	
1.999	1	*(dividing both sides by 5.551)*
2	1	*(rounding)*

Table 11.9.3

Now that we have found the ratio of moles, we can finally determine the empirical formula. Since there are 2 moles of hydrogen for every 1 mole of oxygen, we then get that the empirical formula is H_2O_1, but we ignore the subscript 1, so it is just H_2O.

We are also given that its molecular formula is identical to its empirical formula, which means the molecular formula of the compound is H_2O. So… Can you figure out what compound it is?

If not… you're in serious trouble…

It's water!

And we are done.

Pause and Ponder 11.9

Albite is a type of crystal and is made up of the following elements. Provided in *Table 11.9.4* is also the elements' percentage composition by mass.

Element	Percentage Composition

Sodium	8.7668%
Aluminium	10.288%
Silicon	32.135%
Oxygen	48.810%

Table 11.9.4

Write down the Empirical Formula of Albite, given that it is in the form $Na_wAl_xSi_yO_z$, where w, x, y, z are integer constants to be determined.

[Click to Go to Solution]

11.10 Molecular Formula from Empirical Formula

The Molecular Formula is always a **constant multiple** of the Empirical Formula.

For example, hydrogen peroxide has the Molecular Formula H_2O_2 and the Empirical Formula HO. We can see that its Molecular Formula is exactly 2 times the Empirical Formula.

$$H_2O_2 = 2HO$$

Now, let's say we wanted to determine the Molecular Formula of hydrazine, given that its empirical formula is NH_2, and that its relative molecular mass is 32.0 amu.

Firstly, we can find the relative molecular mass of the Empirical Formula NH_2.

$$\begin{aligned} M_r &= 14.01\ amu + 2 \times 1.008\ amu \\ &= 16.03\ amu\ (4sf) \end{aligned}$$

We know that the Molecular Formula is a constant multiple of the Empirical Formula, and hence the relative molecular mass of the Molecular Formula is also a constant multiple of that of the Empirical Formula.

$$\begin{aligned} &Let\ n\ be\ the\ number\ multiple\ in\ subscript\ of\ the\ empirical\ formula. \\ n \times 16.03\ amu &= 32.0\ amu \\ n &= 1.996 \\ &= 2\ (nearest\ whole\ no.) \end{aligned}$$

Hence, the molecular formula of hydrazine is $(NH_2)_2$, or N_2H_4.

11.11 Avogadro's Law and Molar Volume of Gases

Avogadro's law states that equal volumes of **gases** contain the **same number of molecules** (and hence the same number of moles of molecules), under the same temperature and pressure.

Molar Volume of Gases is a constant and refers to the volume occupied by **1 mole of gas**.
At room temperature and pressure (r.t.p.), the Molar Volume is $24\ dm^3$ or $24000\ cm^3$.

Thus by Avogadro's Law, the number of molecules of oxygen in 20°C at 1 atm (pressure) is equal to the number of nitrogen molecules at 20°C at 1 atm and so on.

We can find a formula to express volume of gas in terms of the number of moles.

Each 1 mole of gas occupies 24 dm^3 of volume, so we have:

$$Volume\ of\ gas\ =\ 24\ dm^3\ \times\ No.\ of\ moles$$

Let's say we wanted to find the volume occupied by 0.52 mol of Sulfur Dioxide at r.t.p..

We just use the formula $Volume\ =\ 24\ dm^3\ \times\ 0.52$

$$=\ 12.48\ dm^3$$

$$=\ 12\ dm^3\ (2sf)$$

Note: 1 cubic decimeter (dm^3) = 1000 cubic centimetres (cm^3) = 1 litre (l). This is useful for converting between units.

Topic 12: Chemical Calculations

12.1 Balanced Chemical Equations

A Balanced Chemical Equation can tell us a lot about the reaction that takes place:

1. **Identities** of Reactants and Products
2. **Ratio** of the Amounts of Reactants and Products
3. **State** of Reactants and Products

However, an equation does not tell us:

1. The **speed** (i.e. how fast or how slow) of a reaction
2. Whether the reaction **gives out heat** or **takes in heat**
3. The **type of bonding** *(which can often be guessed, but you can't be certain)* of the products and reactants

12.2 Stoichiometry

The relationship between the amounts of reactants and products is called the **stoichiometry** *(pronounced stoy-kee-oh-muh-tree)* of the reaction.

Stoichiometry is based on the **Law of Conservation of Matter**, which says that **matter** can **neither be created nor destroyed**. Hence, the **mass of reactants** should theoretically **equal** the **mass of products**.

12.3 Ratio of Moles vs Ratio of Particles

Consider the reaction in *Figure 12.3*.

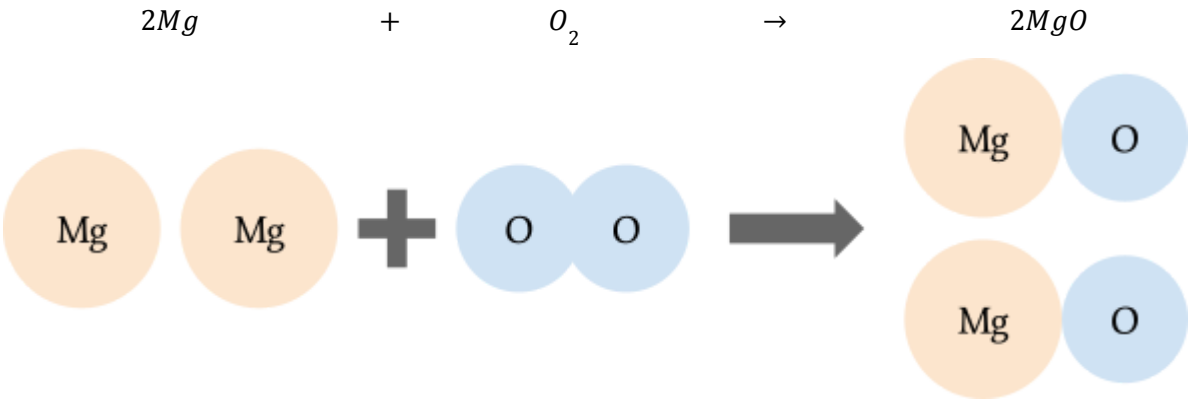


Figure 12.3 - A Simple Reaction

Here, we see 2 atoms of magnesium reacting with 1 molecule of oxygen gas to form 2 molecules of magnesium oxide.

If we multiply each of those numbers by Avogadro's constant, we have that $2N_A$ atoms of magnesium reacts with $1N_A$ molecules of oxygen gas to form $2N_A$ molecules of magnesium oxide.

By definition of the mole, we can then convert these in terms of the **mole**.

We get that 2 moles of magnesium react with 1 mole of oxygen gas to form 2 moles of magnesium oxide.

Hence, we can conclude that the ratio of the moles is **IDENTICAL** to the ratio of the particles (atoms, molecules, ions, etc.) in a chemical equation.

12.3.1 Example Questions

Worked Example 2, CM2131 Topic 3 Notes Modified

Iron is obtained from Iron Ore, which contains Iron (III) Oxide. The reaction of Iron (III) Oxide with Carbon Monoxide, produces Iron and Carbon Dioxide.

a) Write down the Balanced Equation of this reaction, with state symbols.
b) Calculate the mass of iron ore needed to make 150 grams of iron.
c) Calculate the volume of Carbon Dioxide formed when 98.0 grams of iron ore reacts, given that the procedure was conducted at room temperature and pressure.

Solution

a) This part is just balancing an equation. We first identify the products and reactants and write down the imbalanced equation:

$$Fe_2O_3 + CO \rightarrow Fe + CO_2$$

We now just balance the equation to obtain:

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

We add state symbols:

$$Fe_2O_3\,(s) + 3CO\,(g) \rightarrow 2Fe\,(s) + 3CO_2\,(g)$$

b) For this part, we first calculate the number of moles of iron present in 150 grams of iron.

Using $No.\ of\ Moles = \frac{Mass}{Molar\ Mass}$, we get:

$$No.\ of\ Moles = \frac{150g}{55.85g/mol}$$
$$= 2.685\ mol$$ [See Note]

We then use the mole ratio. Since one mole of Fe_2O_3 corresponds to two moles of Fe, we have $No.\ of\ Moles\ of\ Fe_2O_3 = 2.685\ mol \div 2$
$$= 1.342\ mol$$ [See Note]

We then find the mass of Fe_2O_3 needed using $Mass = No.\ of\ Moles \times Molar\ Mass$.

$$Mass = 1.342\ mol \times (2 \times 55.85 + 3 \times 16.00)\ g/mol$$
$$= 214g\ (3\ sf)$$

Note: Only round the final answer to 3s.f. For working, leave to 1 more than the precision of the final answer (in this case, 4s.f.), then TRUNCATE the number (just delete everything after, don't round).

c) No. of moles in 98.0 grams of $Fe_2O_3 = \frac{98.0g}{(2 \times 55.85 + 3 \times 16.00)g/mol}$
$$= 0.6136\ mol$$

Using mole ratio, No. of moles of CO_2 produced $= 0.6136\ mol \times \frac{3}{1}$
$$= 1.840\ mol$$

Using the Molar Volume of Gas at r.t.p., we have:

$$Volume = 1.840\ mol \times 24\ dm^3/mol$$
$$= 44.2\ dm^3\ (3\ sf)$$

Worked Example 3, CM2131 Topic 3 Notes Modified

Calcium Carbonate decomposes to form Calcium Oxide and Carbon Dioxide on heating. If 5.00 g of Calcium Carbonate is heated, find:

a) the Balanced Chemical Equation of this reaction
b) the mass of Calcium Oxide formed
c) the volume of Carbon Dioxide produced at r.t.p.

Solution

a) This part is just balancing an equation. We first identify the products and reactants and write down the imbalanced equation:

$$CaCO_3 \rightarrow CaO + CO_2$$

This equation is already balanced. We add state symbols:

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

b) $$Moles\ of\ CaCO_3 = \frac{5.00g}{(40.08 + 12.01 + 3 \times 16.00)\ g/mol}$$
$$= 0.04995\ mol$$

Using Mole Ratio, $Moles\ of\ CaO = Moles\ of\ CaCO_3 \times \frac{1}{1}$
$$= 0.04995\ mol$$

$$Mass\ of\ CaO = 0.04995\ mol \times (40.08 + 16.00)\ g/mol$$
$$= 2.80g\ (3\ sf)$$

c) Using Mole Ratio, $Moles\ of\ CO_2 = Moles\ of\ CaCO_3 \times \frac{1}{1}$
$$= 0.04995\ mol$$

$$Volume\ of\ CO_2 = 0.04995\ mol \times 24\ dm^3/mol$$
$$= 1.20\ dm^3\ (3\ sf)$$

12.4 Ratio of Volumes of Gases in a Reaction

Since the Ratio of Moles is the same for gases, or rather, any substance in general, the ratio of volumes is the same for gases since the volume is a constant multiple of the number of moles.

For example, take the equation:

$$N_2\,(g) + 2O_2\,(g) \rightarrow 2NO_2\,(g)$$

We know that 1 mole of nitrogen gas reacts with 2 moles of oxygen gas to form 2 moles of nitrogen dioxide gas.

Multiplying each value by $24\ dm^3/mol$ we see that 24 dm^3 of nitrogen gas reacts with 48 dm^3 of oxygen gas to form 48 dm^3 of nitrogen dioxide gas.

Now factoring out a 24, we see that 1 dm^3 of nitrogen gas reacts with 2 dm^3 of oxygen gas to form 2 dm^3 of nitrogen dioxide gas.

Now, we can conclude that the ratio of volumes for gases is the same as the ratio of moles or the ratio of particles.

12.5 Limiting Reactant

Usually, the reactants in a reaction are not present in the **exact stoichiometric ratios**, and instead one (or more) is present **in excess**.

The **limiting reactant** is the one that is **used up first** by the chemical reaction. Usually, we want the limiting reactant to be the one that is **most expensive**, to ensure that minimum wastage occurs.

Consider this equation:

$$2NO\ (g)\ +\ O_2\ (g) \rightarrow 2NO_2\ (g)$$

Suppose you are given 4 moles of NO and 3 of O_2. However, by the **stoichiometric ratio**, only 2 moles of O_2 is needed for every 4 moles of NO. Hence, one mole of O_2 is present **in excess**.

Hence, we call NO the **limiting reactant**, since it is used up (first); and we call O_2 the **excess reactant.**

In Stoichiometric Calculations, the first thing we need to do is to **identify the limiting reactant.** This is because the limiting reactant controls the amount of products produced by the reaction.

12.5.1 Example Question

Worked Example 9, CM2131 Topic 3 Notes

One of the important commercial processes is to convert N_2 from the air into nitrogen-containing compounds. This is based on the reaction of N_2 and H_2 to form ammonia, NH_3, known as the Haber process. The equation is shown below.

$$N_2\ (g)\ +\ 3H_2\ (g) \rightarrow 2NH_3\ (g)$$

How many moles of NH_3 can be formed from 3.0 mol of N_2 and 6.0 mol of H_2?

Solution

To completely react 6.0 mol of H_2, by the mole ratio, only 2.0 mol of N_2 is needed.

But since there is 3.0 mol of N_2, we say N_2 is present in **excess**, while H_2 is the **limiting reactant**.

We always use the mole ratio of the **limiting reactant** to determine the amount of product.

Since the mole ratio of H_2 to NH_3 is $3:2$, when 6.0g of H_2 reacts, 4.0g of NH_3 is formed.

12.6 Percentage Yield

In Stoichiometric Calculations, we assume that all reactants are 100% pure and all reactants are completely turned into products. The amount of product that should be produced by these calculations is known as the **theoretical yield**.

However, the actual amount obtained is **usually** less, and strictly less, than the **theoretical yield,** and this amount is known as the **actual yield**.

The reason for this loss in substance is usually from the following:

1. Some product is stuck to the apparatus (e.g. stuck to filter paper during filtration)
2. Evaporation of volatile product/reactant
3. Presence of sparingly soluble precipitate (causing some product to dissolve)
4. Occurrence of a **side reaction**

We calculate Percentage Yield as such:

$$Percentage\ Yield\ = \frac{Actual\ Yield}{Theoretical\ Yield} \times\ 100\%$$

This value can be used to describe **how well a reaction went**, although this definition is quite loose.

12.7 Percentage Purity

The purity of the reactants can also contribute to loss of product, and we measure the percentage purity of a substance using:

$$Percentage\ Purity\ = \frac{Amount\ of\ Pure\ Substance\ in\ Sample}{Amount\ of\ Substance\ used\ in\ Reaction} \times\ 100\%$$

Topic 13: Concentration of Solutions

13.1 Recap: Solute and Solvent

The solute refers to **the substance that is dissolved**.
The solvent refers to **the substance that dissolves the solute**.

Tip: *If you still can't remember which is which, the way I remember it personally is that in chromatography, we say "solvent front". Then, you can deduce that the solvent is the one that does the dissolving, and you can then deduce the solute is the substance to be dissolved.*

13.2 Concentration of Solution

The concentration of a solution is the **amount of solute dissolved per unit volume of solvent**.

As more solute is introduced into the solvent, the resulting solution is **more concentrated**.

This is very logical - for example, if you add more sugar to water, the water gets sweeter (more sugar concentration).

When no more solute can be dissolved in the solvent, we say that the solution is **saturated**.

There are 2 main ways to measure the **concentration** of a solution - **Mass Concentration** and **Molar Concentration.**

13.2.1 Mass Concentration

Mass concentration expresses the concentration of a solution in terms of **mass of solute per unit volume of solvent**, and can be expressed with:

$$Mass\ Concentration = \frac{Mass\ of\ Solute}{Volume\ of\ Solution}$$

For example, if we added 50g of sugar to water, such that the resulting solution has volume 1 dm^3, we will get that the mass concentration is $50 g/dm^3$.

13.2.2 Molar Concentration

The molar concentration, also known as **molarity**, of a solution, is the **number of moles of solute present per unit volume of solvent**.

$$Molarity\ (mol/dm^3) = \frac{No.\ of\ Moles\ of\ Solute\ (mol)}{Volume\ of\ Solution\ (dm^3)}$$

The SI Unit for Molarity is mol/m^3, but often, we use mol/L (moles per litre) instead, which is equivalent to mol/dm^3.

Pause and Ponder 13.2.2

Can you derive a formula for Molarity in terms of Mass Concentration and Molar Mass?

[Click to Go to Solution]

13.2.2.1 The Molar

No, not the tooth.

The molar is a SI-derived unit, and is denoted by M. It is defined as:

$$1\ M\ =\ 1\ mol/L$$

The Molar is used to measure **concentration** of solution.

13.3 Dilution of Solutions

Dilution of solutions, as the name suggests, is the process of **adding water** to a concentrated solution to obtain a **dilute solution**.

Note: For very strong Acids/Bases, we should always add the acid/base to water, not the other way around, so that the water can "absorb" some of the heat generated in the process.

There is just one main important fact to memorise for this process - The amount of solute does not change when dilution of solutions occurs.

Logically, this is clearly true. When you add water to a sugar solution to dilute it, the amount of sugar never changes, right? Hence, we have:

$$No.\ of\ moles_{conc} = No.\ of\ moles_{dil}$$

And since $No.\ of\ moles = Volume \times Molarity$ we can hence derive the following equation:

$$V_{conc}C_{conc} = V_{dil}C_{dil}$$

Where:

- V_{conc} is the volume of concentrated solution;
- C_{conc} is the molarity of concentrated solution;

- V_{dil} is the volume of diluted solution;
- C_{dil} is the molarity of diluted solution.

Pause and Ponder 13.3

Let's say Ms Woo tells you to prepare a 2.0 M sucrose solution, but she only gives you 420ml of a 6.9 M sucrose solution in a huge beaker. She is also evil and decides to only allow you to use the huge beaker, and doesn't give you a smaller beaker.

She expects you to know the molar volume of water by heart, but you don't since she never bothered to mention it. Ever. Fortunately, your friend whispers to you that it is $18.0 cm^3/mol$.

So how much water, **in moles**, do you have to add to the 6.9 M sucrose solution to obtain a 2.0 M sucrose solution?

[Click to Go to Solution]

13.3.1 Volumetric Flask

Figure 13.3.1.1 - A Volumetric Flask

A volumetric flask is like a conical flask, but with a marking to denote a specific volume. For example, a 100ml volumetric flask will have a marking for 100ml, usually with a single line running across the flask, as in *Figure 13.3.1.1*.

13.3.2 Procedure to Dilute Solution

Let's say we wanted to obtain 250 ml of $0.100\ mol/dm^3\ CuSO_4$ from a stock solution (i.e. the solution provided) containing $1.00\ mol/dm^3\ CuSO_4$.

We first shake the stock solution to **ensure that it is well-mixed**.

Using $V_{conc}C_{conc} = V_{dil}C_{dil}$:

$$V_{conc}(1.00\ mol/dm^3) = 250\ ml \times 0.100\ mol/dm^3$$
$$V_{conc} = 25.0ml$$

Hence, we use a **pipette** to withdraw 25.0ml of concentrated solution and transfer it into a 250ml **volumetric flask**.

We then slowly add **deionized** water (DI water) until the volume of solution reaches the 250ml mark.

13.4 Concentration of Ions

Let's consider the following ionic equation:

$$K_2SO_4\ (aq)\ 4 \rightarrow 2K^+\ (aq) + SO_4^{2-}\ (aq)$$

This means, for every 1 mole of K_2SO_4, 2 moles of K^+ and 1 mole of SO_4^{2-} is produced after dissociation.

Let's say that the molarity of K_2SO_4 is 1M. Can we calculate the molarity of the ions? Yes!

Since the volume of solution remains the same after dissociation, but the number of moles doubles, we have that the molarity of K^+ is simply double that of K_2SO_4, which is 2M.

Similarly, the molarity of SO_4^{2-} is then 1M.

Topic 14: Volumetric Analysis

Disclaimer: These notes… suck.

14.1 What is Volumetric Analysis?

Volumetric Analysis, sometimes referred to as **VA** for short, is the analysis of composition by measuring the volume of one solution needed to react with the given volume of another solution.

To do Volumetric Analysis, we use a method called **"titration"**.

14.2 Titration

In titration, we add a solution of **accurately known molar concentration** (called the "standard solution") into another solution of unknown concentration, gradually, until the reaction is **complete**.

If we know the volumes of the standard and unknown solutions used in the titration, as well as the stoichiometric equation and the concentration of the standard solution, we can calculate the concentration of the unknown solution.

14.2.1 Equivalence Point

The equivalence point, in an acid-base reaction, is the point at which the **acid has completely reacted** (been neutralised) **by the base**.

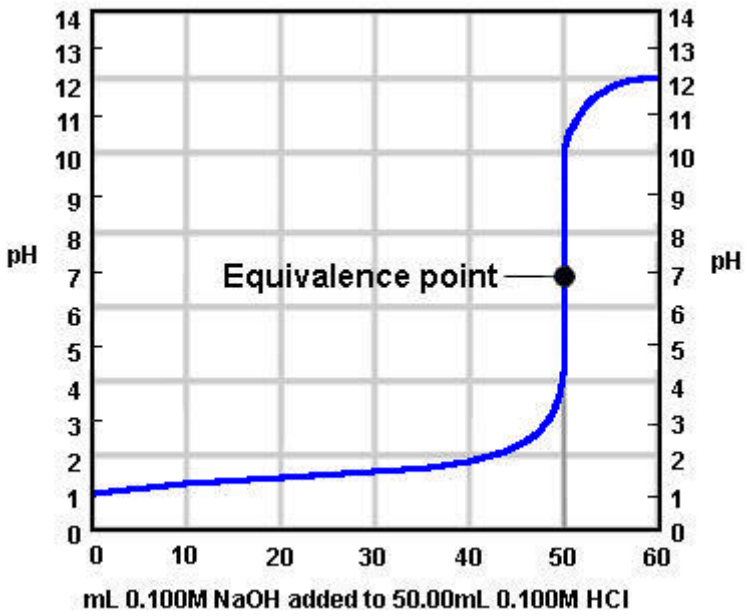


Figure 14.2.1.1 - Graph to Determine Equivalence Point

In an acid-base titration, up to the equivalence point, either an acid or a base is present **in excess**. After the equivalence point, the other one becomes present **in excess**.

We can use a **pH indicator** to determine when this change has occurred. The point at which this change occurs is called the **end point** of the reaction.

14.2.2 Setup

The Setup for titration should look something like:

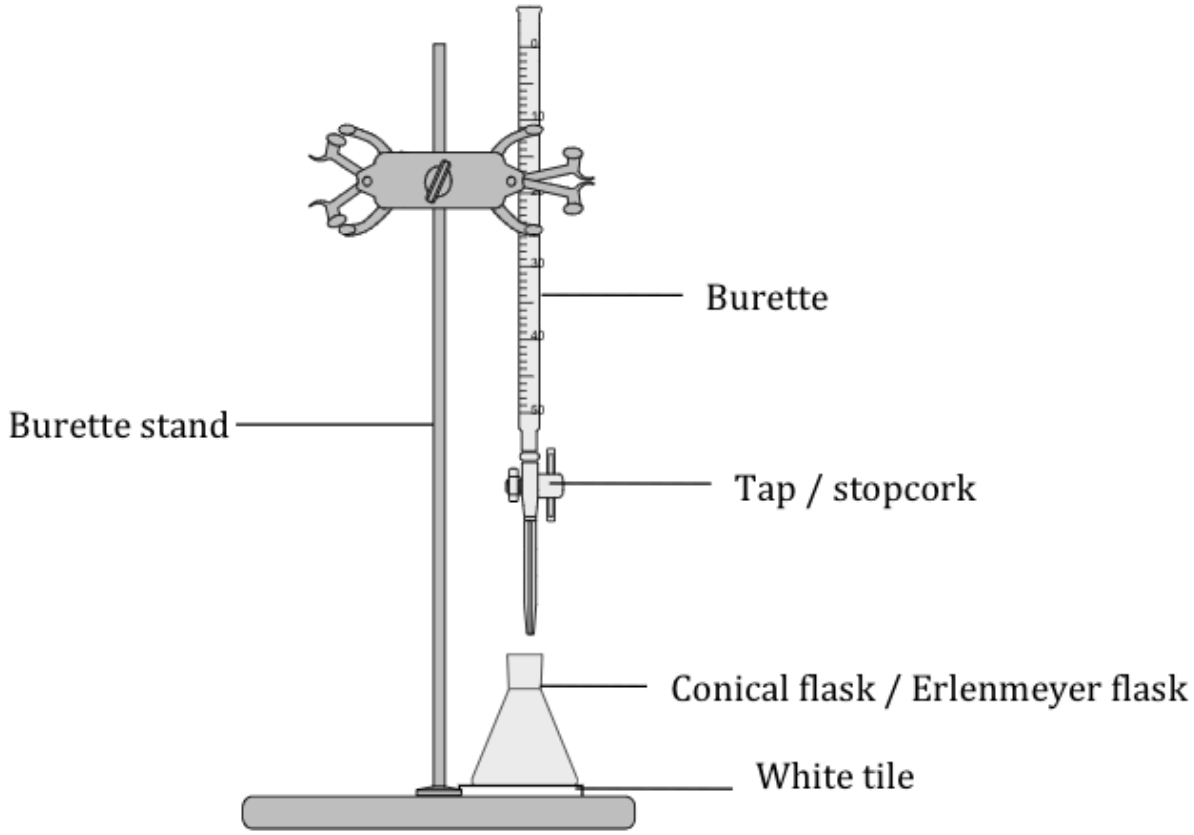


Figure 14.2.2.1 - Setup for Titration

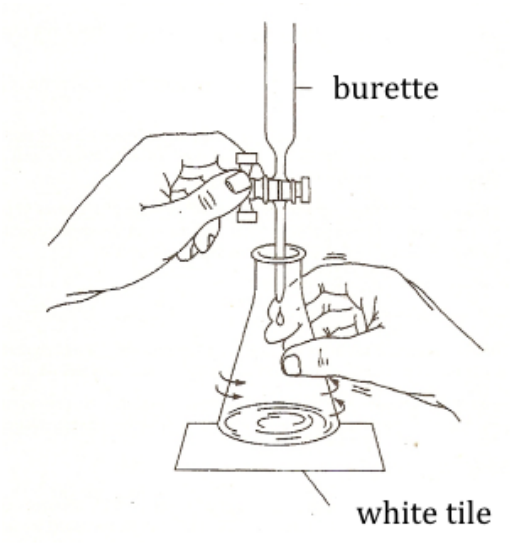


Figure 14.2.2.2

In *Figure 14.2.2.1*, you can see what a standard titration set-up looks like.

While carrying out the titration, you should have your non-dominant hand to open and close the tap which releases the solution from the burette, while the **dominant hand swirls and mixes the two solutions well** for the reaction to take place, like in *Figure 14.2.2.2*.

14.2.3 Titration Experiment Techniques and Details

Here is the apparatus list required for titration:

1. a $50\ cm^3$ burette
2. a burette stand or retort stand
3. a filter funnel
4. a pipette (usually $25cm^3$)
5. a pipette filler
6. 2 conical flasks
7. a white tile
8. a suitable pH indicator
9. a wash bottle containing deionised water (DI Water)

14.2.3.1 The Burette

The burette is a long glass tube which has markings from $0cm^3$ to $50\ cm^3$ (usually). However, one thing to note when you are using a burette is that the zero mark is **NOT** at the bottom. Instead, **the zero mark is at the top of the burette**.

A nozzle or tap at the bottom of the burette allows liquid to flow through.

Measurements using a burette should be to the nearest $0.\ 05cm^3$.

14.2.3.2 Procedure for Preparation of a Burette

Step 1: Wash the burette with **tap water**.

Step 2: Rinse the burette with **DI water.**

Step 3: Rinse the burette with the **solution to be put into the burette** (this ensures the concentration of the solution is not affected by the water).

Step 4: Place the burette on the **burette stand**.

Step 5: Bring the setup down to **fill the burette** and place a filter funnel on the top of the burette.

Step 6: Pour in solution through the filter funnel, into the burette, until the level of the solution is just above the zero mark.

Step 7: Remove the filter funnel and allow some solution to drain out to ensure all gas bubbles are flushed out.

You're now ready to use the burette for titration!

14.2.3.3 The Pipette

A Pipette is an apparatus made of glass that is used to **measure a fixed amount of liquid**. Most pipettes for titration are $20\ cm^3$ or $25\ cm^3$.

14.2.3.4 Procedure for Preparation of a Pipette

Step 1: Rinse the pipette with **DI water once** and the solution that it will be filled with twice.

Step 2: Insert the pipette into a pipette filler. Press "A" and squish the pipette filler.

Step 3: Dip the tip of the pipette into the solution and **hold down** "S" until just above the **graduation mark**.

Step 4: Hold down "E" until the bottom of the meniscus of the solution in the pipette is exactly on the graduation mark.

14.2.3.5 Procedure for Preparation of Conical Flask

Step 1: Wash the Conical Flask with **tap water.**

Step 2: Rinse the Conical Flask with **DI water**.

Step 3: Shake off all the **DI water** from the conical flask, which should be dry at the start of a titration experiment.

14.3 pH Indicators

Here is a table of common pH indicators:

Indicator Name	Colour in Strong Acid	pH where colour change is observed	Colour in Strong Base
Methyl Orange	Red	3 - 5	Yellow
Screened Methyl Orange	Purple	3 - 5	Green
Bromothymol Blue	Yellow	6 - 8	Blue
Thymol Blue	Yellow	8 - 10	Blue
Phenolphthalein	Colourless	8 - 10	Pink
Thymolphthalein	Colourless	8 - 10	Blue

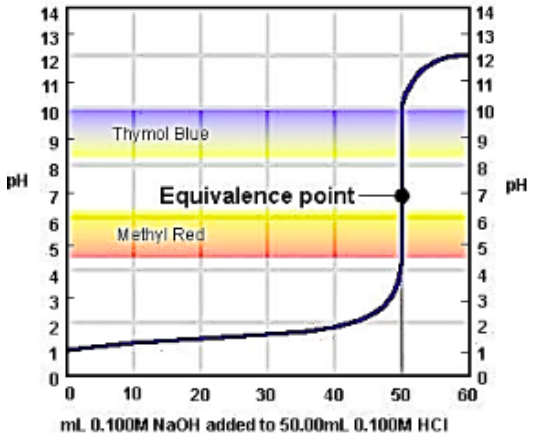
Table 14.3.1 Common pH Indicators

14.4 Acid-Base Titration Curves

But if there are so many pH indicators to choose from, how do we know which one we should use?

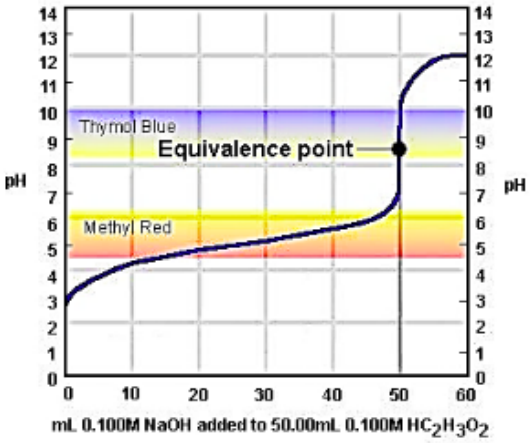
Let's just take the examples of Thymol Blue and Methyl Red.

We can plot the pH of the solution against the volume of $NaOH$ added to Acid for both Acetic Acid and Hydrochloric Acid in a graph known as the **titration curve**.



strong base and strong acid

Figure 14.4.1



strong base and weak acid

Figure 14.4.2

Refer to *Figure 14.4.1* above. As you can see, for a strong acid and a strong base, both of these pH indicators can be used, since the slope of the titration curve is so steep.

Refer to *Figure 14.4.2* above. However, for a strong base and a weak acid, the titration curve is not as steep and we can see that Methyl Red is not appropriate for this type of titration since its **end point** is far away from the **equivalence point**. On the contrary, Thymol Blue is a good indicator to use.

If we keep making these types of observations, we can create this table:

Type of Titration	Suitable pH indicators
Strong Acid - Strong Base	All
Strong Acid - Weak Base	Methyl Orange, Screened Methyl Orange
Weak Acid - Strong Base	Phenolphthalein, Thymol Blue, Thymolphthalein

Weak Acid - Weak Base	None

Table 14.4.3 Suitable pH Indicators

14.5 Requirements of a Reaction Suitable for Titration

Firstly, the reaction must be **fast**, like an acid-base reaction, so that the titration can (theoretically) be performed in a short time.

Secondly, the reaction must essentially **go to completion,** which means there are no **reactants** left[verify].

Thirdly, the reaction must not have **any side reactions**.

Lastly, it must be possible (e.g. through a pH indicator) to determine the **equivalence point** of the reaction.

Topic 15: Salt Preparation

Best Topic! Very fun. Easiest way to prepare sodium chloride!

15.1 Recap - What is a Salt?

A salt is an **ionic compound** and a salt is formed when a metallic ion or an ammonium (NH_4^+) ion **replaces one or more** hydrogen (H^+) ions of an acid.

For example, take the salt zinc sulfate. It is formed by neutralisation between zinc hydroxide and sulfuric acid.

$$Zn(OH)_2\ (s)\ +\ H_2SO_4\ (aq) \rightarrow 2H_2O\ (l)\ +\ ZnSO_4\ (aq)$$

15.2 Recap - Water of Crystallisation

Many salts combine with **water molecules** to form **crystals**. These water molecules are then called the **water(s) of crystallisation**, and these crystals are called the **hydrated salts** (hydrated: with water).

We can then remove the water of crystallisation by **heating**, and in doing so, we get an **anhydrous salt**, which often takes the form of a powder.

The opposite also works - we can add water to an anhydrous salt in order to obtain a **hydrated salt**.

15.3 Common Colours of Salts

Some ions tend to form salts with a unique colour.

Ion	Colour of	
	Aqueous Solution	Precipitate
Iodide (I^-)	Colourless	Yellow
Chloride (Cl^-)	Colourless	White
Aluminium (Al^{3+}), Zinc (Zn^{2+}), Lead (II) (Pb^{2+})	Colourless	White
Copper (II) (Cu^{2+})	Blue or Green	Blue except green $CuCO_3$

Iron (II) (Fe^{2+})	Pale Green	Dirty Green
Iron (III) (Fe^{3+})	Yellow	Reddish Brown

Table 15.3.1 - Colours of Salts

15.4 Reminder - Solubility Rules

If you still don't remember your Solubility Rules, time to revisit **Topic 9: Solubility Rules**!

15.5 Preparation of Soluble Salts

If you are trying to prepare a soluble salt, such as Zinc Sulfate ($ZnSO_4$), you should consider using one of the following methods corresponding to the metal in the table below.

Let's consider the reactivity series of metals.

<table>
<tr><th>Metal</th><th>Method to Prepare</th></tr>
<tr><td>Potassium (K)</td><td rowspan="2">Acid + Alkali
Acid + Soluble Carbonate</td></tr>
<tr><td>Sodium (Na)</td></tr>
<tr><td>Calcium (Ca)</td><td>Acid + Insoluble Metal Oxide
Acid + Insoluble Metal Carbonate</td></tr>
<tr><td>Magnesium (Mg)</td><td rowspan="5">Acid + Metal
Acid + Insoluble Metal Oxide
Acid + Insoluble Metal Carbonate</td></tr>
<tr><td>Aluminium (Al)</td></tr>
<tr><td>Zinc (Zn)</td></tr>
<tr><td>Iron (Fe)</td></tr>
<tr><td>Tin (Sn)</td></tr>
<tr><td>Lead (Pb)</td><td></td></tr>
<tr><td>Copper (Cu)</td><td rowspan="5">Acid + Insoluble Metal Oxide
Acid + Insoluble Metal Carbonate</td></tr>
<tr><td>Mercury (Hg)</td></tr>
<tr><td>Silver (Ag)</td></tr>
<tr><td>Gold (Au)</td></tr>
<tr><td>Platinum (Pt)</td></tr>
</table>

Table 15.5.1

15.5.1 Acid + Insoluble Base/Metal/Insoluble Carbonate

For this method, the base/metal/carbonate must be **insoluble** and must be **present in excess**. This is because the insoluble base/metal/carbonate can be easily filtered away.

Let's say we want to prepare Magnesium Sulfate ($MgSO_4$).

We see that $MgSO_4$ is a **soluble salt**, and that Mg can be prepared with an **Insoluble Base, Metal or Carbonate**.

We then check that Mg, $Mg(OH)_2$ and $MgCO_3$ are in fact, all insoluble, and hence can all be used.

Hence, we can use any of these three reactions:

$$Mg\ (s)\ +\ H_2SO_4\ (aq) \rightarrow H_2\ (g)\ +\ MgSO_4\ (aq)$$
$$Mg(OH)_2\ (s)\ +\ H_2SO_4\ (aq) \rightarrow 2H_2O\ (l)\ +\ MgSO_4\ (aq)$$
$$MgCO_3\ (s)\ +\ H_2SO_4\ (aq) \rightarrow H_2O\ (l)\ +\ CO_2\ (g)\ +\ MgSO_4\ (aq)$$

Let's use the third reaction as an example.

Step 1: Measure 10ml of $H_2SO_4\ (aq)$ using a measuring cylinder and add the H_2SO_4 into a beaker.

Step 2: Consistently add $MgCO_3\ (s)$ powder, stirring constantly, until no more effervescence forms and when no more $MgCO_3\ (s)$ dissolves.

Step 3: Filter the mixture. Dispose the $MgCO_3\ (s)$ residue left on the filter paper and collect the filtrate, which will be $MgSO_4\ (aq)$.

Step 4: Begin crystallisation by heating $MgSO_4$ solution in an evaporating dish to evaporate some of the water.

Step 5: Test for saturation constantly using a clean, dry, cold glass rod, dipping it into the filtrate. If the filtrate is saturated, tiny crystals will form on the glass rod. When the filtrate is saturated, stop heating the filtrate.

Step 6: Leave the filtrate to cool and crystallise on the evaporating dish.

Step 7: Wait for the crystals to precipitate out and filter to collect the crystals.

Step 8: Wash the crystals under cold DI water to remove impurities.

Step 9: Dry the crystals between filter papers.

Now we have finished the preparation procedure and we have some saturated $MgSO_4$ crystals.

15.5.2 Titration (Acid + Alkali/Soluble Carbonate)

Let's say we wanted to, now, prepare Sodium Nitrate ($NaNO_3$).

We see that $NaNO_3$ is a **soluble salt**, and that Na can be prepared with a **Soluble Base or Carbonate**. (*We don't use the metal because Na is extremely reactive and explosive.*)

We then check that $NaOH$ and Na_2CO_3 are in fact, both soluble, and hence can both be used, using titration.

The two possible reactions to consider are:

$$NaOH\ (aq)\ +\ HNO_3\ (aq) \rightarrow H_2O\ (l)\ +\ NaNO_3\ (aq)$$
$$Na_2CO_3\ (aq)\ +\ 2HNO_3\ (aq) \rightarrow H_2O\ (l)\ +\ CO_2(g)\ +\ 2NaNO_3\ (aq)$$

Let's use the first one.

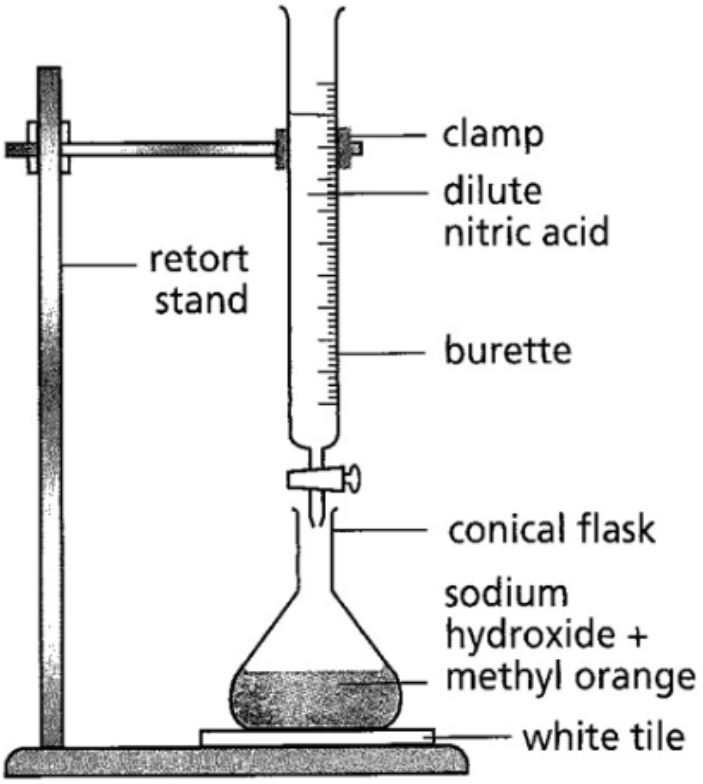


Figure 15.5.2.1 - Titration Setup

Step 1: We start by performing titration. Fill a burette with **dilute** HNO_3. Record the initial burette reading as $V_{initial}$.

Step 2: **Pipette** 25.0 ml of $NaOH$ solution into a conical flask.

Step 3: Add 2 drops of **Methyl Orange** indicator to the $NaOH$ solution. The solution should turn yellow.

Step 4: While swirling the conical flask with your dominant hand, slowly add dilute HNO_3 into the conical flask until the solution turns orange. This is the **end point**. Record the final burette reading as V_{final}.

Step 5: Find the amount of acid that was used using $V = V_{initial} - V_{final}$.

Step 6: Pipette 25.0 ml of $NaOH$ solution into a beaker, then add V ml of dilute HNO_3 from the same burette from earlier. We now have $NaNO_3$ solution.

Step 7: Begin crystallisation by heating $NaNO_3$ solution in an evaporating dish to evaporate some of the water.

Step 8: Test for saturation constantly using a clean, dry, cold glass rod, dipping it into the filtrate. If the solution is saturated, tiny crystals will form on the glass rod. When the solution is saturated, stop heating the solution.

Step 9: Leave the solution to cool and crystallise on the evaporating dish.

Step 10: Wait for the crystals to precipitate out and filter to collect the crystals.

Step 11: Wash the crystals under cold DI water to remove impurities.

Step 12: Dry the crystals between filter papers.

Now we have hydrated $NaNO_3$ crystals!

Pause and Ponder 15.5.2

Why do we have to use a new setup instead of just using the solution obtained after titration? (i.e. Why can't we just perform crystallisation on Step 5?)

[Click to Go to Solution]

15.6 Preparation of Insoluble Salts

To prepare an insoluble salt, we use a type of reaction known as a **precipitation reaction**.

The general formula for preparing an insoluble salt AY is:

$$AX\ (aq)\ +\ BY\ (aq) \rightarrow AY\ (s)\ +\ BX\ (aq)$$

Let's say we want to prepare the salt Barium Sulfate ($BaSO_4$).

We first check that $BaSO_4$ is insoluble in water.

Hence, we use the precipitation method. We "plug in" the chemical formula $BaSO_4$ into AY:

$$BaX(aq)\ +BSO_4\ (aq) \rightarrow BaSO_4\ (s)\ +BX\ (aq)$$

Note: The blue chemicals are unknowns to be determined, not the chemical symbols.

We need to find X such that BaX is soluble. One such X is $(NO_3)_2^{2-}$.

As for B, we need BSO_4 to be soluble. One such B is Fe^{2+}.

Hence, we have the reaction:

$$Ba(NO_3)_2\,(aq)\ +\ FeSO_4\,(aq) \rightarrow BaSO_4\,(s)\ +\ Fe(NO_3)_2\,(aq)$$

Let's carry out this procedure.

Step 1: Pour 10.0 ml of $Ba(NO_3)_2$ solution into a beaker.

Step 2: Keep adding $FeSO_4$ solution until no more precipitate forms.

Step 3: Filter to collect the precipitate, which is $BaSO_4\,(s)$.

Step 4: Wash the precipitate with a small amount of DI water to remove impurities.

Step 5: Dry the precipitate between filter papers.

And we're done! *Fun!*

Topic 16: Structure and Bonding (Metallic)

16.1 The Different Types of Structure in Matter

Figure 16.1.1 is a flowchart on the different possible types of structure in matter.

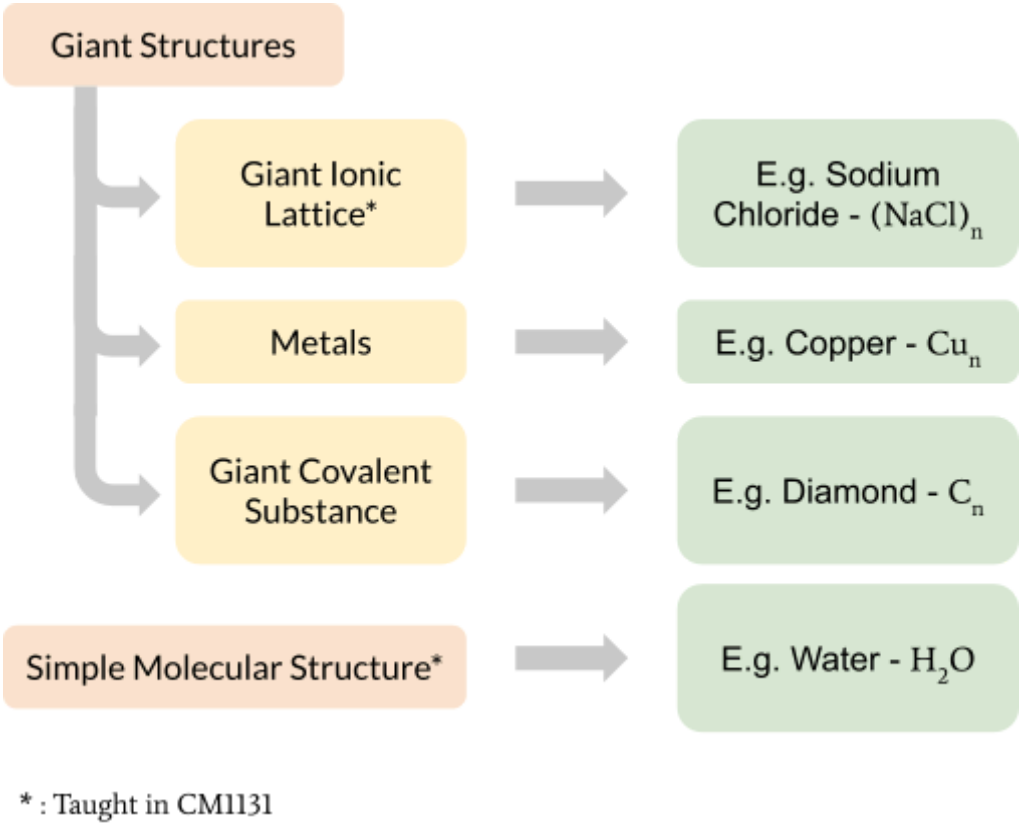


Figure 16.1.1 - The Different Types of Bonding

For Metallic Bonding, continue to ***Section 16.2****.*
For Giant Covalent Bonding, go to ***Topic 17: Structure and Bonding (Giant Covalent)****.*

Notice how in sodium chloride, the formula included is $(NaCl)_n$. This is due to the fact that it has a **giant ionic lattice**, as the subscript $_n$ simply denotes that the **empirical unit** is repeated over and over again.

The same goes for other giant structures. For example, diamond, which has a **giant covalent structure,** is reflected as C_n.

On the other hand, since water has a **simple molecular structure**, there is no subscript $_n$ in its formula.

16.2 Metallic Structure and Bonding

Substances with metallic structures include **pure metals** like aluminium, and **alloys**, which are **mixtures of different metals**.

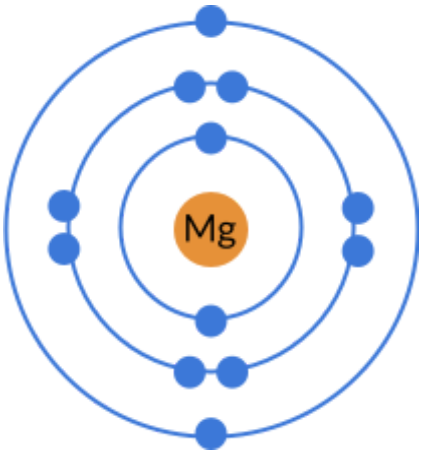


Figure 16.2.1

Let's consider the electronic configuration for Magnesium, which is **2.8.2**.

Let's focus on the two valence electrons. These electrons are **not strongly attracted** to the **positively** charged nucleus, as they are shielded by the inner electron shells. This makes it easy for metals to form ions.

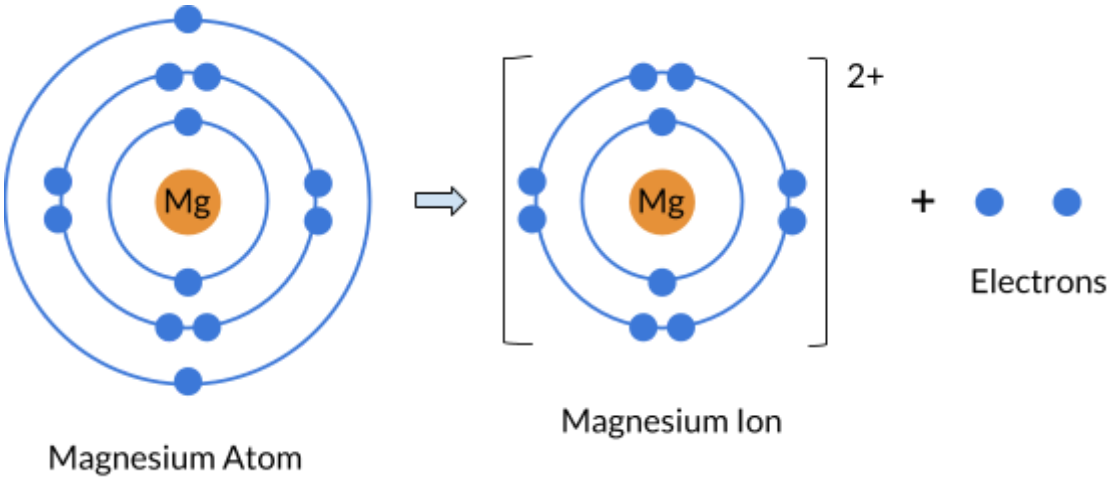


Figure 16.2.2 - Formation of Metal Ion (Taken from Topic 5 Notes)

Hence, metal atoms tend to lose their valence electrons, forming positively charged metal ions. The valence electrons then do not belong to any metal ion and are said to be **delocalised**.

Hence, the electrons from the magnesium atom in *Figure 16.2.2* are **delocalised** and have become **mobile**.

As a result, the structure of a metal can be described as "**positive metal ions surrounded by a sea of delocalised mobile electrons"**, as in *Figure 16.2.3* below.

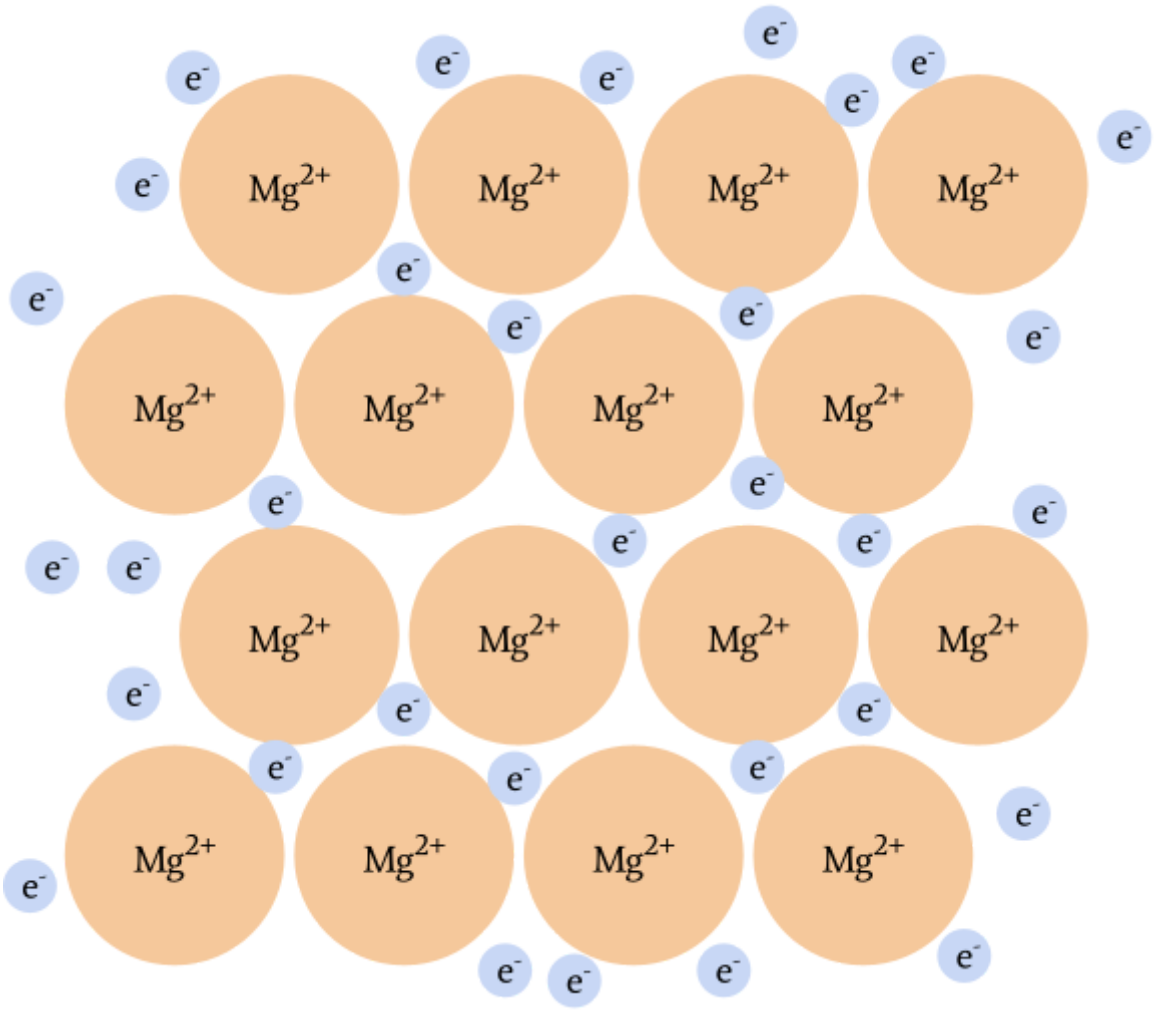


Figure 16.2.3 - Magnesium Metallic Lattice

This arrangement of metal ions and electrons is called a **metallic lattice**.

Hence, **Metallic Bonding** is the **strong electrostatic force of attraction between the positively charged metal ions** and **the negatively charged sea of delocalised mobile electrons**.

Note: Delocalised means that the electron(s) are not associated with any one atom and are simply "floating around". On the contrary, localised means that the electron(s) are associated with an atom.

16.2.1 Density of Metals

Metals have a **high density** because they have a closely packed, regular structure (Refer to *Figure 16.2.1* below) held together by strong metallic bonds.

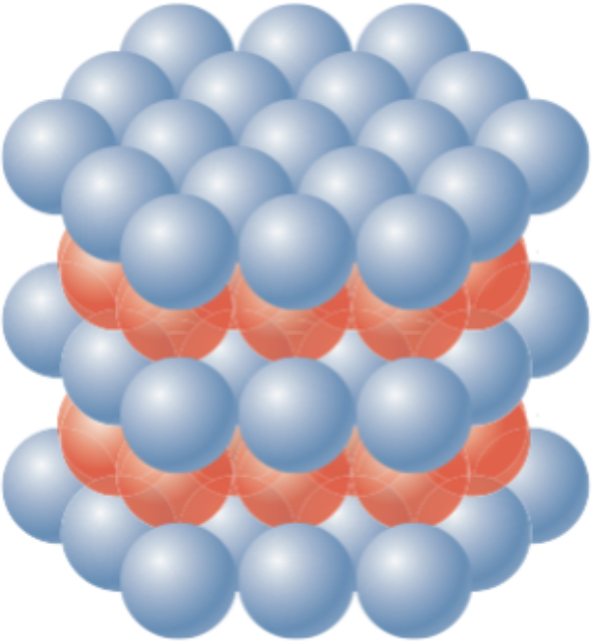


Figure 16.2.1.1 - Metallic Lattice

The majority of metals have a higher density than non-metals. (Except Group 1 Metals, which are sometimes known as the "light metals" due to their relatively low density.)

The least dense solid metal is **lithium**, *(FYI:* $0.534\ g/cm^3$*)*, while the densest solid metal is **osmium** *(FYI:* $22.6\ g/cm^3$*)*.

This shows that metals have a **wide variation** in terms of density.

16.2.2 Melting and Boiling Points of Metals

The strong **electrostatic forces of attraction** between the **positively charged metal ions** and the **negatively charged delocalised sea of electrons** require a lot of energy to overcome, so **metals have a high melting and boiling point**.

16.2.3 Conductivity of Electricity of Metals

The **delocalised electrons** will move towards the positively charged end when a potential difference is applied, as shown in *Figure 16.2.3.1* below.

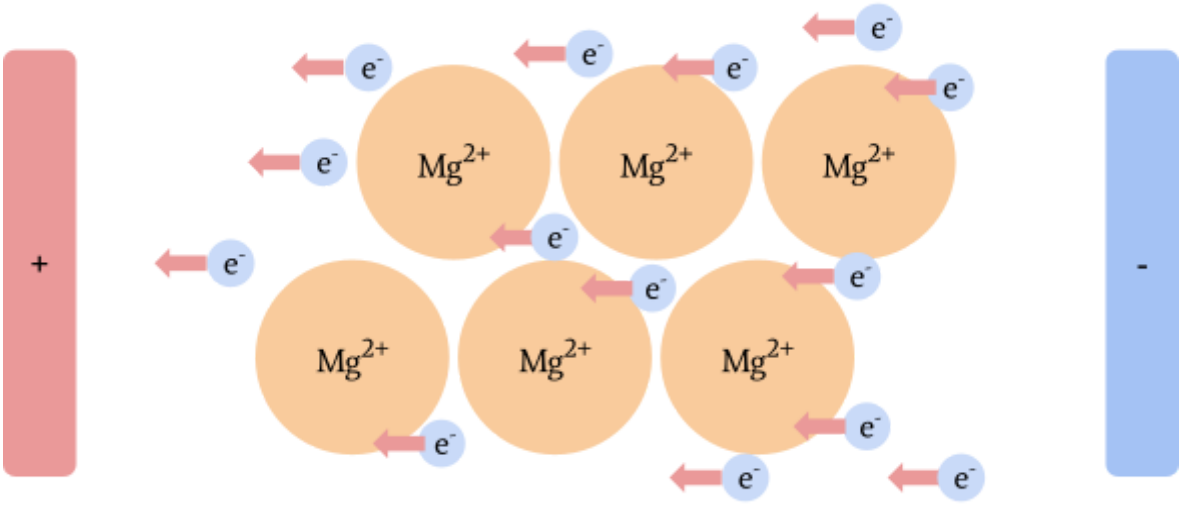


Figure 16.2.3.1

The movement of this electrical charge, denoted by the red arrows in this case, is called the **electric current**.

Since metals have this sea of delocalised electrons that can carry the electrical charge, metals are hence **good conductors of electricity**.

16.2.4 Conductivity of Heat of Metals

Recall that metal atoms are **closely packed together**. Hence when one end of the metal is heated, the heat can be transferred from one atom to the next by vibrations.

The mobile electrons also gain energy, and collide with each other more, transferring some of the heat energy too.

Based on these two reasons, metals are able to conduct heat quickly.

16.2.5 Malleability and Ductility of Metals

As we can recall from CM1131, metals are **malleable** and **ductile**.

This is possible because the metallic bonding remains largely **undisrupted** when a force is applied to it, as seen in *Figure 16.2.5.1*.

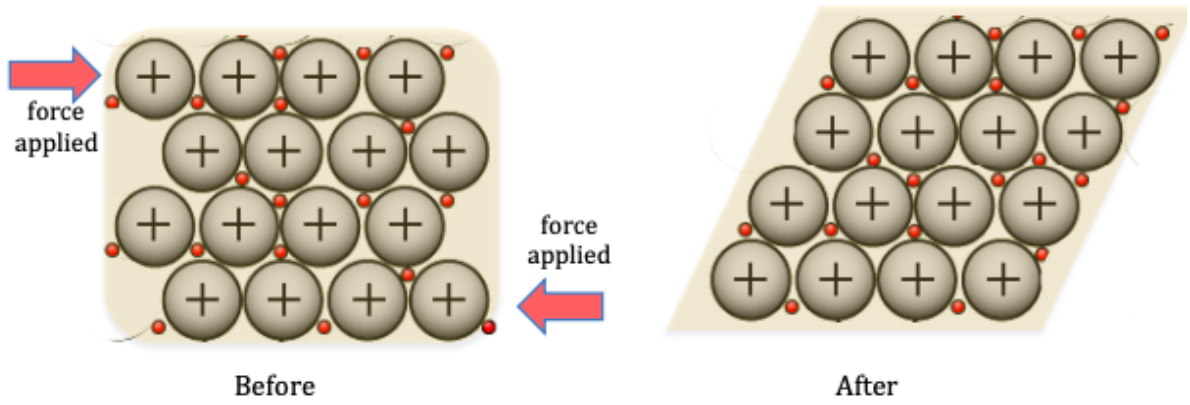


Figure 16.2.5.1

In Summary, the layers of metal cations are able to **slide over one another** through the sea of delocalised electrons to take up **new positions without disrupting the metallic bonding**.

16.3 Alloys

Consider *Figure 16.3.1* and *Figure 16.3.2* below.

Figure 16.3.1 - A Pure Metal

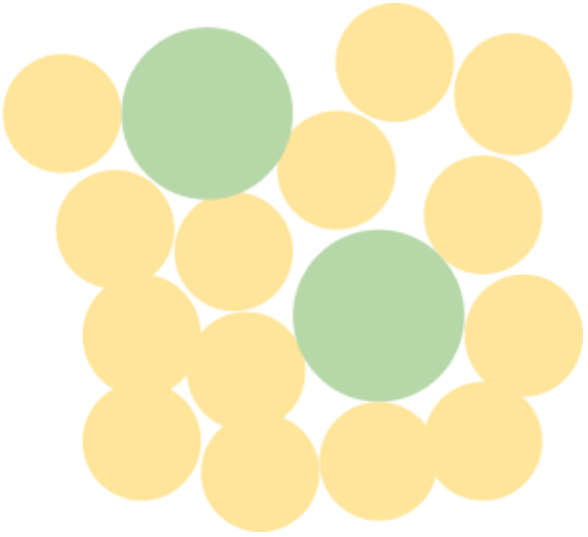


Figure 16.3.2 - An Alloy

The metal atoms in a pure metal are all of the same size. Hence, the metal cations are able to **slide over one another**, but in an alloy, the metal atoms are not all of the same size, which means the metal cations are **not able to slide over one another as easily** as the orderly arrangement is disrupted.

Hence, alloys tend to be **stronger and harder** than pure metals.

Alloy	Composition	Properties	Uses
Brass	Copper and Zinc	Doesn't Corrode Easily, Looks Like Gold	Coins and Musical Instruments

Stainless Steel	Iron, Chromium, Nickel, Carbon	Resistant to Corrosion, Strong	Cutlery and Utensils
Solder *(Yes, the thing in DV D&E)*	Tin, Lead	Low Melting Point	Joining Metals (i.e. Soldering)
Pewter	Tin, Antimony, Copper	Bright, Shiny, Looks Like Silver	Decorative Ornaments

Table 16.3.1 - Common Alloys and Their Properties

Topic 17: Structure and Bonding (Giant Covalent)

17.1 Giant Covalent Structure

Covalent substances may exist either as Simple Molecular *(Covered in **CM1131 Topic 7: Structure and Bonding (Simple Molecular)**)*, or as Giant Covalent Structures.

A Giant Covalent Structure is when a giant network of atoms are covalently bonded to each other. This will cause it to display **very different** properties from Simple Molecular Structures.

The three main Giant Covalent Structure substances to know in CM2131 are:

1. Diamond
2. Graphite
3. Silicon Dioxide (Sand; also known as Silica)

17.2 Allotropes

Allotropes are different forms of the same element, in the same physical state. For example, coal and charcoal are allotropes of carbon, since they are both solid forms of carbon.

17.2.1 Allotropes of Carbon

Carbon has many allotropes, mainly:

1. Diamond
2. Graphite
3. Fullerene
4. Amorphous Carbon (Coal, Charcoal, etc.)

17.3 Diamond

Diamond is an **allotrope** of carbon, meaning it is made up of only carbon atoms.

In diamond, every carbon atom is covalently bonded to 4 other carbon atoms in a tetrahedral manner, forming a **giant covalent structure**. *Figure 17.3.1* below shows the structure of diamond.

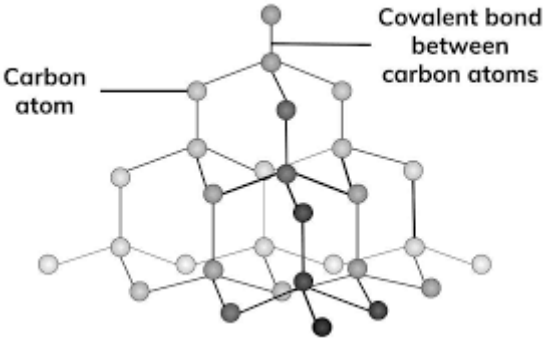


Figure 17.3.1 - Structure of Diamond

The Carbon-Carbon bonds present in diamond are very strong, and the combined strength of all these C-C bonds is what gives diamond its unique properties:

1. Diamond has high **hardness**
2. Diamond has a **very high melting/boiling point** *(FYI: 3550°C/4827°C respectively)*
3. Diamond has **very low chemical reactivity**

17.3.1 Explanations for Properties of Diamond

Diamond has **high melting/boiling point** because **a large amount of energy** is required to **overcome the strong covalent bonds** between the carbon atoms **in the Giant Covalent Structure.**

Diamond **does not conduct electricity** due to an **absence of delocalised electrons** as all the valence electrons are used for covalent bonding and hence **an absence of mobile charge carriers**.

17.4 Graphite

Graphite is another **allotrope** of carbon. Like diamond, it is only made of carbon atoms.

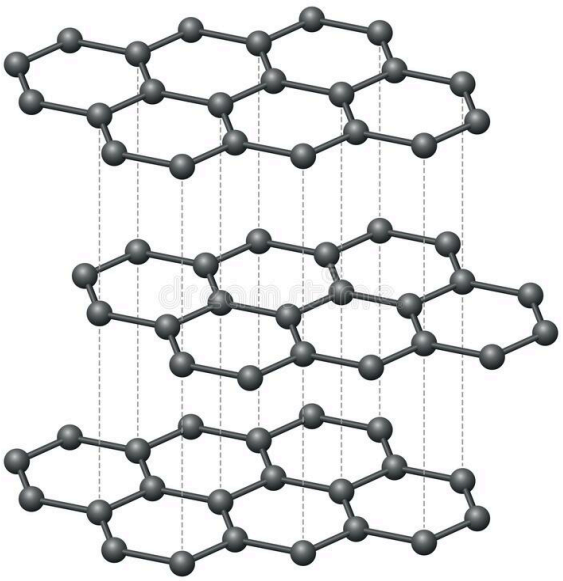


Figure 17.4.1 - Structure of Graphite

As you can see from *Figure 17.4.1*, graphite consists of several flat parallel layers of carbon layered on top of one another.

In each layer, each carbon atom forms strong covalent bonds with 3 other carbon atoms, forming a regular hexagonal structure.

Between the layers, however, there is no strong bonding, and hence the layers can be easily separated from each other. The layers of carbon, held together by weak forces of attraction lie on top of one another.

This also means that the layers are able to slide over one another easily, which gives graphite its **softness** and **lubricating power**.

17.4.1 Explanations for Properties of Graphite

Graphite has high **melting/boiling** point because even though the forces of attraction between layers are weak and can be easily overcome, the **strong covalent bonds between the carbon atoms on each layer** still require **a lot of energy** to overcome to induce a change in state.

Graphite is a **good conductor of electricity** because each carbon atom has one electron that is not involved in **covalent bonding**, and that electron becomes delocalised. (remember to always check for delocalised electrons if structure is given, e.g. in boron nitride) All the delocalised electrons can act as **mobile charge carriers** that can move freely from one carbon atom to the next, giving graphite its electrical conductivity.

Graphite is **soft and slippery** because when a force is applied, the weak forces of attraction between the layers of carbon atoms is easily overcome and the layers slide over one another.

17.5 Silicon Dioxide

Silicon Dioxide or Silicon (IV) Oxide, more commonly known as **silica** or **sand**, has a **giant covalent structure**.

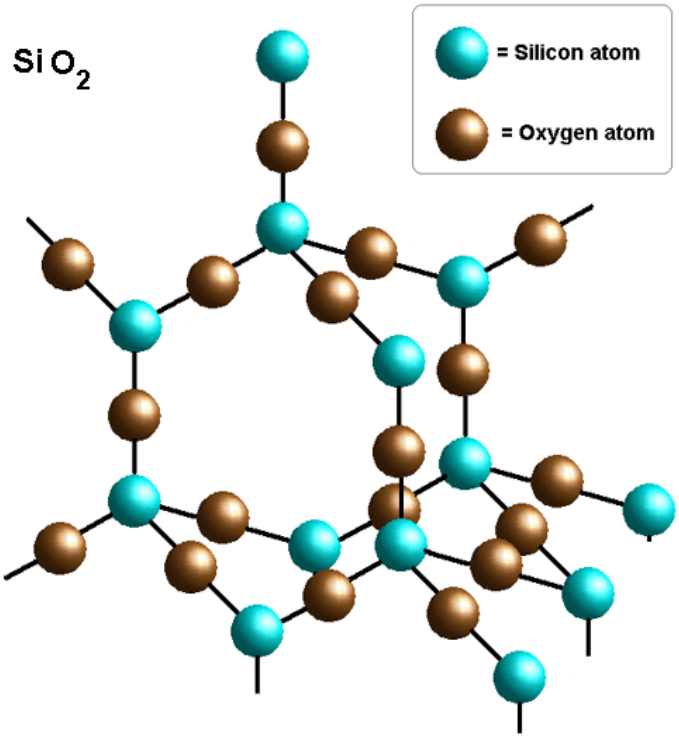


Figure 17.5.1 - Structure of Silica

As you can see from *Figure 17.5.1* above, each silicon atom is covalently bonded to 4 oxygen atoms in a tetrahedral manner, and each oxygen atom is covalently bonded to 2 silicon atoms.

Pause and Ponder 17.5

Since Silicon Dioxide has a similar structure to diamond, determine its electrical conductivity and whether it has a high or low melting/boiling point. Explain why.

[Click to Go to Solution]

17.6 Solubility of Giant Covalent Substances

Giant Covalent Substances tend to be **insoluble in all solvents**.

17.7 Summary of All Types of Bonding

Table 17.7.1 below is a summary/comparison on all the types of bonding.

Structure	Ionic	Metallic	Simple Molecular	Giant Covalent	
Examples	Sodium Chloride $(NaCl)_n$	Aluminium $(Al)_n$	Water H_2O	Graphite C_n	Diamond and Silicon Dioxide C_n & $(SiO_2)_n$
Particles Present	Cations and Anions	Atoms (cations and electrons)	Molecules	Carbon Atoms	Atoms
Forces of Attraction	Electrostatic forces of attraction between oppositely charged ions	Electrostatic forces of attraction between metal cations and sea of delocalised electrons	Covalent Bonds within Molecule Inter-molecular forces of attraction between molecules	Covalent Bonds within Carbon Layer Forces of attraction between layers of atoms	Covalent Bonds throughout the Giant Covalent Structure between atoms
Melting and Boiling Points	High	High	Low	High	High

Physical State at r.t.p.	Solid	Solid	Liquid or Gas except Solid Iodine	Solid	Solid
Solubility in Water	Soluble	Insoluble	Insoluble	Insoluble	Insoluble
Solubility in Organic Solvent	Insoluble	Insoluble	Soluble	Insoluble	Insoluble
Electrical Conductivity	No in solid state; Yes in aqueous /molten states	Yes	No	Yes	No

Table 17.7.1 - Summary Table on Bonding

Topic 18: Redox Reactions

18.1 Oxidation

There are 4 ways for an atom, ion or molecule to **oxidise** (increase oxidation stage/number).

1. Loss of Electron
2. Loss of Hydrogen
3. Gain of Oxygen
4. Increase in Oxidation State

18.2 Reduction

Reduction can be thought of as the "inverse" of Oxidation. There are 4 ways for reduction to occur:

1. Gain of Electron
2. Gain of Hydrogen
3. Loss of Oxygen
4. Decrease in Oxidation State

18.3 Assigning Oxidation State and Oxidation Numbers

In general, the term **oxidation state** is used for ionic compounds like MgO, but the term **oxidation number** is used for covalent compounds like HCl, when **no electrons are transferred**.

Consider the reaction $2\ Mg(s)\ +\ O_2\ (g) \rightarrow 2\ MgO\ (s)$.

In this example, we can see that each Mg atom has lost 2 electrons to form Mg^{2+}, and each O atom has gained 2 electrons to form O^{2-}.

This means that the Mg an atom's oxidation state increases from 0 to +2. Oxidation has occurred.
On the other hand, the O the atom's oxidation state decreases from 0 to -2. Reduction has occurred.

In such a reaction, when both **oxidation** and **reduction** occurs, we call it a **Redox Reaction** (Reduction-Oxidation Reaction).

18.3.1 Electronegativity

Electronegativity is a measure of the ability of an atom to **attract electrons**. The larger the electronegativity value, the better the atom is at pulling electrons to it.

H 2.1																	He ---
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne ---
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.2	S 2.5	Cl 3.0	Ar ---
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La-Lu 1.1-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn ---
Fr 0.7	Ra 0.9	Ac-No 1.1-1.7															

Figure 18.3.1.1 - Pauling's Electronegativity Table

The highest electronegativity number is that of fluorine, 4.0, and the lowest is Francium, of 0.7.

Electronegativity increases across a period from left to right due to the increase in **nuclear charge**, and decreases down a group, due to the increase in **atomic radius**.

18.3.2 Oxidation Number

Sometimes, like in HCl, electrons are not transferred, since it is a covalent bond.
In these cases, we use the concept of **oxidation number**.

Let's refer to *Figure 18.3.1.1*. H has an electronegativity value of 2.1, while Cl has an electronegativity value of 3.0.

Since Cl has a greater electronegativity value than H, the electrons that are shared between H and Cl are pulled towards the Cl atom. (Denoted by the red arrow in *Figure 18.3.2.1* below)

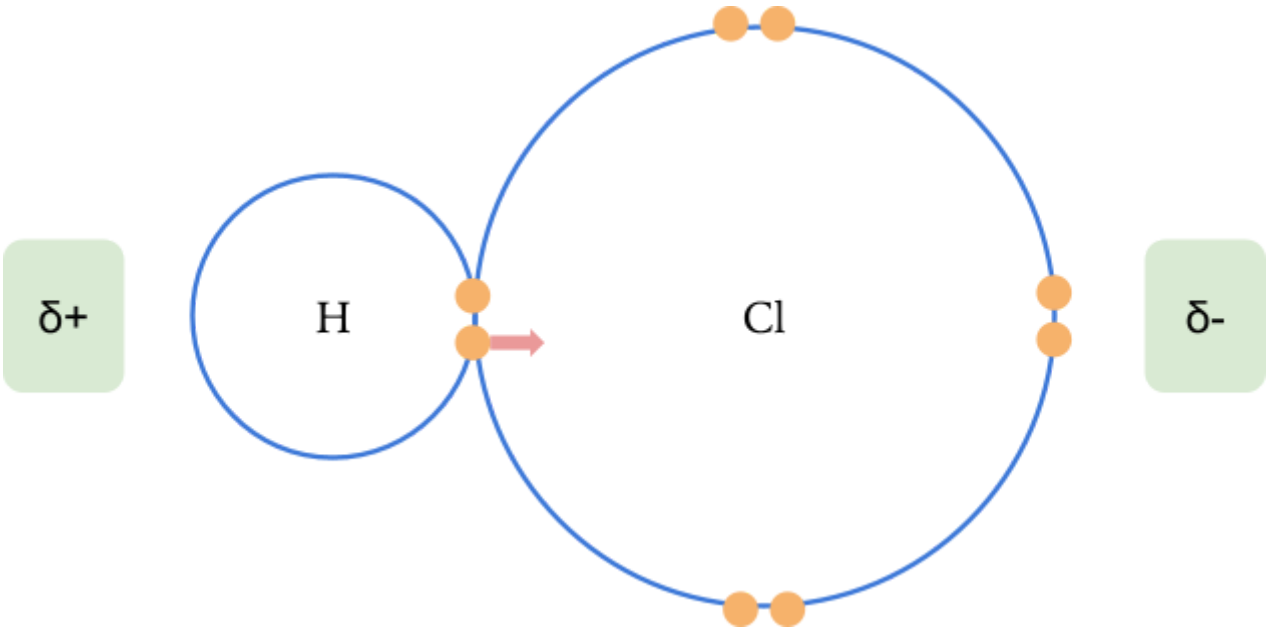


Figure 18.3.2.1 - Electronegativity in HCl

Now, we can determine the oxidation number of H and Cl.

The two parts of an oxidation number are its **sign** and its **value**.

A positive oxidation number indicates a **loss in electron control**, while a negative oxidation number indicates a **gain in electron control**.

The **value** of the oxidation number indicates the **number of electrons** for which **control has shifted**.

In the example, H **loses control** of **one electron,** so it has oxidation number +1.
Cl **gains control** of **one electron**, so it has oxidation number -1.

18.3.3 Rules to Follow to Assign Oxidation Number

Rule 1: Atoms of an element (e.g. Na in Na metal, C in diamond, O in O_2) all have oxidation numbers of **zero**.

Rule 2: In a simple ion (i.e. ion of a singular atom), the oxidation number is **equal** to the **charge of the ion**.

- Mg^{2+} has oxidation number +2.

Rule 3: In a neutral compound, all the oxidation numbers of the constituent atoms must add up to zero.

Rule 4: In a polyatomic ion, all the oxidation numbers of the constituent particles must add up to the charge on that ion.

- SO_4^{2-} has a sum of oxidation numbers of -2.

Rule 5: The usual oxidation number of an element is the charge on its most common ion.

Element	Usual Oxidation State	Exceptions
$Na,\ K$	+1	None
F	-1	None
O	-2	Peroxides (O_2^{2-}): -1 OF_2: +2
H	+1	Metal Hydrides (e.g. NaH): -1
Cl	-1	When combined with O, F: +1

Table 18.3.3.1

We should always try to assign oxidation numbers to the elements in *Table 18.3.3.1* first, as the other elements can be unpredictable sometimes.

Pause and Ponder 18.3.3

Why does oxygen have oxidation state +2 in Oxygen Difluoride?

[Click to Go Solution]

18.4 Examples

Worked Example 1, CM2131 Topic 9 Notes

Find the oxidation states of each element in $K_2Cr_2O_7$.

Solution

K has oxidation state +1.
O has oxidation state -2.

These two are determined from **Rule 5**.

Let the oxidation state of Cr be x.

Since the compound $K_2Cr_2O_7$ is neutral,

$$2(+1) + 2x + 7(-2) = 0$$
$$2x = +12$$
$$x = +6$$

Hence, the oxidation states of K, O and Cr are +1, -2 and +6 respectively.

Worked Example 2, CM2131 Topic 9 Notes

Determine the oxidation state of Mn in MnO_4^{2-}.

Solution

O has oxidation state -2.

Since MnO_4^{2-} has charge -2, the sum of oxidation states must be -2.

Let x be the oxidation state of Mn.

$$x + 4(-2) = -2$$
$$x = +6$$

Hence the oxidation state of Mn is +6.

18.5 Identifying Redox Reactions

We can define redox reactions in 4 ways.

18.5.1 In Terms of Gain/Loss of Oxygen

Consider $Fe_2O_3\,(s) + 3\,CO\,(g) \rightarrow 2\,Fe\,(s) + 3\,CO_2\,(g)$.

Fe_2O_3 loses oxygen to form $2\,Fe$. It is **reduced**.
$3\,CO$ gains oxygen to form $3\,CO_2$. It is **oxidised**.

Since **oxidation and reduction occurred**, this is a **Redox Reaction**.

18.5.2 In Terms of Gain/Loss of Hydrogen

Consider $H_2S\ (g)\ +\ Cl_2\ (g)\ \rightarrow\ 2\ HCl\ (g)\ +\ S\ (s)$.

H_2S loses hydrogen to form S. It is **oxidised**.
Cl_2 gains hydrogen to form 2 HCl. It is **reduced**.

Since **oxidation and reduction occurred**, this is a **Redox Reaction**.

18.5.3 In Terms of Gain/Loss of Electrons

Consider $Cu\ (s)\ +\ 2\ AgNO_3\ (aq)\ \rightarrow\ Cu(NO_3)_2\ (aq)\ +\ 2\ Ag\ (s)$.

Cu loses 2 electrons to form Cu^{2+} in $Cu(NO_3)_2$. It is **oxidised**.
Ag^+ in $AgNO_3$ gains 1 electron to form Ag. It is **reduced**.

Since **oxidation and reduction occurred**, this is a **Redox Reaction**.

18.5.4 In Terms of Increase/Decrease in Oxidation State

Consider $Mg\ (s)\ +\ 2\ HCl\ (aq)\ \rightarrow MgCl_2\ (aq)\ +\ H_2\ (g)$.

Mg in $Mg\ (s)$ has oxidation state 0, and in $MgCl_2$, It has oxidation state +2. It is **oxidised**.
H in HCl has oxidation state +1, and in H_2, it has oxidation state 0. It is **reduced**.

Since **oxidation and reduction occurred**, this is a **Redox Reaction**.

18.6 Oxidising and Reducing Agents

An Oxidising Agent (or oxidant) is a substance that causes another substance to be **oxidised**. In the process, it itself is **reduced**.

A Reducing Agent (or reductant) is a substance that causes another substance to be **reduced**. In the process, it itself is **oxidised**.

Consider the following reaction:

$$CuO\ +\ Mg \rightarrow Cu\ +\ MgO$$

In this reaction, CuO is the **oxidising agent** as it gives Mg an oxygen atom, which causes Mg to oxidise and form MgO. In doing so, CuO itself loses an oxygen atom to form Cu and is hence **reduced**.

18.6.1 Strong and Weak Reducing Agents

Table 18.6.1.1 below is a part of the reactivity series of metals.

Strongest Reducing Agent ↓	Magnesium (Mg)
	Aluminium (Al)
	Zinc (Zn)
	Iron (Fe)
	Tin (Sn)
	Lead (Pb)
	Copper (Cu)
Weakest Reducing Agent	Mercury (Hg)
	Silver (Ag)

Table 18.6.1.1

Since Mg is very reactive, it tends to lose electrons more and hence it is more easily **oxidised**. This means that it is a **stronger reducing agent** than a less reactive metal like Ag.

Worked Example 8, CM2131 Topic 9 Notes Modified

Consider the reaction below:

$$Zn\,(s) + CuSO_4\,(aq) \rightarrow Cu\,(s) + ZnSO_4\,(aq)$$

Explain why this reaction is possible.

Solution

Zn is a more reactive metal than Cu, and thus acted as a reducing agent, allowing this reaction to occur.

Tip: The ionic equation is: $Zn\,(s) + Cu^{2+}\,(aq) \rightarrow Cu\,(s) + Zn^{2+}\,(aq)$.

You can think of it as Zinc being more reactive and thus being more "powerful", pushing two

of its electrons to Copper.

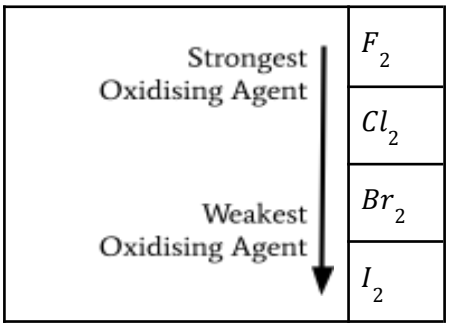


Table 18.6.1.2

Similarly, the different **halogens** have a similar behaviour. Since Fluorine is **most reactive** and most easily gains electrons, it is **reduced** most easily. Hence, it is a stronger oxidising agent than Iodine, for example.

18.7 Test for Oxidising Agents

One test we need to know for oxidising agents is aqueous potassium iodide, KI.

When using KI, the I^- ion causes the solution to appear colourless. When $2I^-$ loses electrons to form I_2 and is oxidised, the solution turns **brown**.

Hence, we can confirm that there is an **oxidising agent** if we observe the colourless solution turning brown.

18.8 Tests for Reducing Agents

Two tests we need to know for reducing agents involve aqueous potassium permanganate, $KMnO_4$ and potassium dichromate $K_2Cr_2O_7$.

When using $KMnO_4$, the MnO_4^- ion causes the solution to appear purple. When MnO_4^- loses oxygen atoms and is reduced to Mn^{2+}, the solution appears colourless.

The half-equation (i.e. equation with electrons) for this reaction is:

$$MnO_4^- + 8H^+ + 5e \rightarrow Mn^{2+} + 4\,H_2O$$

Note: The electrons $(5e)$ are there to balance the charges on both sides of the half-equation.

We can hence confirm that there is a **reducing agent** if we observe the purple solution turning colourless.

When using $K_2Cr_2O_7$, the $Cr_2O_7^{2-}$ ion causes the solution to appear yellow/orange. When $Cr_2O_7^{2-}$ loses oxygen atoms and is reduced to $2Cr^{3+}$, the solution appears green.

The half-equation for this reaction is

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2\,Cr^{3+} + 7\,H_2O$$

Note: The electrons $(6e)$ are there to balance the charges on both sides of the half-equation.

We can confirm that there is a **reducing agent** if we observe the yellow-orange solution turning green.

18.9 Types of Redox Reactions

There are 5 types of reactions that can be classified as redox reactions:

1. Combination
2. Combustion
3. Decomposition
4. Disproportionation
5. Displacement of Metal/Hydrogen/Halogen

18.9.1 Combination

A Combination reaction is one where **two or more reactants** combine to form **one** product. For example:

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

In the above reaction, Na starts with an oxidation number 0 and gets **oxidised** to oxidation number +1 as Na^+ is formed.

On the other hand Cl starts with an oxidation number 0 and gets **reduced** to oxidation number -1 as Cl^- is formed.

Since both **oxidation and reduction** occurred, **combination reactions** are **Redox Reactions**.

18.9.2 Combustion

A Combustion reaction is one where **a reactant** reacts with **oxygen gas**.

$$2\ Zn\ (s)\ +\ O_2\ (g) \rightarrow 2\ ZnO\ (s)$$

In the above reaction, O starts with an oxidation number 0, and then gets **reduced** to oxidation number -2 as O^{2-} is formed.

Zn on the other hand, starts with an oxidation number 0, and then gets **oxidised** to oxidation number +2 as Zn^{2+} is formed.

Since both **oxidation** and **reduction** occurred, **combustion reactions** are **Redox Reactions**.

18.9.3 Decomposition

A decomposition reaction is the "inverse" of a combination reaction. It is the breaking down of a single **reactant** into **multiple products**.

$$2\ HgO\ (s) \rightarrow\ 2\ Hg\ (l)\ +\ O_2\ (g)$$

In the above reaction, Hg starts as Hg^{2+}, which has an oxidation number **+2**, and is then **reduced** to oxidation number **0** as Hg is formed.

On the other hand, O starts off as O^{2-}, which has an oxidation number **-2**, and is then **oxidised** to oxidation number **0** as O_2 is formed.

Since both **oxidation** and **reduction** has occurred, **decomposition reactions** are **Redox Reactions.**

18.9.4 Disproportionation

A **disproportionation** reaction is defined as a type of **Redox Reaction** where one element is **simultaneously oxidised and reduced**.

For example:

$$2\ H_2O_2\ (l) \rightarrow 2\ H_2O\ (l)\ +\ O_2\ (g)$$

*(Hey look its BL1131 this is the reaction catalysed by **catalase** the thingy in your liver)*

In the above reaction, let's calculate the oxidation numbers of every element.

In H_2O_2, H has the oxidation number +1, while O has oxidation number -1 (remember the peroxide rule).

In H_2O, H has oxidation number +1, while O has oxidation number -2. O has been **reduced**.

In O_2, O has oxidation number 0. O has been **oxidised**.

Since O was simultaneously **oxidised** and **reduced**, this is a **disproportionation reaction** (which means it is also a **Redox Reaction**).

18.9.5 Displacement

Displacement is when an **element** in a compound is **displaced** ("forced out of its original place") by another more reactive element.

Here is the reactivity series of metals, again:

Potassium (K)
Sodium (Na)
Calcium (Ca)
Magnesium (Mg)
Aluminium (Al)
Zinc (Zn)
Iron (Fe)
Tin (Sn)
Lead (Pb)
Hydrogen (H)
Copper (Cu)
Mercury (Hg)
Silver (Ag)
Gold (Au)
Platinum (Pt)

Table 18.9.5.1

18.9.5.1 Displacement of Hydrogen

Displacement of Hydrogen occurs in many acids. In an acid-metal reaction like:

$$2\ Al\ (s)\ +\ 6\ HCl\ (aq)\ \rightarrow\ 2\ AlCl_3\ (aq)\ +\ 3\ H_2\ (g)$$

The Hydrogen in HCl is "kicked out" by the Al, and forms hydrogen gas.

Let's refer to *Table 18.9.5.1* above. Displacement of Hydrogen will only occur with elements that are **more reactive** than hydrogen (those above H in the table).

In the above reaction, Al is **oxidised** from oxidation number **0** to **+3**, while
H is **reduced** from oxidation number **+1** to **0**.

Hence, **Displacement of Hydrogen** is a **Redox Reaction**.

18.9.5.2 Displacement of Metals

A Metal can also be displaced by a **More Reactive Metal** (refer to *Table 18.9.5.1* above), like in the following reaction:

$$CuSO_4\ (aq)\ +\ Zn\ (s) \rightarrow ZnSO_4\ (aq)\ +\ Cu\ (s)$$

In the above reaction, Cu is **reduced** from oxidation number **+2** to **0**, while
Zn is **oxidised** from oxidation number **0** to **+2**.

Hence, **Displacement of Metals** are **Redox Reactions**.

18.9.5.3 Displacement of Halogen

Here is the Reactivity Series of Halogens:

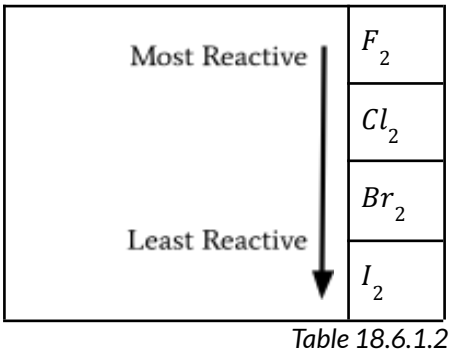


Table 18.6.1.2

A more reactive halogen can displace a less reactive one:

$$Cl_2\ (g)\ +\ 2\ KBr\ (s) \rightarrow 2\ KCl\ (s)\ +\ Br_2\ (l)$$

In the above reaction, Cl was **reduced** from oxidation number **0** to **-1**.
Br was **oxidised** from oxidation number **-1** to **0**.

18.10 Half Equations

Half Equations clearly break down a reaction further into smaller steps by making use of the symbol e to denote **electrons**.

For example, consider the following reaction:

$$4\ Na\ + O_2 \rightarrow 2\ Na_2O$$

We can break down this into two half-equations, by showing the ionisation of Na and O:

$$Na \rightarrow Na^+ + e$$
$$O_2 + 4e \rightarrow 2\ O^{2-}$$

Now, to balance this equation, we just multiply the first equation by 4 so that there are 4 electrons produced (since the second equation requires 4 electrons)

$$4\ Na \rightarrow 4\ Na^+ + 4e$$

We can now add these two equations together:

$$4\ Na + O_2 + 4e \rightarrow 4\ Na^+ + 4e + 2\ O^{2-}$$

Cancelling the electrons and combining the oppositely charged Na^+ and O^{2-} ions (to form an ionically bonded substance Na_2O):

$$4\ Na\ + O_2 \rightarrow 2\ Na_2O$$

We now have back the original equation.

Topic 19: Reactivity Series of Metals

Again… ;-;

19.1 What is the Reactivity Series of Metals?

The **Reactivity Series** (also known as **Activity Series)** ranks metals in order of their **tendency** to **lose electron(s**) to form a **positive ion**. Chemists can use this Reactivity Series to predict what happens in chemical reactions.

	Metal	Category	Physical Description
Reactivity Decreases ↓	Potassium (K)	Very Reactive Metals	Silvery
	Sodium (Na)		
	Calcium (Ca)		
	Magnesium (Mg)	Reactive Metals	
	Aluminium (Al)		
	Carbon (C)		
	Zinc (Zn)	Reactive Metals	Silvery/Grey
	Iron (Fe)		Grey
	Tin (Sn)		Silvery
	Lead (Pb)		Grey/Black
	Hydrogen (H)		
	Copper (Cu)	Unreactive Metals	Reddish-Brown
	Silver (Ag)		Silvery
	Gold (Au)		Golden Yellow
	Platinum (Pt)		Silvery

Table 19.1.1 - Reactivity Series of Metals and Physical Descriptions of Metals

When an element is **more reactive**, it is more likely to **ionise** and hence reacts quickly and often vigorously with other elements.

Of course, these "other elements" also include Oxygen. When a metal reacts with oxygen, it oxidises, which means it corrodes. Hence, a more reactive metal will **corrode more easily.**

19.2 Some Common Reactions of Metals

Here are 4 common reactions of metals:

19.2.1 Metal + Cold Water

$$Metal + Water \rightarrow Metal\ Hydroxide + Hydrogen\ Gas$$

Only the most reactive metals can react with cold water.

Reactions with cold water are **highly exothermic**, which means **a lot of heat is produced**.

$$2\ K\ (s)\ +\ 2\ H_2O\ (l) \rightarrow 2\ KOH\ (aq)\ +\ H_2\ (g)$$

Potassium has a very violent and explosive reaction with cold water. The Hydrogen Gas produced is ignited by the heat produced and burns with a lilac (purple) flame.

Figure 19.2.1.1 - Potassium reacting with Water

$$2\ Na\ (s)\ +\ 2\ H_2O\ (l) \rightarrow 2\ NaOH\ (aq)\ +\ H_2\ (g)$$

Sodium has a violent reaction with water. The Hydrogen Gas produced is ignited by the heat produced and burns with a yellow flame.

Figure 19.2.1.2 - Sodium reacting with Water

$$Ca\ (s)\ +\ 2\ H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)\ +\ H_2\ (g)$$

Calcium reacts **readily** with water.

$$Mg\ (s)\ +\ 2\ H_2O\ (l) \rightarrow Mg(OH)_2\ (aq)\ +\ H_2\ (g)$$

Very slow reaction as bubbles of hydrogen gas slowly form; Magnesium Hydroxide produced is very sparingly soluble in the water.

All other metals have **No Reaction** with cold water.

19.2.2 Metal + Steam

$$Metal\ +\ Steam\ \rightarrow\ Metal\ Oxide\ +\ Hydrogen\ Gas$$

Reactions with steam are often even more **vigorous** than reactions with cold water. This is simply because the temperature of steam is higher, and hence there is more energy in a reaction involving steam.

Potassium, Sodium, Calcium will react explosively with steam.

More metals also react with steam:

$$Mg\ (s)\ +\ H_2O\ (g) \rightarrow MgO\ (s)\ +\ H_2$$

Magnesium reacts violently with steam with a **bright white flame**.

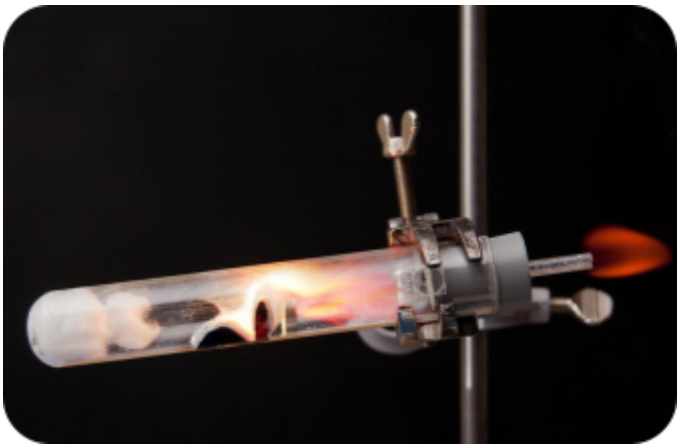


Figure 19.2.2.1 - Magnesium Reacting with Steam

$$Zn\,(s)\ +\ H_2O\,(g)\ \rightarrow\ ZnO\,(s)\ +\ H_2\,(g)$$

Zinc reacts readily with steam to form Zinc Oxide, which is **yellow** when heated and **white** when cold.

$$3\,Fe\,(s)\ +\ 4\,H_2O\,(g) \rightarrow Fe_3O_4\,(s)\ +\ 4\,H_2\,(g)$$

Iron reacts very slowly with steam, and the Iron must be **heated constantly** for the reaction to occur.

19.2.3 Metal + Acid

$$Metal\ +\ Acid\ \rightarrow\ Salt\ +\ Hydrogen\,Gas$$

Most Metals can react with dilute acid (H_2SO_4, HCl, HNO_3, etc.) to form a salt and hydrogen gas.

Metal	Description of Reaction with Dilute Acid
K	Explosive
Na	Explosive
Ca	Violent
Mg	Very Fast
Zn	Moderately Fast
Fe	Moderately Fast
Pb	Very Slow
Cu	No Reaction
Ag	No Reaction
Au	No Reaction
Pt	No Reaction

Table 19.2.3.1 - Reaction of Metals with Dilute Acid

However, recall that copper and silver can react with **concentrated nitric acid** in the following reaction, releasing **Nitrogen Dioxide**.

$$4HNO_3\,(aq)\ +\ Cu\,(s) \rightarrow 2H_2O\,(l)\ +\ 2NO_2\,(g)\ +\ Cu(NO_3)_2\,(aq)$$

$$2HNO_3\,(aq) + Ag\,(s) \rightarrow H_2O\,(l) + NO_2\,(g) + AgNO_3\,(aq)$$

On the other hand, Gold and Platinum will not react with concentrated nitric acid.

19.2.4 Metal + Oxygen

$$Metal + Oxygen \rightarrow Metal\ Oxide$$

Potassium, Sodium, Lithium, Calcium and Magnesium, when burnt in air with oxygen, will form different coloured flames and a metal oxide.

Element	Colour of Flame	Flame
Potassium (K)	Lilac	 *Figure 19.2.4.1*
Sodium (Na)	Bright Yellow	 *Figure 19.2.4.2*
Lithium (Li)	Red	 *Figure 19.2.4.3*

Calcium (Ca)	Brick Red	 *Figure 19.2.4.4*
Magnesium (Mg)	Brilliant White	 *Figure 19.2.4.5*

Table 19.2.4.6

Other metals also react with oxygen in the air in different ways.

$$4\,Al\,(s)\,+\,3\,O_2\,(g)\,\rightarrow\,2\,Al_2O_3\,(s)$$

Aluminium reacts quickly with oxygen gas to form a **protective layer** of **Aluminium Oxide**.

$$2\,Zn\,(s)\,+\,O_2\,(g)\,\rightarrow\,2\,ZnO\,(s)$$

Zinc reacts fairly quickly with oxygen to form **Zinc Oxide**.

$$4\,Fe\,(s)\,+\,3O_2\,(g)\,\rightarrow\,2Fe_2O_3\,(s)$$

Iron reacts slowly with oxygen at room temperature, but reacts quickly when heated. **Iron (III) Oxide is formed.** Do you recognise this chemical? It is the chemical formula for rust. Hence the above reaction is also known as **rusting**.

$$2\,Sn\,(s)\,+\,O_2\,(g)\,\rightarrow\,2SnO\,(s)$$
$$2\,Pb\,(s)\,+\,O_2\,(g)\,\rightarrow\,2PbO\,(s)$$
$$2\,Cu\,(s)\,+\,O_2\,(g)\,\rightarrow\,2CuO\,(s)$$

Tin, Lead and Copper all react with oxygen **only when heated**, forming Tin (II) Oxide, Lead (II) Oxide and Copper (II) Oxide respectively.

Silver, Gold and Platinum **do not react** with oxygen.

19.3 Reactivity of Aluminium

When aluminium is exposed to air, it quickly reacts with oxygen to form a **protective layer** (as seen in *Figure 19.3.1* below) of aluminium oxide through $4\ Al\ (s)\ +\ 3\ O_2\ (g)\ \rightarrow\ 2\ Al_2O_3\ (s)$.

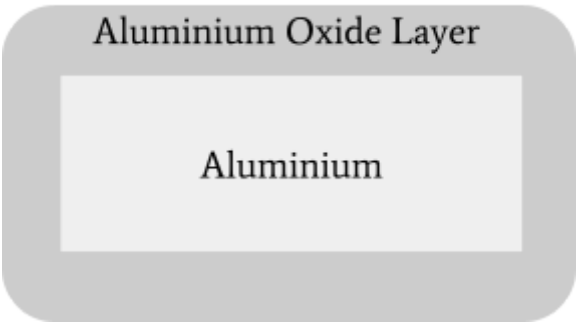


Figure 19.3.1

This protective layer means that aluminium doesn't often actually come into contact with water, or steam, which makes it seem that aluminium does not react with these substances.

Chromium (Cr) also has this property, forming Chromium (III) Oxide through:

$$4\ Cr\ (s)\ +\ 3\ O_2\ (g)\ \rightarrow\ 2\ Cr_2O_3\ (s)$$

19.4 Effect of Heat on Metal Carbonates

K_2CO_3 and Na_2CO_3 are not affected by strong heating.

$CaCO_3$, $MgCO_3$, $ZnCO_3$, $FeCO_3$, $PbCO_3$ and $CuCO_3$ decompose to form a metal oxide and carbon dioxide. For example, $CaCO_3\ (s)\ \rightarrow\ CaO\ (s)\ +\ CO_2\ (g)$.

Ag_2CO_3 decomposes as such:

$$2Ag_2CO_3\ (s) \rightarrow 2Ag_2O\ (s)\ +\ 2CO_2\ (g)$$
$$2Ag_2O\ (s) \rightarrow 4Ag\ (s)\ +\ O_2\ (g)$$

Combining the equations, we get:

$$2Ag_2CO_3\,(s) \rightarrow 4Ag\,(s) + 2CO_2\,(g) + O_2\,(g)$$

Hence, Silver Carbonate first decomposes into Silver Oxide and Carbon Dioxide, then the Silver Oxide further decomposes into Silver and Oxygen Gas.

Topic 20: Qualitative Analysis

20.1 What is Qualitative Analysis?

Qualitative Analysis (QA) will be used to **identify cations and anions** present in an **unknown chemical** in CM2131.

20.2 Colours of Ions in Solution

Ion	Colour	Image
Mn^{7+} found in MnO_4^- Manganate (VII) Ion	Deep Purple	 *Figure 20.2.1*
Ni^{2+} Nickel (II) Ion	Green	 *Figure 20.2.2*
Cr^{3+} Chromium (III) Ion	Green	 *Figure 20.2.3*

Co^{2+} Cobalt (II) Ion **Note: Remember, Co is Cobalt; Cu is copper!**	Orange	 *Figure 20.2.4*
Cu^{2+} Copper (II) Ion	Blue	 *Figure 20.2.5*
Fe^{2+} Iron (II) Ion	**Very Pale** Green (almost colourless)	 *Figure 20.2.6*
Fe^{3+} Iron (III) Ion	Yellowish-Brown / Pale Brown	 *Figure 20.2.7*

Zn^{2+}	Colourless	 *Figure 20.2.8*
Mn^{2+}	Light Pink / Colourless	 *Figure 20.2.9*

Table 20.2.10

20.3 Recall: Reactions of Cations with Hydroxide

Let's recall our solubility rules - most cations produce **precipitates** when reacted with hydroxide ions, except Na^{+}, K^{+}, $NH_4{}^{+}$.

Note: Other Group 1 Cations and Strontium, Sr^{2+}, and Barium, Ba^{2+}, in fact, also do not produce precipitates when reacted with a hydroxide, but those are outside the scope of this topic in CM2131.

Scope: Zn^{2+}, Al^{3+}, Pb^{2+}, Ca^{2+}, Mg^{2+}, $NH_4{}^{+}$, Cu^{2+}, Fe^{2+} and Fe^{3+}

20.4 Test for Identifying Cations

Firstly, we use either $NaOH$ or NH_3 and dissolve it in water to obtain an aqueous solution and dissociating to form OH^{-} ions.

$$NaOH\ (s) \rightarrow Na^{+}\ (aq)\ +\ OH^{-}\ (aq)$$
$$NH_3\ (g)\ +\ H_2O\ (l) \rightarrow NH_4{}^{+}\ (aq)\ +\ OH^{-}\ (aq)$$

We can use two main observations to determine the cation:

1. Colour of precipitate
2. Whether the precipitate is **soluble in excess** $NaOH/NH_3$ **solution**

We can also observe whether NH_3 is liberated on addition of $NaOH\ (aq)$ to identify an ammonium salt.

$$NH_4NO_3\ (s)\ +\ NaOH\ (aq) \rightarrow NaNO_3\ (s)\ +\ H_2O\ (l)\ +\ NH_3\ (g)$$

As seen from the above reaction, an ammonium salt like NH_4NO_3 will produce ammonia gas, $NH_3\ (g)$ when $NaOH\ (aq)$ is added to it.

20.4.1 Adding NaOH/NH_3 Dropwise

The first step is to add aqueous $NaOH$ or NH_3 dropwise.

For a metal M^{n+}, a precipitate forms as in the following ionic equation:

$$M^{n+}\ (aq)\ +\ nOH^-\ (aq) \rightarrow M(OH)_n\ (s)$$

Note: n simply is the charge of the Metal Ion M, these are both variables and are not chemical formulae.

We can observe the colour of precipitate formed.

Colour of Precipitate	Possible Ion(s) with $NaOH$	Possible Ion(s) with NH_3
White	Zn^{2+}, Pb^{2+}, Al^{3+}, Ca^{2+}, Mg^{2+}	Zn^{2+}, Pb^{2+}, Al^{3+}, Mg^{2+}
Light Blue	Cu^{2+}	Cu^{2+}
Green	Fe^{2+}	Fe^{2+}
Yellow-Brown	Fe^{3+}	Fe^{3+}
No Precipitate	NH_4^+	NH_4^+, Ca^{2+}

Table 20.4.1.1 - Precipitates when Adding NaOH/NH_3 Dropwise

If we can determine the cation directly, good!
Otherwise, we proceed to **Section 20.4.2: Adding NaOH/NH_3 in Excess**.

20.4.2 Adding NaOH/NH_3 in Excess

When we add $NaOH$ or NH_3 in excess, we likely can differentiate between the cations Zn^{2+}, Pb^{2+}, Al^{3+}, Mg^{2+} and Ca^{2+}, by observing if the precipitate formed originally is soluble in excess $NaOH$ and NH_3.

Precipitate of Ion	Solubility in Excess $NaOH$	Solubility in Excess NH_3
Zn^{2+}	Soluble	Soluble
Pb^{2+}	Soluble	Insoluble
Al^{3+}	Soluble	Insoluble
Mg^{2+}	Insoluble	Insoluble
Ca^{2+}	Insoluble	No Precipitate Originally

Table 20.4.2.1 - Solubility of Precipitates

However, you might notice a problem. Pb^{2+} and Al^{3+} have the same properties in both these tests, so how should we differentiate between them?

Well, we can make use of the solubility rule - "All halides are soluble except Lead (II) and Silver Halides".

Hence, if we react these compounds with HCl (remember, Cl is a halide), the following reactions occur (using the oxides PbO and Al_2O_3 as an example):

$$PbO\ (s)\ +\ 2HCl\ (aq) \rightarrow H_2O\ (l)\ +\ PbCl_2\ (s)$$
$$Al_2O_3\ (s)\ +\ 6HCl\ (aq) \rightarrow 3H_2O\ (l)\ +\ 2AlCl_3\ (aq)$$

As you can see, the lead compound forms a **precipitate** of $PbCl_2\ (s)$ while the aluminium compound forms a **soluble salt** $AlCl_3\ (aq)$**.** Hence, we can use the HCl test to differentiate between lead and aluminium compounds.

20.4.3 Format to Record Down Observations

We should record the observations in the following format:

- A [colour] precipitate is formed which dissolves in excess [$NaOH/NH_3\ (aq)$] to give a [colour] solution.
- A [colour] precipitate is formed which is insoluble in excess [$NaOH/NH_3\ (aq)$].

20.5 Tests for Identifying Anions

The Anions tested are: ${CO_3}^{2-},\ {Cl}^{-},\ {I}^{-},\ {NO_3}^{-}$ and ${SO_4}^{2-}$.

20.5.1 Test for Carbonate

Recall the reaction for an acid with a carbonate.

$$Acid\ +\ Carbonate \rightarrow Salt\ +\ CO_2 + H_2O$$

Hence, to test for a carbonate, we just add some dilute acid, such as dilute HCl to the unknown substance.

Then, if the substance is a carbonate, CO_2 gas should form. To test for this, we bubble the unknown gas into limewater. If a white precipitate forms, we can conclude that the original substance was indeed a carbonate.

20.5.2 Test for Chloride

When we react a chloride with aqueous silver nitrate, $AgNO_3\ (aq)$, the following reaction occurs (using $NaCl$ as an example):

$$AgNO_3\ (aq)\ +\ NaCl\ (aq) \rightarrow AgCl\ (s)\ +\ NaNO_3\ (aq)$$

Notice the state symbol for $AgCl$ is (s) instead of (aq). This is because $AgCl$ is **insoluble in water**. Hence, if a white precipitate (of $AgCl$) is formed, we can conclude that the original substance was a **chloride**.

But wait! Let's consider this reaction:

$$2AgNO_3\ (aq)\ +\ Na_2CO_3\ (aq) \rightarrow Ag_2CO_3\ (s)\ +\ 2NaNO_3\ (aq)$$

In this example, the original substance, Na_2CO_3, was not a chloride, yet a white precipitate of $Ag_2CO_3\ (s)$ was formed, so our test identifies it as a chloride! Uh oh! In general, all carbonates will be false positives of this test *(which means it returns a positive result, even though it shouldn't)*, if we don't add another step.

So what should we do first? We should first add some dilute nitric acid, HNO_3, before adding the silver nitrate.

If we again test Na_2CO_3, we see the following reaction occur first:

$$Na_2CO_3\,(aq)\ +\ 2HNO_3\,(aq) \rightarrow CO_2\,(g)\ +\ H_2O\,(l)\ +\ 2NaNO_3\,(aq)$$

There is then no formation of white precipitate when $AgNO_3\,(aq)$ is added.
Ah ha! We have eliminated the false positive case.

However, we should still test if the test still works in identifying chlorides.

$NaCl\,(aq)$ will not react with $HNO_3\,(aq)$, which is just a spectator molecule, but the original reaction $AgNO_3\,(aq)\ +\ NaCl\,(aq) \rightarrow AgCl\,(s)\ +\ NaNO_3\,(aq)$ still occurs, forming the white precipitate.

Hence, the test for chloride still works.

In summary, we should **add dilute** $HNO_3\,(aq)$**, then add** $AgNO_3\,(aq)$**, and if the original substance was a chloride, a white precipitate of** $AgCl\,(s)$ **should form.**

20.5.3 Test for Iodide

Just like the Test for Chloride, we add $AgNO_3\,(aq)$ to the substance. If a **yellow precipitate** of $AgI\,(s)$ forms, we know that the original substance was a iodide (using NaI as an example):

$$NaI\,(aq)\ +\ AgNO_3\,(aq) \rightarrow AgI\,(s)\ +\ NaNO_3\,(aq)$$

However, just like the Test for Chloride, the carbonates still act as a false positive, as a precipitate forms.

$$Na_2CO_3\,(aq)\ +\ 2AgNO_3\,(aq) \rightarrow Ag_2CO_3\,(s)\ +\ 2NaNO_3\,(aq)$$

Hence, the proper procedure would be to add **dilute nitric acid, then silver nitrate and finally observe for a precipitate**.

20.5.4 Test for Nitrate

The test for nitrate involves adding $NaOH\ (aq)$ solution, then adding a piece of **aluminium foil**. After warming the solution, we can test the gas given off to see if it is **ammonia gas**.

If the original substance was a nitrate, a damp red litmus paper can be observed turning blue, which is the ammonia test.

Extension 20.5.4.1 - Why is Ammonia Gas Formed? (FYI)

These are the equations[verify] (taking $NaNO_3$ as an example):

Firstly, dissociation occurs:

$$3NaNO_3\ (aq)\ +\ 8Al\ (s)\ +\ 5NaOH\ (aq) \rightarrow 8Na^+\ (aq)\ +\ 3NO_3^-\ (aq)\ +\ 5OH^-\ (aq)\ +\ 8Al\ (s)$$

The Al acts as a reducing agent for NO_3, reducing N in NO_3^- from 5+ to 3- in NH_4^+, while it itself is oxidised from 0 to 3+ in $[Al(OH)_4]^-$:

$$8Al\ +\ 3NO_3^-\ +\ 5OH^-\ +\ 21H_2O \rightarrow 3NH_4^+\ +\ 3OH^-\ +\ 8[Al(OH)_4]^-$$

Then, $NH_4^+\ +\ OH^- \rightleftharpoons H_2O\ +\ NH_3$, so we have:

$$8Al\ +\ 3NO_3^-\ +\ 5OH^-\ +\ 21H_2O \rightarrow 3NH_3\ +\ 3H_2O\ +\ 8[Al(OH)_4]^-$$

Cancelling $3H_2O$ on both sides:

$$8Al\ +\ 3NO_3^-\ +\ 5OH^-\ +\ 18H_2O \rightarrow 3NH_3\ +\ 8[Al(OH)_4]^-$$

Hence, the overall equation is:

$$3NaNO_3\ (aq)\ +\ 8Al\ (s)\ +\ 5NaOH\ (aq)\ +\ 18H_2O\ (l) \rightarrow 8NaAl(OH)_4\ (aq)\ +\ 3NH_3\ (g)$$

Hence, we conclude that ammonia gas, NH_3, is indeed released in this reaction, also known as Devarda's Test.

20.5.5 Test for Sulfate

The test for sulfate is done by adding **barium chloride** to the solution. For example, taking good ol' **Copper (II) Sulfate** as an example:

$$CuSO_4\ (aq)\ +\ BaCl_2\ (aq) \rightarrow CuCl_2\ (aq)\ +\ BaSO_4\ (s)$$

Just like in the above reaction, if a white precipitate (of $BaSO_4$) is formed, we can conclude the original substance was a **sulfate**.

But wait! Carbonates are back to wreck havoc! Consider the example:

$$Na_2CO_3\,(aq) + BaCl_2\,(aq) \rightarrow 2NaCl\,(aq) + BaCO_3\,(s)$$

A white precipitate of $BaCO_3$ forms. This is a false positive!

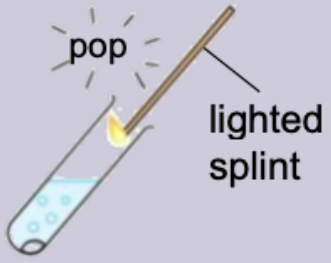
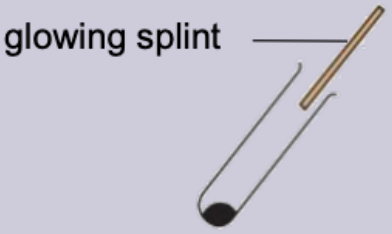
Hence, we should again add dilute acid at the start of the test. However, this time, we add dilute hydrochloric acid instead.

$$2HCl\,(aq) + Na_2CO_3\,(aq) \rightarrow 2NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)$$

In Summary, to test for sulfate, we **add hydrochloric acid, then add barium chloride. If a white precipitate forms, we conclude a sulfate is present**.

20.6 Tests for Identifying Gases

Here is a summary table on identifying gases:

Gas to be Tested	Description of Gas	Description of Test	Diagram
Hydrogen Gas, H_2	Colourless, Odourless	The lighted splint extinguishes with a 'pop' sound.	 *Figure 20.6.1*
Oxygen Gas, O_2	Colourless, Odourless	The glowing splint is rekindled (catches on fire again).	 *Figure 20.6.2*

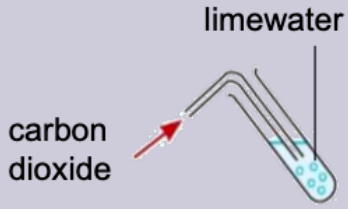
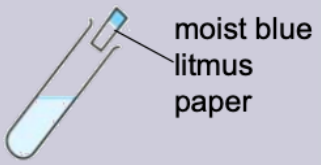
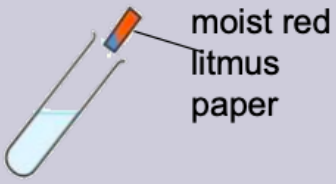
Carbon Dioxide, CO_2	Colourless, Odourless	A white precipitate forms in the limewater.	 *Figure 20.6.3*
Chlorine Gas, Cl_2	Greenish-Yellow Gas; Pungent-Smelling	The blue litmus paper turns red before becoming bleached (i.e. turns white).	 *Figure 20.6.4*
Sulfur Dioxide, SO_2	Colourless; Pungent-Smelling	The purple filter paper soaked with Potassium Manganate (VII) turns colourless.	 *Figure 20.6.5*
Ammonia Gas, NH_3	Colourless; Pungent-Smelling	The moist red litmus paper turns blue.	 *Figure 20.6.6*

Table 20.6.7

Pause and Ponder 20.6

Given that limewater is calcium hydroxide, what is the precipitate that forms when carbon dioxide is bubbled?

[Click to Go to Solution]

20.6.1 Recap: Glowing Splint vs Lighted Splint

A lighted splint refers to one that is **still on fire**, as in *Figure 20.6.1.1* below.

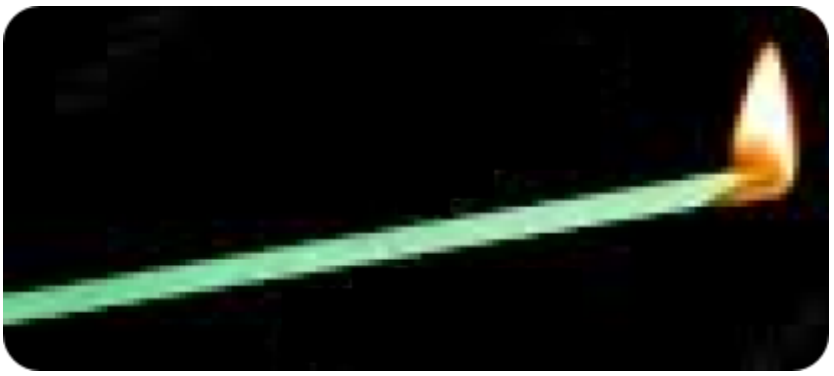


Figure 20.6.1.1 - A Lighted Splint

A glowing splint is one that was blown out (or shaken) such that it is no longer on fire visibly (or barely visibly), but a tiny ember can still be seen, as shown in *Figure 20.6.1.2* below.

Figure 20.6.1.2 - A Glowing Splint

20.7 Tests for Identifying Water

Drink it. (For legal reasons that was a joke please don't randomly start drinking chemicals during Practical Test)

There are two main ways to test for presence of water. Firstly, we can test with anhydrous Copper (II) Sulfate. The equation is:

$$\underset{\textit{(grey powder)}}{CuSO_4} + 5H_2O \rightarrow \underset{\textit{(blue crystal)}}{CuSO_4.5H_2O}$$

Hence, if we observe anhydrous Copper (II) Sulfate turning blue, we know water is present.

Secondly, we can test with anhydrous Cobalt (II) Chloride paper. If the Cobalt (II) Chloride turns pink, as shown in *Figure 20.7.1* below, we know water is present.

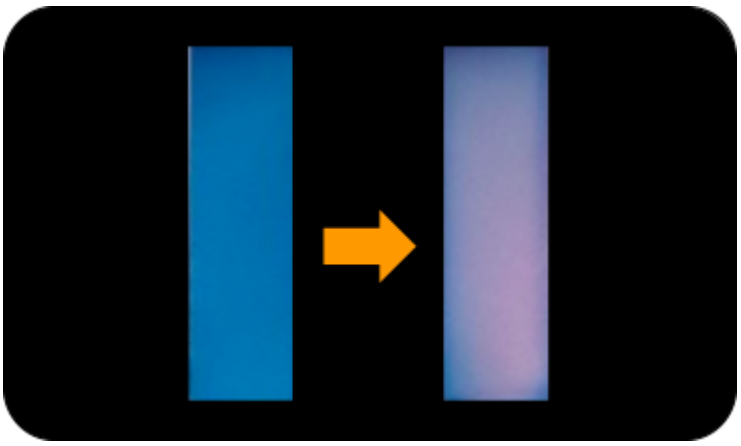


Figure 20.7.1 - Cobalt (II) Chloride Paper Changing Colour

Note: These two tests are only for testing for PRESENCE of water, and doesn't tell us anything about its purity.

Pause and Ponder 20.7

Given that hydrated Cobalt (II) Chloride contains 12 water molecules as waters of crystallisation per 2 formula units, write the equation of the reaction that turns Cobalt (II) Chloride Paper pink.

[Click to Go to Solution]

Topic 21: Speed of Reaction

21.1 Measuring Speeds of Reactions

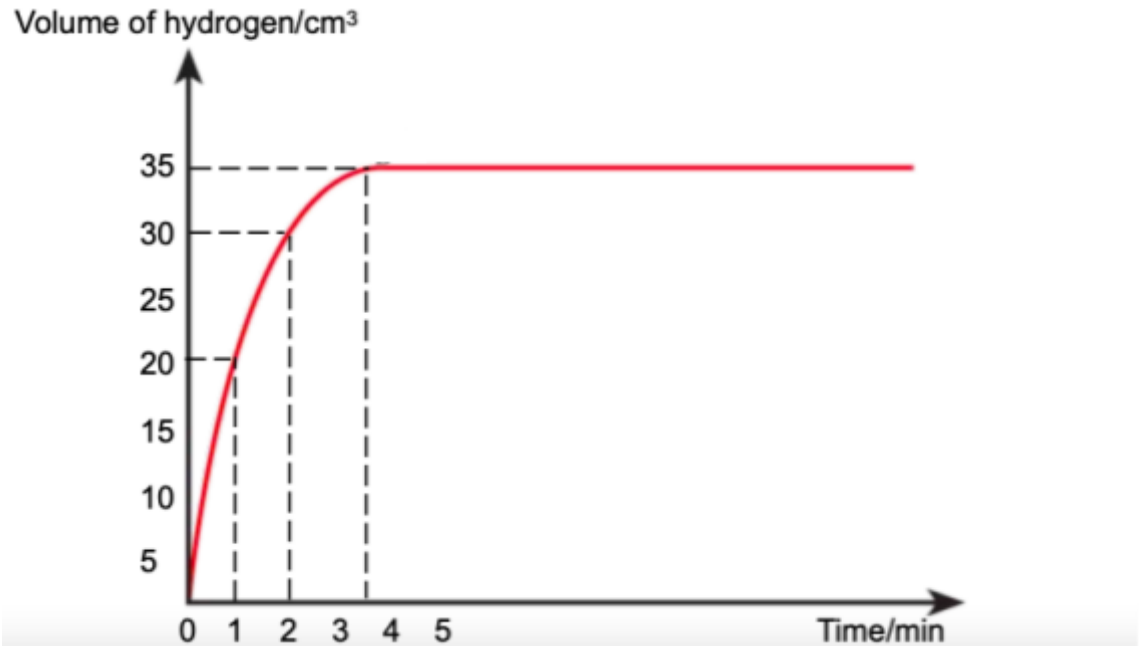
Different reactions have a **wide variety** of speeds. For example, when you drop sodium metal into water, the resulting reaction occurs **extremely quickly**. On the other hand, rusting of iron (oxidation of iron) occurs **extremely slowly**.

We can measure the speed of reaction by measuring:

1. Amount of **product formed** over a certain **period of time**; or
2. Amount of **reactants used up** over a certain **period of time**,

…then dividing by the time taken.

However, let's consider the following reaction: $Mg\ (s)\ +\ 2HCl\ (aq) \rightarrow MgCl_2\ (aq)\ +\ H_2\ (g)$.
Here is the graph of volume of hydrogen gas produced against time.



Graph 21.1.1 - Volume of H_2 Gas Produced Against Time

Let's consider the average speed of reaction over the period of 0 to 1 minutes.
$20\ cm^3$ of hydrogen gas was produced over this 1-minute period, so the average speed of reaction was $20 cm^3/min$.

Similarly, we can construct *Table 21.1.2*:

Period of Time	Speed of Reaction (cm^3/min)
0 to 1 minutes	20

1 to 2 minutes	10
2 to 3.5 minutes	3.3 $(2sf)$

Table 21.1.2 - Average Speeds of Reaction

As we can conclude, this reaction does not occur at a constant speed. This is true for **basically all reactions** in general.

21.1.1 Speed of Reaction by Gradient

Refer to *Graph 21.1.1.2* below. Let's say we wanted to calculate the speed of reaction at a specific time, say at P.

First, we draw a **tangent line** to the graph at point P.

MA2131 Reminder:

A line l is **tangent** to a curve Γ at point P if l touches Γ at **exactly one point,** P.

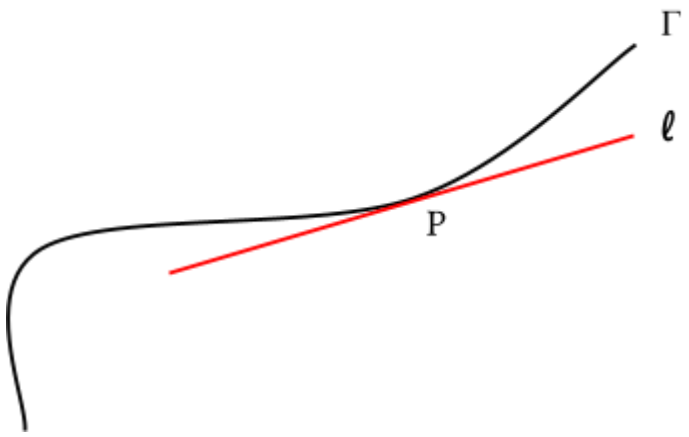
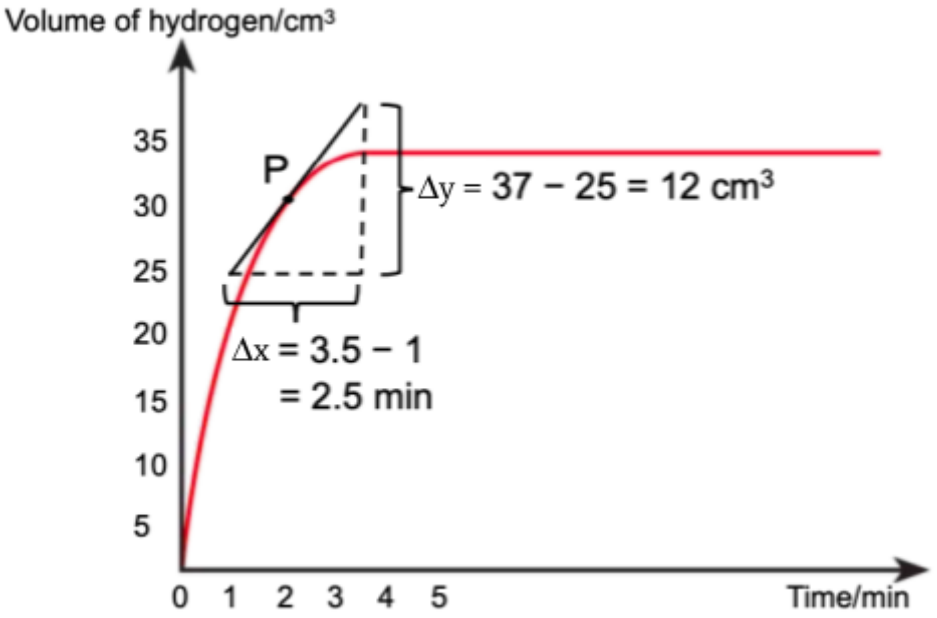


Figure 21.1.1.1



Graph 21.1.1.2 - Calculating Speed of Reaction by Gradient

Of course, the tangent line doesn't have to be perfect.

Now, we can calculate the gradient, m, of the graph at that point, using $m = \frac{\Delta y}{\Delta x}$, by calculating the gradient of the **tangent line**.

In our example, the gradient at point P is $m = \frac{\Delta y}{\Delta x}$

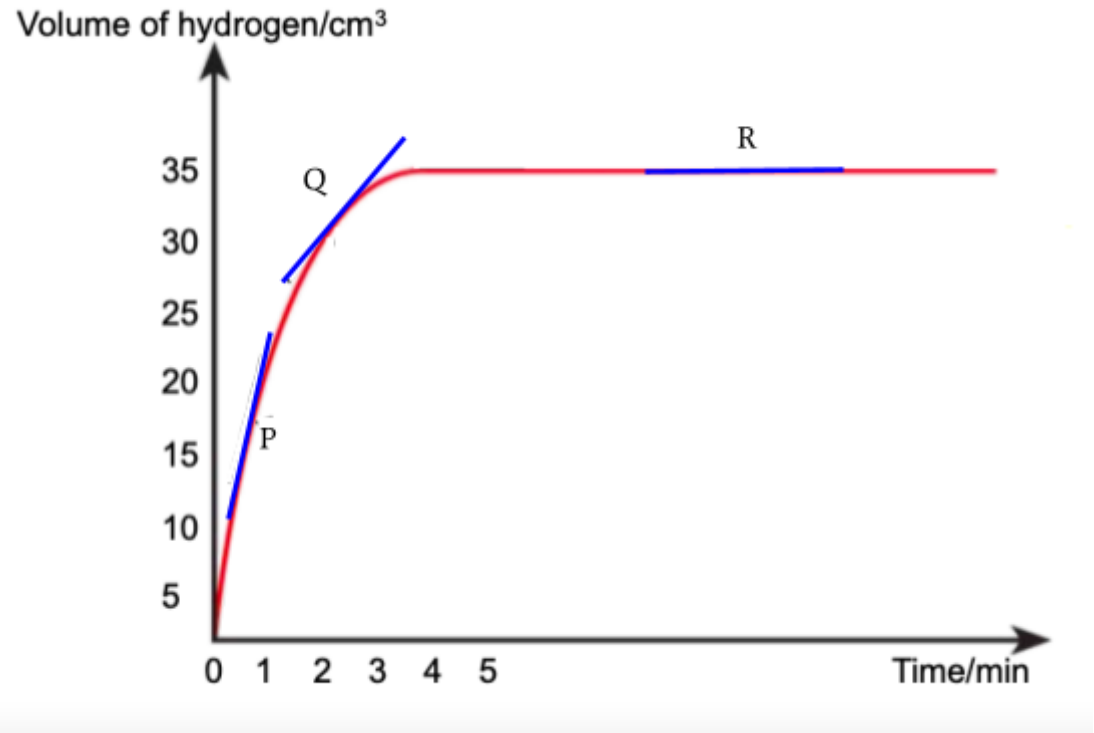
$$= \frac{12\ cm^3}{2.5\ min}$$

$$= 4.8 cm^3/min$$

Don't forget units! Now, this gradient we just calculated is the **speed of reaction at point P**.

21.1.2 Comparing Speeds of Reaction

Consider *Graph 21.1.2.1* below. How can we compare the speeds of the reaction at points P, Q and R?

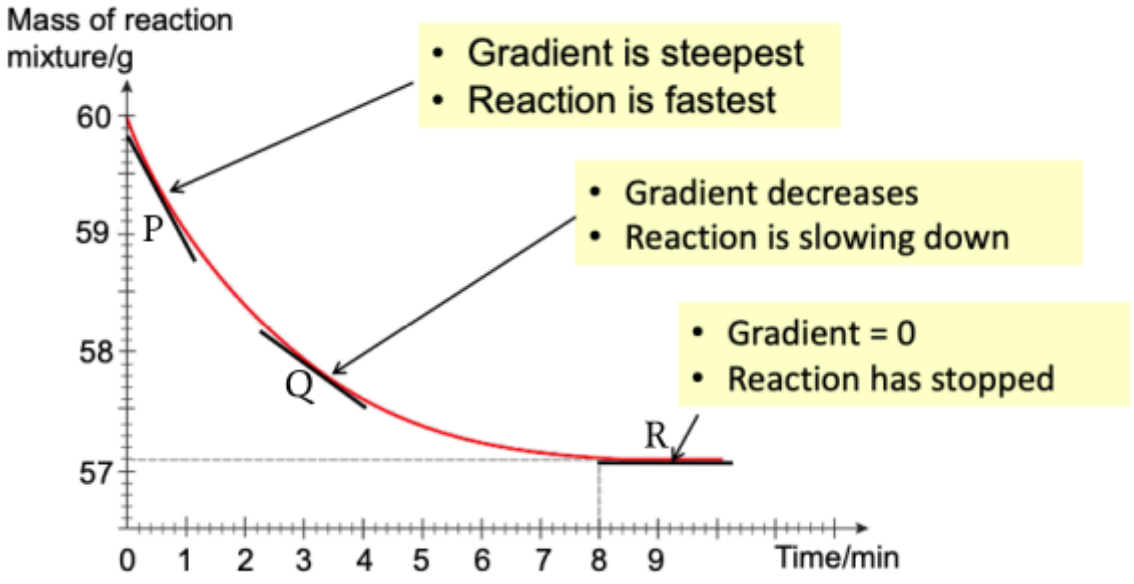


Graph 21.1.2.1

The blue lines are the **tangents** to the graph at points P, Q and R respectively. We see that the tangent at P has a **steep gradient**, which means the reaction is occuring **quickly** at that point.

The tangent at Q has a **less steep gradient**, which means the reaction is occuring **less quickly** at that point.

The tangent at R has a **gradient of 0** (i.e. it's flat and parallel to the x-axis) which means the reaction has **stopped**.



Graph 21.1.2.2

Similarly, refer to *Graph 21.1.2.2* above. It describes the **total mass of reaction mixture against time**.

At P, the gradient is again, the steepest, which means the reaction is occuring at the **highest speed**.

At Q, the gradient is less steep, which means the reaction is occuring at a **slower speed**.

At R, the gradient is **0**, which means the reaction **has stopped**.

21.2 The Collision Theory

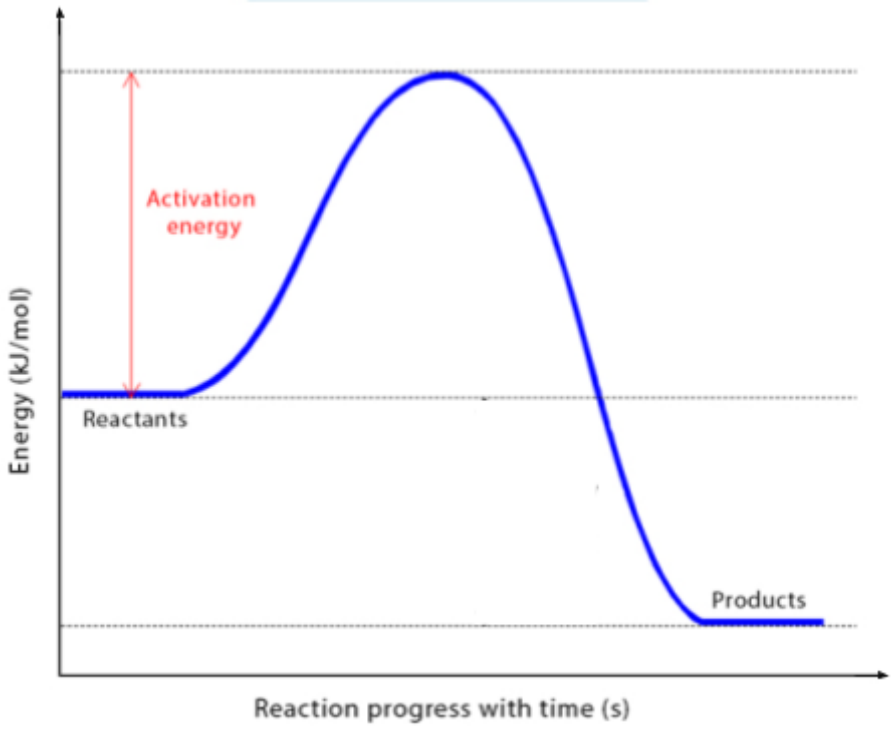
The **Collision Theory** builds on the **Kinetic Particle Theory** and says that a reaction occurs as a result of **collisions of particles** colliding with **sufficient energy** and at the **right orientation**.

21.2.1 Activation Energy

BL1131 Enzymes :0

The Activation Energy of a reaction is defined as the **minimum amount of energy required to start a reaction**, and is sometimes denoted $\mathrm{E_A}$.

Refer to *Graph 21.2.1.1* below.



Graph 21.2.1.1

You can think of this as trying to push a watermelon up a hill. If you push it by less than the activation energy, it will just slide back down.

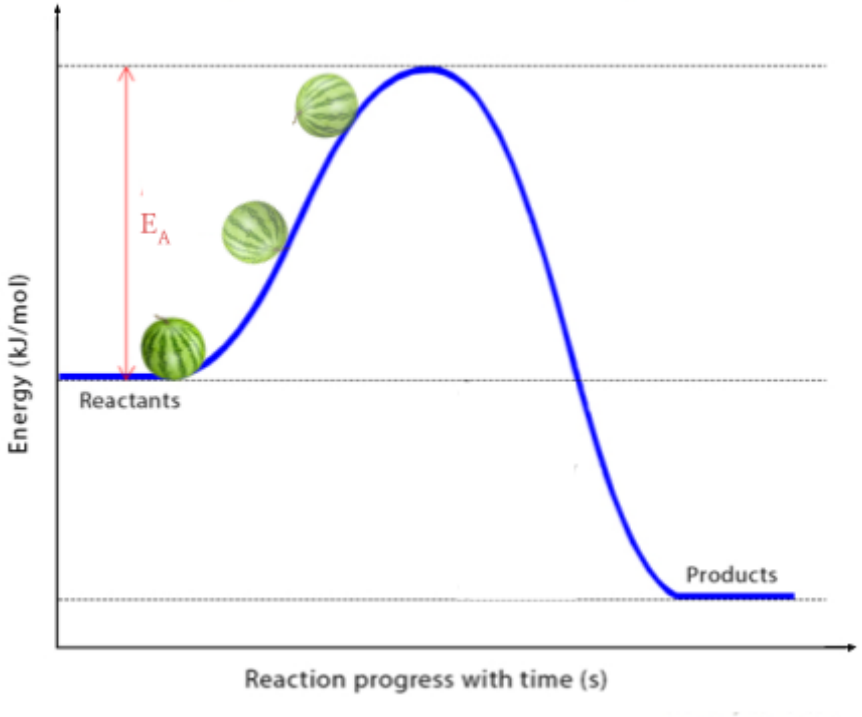


Figure 21.2.1.2 - Watermelon Falls Back (Oh No!)

However, if we give it a push with more energy than the activation energy, watch what happens:

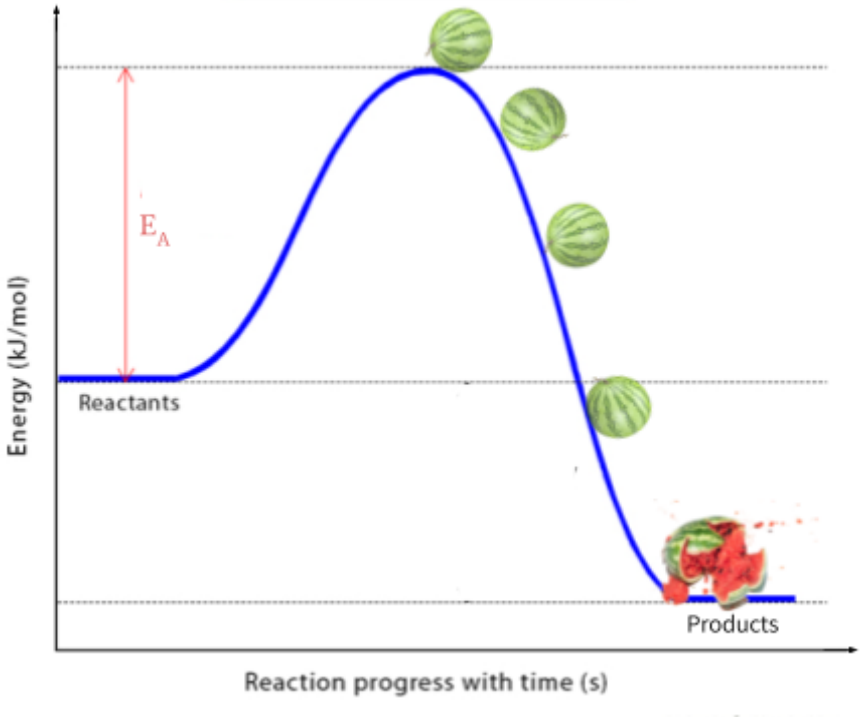


Figure 21.2.1.3 - NOOOOO WATERMELON DIED

Aww man… Who's gonna clean the watermelon up…

Watermelons aside, as we can see from *Figure 21.2.1.3* above, only when we give the reaction enough energy (i.e. greater than the activation energy), can the reaction **form products** *(in the watermelon example, the products are the leftover chunks of watermelon).*

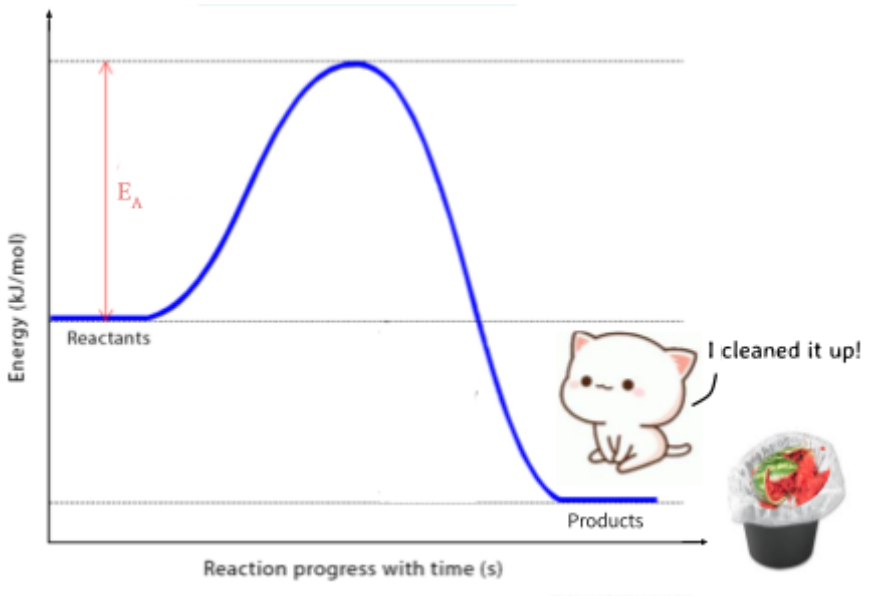


Figure 21.2.1.4 - Yay!

Yay the watermelon was cleaned up! :D Aaaanyway…

21.2.2 Orientation of Particles

Let's consider the reaction $NO\ (g)\ +\ O_3\ (g)\ \rightarrow\ NO_2\ (g)\ +\ O_2\ (g)$.

Figure 21.2.2.1 - NO and O_3

Here they come hurling at each other!

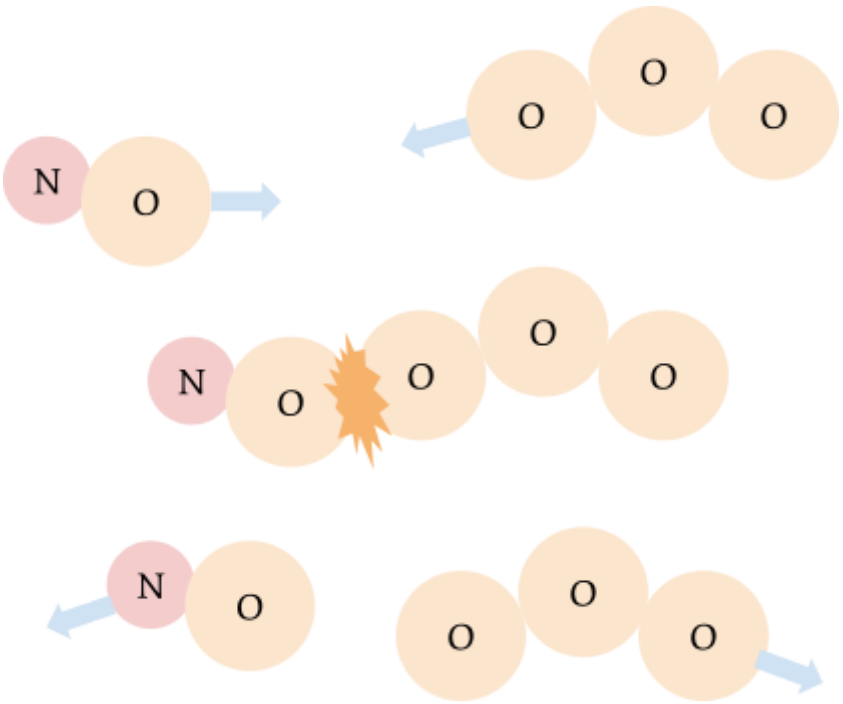


Figure 21.2.2.2 - A Failed Reaction

Let's say we did, in fact, give the reaction enough energy, higher than the required Activation Energy. But why did the reaction still fail?

This is because the ozone and the nitrogen dioxide particles collided in the wrong orientations. Let's try again, but with a different orientation.

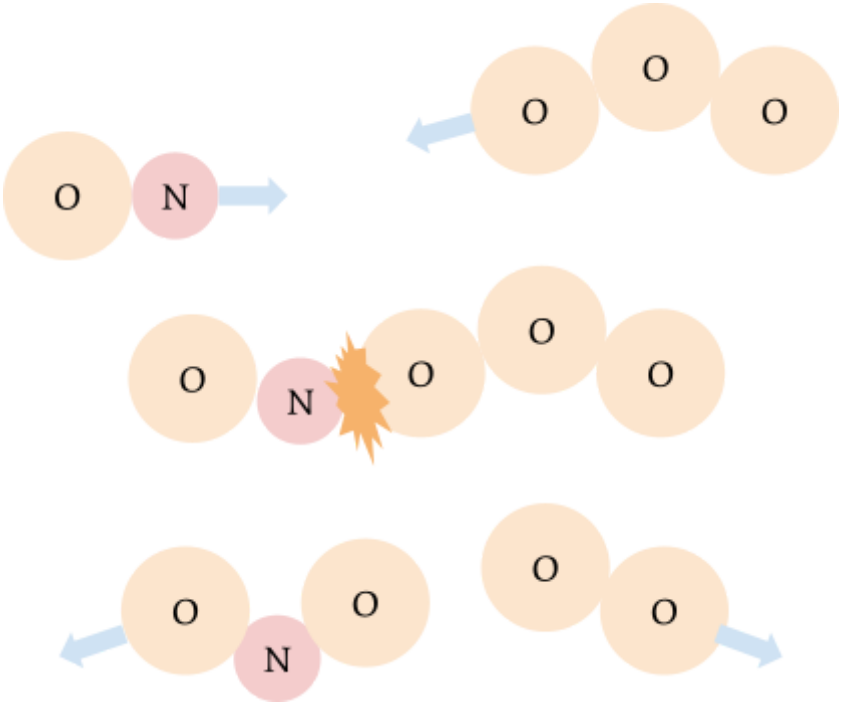


Figure 21.2.2.3 - A Successful Reaction

Oh hey! This time, the reaction actually succeeded in forming oxygen gas and nitrogen dioxide!

21.3 Factors Affecting Rate of Collisions

There are many factors affecting rate of collisions, namely:

1. Concentration of Reactants
2. Pressure (for Gases)
3. Size of Substance
4. Temperature
5. Presence/Absence of Catalyst

21.3.1 Concentration of Reactants and Pressure

Consider the dilute and concentrated solutions below in *Figure 21.3.1.1*:

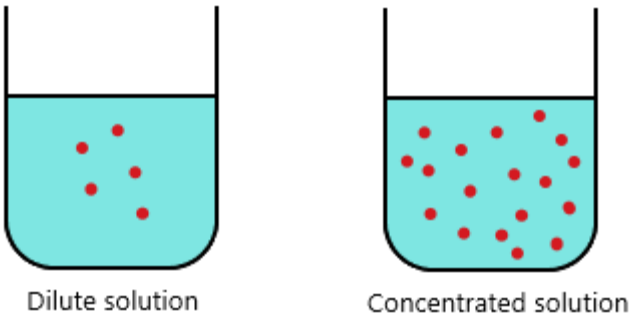


Figure 21.3.1.1 - Dilute and Concentrated Solutions

As we can observe, the number of particles occupying a certain volume of solution will **increase.** This makes **collisions between particles more frequent**, and hence **effective** (i.e. successful collisions that produce products) **collisions are more frequent,** hence **increasing speed of reaction**.

21.3.2 Gas Pressure

Refer to *Figure 21.3.2.1* below.

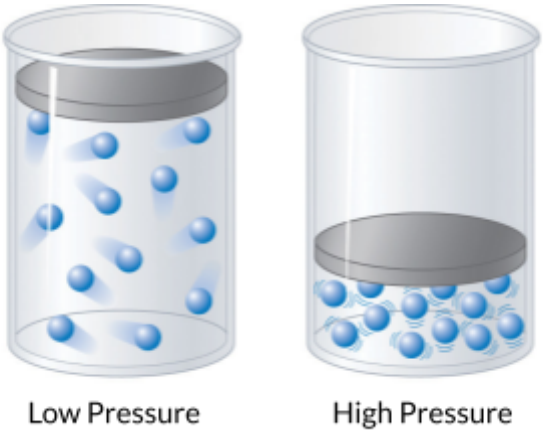


Figure 21.3.2.1 - Low Pressure Gas and High Pressure Gas

As the particles are **closer together** in a **gas at higher pressure**, **collisions** between particles are **more frequent**, hence **effective collisions** are **more frequent**. This induces a **higher rate of reaction**.

21.3.3 Size of Substance

Note: *The School Notes use the term "Particle Size", which could be misleading as the size of the particles themselves don't play as large of a part in this context. Instead, the "Particle Size" refers to the size of the chunks that the solids are broken into (e.g. big chunk of crystal vs tiny crystals)*

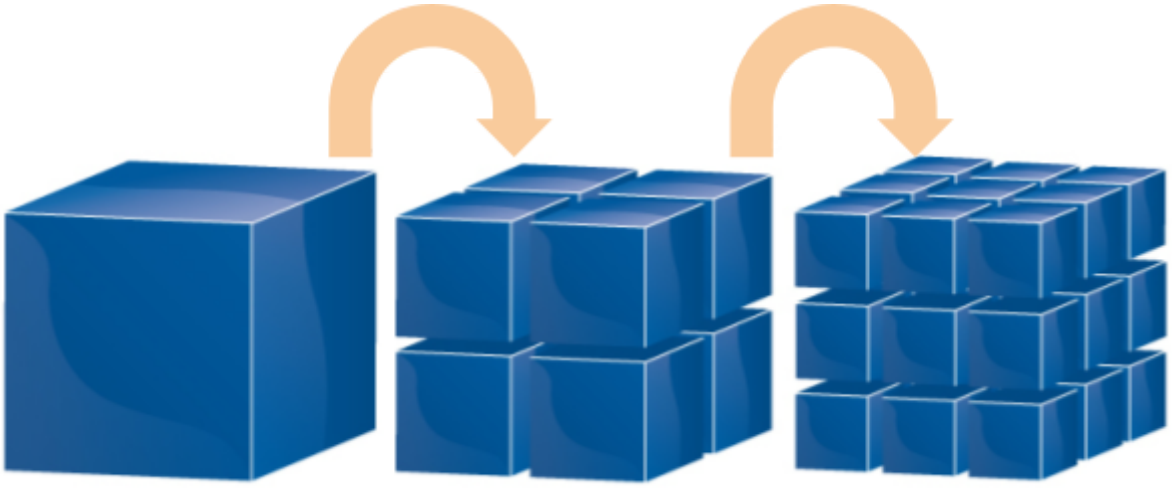


Figure 21.3.3.1 - Different Particle Sizes

Oooooh BL1131 Agar Cube Experiment (ahem i wonder who ate the agar cubes)

As we can see from *Figure 21.3.3.1* above, as we follow the orange arrows, the Size of the Substance decreases. This increases the **Surface Area (SA)** of substance.

However, this is just an analogy. On a particulate scale, this is what it roughly looks like:

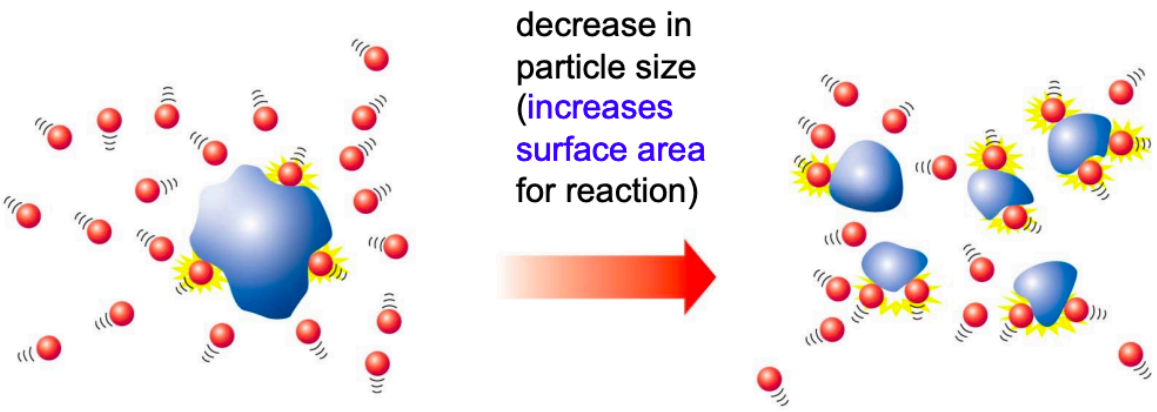


Figure 21.3.3.2 - Different Particle Sizes on a Particulate Scale

As the Particle Size **decreases**, the Surface Area for particles to collide into **increases**, making **collisions more frequent**, and hence making **effective collisions more frequent.** This induces an increase in **Speed of Reaction**.

21.3.4 Temperature

As temperature increases, the particles gain **kinetic energy** and move around faster. Thus, more particles have energies greater than the **activation energy** E_A.

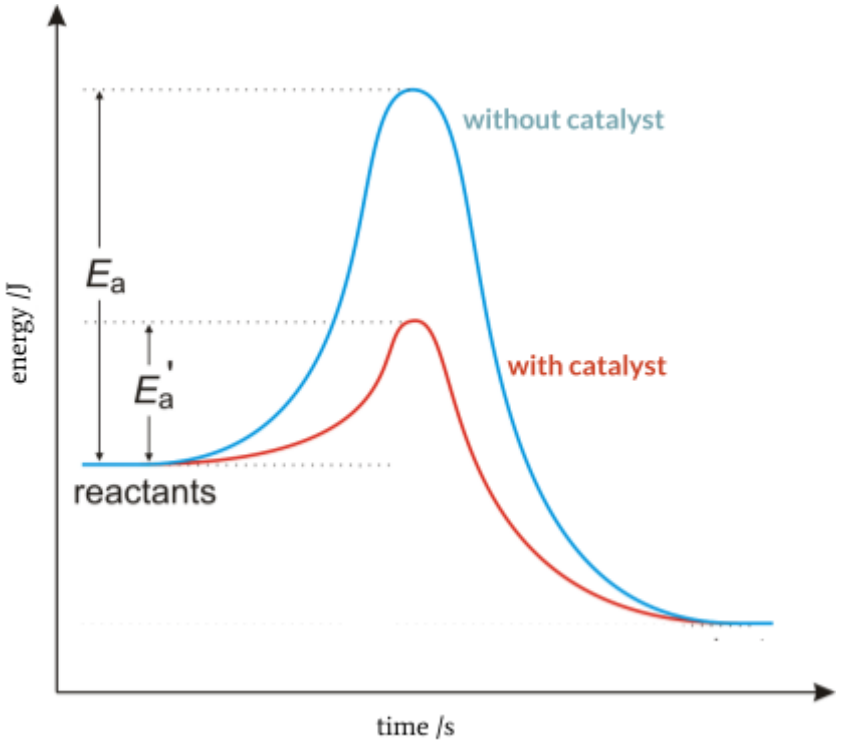
This causes **effective collisions** to be **more frequent**, and hence **a higher speed of reaction is achieved**.

21.3.5 Presence and Absence of Catalyst

A **catalyst** in chemistry is defined as a chemical that **increases the speed of a chemical reaction AND remains chemically unchanged** throughout the reaction.

Catalysts have the following properties:

1. **Lowers** the **Activation Energy** of a reaction;
2. **Chemical Properties do not change** during the reaction;
3. **Selective** in its action - Different Catalysts speed up different reactions;
4. Impurities can **prevent** the catalyst from working



Graph 21.3.5.1 - Activation Energies with and without Catalyst

As we can see from *Graph 21.3.5.1* above, the activation energy with a catalyst, E'_A, is significantly lower than the activation energy without a catalyst, E_A.

21.4 Examples of Catalysts

Consider the decomposition of hydrogen peroxide:

$$2H_2O_2\,(aq) \rightarrow\ 2H_2O\,(l)\ +\ O_2\,(g)$$

This reaction is catalysed by Manganese (IV) Oxide.

Pause and Ponder 21.4

What is the empirical formula for Manganese (IV) Oxide?

[Click to Go to Solution]

We can add 10 cm^3 of H_2O_2 solution to two test tubes labelled A and B, and then add half a spatula of Manganese (IV) Oxide to test tube B.

After conducting the burning splint test for oxygen on the two test tubes, we observe that the glowing splint in B rekindles and **burns more brightly**, which means **more oxygen** was released in B.

This is an example of how a catalyst can increase the rate at which a reaction occurs.

21.4.1 Examples of Catalysts

Table 21.4.1.1 below shows some examples of catalysts and their uses.

Catalyst	Use of Catalyst
Iron	Manufacturing of Ammonia in the **Haber Process**
Platinum, Rhodium	Catalytic Converters, which convert the toxic exhaust gases from cars into less toxic pollutants
Aluminium Oxide, Silicon Dioxide	Producing of Hydrogen in the **Cracking Process**

Table 21.4.1.1 - Uses of Catalysts

21.4.2 Biological Catalysts

See **BL1131 Topic 5: Biological Nutrients (II) - Enzymes**.

As covered in BL1131, **enzymes** are also catalysts for reactions. For example, in the decomposition of hydrogen peroxide mentioned above, we can also use the enzyme **catalase** instead of **Manganese (IV) Oxide**, MnO_2.

However, **catalase** is way less effective than **Manganese (IV) Oxide**, reducing the activation energy by 16% less.

This shows that different catalysts also have **different efficiencies**.

CM3131

Foundations in Chemistry III

Topic 22: Chemical Bonding

Please refer to the following sections if you require a recap:

- Topic 6: Structure and Bonding (Ionic) **[CM1131]**
- Topic 7: Structure and Bonding (Simple Molecular) **[CM1131]**
- Topic 16: Structure and Bonding (Metallic) **[CM2131]**
- Topic 17: Structure and Bonding (Giant Covalent) **[CM2131]**

22.1 Formation of Chemical Bond

In a Chemical Reaction between 2 atoms, their valence electrons are **reorganised** so a **net attractive force,** the chemical bond, is formed **between the reacting atoms**.

Energy is **required** to **break chemical bonds**, whereas **energy** is **released** when a **chemical bond is formed**.

Generally, chemical bonds can be classified into 3 types: Covalent, Ionic and Metallic.

22.2 Ionic Bond and Lattice Energy

The ionic bond refers to the **electrostatic force of attraction** between oppositely charged ions.

Consider the following:

$$Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$$

To measure the strength of the ionic bond, we use its **lattice energy**, which is defined as **the energy released from the formation of one mole of solid crystalline ionic compound from its gaseous ions**.

Lattice energy is denoted ΔH_{LE}, and is directly proportional to the product of the charges of the ions divided by the internuclear distance.

$$\Delta H_{LE} \propto \frac{q_1 q_2}{r}$$

Where q_1 and q_2 denote the charges on the ions, and r is the internuclear distance.

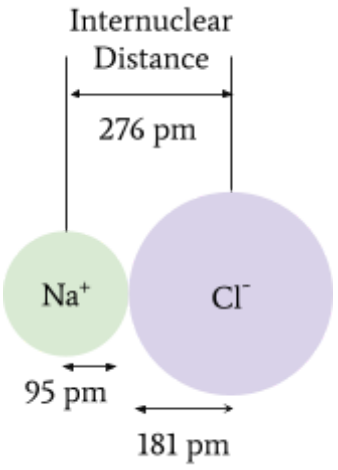


Figure 22.2.1 - Internuclear Distance

As seen from *Figure 22.2.1* above, the internuclear distance is simply the sum of the two ionic radii.

Pause and Ponder 22.2

Given that the ΔH_{LE} value of sodium chloride is $-786\ kJ/mol$, deduce if all ΔH_{LE} values of ionic substances should be negative and explain why.

[Click to Go to Solution]

22.3 Covalent Bond and Bond Energy

A covalent bond refers to the **overall electrostatic attraction** between the nuclei and the shared pair(s) of electrons, which acts as a "glue".

Similar to Ionic Bonds, we have a value describing how strong a covalent bond is, called Bond Energy and denoted by $\Delta H_{B.E.}$.

The Bond Energy is defined as the **amount of energy needed to break a covalent bond in one mole of substance**. Notice that this is actually opposite to the definition we use in the ionic bond, where $\Delta H_{L.E.}$ was the amount of energy released when an ionic bond forms.

For example, in gaseous chlorine, Cl_2, there is a covalent bond holding the two Cl atoms together. When this bond is broken in $Cl_2\ (g) \rightarrow 2\ Cl\ (g)$, the amount of energy required is around $\Delta H_{B.E.} = +242 kJ/mol$.

Extension 22.3.1 - Radicals (FYI)

In $Cl_2\ (g)\ \rightarrow 2\ Cl\ (g)$, we note that the Cl atoms produced do not have a complete octet.

Hence, we can also write $Cl_2\ (g)\ \rightarrow 2\ Cl\cdot\ (g)$, where the $\cdot$ denotes a free electron. Here, we say $Cl\cdot$ is a radical, where it has an incomplete octet, and is highly unstable.

Hence, this reaction is actually very highly reversible. This is useful in organic synthesis.

Bond Length is the **distance** between the nuclei of the atoms in a covalent bond. In general, we measure Bond Length in a unit called an Ångström, or Angstrom, denoted Å.

1 Å is defined to be 10^{-10} m, or equivalently, 0.1nm.

In general, if the bond length is shorter, this indicates a **stronger electric attractive force** between the atoms, which corresponds to a **stronger covalent bond** and hence **higher bond energy**.

22.3.2 Factors Affecting Bond Length

Size of Atom: As the size of the atoms decrease, the bond length decreases and the bond energy increases.

We can use halides as an example, since they are diatomic.

Factor	$Cl\ -\ Cl$ Bond	$Br\ -\ Br$ Bond	$I\ -\ I$ Bond
Bond Length / Å	2.00	2.28	2.66
Bond Energy / kJ mol^{-1}	239	193	151

FYI: F_2 is an exception to this rule, since it has a bond energy of 154 kJ/mol because it has a very high electronegativity and this causes the fluorine atoms to repel each other.

Bond Multiplicity: As the multiplicity (number) of bonds increases, the atoms are held more tightly and closer together, so the bond length decreases, and the bond energy increases.

Carbon is a good example for this, as it often makes 1, 2 and 3 bonds.

Factor	$C\ -\ C$ Single Bond	$C\ =\ C$ Double Bond	$C\ \equiv\ C$ Triple Bond
Bond Length / Å	1.54	1.34	1.20
Bond Energy / kJ mol^{-1}	348	614	839

22.4 Metallic Bond

Metals are described as a **lattice of cations surrounded by a sea of delocalised electrons**. In this model, we say a metallic bond is the **electrostatic force of attraction between the positively charged cations** and the **negatively charged sea of delocalised electrons**.

Also, recall that alloys are **mixtures** with at least one metal. (It doesn't have to be purely metallic - Steel is a mixture of **Carbon** and Iron, where carbon is a non-metal!)

22.4.1 Factors Affecting Strength of Metallic Bond

Number of Valence Electrons: The greater the number of valence electrons, the higher the nuclear charge of the metal cations. This leads to stronger electrostatic forces of attraction, that is, a stronger metallic bond.

For example, we can look at atoms in the same period. In general (of course, there are many, many exceptions), the above statement does hold true.

Ionic Radius: The smaller the cation size, the stronger the metallic bond.

For example, sodium and potassium have the same number of valence electrons (1 each), and the resulting melting point of sodium is 97.8 °C whereas the melting point of potassium is 6.5 °C. This means sodium has stronger metallic bonds than potassium.

22.4 Electronegativity and Electron Affinity

Please refer to **Section 18.3.1: Electronegativity** if you require a recap.

The **Electronegativity** is a measure of the ability of an atom to **attract electrons**. The larger the electronegativity value, the better the atom is at pulling electrons to it.

An atom with very low electronegativity is said to have be highly **electropositive**, such as group 1 metals, since they tend to lose electrons and not gain electrons.

The **Electron Affinity** is a measure of the change in energy when an electron is added to a gaseous atom.

Unlike **electronegativity**, which is an **estimated value** that **cannot** be **directly measured**, we can **experimentally determine electron affinity.**

22.5 Ionic vs Covalent - Lies and Deception

Well, not quite lies and deception. More like, simplification.

Ionic and Covalent are simply two ends of a spectrum, and most bonds show both ionic and covalent character (i.e. are a mix of both).

In many bonds, electrons are **not shared equally** (like in covalent bonds), and are **not completely transferred** (like in ionic bonds).

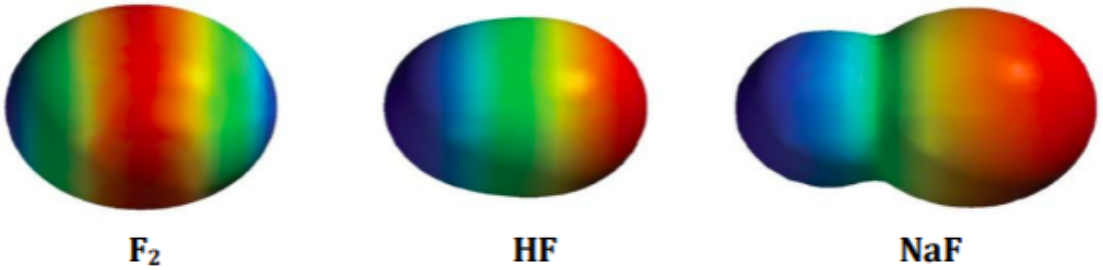


Figure 22.4.1 - Ionic vs Covalent

Consider these three bonds. In *Figure 22.4.1* above, as the colour shifts from blue to red, for simplicity, we can think of it as representing a higher electron density. (i.e. electrons are pulled more towards the red regions, and away from the blue)

FYI: In reality, we are using Schrödinger's atomic model, where electrons are described by waves describing the probability that an electron is found at a certain location, and not as particles like in Bohr's atomic model that we are used to.

In Fluorine, this bond is **non-polar covalent** (purely covalent). The electrons are **shared equally** between the two fluorine atoms, as can be seen by the symmetry.

In Hydrogen Fluoride, the bond is called **polar covalent**. The electrons are not shared equally, but are not transferred completely. Fluorine **pulls** the electrons **more strongly**, since it has a **higher electronegativity** than Hydrogen.

This can be seen as the fluorine being the red end, which means there is a **higher electron density around fluorine** than **around hydrogen**.

In Sodium Fluoride, the bond is **ionic**. The electrons are not shared but transferred completely from sodium to fluorine.

22.5 Classifications

A polar covalent bond is a bond with **more electron density** on one of the atoms than the other.

The **greater** the **difference** in **electronegativity** between the two atoms, denoted ΔEN, the more polar the covalent bond is said to be.

In general, if $\Delta EN < 0.5$, the bond is said to be **non-polar covalent**. If $0.5 < \Delta EN < 2.0$, the bond is said to be **polar covalent**. If $\Delta EN > 2.0$, the bond is said to be **ionic**.

For example, in water, hydrogen has an electronegativity of 2.2, while oxygen has an electronegativity of 3.44. $\Delta EN = 3.44 - 2.2 = 1.24$, and $0.5 < 1.24 < 2.0$, so we say that water is a **polar covalent molecule**.

We denote the regions of low and high electron density with δ+ and δ- respectively, as we can see in *Figure 22.5.1* below.

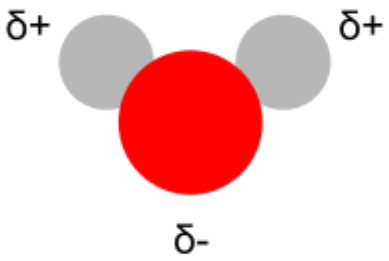


Figure 22.5.1 - Water

Topic 23: Lewis Structure

23.1 Subshell Notation

Let's recall our old electronic configuration with the Shell Model, see **5.5.2 - Electronic Configuration (Shell Model)** if you need a recap.

In this notation, Argon (atomic number 18) has electronic configuration 2.8.8, Potassium (atomic number 19) has 2.8.8.1 and Calcium (atomic number 20) has 2.8.8.2.

Everything seems fine up to the Calcium, where we keep adding **one electron to the outermost shell**. But what happens beyond Calcium? Scandium (atomic number 21), has the electronic configuration 2.8.9.2. What???

The pattern breaks! Why not 2.8.8.3? To answer this, we need to use the idea of sub-shells.

The *n*-th shell is divided into subshells, named s, p, d and f, with *n* attached before it. This number before is the **principal quantum number**. Sometimes, not all subshells are present.

In the first electron shell, we have the 1s subshell.
In the second electron shell, we have the 2s and 2p subshells.
In the third electron shell, we have the 3s, 3p and 3d subshells.
In the fourth electron shell, we have the 4s, 4p, 4d and 4f subshells.

Each s shell can contain 2 electrons, each p shell can contain 6 electrons, each d shell can contain 10 electrons and each f shell can contain 14 electrons.

Extension 23.1.1 - Why 2, 6, 10, 14? (FYI)

This is covered in **CM2231 Chemistry Olympiad Training I.** Electrons are actually found in orbitals, which describe where the electrons will move within a subshell.

Each specific electron has 4 numbers assigned to it, called the **quantum numbers**. These are:

- **n**, the principal quantum number;
- **ℓ**, the angular momentum quantum number;
- **m**, the magnetic quantum number, and;
- **s**, the spin quantum number.

n describes the shell number that the electron is in. For example, an electron in the 3s subshell will have principal quantum number 3.

ℓ describes the shape of the orbital (think of it as 0 meaning s, 1 meaning p, and so on) that forms, and can take values from 0 to n - 1. For example, in the 3rd shell, ℓ can take the values 0, 1 or 2, corresponding to the 3s, 3d and 3p subshells.

m describes the orientation of the orbital, and can take values from -ℓ to ℓ.

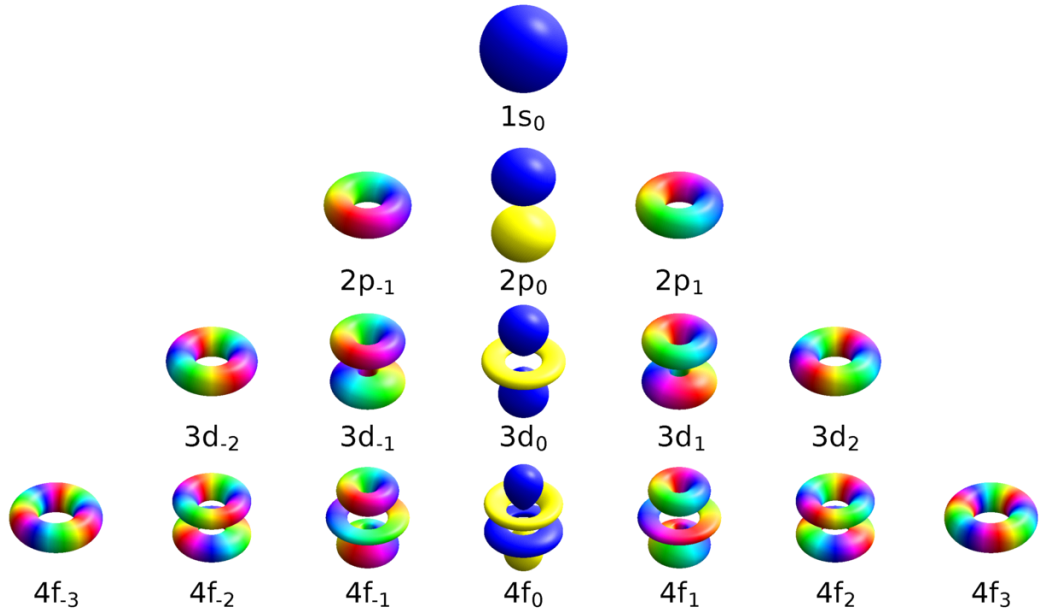


Figure 23.1.1 - Orbitals

Figure 23.1.1 shows the different orbitals up to the 4th shell. Here, the first number is the principal quantum number n, the letter corresponds to the angular momentum quantum number ℓ and the subscript number refers to the magnetic quantum number m.

In each orbital, we can have 2 electrons. This is because of the spin of electrons.

s describes the spin of the electron, and can either take the value +½ (up) or -½ (down).

Now, the **Pauli Exclusion Principle** states that no two electrons can have the same quantum principal numbers.

As such, if we follow these rules, we see that in the ns subshell, where n is the principal quantum number, s subshell implies ℓ = 0, which forces m = 0, and there are only 2 possibilities, (n, 0, 0, ½) or (n, 0, 0, -½). As such, an s subshell has 2 electrons.

Similarly, in np subshells, p implies ℓ = 1. This means m = -1, 0 or 1. This corresponds to 6 possibilities, (n, 1, -1, ½), (n, 1, -1, -½), (n, 1, 0, ½), (n, 1, 0, -½), (n, 1, 1, ½), (n, 1, 1, -½). Hence, p subshells have 6 electrons.

We repeat similar arguments for the p and f orbitals.

23.1.2 Filling the Subshells - Periodic Table Method

There are two ways to think of and remember how you fill the subshells. One way is to remember which positions on the periodic table corresponds to what subshell, as shown below.

We split the periodic table into 4 blocks, the s-block, in blue, the p-block, in red, the d-block, in orange and the f-block in peach.

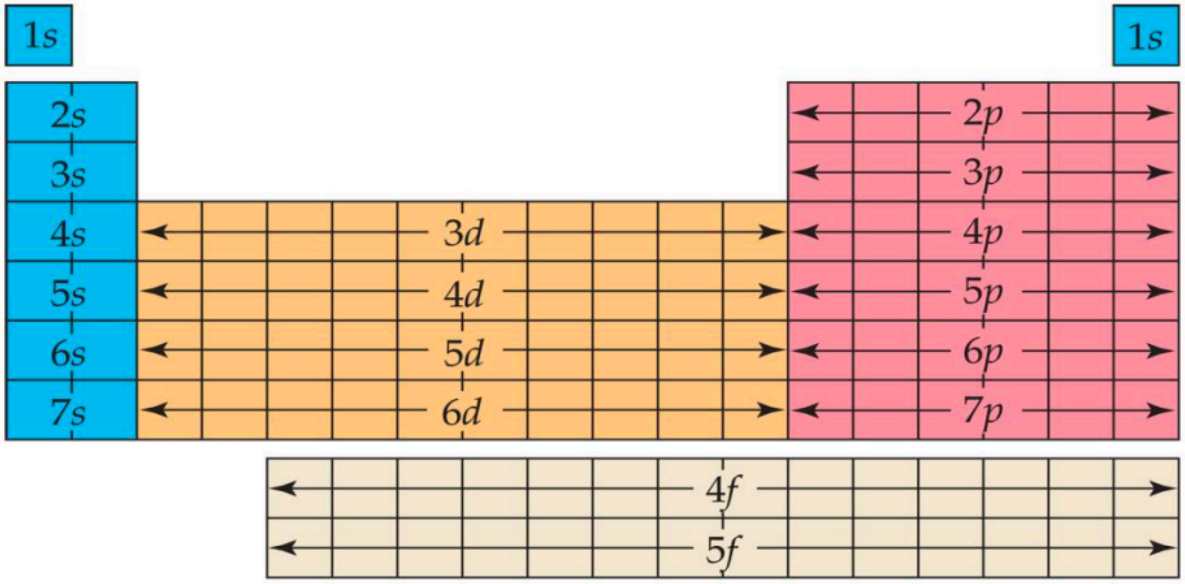


Figure 23.1.2 - spdf Subshells

Remember that the d blocks are always 1 number behind the s and p, while the f blocks are always 2 behind. That is, 3d is with 4s and 4p, and 4f is with 6s, 5d and 6p.

As in *Figure 23.1.2*, as we go down the periodic table by increasing atomic number, we will add 1 electron of the specific subshell to our atom. (Yes, with many exceptions!)

We will write each subshell with the number of electrons in that subshell in superscript. For example, if there are 4 electrons in the 3d subshell, we would write $3d^4$.

By writing out all the subshells containing electrons, we obtain the **electronic configuration** of an atom. For example, Calcium has electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$.

23.1.3 Filling the Subshells - Aufbau Principle Method

On the other hand, we can also think of the order of filling subshells by remembering the **Aufbau Principle**.

In simple terms, we draw the diagram in *Figure 21.1.3*, and fill the shells in order drawn by the arrows. That is, 1s, 2s, 2p, 3s, 3p, 4s then 3d, and so on.

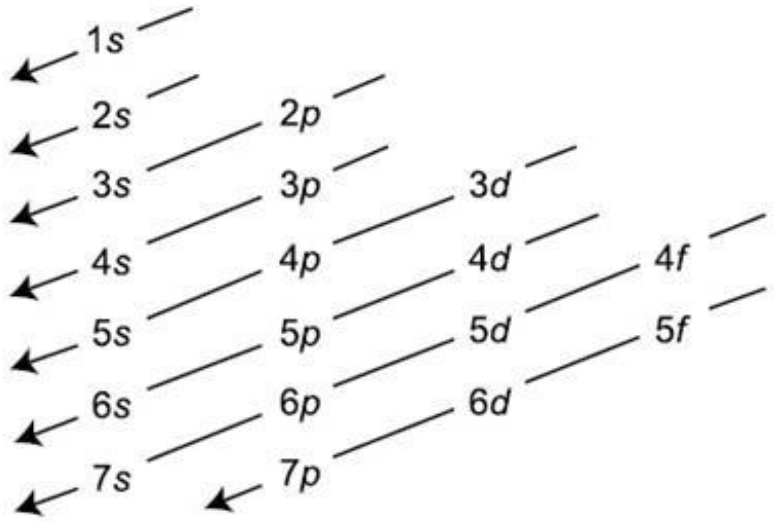


Figure 23.1.3 - Aufbau Principle

Here, we should remember that s, p, d and f subshells have 2, 6, 10, and 14 electrons, and we will reach the same electron configuration as the previous method.

However, in either case, for larger atoms, this can get quite long.

23.1.4 Condensed Electron Configuration

Here, we use the closest noble gas that is before an atom. For example, if we start with calcium again, we see that the closest noble gas before it is Argon.

Now, we start from argon, and write out the electrons after Ar, which will be $4s^2$.

Hence, the condensed electron configuration of Calcium is [Ar] $4s^2$.

23.2 Drawing a Lewis Structure

Behind the Lewis Structure, there are 2 guiding rules:

- **Stable Octet**: Atoms try to obtain the stable octet valence electron configuration.
- **Minimising Charge**: Atoms try to obtain a neutral charge by bonding.

Here, charge both refer to the actual charge on a particle, as well as something known as the **formal charge**, which will be covered later.

23.2.1 Formal Charge

The formal charge is the charge of any atom in a molecule if all the atoms had the same electronegativity. This means, idealising covalent bonds such that each atom is assigned one of the electrons of the covalent bond because it is shared equally.

As such, the formula for the formal charge is given by $V - N - \frac{1}{2}B$, where:

- V is the number of valence electrons,
- N is the number of non-bonding electrons, and
- B is the number of bonding electrons

Note: These abbreviations are just used for brevity and may not be appropriate for exams.

Equivalently, it is easier to think of $\frac{1}{2}B$ as simply the number of covalent bonds an atom is attached to.

For example, we consider a molecule like ammonia. Recall the dots on top of the nitrogen denote non-bonding electrons (called a **lone pair**), and the dashes denote covalent bonds. What are the formal charges around each of the atoms?

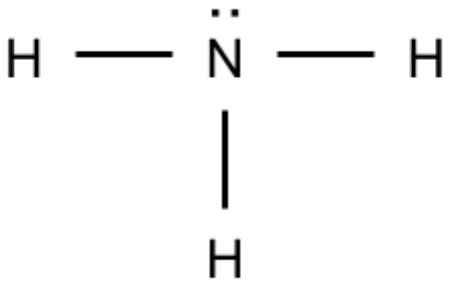


Figure 23.2.1 - Lewis Structure of Ammonia

We can find the formal charge of the hydrogens by $V - N - \frac{1}{2}B = 1 - 0 - \frac{2}{2} = 0$.
Similarly, we can find the formal charge of the nitrogen by $V - N - \frac{1}{2}B = 5 - 2 - \frac{6}{2} = 0$.

You can also think of the calculation of formal charge as "Group Number minus Dots minus Lines". Convince yourself that this is equivalent!

23.2.2 Lewis Structure Rules

The correct Lewis Structure (of a neutral molecule) is the structure in which all the atoms have formal charges closest to 0, and any negative formal charge present is on the most **electronegative** atom.

The **least electronegative** atom is usually placed in the centre.

In a polyatomic ion, the formal charges of the atoms must sum to the charge on the whole ion.

For example, consider the ammonium ion shown in *Figure 23.2.2* below.

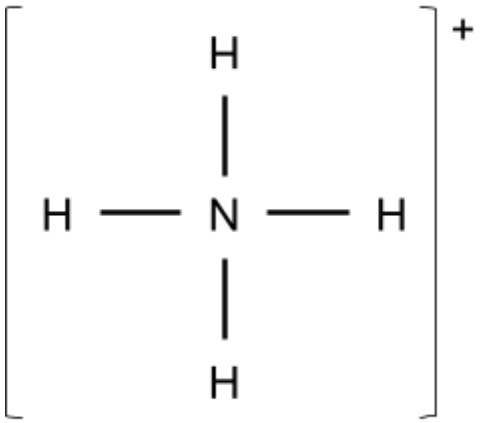


Figure 23.2.2 - Lewis Structure of Ammonium Ion

Here, we calculate the formal charge of the hydrogens to be 0, while nitrogen has +1. We can check that the sum of formal charges is then +1 + 0 + 0 + 0 + 0 = +1, which is indeed the charge of the ammonium ion.

23.3 Resonance Structures

For many molecules, there is not a single correct Lewis Structure. For example, consider the carbonate ion. You could find that this is a Lewis Structure for CO_3^{2-}:

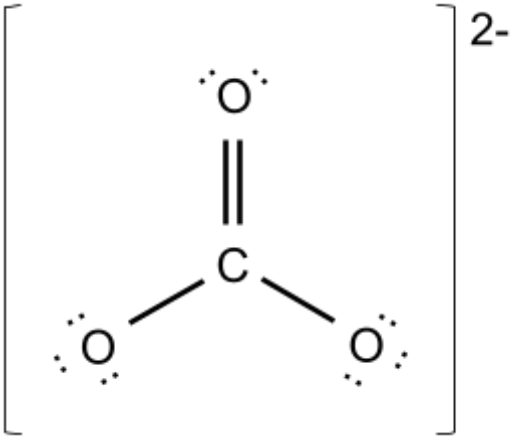


Figure 23.3.1 - A Lewis Structure of the Carbonate Ion

However, the double bond might not be in that position! If the Lewis Structure was like *Figure 23.3.1*, then we would expect one of the three bonds to be measurably shorter than the other two, because as we discussed, Bond Multiplicity affects Bond Length!

However, if we experimentally determine the bond lengths, we realise that all the bond lengths are actually about the same (around 1.28Å). This suggests that there is something that we haven't accounted for.

In fact, the carbonate instead could take one of the two structures in *Figure 23.3.2* below:

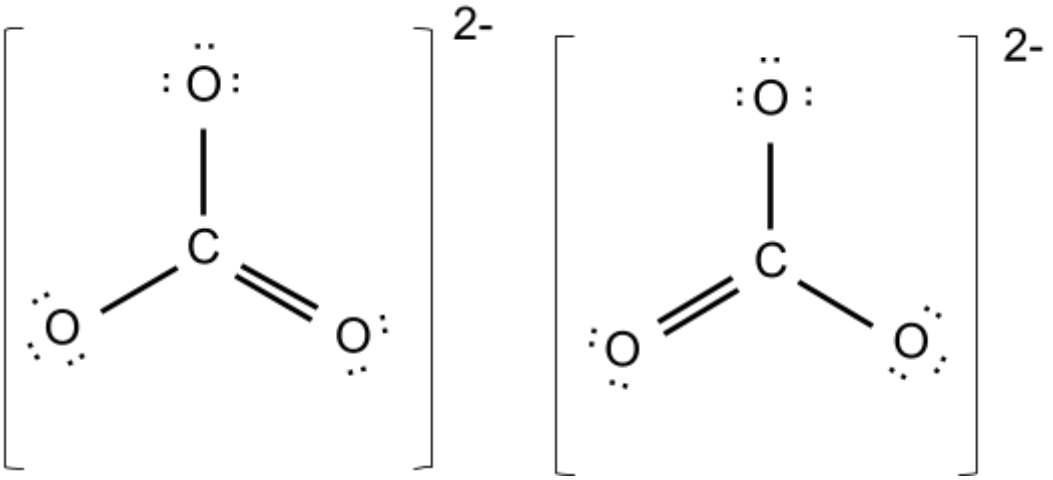


Figure 23.3.2 - Alternate Lewis Structures of Carbonate

These three structures, are all technically correct, but on their own do not fully describe the carbonate ion. As such, we say that these structures are **resonance structures**. We denote this by drawing all 3 possibilities and drawing double-headed arrows between them.

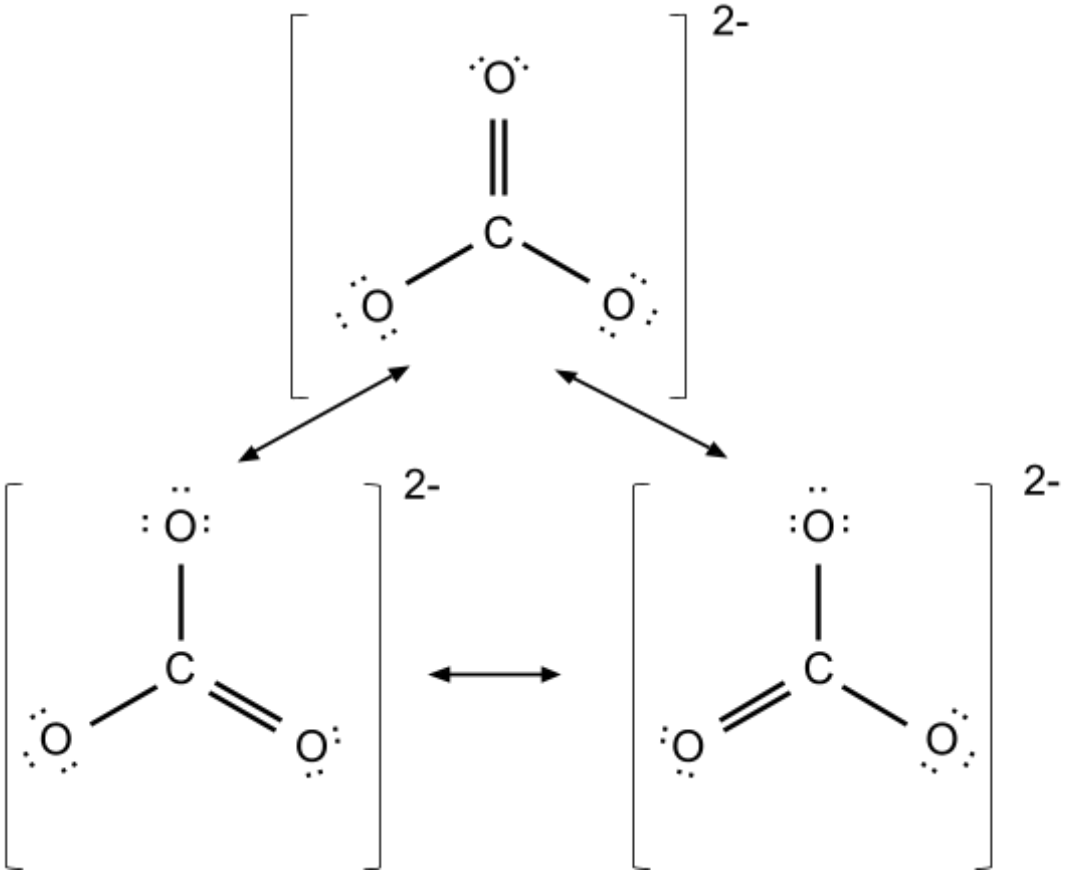


Figure 23.3.3 - Resonance Structures of Carbonate

Technically, *Figure 23.3.3* better reflects the true structure of CO_3^{2-}, which is a blend of all resonance structures. As such, all of the bond angles in carbonate is **120°**, due to symmetry.

23.4 Exceptions to the Octet Rule

Well, everything has exceptions in chemistry. Here are 3 exceptions to the Octet Rule.

23.4.1 Covalent Molecule with Odd Number of Electrons

For example, consider Nitrogen Monoxide (NO). Nitrogen has 5 valence electrons, while oxygen has 6. This makes for a total of 11 valence electrons, so unfortunately one electron will be very lonely *(like me)* and will not be paired. This electron is called an **unpaired electron**.

Figure 23.3.4 shows the lewis structure of Nitrogen Monoxide:

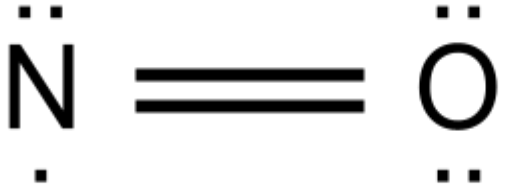


Figure 23.4.1.1 - Lewis Structure of NO

We can check that both Nitrogen and Oxygen have formal charges of 0. However, nitrogen has only 7 valence electrons, and does not have a full octet. This shows that sometimes, the Octet Rule is violated to minimise formal charge.

Pause and Ponder 23.4.1

We know that Nitrogen Monoxide, NO, does not have a complete octet around Nitrogen. Given that the Nitrosonium Ion, NO^+, has 2 resonance structures, determine these resonance structures, drawing their Lewis Structures.

[Click to Go to Solution]

23.4.2 Covalent Molecule with More Than an Octet of Valence Electrons

Elements from Period 3 onwards are able to **expand** their valence shell to accommodate more than 8 electrons.

For example, Phosphorus, which is in Period 3, is able to expand its valence shell, and hence a substance like Phosphorus Pentachloride (PCl_5), as shown in *Figure 23.4.2.1* below, can exist.

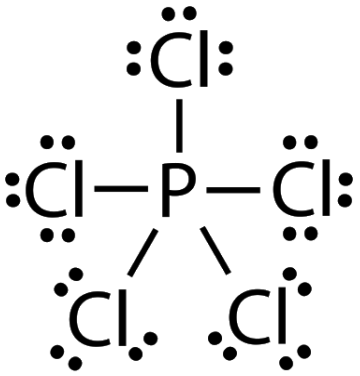


Figure 23.4.2.1 - Lewis Structure of PCl_5

If we check the valence shell of phosphorus, we see that it has 10 valence electrons. This is a result of an expanded valence shell. However, if we check for formal charge, we see that both Phosphorus and Chlorine have formal charges of 0.

Again, this is an example where the Octet Rule is violated to minimise formal charge.

Extension 23.4.2.1 - Hybridisation in PCl_5 (CM4131)

In fact, this is because from the third period onwards, the d orbital becomes available, so these atoms can make space for the extra valence electrons by filling the d orbitals early.

Phosphorus has the electronic configuration $[Ne]\ 3s^2\ 3p^3$. Let's draw its orbital box diagram, only focusing on valence electrons, as in *Figure 23.4.2.1.1* below. (Orbital Box Diagrams are another way of describing electrons, with each arrow head denoting an electron)

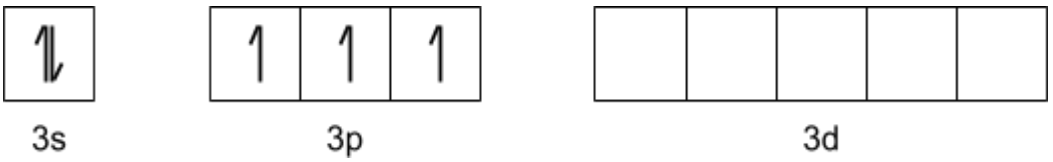


Figure 23.4.2.1.1 - Ground State of Phosphorus

This is called the **ground state** of phosphorus. When it tries to bond, it can promote one of the 3s electrons to a 3d electron, forming the **excited state** of phosphorus, sometimes denoted as P^*. *Figure 23.4.2.1.2* below is the new orbital box diagram:

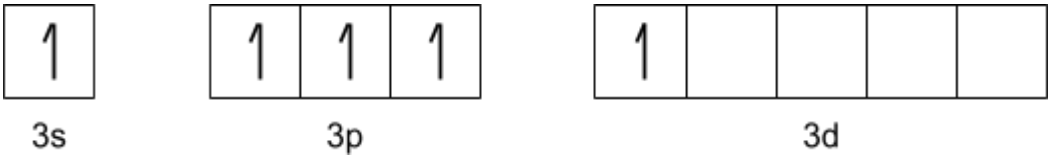


Figure 23.4.2.1.2 - Excited State of Phosphorus

Now, these orbitals **hybridise** (think of it as averaging), to form sp^3d orbitals.

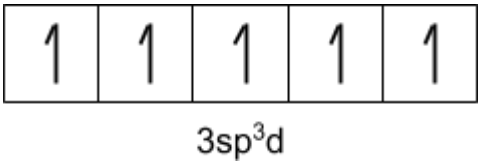


Figure 23.4.2.1.3 - Hybridised Orbitals of Phosphorus

Now that Phosphorus has 5 hybridised $3sp^3d$ orbitals, each of these orbitals can form a σ bond with the 3p orbital of a chloride atom. This allows PCl_5 to form.

In fact, this hybridisation theory in *Figure 23.4.2.1.3* above also explains the trigonal bipyramidal molecular geometry of PCl_5, due to the shape of the sp^3d orbitals.

23.4.3 Covalent Molecule with Less Than an Octet of Valence Electrons

Consider Aluminium Trichloride, $AlCl_3$. Wait. Aluminium is a metal, and chlorine is a non-metal, so shouldn't the bond be ionic? Why is it named "Aluminium Trichloride" and not follow the naming of ionic substances - something like "Aluminium Chloride"?

Aluminium has an electronegativity value of 1.61, while Chlorine has an electronegativity value of 3.16. The difference in electronegativity is 1.55, which is between 0.5 and 2.0. Hence, it has more **polar covalent character** than ionic character, and hence the covalent naming.

Let's draw the Lewis Structure!

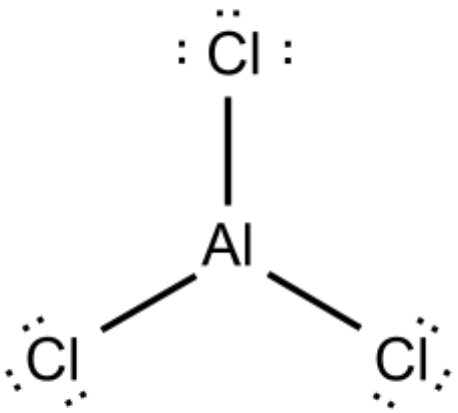


Figure 23.4.3.1 - Lewis Structure of $AlCl_3$

We can check that this is the optimal Lewis Structure, with 0 formal charge on both Aluminium and Chlorine. But wait! Aluminium only has 6 valence electrons in this case. This is another case where the Octet Rule is violated to minimise Formal Charge.

23.5 Dative Bonding

Dative Bonding, also known as **co-ordinate bonding**, is a special type of Covalent Bonding where both electrons are donated by the same atom. Equivalently, a **lone pair** is donated.

When a molecule with a lone pair like NH_3 meets one with an incomplete octet like BF_3, a dative bond may form.

Figure 23.5.1 - Dative Bonding between NH_3 and BF_3

Here, we see the Nitrogen from NH_3 donates its lone pair to form a dative bond with Boron in BF_3, because Boron only has 6 valence electrons. This dative bond is indicated by an arrow pointing from the donor (N) to the recipient (B). (CM4131: Lewis Acid-base Adducts)

Another common example is Al_2Cl_6, formed by $2AlCl_3 \rightarrow Al_2Cl_6$, under certain temperature and pressure conditions. In this case, the chlorine has 3 lone pairs, while the aluminium has an incomplete octet, so the chlorine donates 1 lone pair to the aluminium, forming a dative bond.

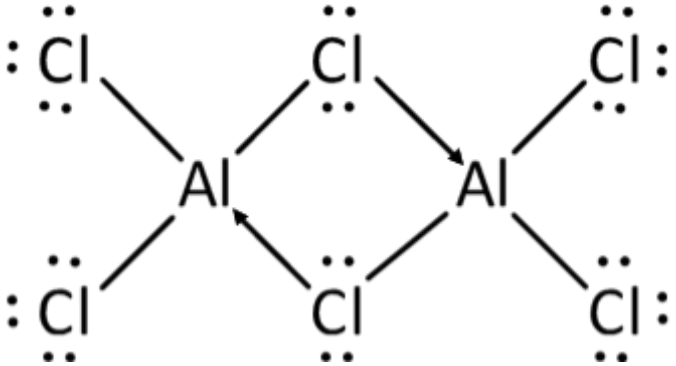


Figure 23.5.2 - Lewis Structure of Al_2Cl_6

Topic 24: VSEPR Theory

24.1 Why VSEPR? Who Cares?

The geometry of a molecule refers to the **three-dimensional arrangement of atoms** in a molecule. The Valence Shell Electron Pair Repulsion (VSEPR) Theory is a model to predict this molecular geometry of molecules and polyatomic ions.

Note: For brevity, the rest of this set of notes will refer to "Valence Shell Electron Pair Repulsion Theory" simply as "VSEPR" or as "VSEPR Theory".

But now, why do we even care about Molecular Geometry?

Well, molecular geometry has a large influence on both **physical and chemical** properties of covalent compounds, like:

- Melting and Boiling Point
- Reactivity
- Odour/Scent

For example, only specific molecules with the correct shape and size can fit properly on the receptor sites in our nose, allowing impulses to be sent to the brain and allowing us to detect the smell of things around us.

24.2 Basis of VSEPR Theory

VSEPR Theory asserts that the electron pairs (both bonding and non-bonding) in the valence shell of an atom repel each other and cause them to be **separated as far as possible**.

Hence, the geometry that the molecule will eventually take up is the one that has the **minimum repulsion between electron pairs** (i.e. the electron pairs are furthest apart).

In VSEPR Theory, **all valence shell electron pairs**, both bonding and non-bonding, are assumed to be the **same**, with the exception that electrons of a **double** or **triple bond** are treated as a **single pair of electron**s.

We also can refer to a bond (single, double or triple), a lone pair, or a single unpaired electron as simply an "**electron domain"**. The number of electron domains determine the molecular geometry in VSEPR.

24.3 Electron Domain Geometries

An electron domain geometry is the geometry of a molecule that is predicted by VSEPR Theory, based on the number of electron domains.

This **may not** reflect the **actual molecular geometry** of a molecule, because some of the electron domains could be lone pairs, which are reflected in the **Electron Domain Geometry**, but not in the **Molecular Geometry,** which only describes atoms.

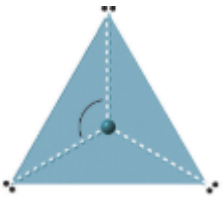
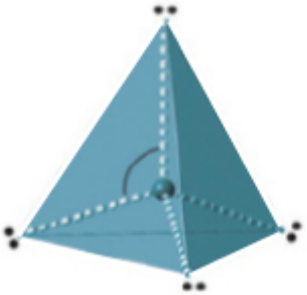
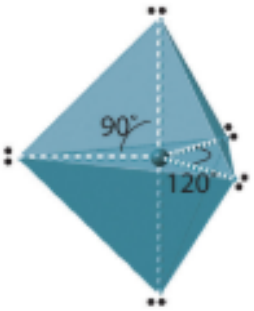
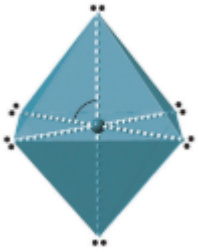
No. of Electron Domains	Electron Domain Geometry	Angles Between Domains
2	 Linear	180°
3	 Trigonal Planar	120°
4	 Tetrahedral	109.5°
5	 Trigonal Bipyramidal	90° and 120°
6	 Octagonal	90°

Table 24.3.1 - Electron Domain Geometries

24.4 Molecular Geometries

In short, we can obtain all the possible molecular geometries of an electron domain geometry by repeatedly adding lone pairs.

24.4.1 Linear

For this case, we can only have one molecular geometry - **Linear** (e.g. CO_2).

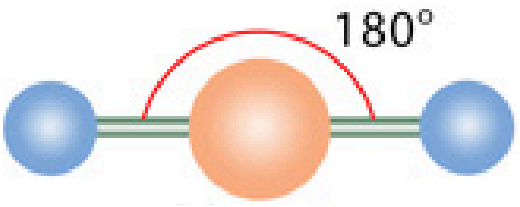


Figure 24.4.1.1 - Linear Molecular Geometry

24.4.2 Trigonal Planar

If there are no lone pairs, we will have the **Trigonal Planar** molecular geometry (e.g. BH_3):

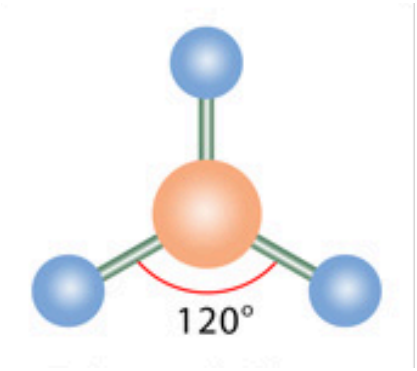


Figure 24.4.2.1 - Trigonal Planar Molecular Geometry

Now, if we add a lone pair, we have the **Bent** or **V-Shape** molecular geometry (e.g. SO_2):

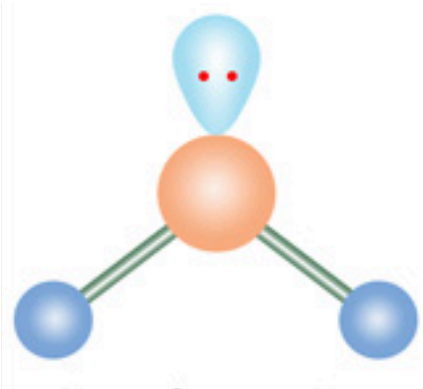


Figure 24.4.2.2 - Bent Molecular Geometry

Note: The light-blue orbital denotes a lone pair, and in reality there is no atom there.

24.4.3 Tetrahedral

If there are no lone pairs, we have the **Tetrahedral** molecular geometry (e.g. CH_4):

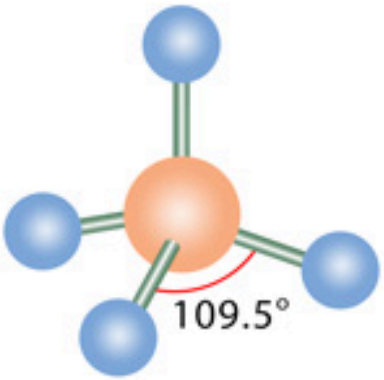


Figure 24.4.3.1 - Tetrahedral Molecular Geometry

If we add 1 lone pair, we have the **Trigonal Pyramidal** molecular geometry (e.g. NH_3):

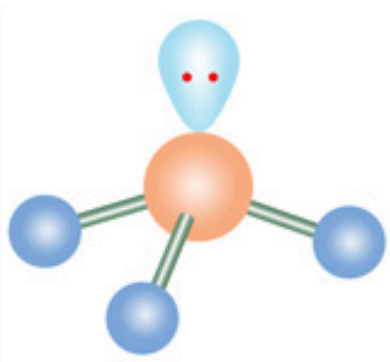


Figure 24.4.3.2 - Trigonal Pyramidal Molecular Geometry

If we have 2 lone pairs, we have the **Bent** or **V-shape** molecular geometry (e.g. H_2O):

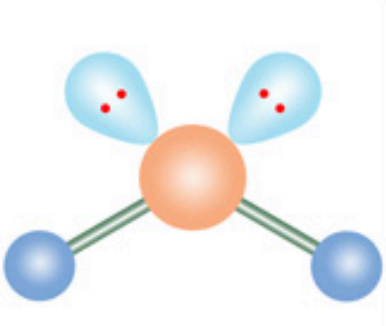


Figure 24.4.3.3 - Bent Molecular Geometry

24.4.4 Trigonal Bipyramidal

We follow the same idea to obtain *Table 24.4.4.1* below.

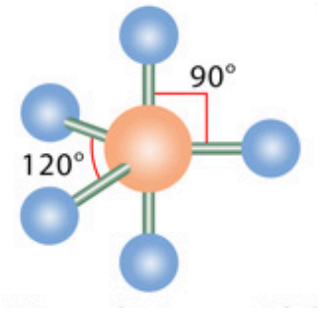
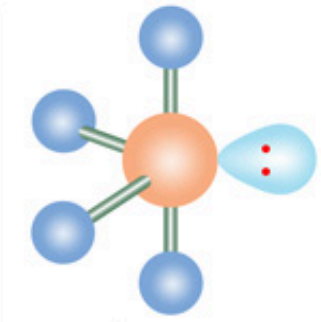
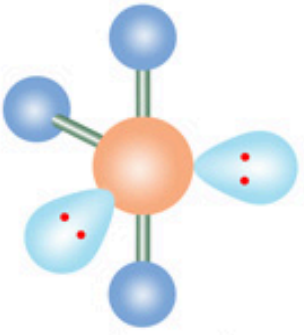
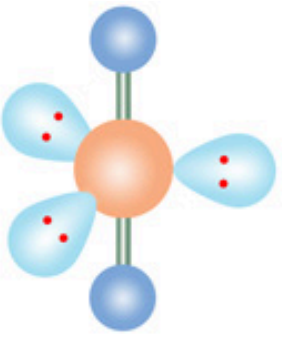
No. of Lone Pairs	Molecular Geometry	Structure
0	Trigonal Bipyramidal (e.g. PCl_5)	
1	Seesaw (e.g. SF_4)	
2	T-Shaped (e.g. ClF_3)	
3	Linear (e.g. XeF_2)	

Table 24.4.4.1 - Molecular Geometries with 5 Electron Domains

The Trigonal Pyramidal Electron Domain Geometry is unique in the sense that not all Electron Domains can be considered "equal". It has two distinct positions: **Axial** and **Equatorial**, denoted by "ax" and "eq" respectively in *Figure 24.4.4.2* below.

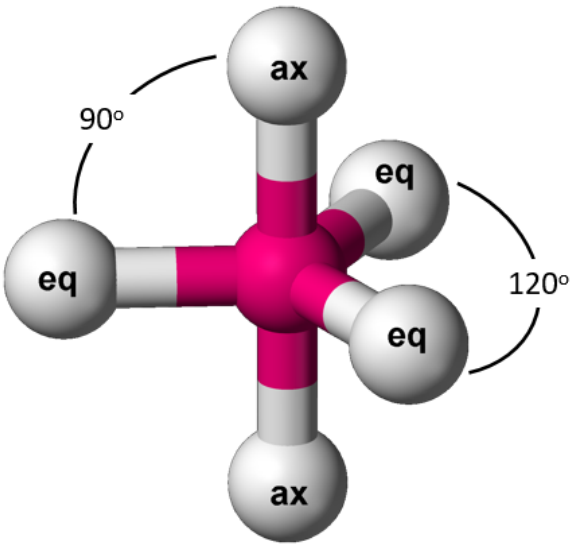


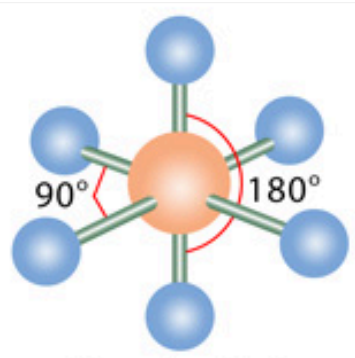
Figure 24.4.4.2 - Axial vs Equatorial

These Axial-Equatorial and Equatorial-Equatorial bonds have different bond angles, with 90° and 120° respectively.

In the Trigonal Bipyramidal Electron Domain Geometry, the Lone Pairs will always tend to take the **Equatorial** positions.

24.4.5 Octahedral

Again, we use the same idea to obtain *Table 24.4.5.1* below.

No. of Lone Pairs	Molecular Geometry	Structure
0	Octahedral (e.g. SF_6)	

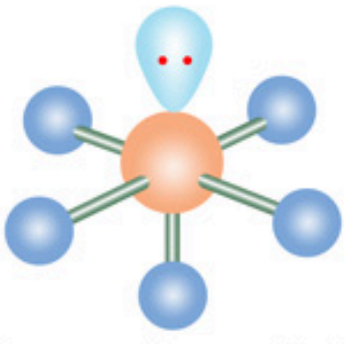
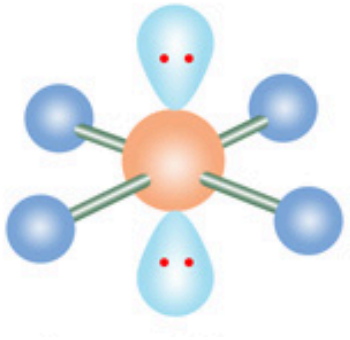
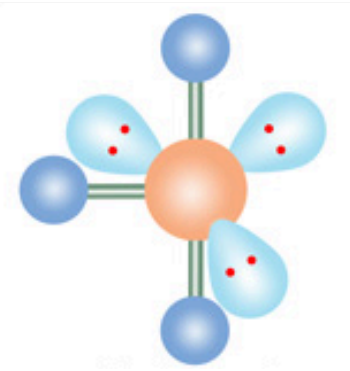
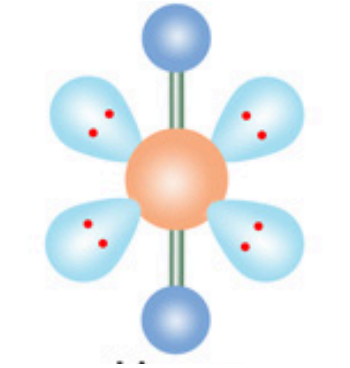
1	Square Pyramidal (e.g. BrF_5)	
2	Square Planar (e.g. XeF_4)	
3	T-Shaped	
4	Linear	

Table 24.4.5.1 - Molecular Geometries with 6 Electron Domains

24.5 Bond Angles

There are a few factors that will affect the Bond Angle within a molecule. For example, Non-Bonding Pairs and Double/Triple Bonds both change the Bond Angle.

24.5.1 Non-Bonding Electron Pairs

Bonding Electrons are **attracted** by the positive nuclei of the **two atoms** that are bonded together, so they have **less spatial distribution**. Think of it as the covalent bond being pulled upon by the two atoms, "locking" in space.

On the other hand, Non-Bonding Electrons are only **attracted** by the positive nucleus of **one atom**, which allow it to move more freely and **occupy a larger space** than Bonding Electrons.

This means that Non-Bonding Electrons are usually further away from the surrounding nuclei, while the Bonding Electrons are held closer to the nuclei, implying that **Non-Bonding Electrons** exert a **greater repulsive force** on the other electron pairs.

In short, VSEPR Theory tells us that Bond Angles **decrease** by approximately 2.5° for each additional **lone pair** present.

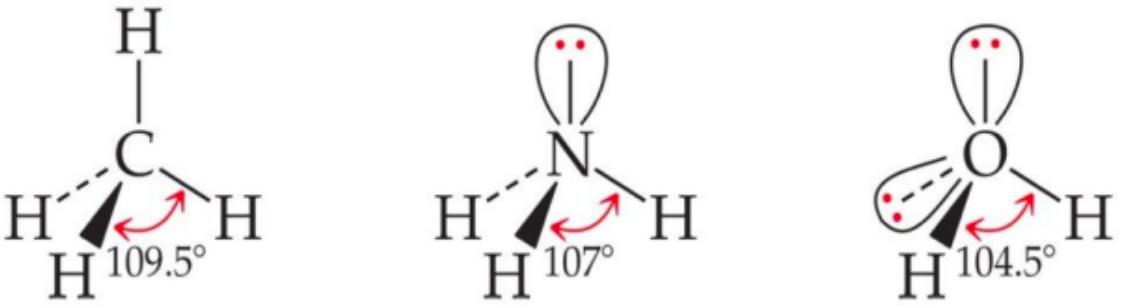


Figure 24.5.1.1 - Bond Angles Varying as the Number of Lone Pairs Change

For example, CH_4, NH_3 and H_2O all have a Tetrahedral Electron Domain Geometry, but have 0, 1 and 2 lone pairs respectively. We can see that the bond angle decreases by 2.5° from CH_4 to NH_3, and from NH_3 to H_2O.

Again, this is due to the slightly higher repulsion **between a Lone Pair and a Bonding Pair**, than **between two Bonding Pairs**, forcing the Bonding Pairs closer when there are more Lone Pairs.

24.5.2 Multiple Bonds

In a Double or Triple Bond, there is a **higher electron density** present in the bond than in a Single Bond. As such, similar to a non-bonding electron pair, this exerts a **greater repulsive force** on adjacent electron domains, which change the bond angles present.

Consider a molecule of phosgene, $COCl_2$, as in *Figure 24.5.2.1* below.

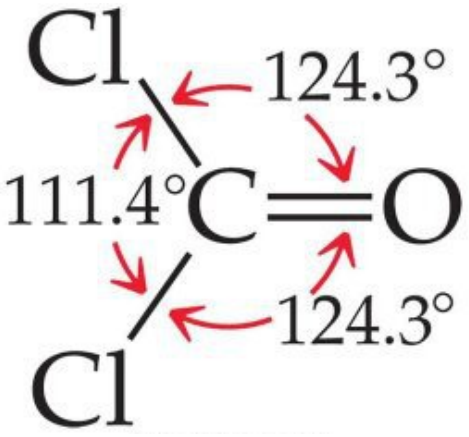


Figure 24.5.2.1 - Phosgene

It has a Trigonal Planar Molecular Structure, because there are three electron domains around the carbon, with no lone pairs.

However, notice that the bonding electrons in the $C = O$ double bond exerts a **greater repulsive force** on the electrons in the $C - Cl$ single bonds, which force the $Cl - C - Cl$ bond angle to decrease below the expected 120°, while increasing the $Cl - C = O$ bond angle to above 120°.

24.6 Dipole Moments

The **Dipole Moment** of a molecule is a way to measure its **polarity**. When there is a **separation of charge** due to a difference in electronegativity, a dipole moment is induced.

In a bond between two atoms, when one atom is more electronegative than the other, that atom will attract the electrons of the bond closer to it. For example, in water, the Hydrogens have an electronegativity of 2.2, while the Oxygen has an electronegativity of 3.44.

Oxygen is significantly more electronegative than Hydrogen, so it attracts the electrons of the two bonds closer to it.

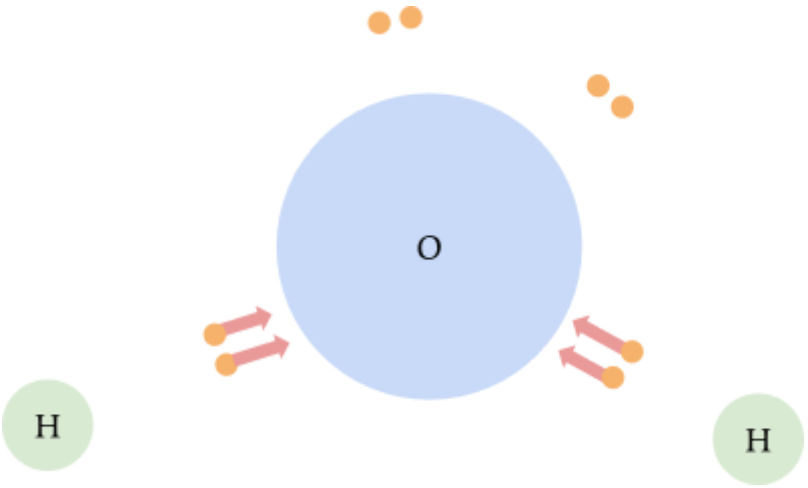


Figure 24.6.1 - Oxygen Attracting Bonding Pairs

In *Figure 24.6.1*, the orange dots denote the Electron Pairs, and the red arrows denote the attractive forces due to the difference in electronegativity.

These red attractive forces can actually be treated as vectors and by vector addition, we will obtain the **Dipole Moment** of the molecule, denoted in purple.

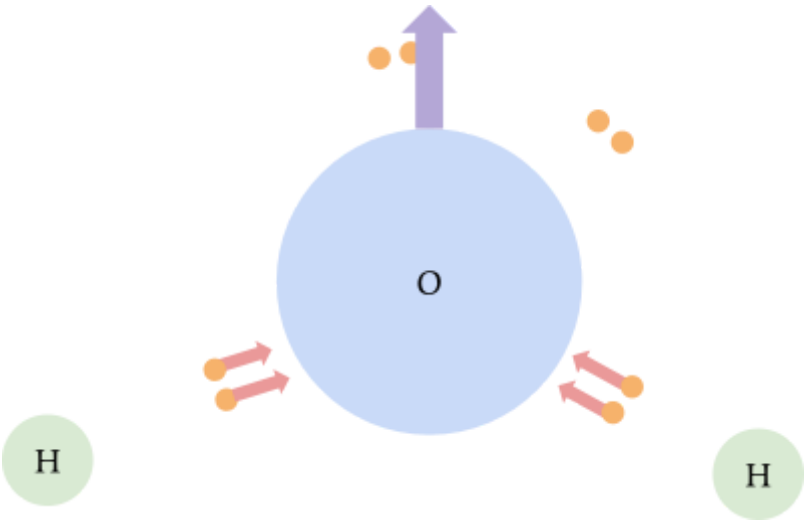
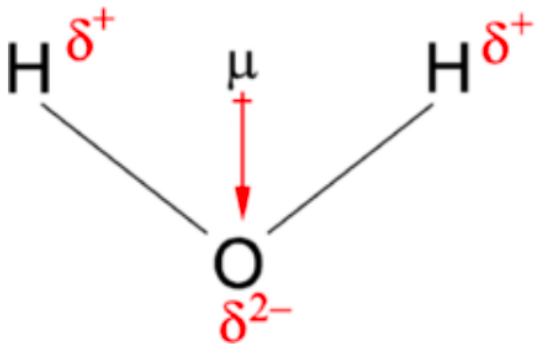


Figure 24.6.2 - Dipole Moment of Water

Notice that the water molecule is drawn this way because VSEPR Theory tells us that water has a Bent Molecular Geometry. We can also omit the lone pairs, and obtain the following:



Zs

Figure 24.6.3 - Dipole Moment of Water

In *Figure 24.6.3*, μ denotes the **Dipole Moment** while δ^+ and δ^{2-} denote partial positive and partial negative charges, because of the **electron poor** and **electron rich** regions respectively.

Note the arrow denoting Dipole Moment points **from positive to negative**.

Molecules with **symmetric** partial charges are **non-polar** because the dipole moments **cancel each other out**, causing **no net dipole moment**.

On the other hand, molecules with **non-symmetric partial charges** are **polar**, because the dipole moments **do not cancel each other out** and there is a **non-zero net dipole moment**.

Pause and Ponder 24.6

Answer the following questions about Boron Trifluoride.

a) State its chemical formula.
b) Calculate the difference in electronegativity and predict the type (Polar Covalent, Non-Polar Covalent, or Ionic) of bonding present.
c) Even though the difference in electronegativity is high, Boron Trifluoride has no net dipole moment. Explain why by drawing its Lewis Structure.
d) Predict if it will have high or low reactivity with Ammonia.

[Click to Go to Solution]

Topic 25: Formulae and Representations of Organic Compounds

25.1 Organic and Inorganic Compounds

An Organic Compound is a compound that contains **Carbon.** Most Organic Compounds will also contain **Hydrogen**, and some can contain other elements like oxygen or nitrogen.

An Organic Compound that only contains **Carbon** and **Hydrogen** is called a **Hydrocarbon.**

25.1.1 Properties of Organic vs Inorganic Compounds

Table 25.1.1.1 below summarises the properties of organic and inorganic compounds in general. Note that r.t.p. stands for "room temperature and pressure".

Property	Organic Compound	Inorganic Compound
Type of Bonding	Usually Covalent	Usually Ionic
Strength of Forces Between Particles	Weak	Strong
Physical State (at r.t.p.)	Gas/Liquid	Solid
Melting Point	Low	High
Boiling Point	Low	High
Flammability	Often Flammable	Non-flammable
Solubility in Water	Low or Insoluble	Often High
Solubility in Organic Solvents	Often High	Low or Insoluble
Electrical Conductivity	Insulator	Conductor
Rate of Reaction	Usually Low	Usually High

Table 25.1.1.1 - Properties of Organic vs Inorganic Compounds

25.2 Functional Groups

Functional Groups are what determines the reactivity of organic molecules. A functional group is simply a **structural feature** made of **a group of atoms** that cause compounds to have **similar physical and chemical properties**.

25.2.1 Hydrocarbons

See **Topic 27: Hydrocarbons** for more information.

There are 3 main groups of hydrocarbons: **alkane, alkene** and **alkyne**.

Alkanes only contain $C - C$ single bonds and have the general formula C_nH_{2n+2}.

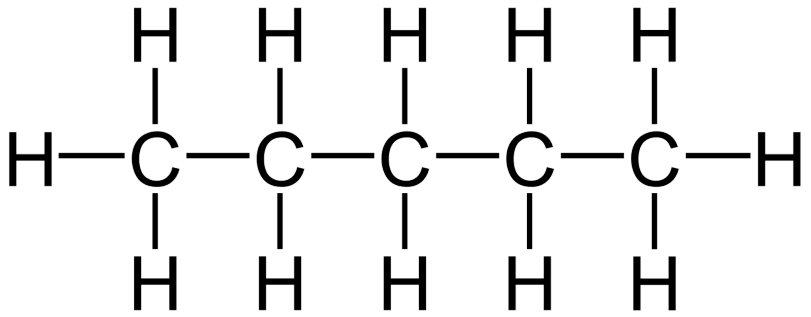


Figure 25.2.1.1 - Pentane

Alkanes are generally **not** regarded as **functional groups**.

Alkenes contain $C = C$ double bond(s) and have the general formula C_nH_{2n}.

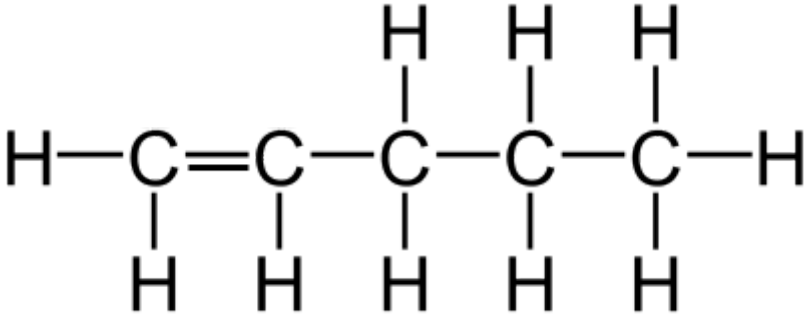


Figure 25.2.1.2 - Pent-1-ene

Alkynes contain $C \equiv C$ triple bond(s) and have the general formula C_nH_{2n-2}.

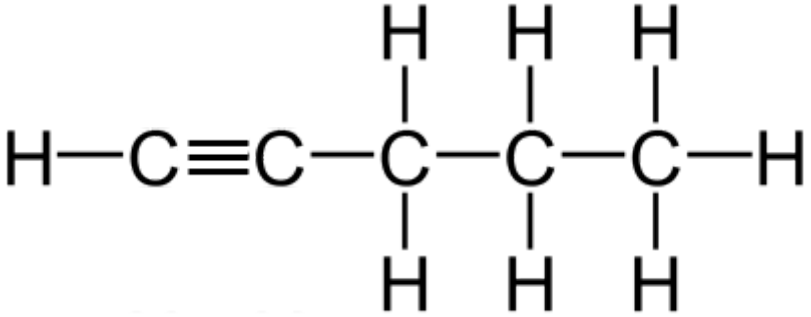


Figure 25.2.1.3 - Pent-1-yne

For **alkenes** and **alkynes**, the general formula will change if there is more than 1 double/triple bond, but we normally treat the **functional group** as having only 1 double/triple bond.

25.2.2 Haloalkane and Aromatic Compounds

A **haloalkane (See Section 27.1.3: Haloalkanes)**, or an **alkyl halide**, is an alkane with one (or more) hydrogens substituted for halogens, like **chloromethane**.

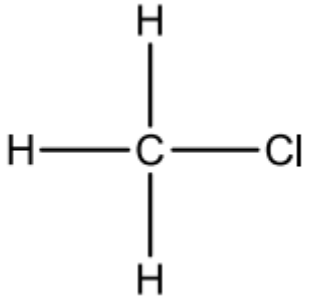


Figure 25.2.2.1 - Chloromethane

Another very famous hydrocarbon is **benzene** (See Section 27.6.3: Benzene), and it looks like:

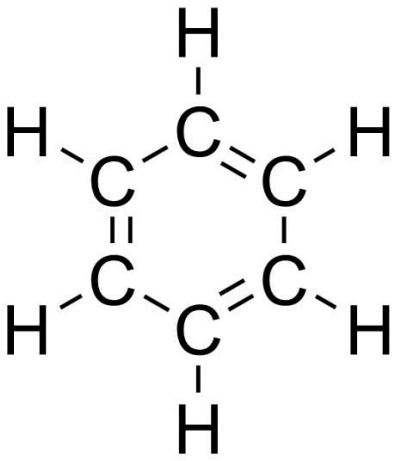


Figure 25.2.2.2 - Benzene

Note: Technically, this is not accurate, as benzene has 2 resonance structures.

Benzene-containing compounds are called **aromatic compounds (See Section 27.6: Aromatic Compounds)**.

25.2.3 Oxygen-Containing Functional Groups

See Topic 29: Organic Compounds Containing Oxygen for more information.

The main functional groups containing oxygen (marked in orange) are:

Alcohol $(-OH)$ See **Section 29.1: Alcohol**

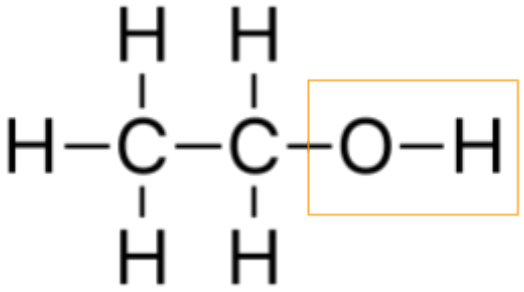


Figure 25.2.3.1 - Ethanol

Aldehyde $(-CHO)$

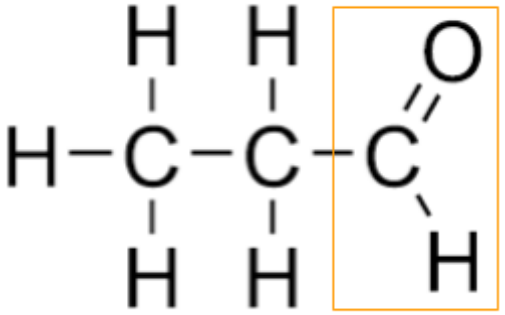


Figure 25.2.3.2 - Propanal

Ketone $(-CO-)$

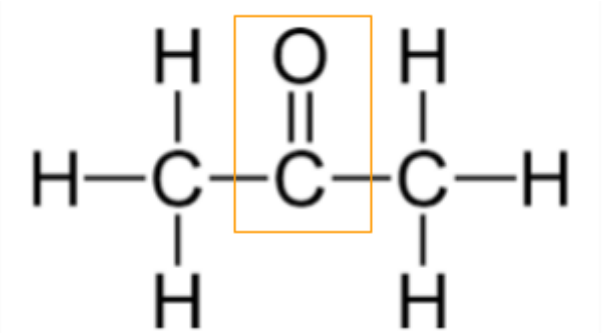


Figure 25.2.3.3 - Propan-2-one

Carboxylic Acid $(-COOH)$ See **Section 29.2: Carboxylic Acid**.

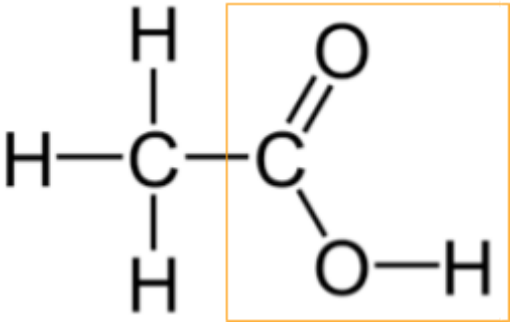


Figure 25.2.3.4 - Ethanoic Acid

Ester $(-COO-)$ See **Section 29.3: Ester**.

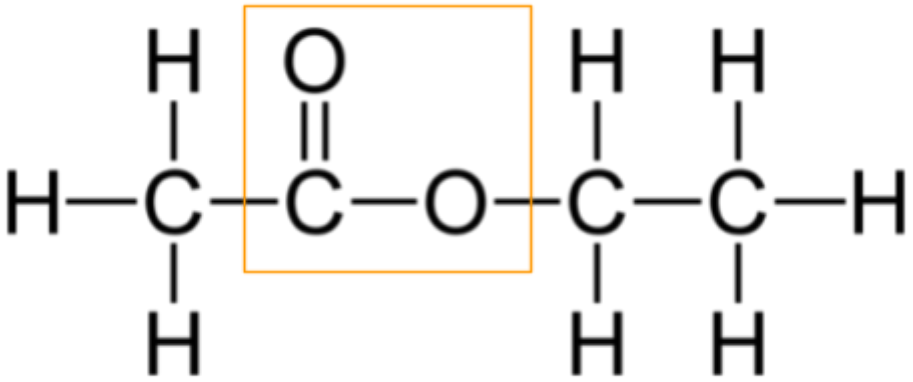


Figure 25.2.3.5 - Ethyl Ethanoate

Ether $(-O-)$

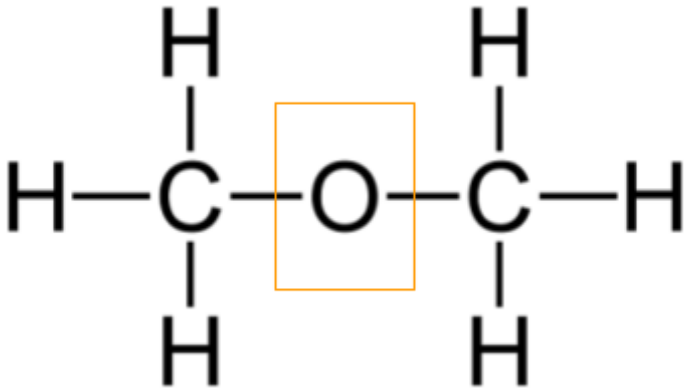


Figure 25.2.3.6 - Dimethyl Ether

25.2.4 Nitrogen-Containing Functional Groups

See **Topic 30: Organic Compounds Containing Nitrogen** for more information.

The main functional groups containing nitrogen (marked in orange) are:

Amide $(-NCO-)$

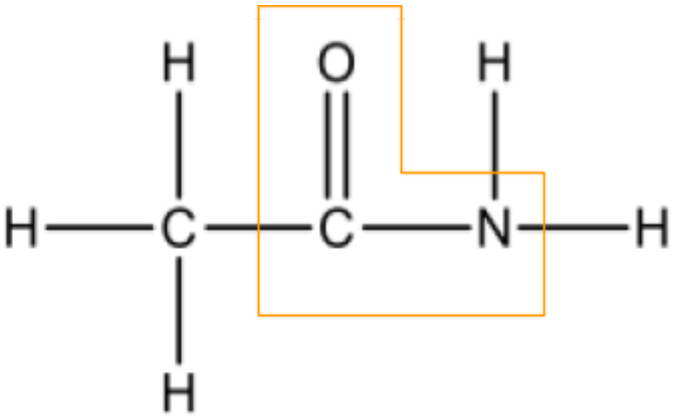


Figure 25.2.4.1 - Ethanamide

Amine $(-CN-)$

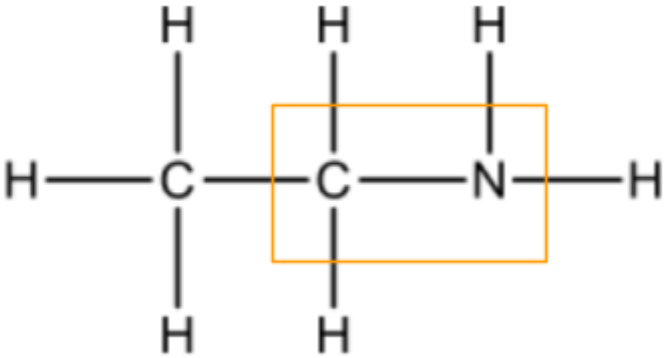


Figure 25.2.4.2 - Ethanamine

Nitrile $(-C \equiv N)$

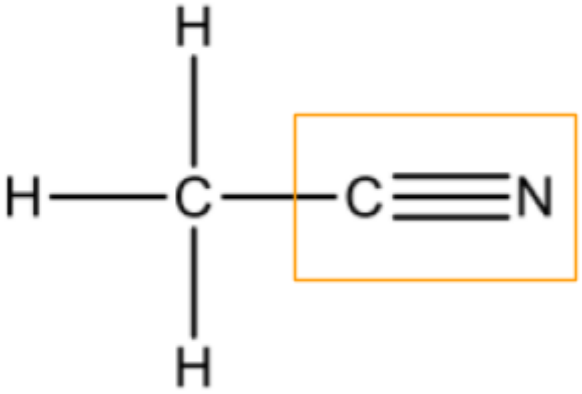


Figure 25.2.4.3 - Ethanenitrile

Note: Nitrile has an inorganic name as well: CN^- is also called cyanide.

25.3 Representations of an Organic Molecule

There are many equivalent ways to represent any organic molecule. We will use hexane as an example molecule.

25.3.1 General, Molecular and Empirical Formulae

For hexane, the General Formula refers to the General Formula for an alkane, C_nH_{2n+2}.

The Molecular Formula reflects the number of each type of atom that makes up each molecule, and for hexane it is C_6H_{14}.

The Empirical Formula reflects the simplest ratio of the number of each type of atom that makes up each molecule, and for hexane it is C_3H_7.

25.3.2 Displayed Structural Formula

Here, we just draw out the entire structure of the molecule, showing **all atoms and covalent bonds** like *Figure 25.3.1.1* below:

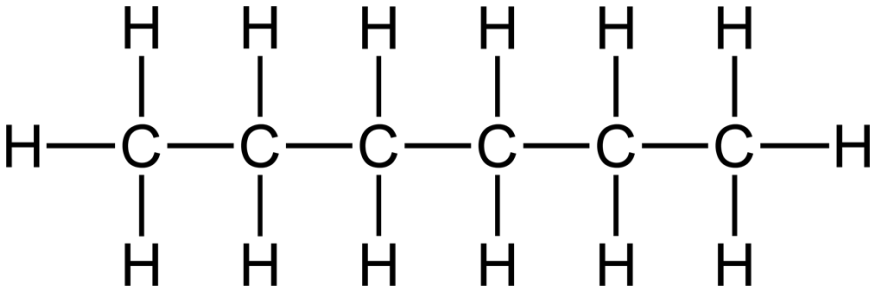


Figure 25.3.2.1 - Hexane

25.3.3 Condensed Formula

The Condensed Formula is the structural formula but without explicitly drawing all the covalent bonds. We simply take each carbon, and count the number of hydrogens around it. For hexane, the first carbon has 3 hydrogens, the second to fifth carbons have 2 and the last carbon has 3.

Hence, we can write hexane as $CH_3CH_2CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_4CH_3$.

25.3.4 Skeletal Formula

For a skeletal formula, we do not explicitly write out carbon atoms, and we do not write out hydrogen atoms that are attached to carbon atoms. Instead, we use zig-zag lines to denote carbon chains, with every junction having an implied carbon, unless another atom is written at the junction.

Figure 25.3.4.1 is, again, our hexane molecule!

Figure 25.3.4.1 - Hexane

Notice the implied carbons at these positions marked in blue, and the implied hydrogens that are bonded to these carbons!

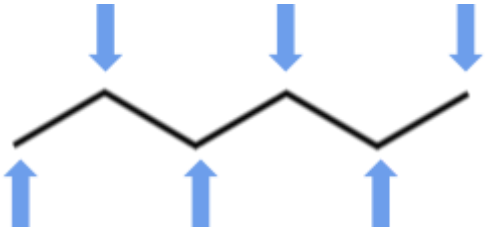


Figure 25.3.4.2 - Carbons and Hydrogens in Hexane

We shall introduce one more example. Consider butanoic acid, $CH_3(CH_2)_2COOH$.

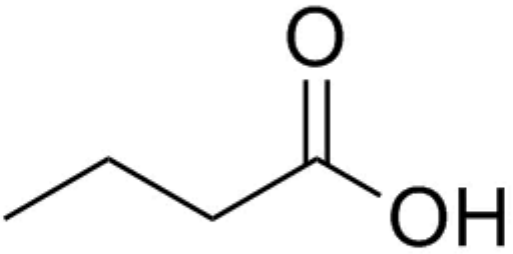


Figure 25.3.4.3 - Butanoic Acid

Here, we see that oxygen atoms are still shown, and double bonds are denoted by two lines. Also, hydrogens that are not bonded to carbons must still be written, like the hydrogen in the $O - H$ bond above. The implied carbons are marked with blue arrows below:

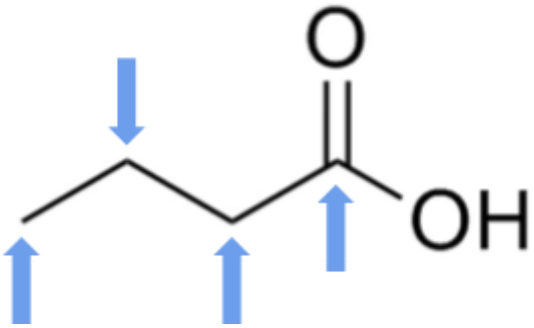


Figure 25.3.4.4 - Carbons in Butanoic Acid

Topic 26: Intermolecular Forces of Attraction

26.1 Intermolecular and Intramolecular Forces

Intermolecular Forces are electrostatic in nature and refer to forces of attraction between molecules, and comprises both **attractive** and **repulsive** components. That is, it is the sum of "pushing" and "pulling" between molecules. These Forces are important because they determine a lot of properties such as melting point of solids and boiling point of liquids.

They decrease in magnitude rapidly (proportional to the inverse square) with increasing distance between molecules, so they are very important for **Solids and Liquids**, where the atoms are packed **close together**.

They are only important for gases at high pressures, where these forces causes the gas to behave differently from an **ideal gas** (i.e. idealised model of gas where the size of molecules and forces between molecules are neglected)

Intramolecular Forces refer to forces present within a molecule such as covalent bonds that hold atoms together.

26.2 Dipole-Dipole Forces (DDF)

Recall that in Section 24.6: Dipole Moments we mentioned net dipole moments being present in polar molecules.

These net dipole moments can also be treated as partial positive δ+ and partial negative δ- charges present on the molecules.

The δ+ and δ- ends of different molecules will attract, generating an attractive force. Occasionally, some δ+ and δ+ ends of different molecules and some δ- and δ- ends of different molecules repel, generating a repulsive force. Overall, however, the **attractive forces will dominate**.

This overall attractive force, called **net attraction**, are the **dipole-dipole forces**, or **DDF**, and they exist only between **polar molecules**.

26.2.1 Factors Affecting Dipole-Dipole Forces

Since Dipole-Dipole Forces are caused by **dipole moments**, it is natural that **as the net dipole moment increases**, the **magnitude of Dipole-Dipole Forces increases**.

Experimentally, we can confirm this by measuring boiling points. As the magnitude of **Dipole-Dipole Forces increase,** more energy is required to overcome these attractions, and hence the boiling point will increase.

26.3 London Dispersion Forces (LDF)

London Dispersion Forces are (relatively) the weakest type of attractive force, and they exist in all **covalent molecules**, both between polar and non-polar molecules.

This is due to something called an **instantaneous dipole**. Usually, in non-polar molecules, the electrons are **symmetrically distributed**, so there is no **net dipole moment**.

However, at any instant in time, the electrons can be distributed **asymmetrically**, which causes one end to have a higher electron density and obtain a δ- charge, while the other end has a lower electron density and obtains a δ+ charge. An **instantaneous dipole** is formed.

The δ+ end can pull electrons from another nearby molecule closer to it, which causes the distribution of electrons to become **asymmetric**, and we say an instantaneous dipole is **induced**. An attractive force between the δ+ end of one molecule and the δ- end of the other will form.

This attractive force is the **London Dispersion Force**, or **LDF**.

26.3.1 Factors Affecting London Dispersion Forces

There are 2 main factors affecting London Dispersion Forces: **Polarizability** and **Shape of Molecule**.

Polarizability refers to the **tendency for an electron cloud to distort** due to an **instantaneous dipole**. In general, the larger the electron cloud, the more easily it is polarised. We can use the **molecular mass** or **molar mass** as an indicator of the size of the electron cloud.

Hence, **as the molar mass or molecular mass increases**, the **strength of London Dispersion Forces** increase.

For example, Xe has a higher boiling point than He because, Xe is **larger** and has a higher **molar mass**, which both indicate a **larger electron cloud** than He. As such, the strength of the **London Dispersion Forces** between Xe atoms are stronger than those between He atoms and require more energy to overcome. Hence, Xe has a higher boiling point.

Shape of Molecule here refers largely to **Surface Area/Compactness**. The **larger the surface area**, the strength of **London Dispersion Forces** increase.

For example, n-Pentane and Neopentane, which both have the Molecular Formula C_5H_{12}, have different boiling points.

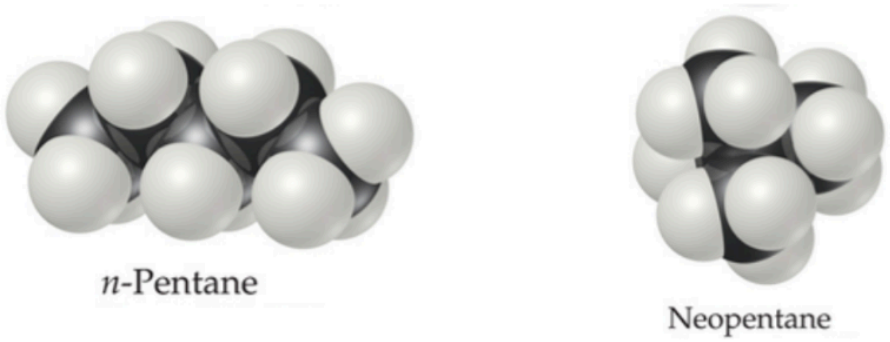


Figure 26.3.1.1 - n-Pentane and Neopentane

n-Pentane has a straight chain and hence has a **larger surface area**, while neopentane has a more spherical shape with a **smaller surface area**. Hence, there is **more contact** between molecules of n-Pentane than between molecules of neopentane.

This means that there are more **London Dispersion Forces** between molecules of n-Pentane, which requires **more energy** to overcome, and hence n-Pentane has a higher boiling point, of 36.1 °C, than Neopentane, with boiling point 9.5 °C.

26.4 Comparing Dipole-Dipole and London Dispersion Forces

In general, when two molecules have **similar molar mass** but have a large **difference** in **net dipole moments**, the London Dispersion Forces have a similar effect, and the **Dipole-Dipole Forces** will dominate in this case.

When two molecules have **similar net dipole moments** but have a large **difference** in **molar masses**, the Dipole-Dipole Forces will have a similar effect. Hence, the **London Dispersion forces** will dominate in this case.

26.5 Hydrogen Bonds

Hydrogen has an electronegativity of 2.2. When it bonds with highly electronegative elements like Fluorine, which has an electronegativity of 3.98, the difference in electronegativity is very large, around 1.78 in this case.

Hence, the bond generated is highly **polar** and has a **very large bond dipole**, and there is a **large** δ+ charge on the hydrogen atom, and a **large** δ- charge on the other atom. The δ+ hydrogen is attracted to a lone pair on the δ- part of another molecule, and an attractive force is formed.

What is unique to hydrogen is that it has a very small atomic radius, and so these bond dipoles are able to approach one another much more **closely**, before the positively charged nucleus of hydrogen begin to have a significant repulsive effect.

This means that hydrogen bonds have very **large bond dipoles** and very **small dipole-dipole distances**, and are **very strong interactions**.

Hydrogen bonds can occur when there are $H - O, H - F$ or $H - N$ bonds (and very rarely, $H - Cl$ and $H - S$ as well).

Hydrogen Bonding between water molecules is shown in *Figure 26.5.1* below, where the hydrogen dots are denoted by the green dashed lines.

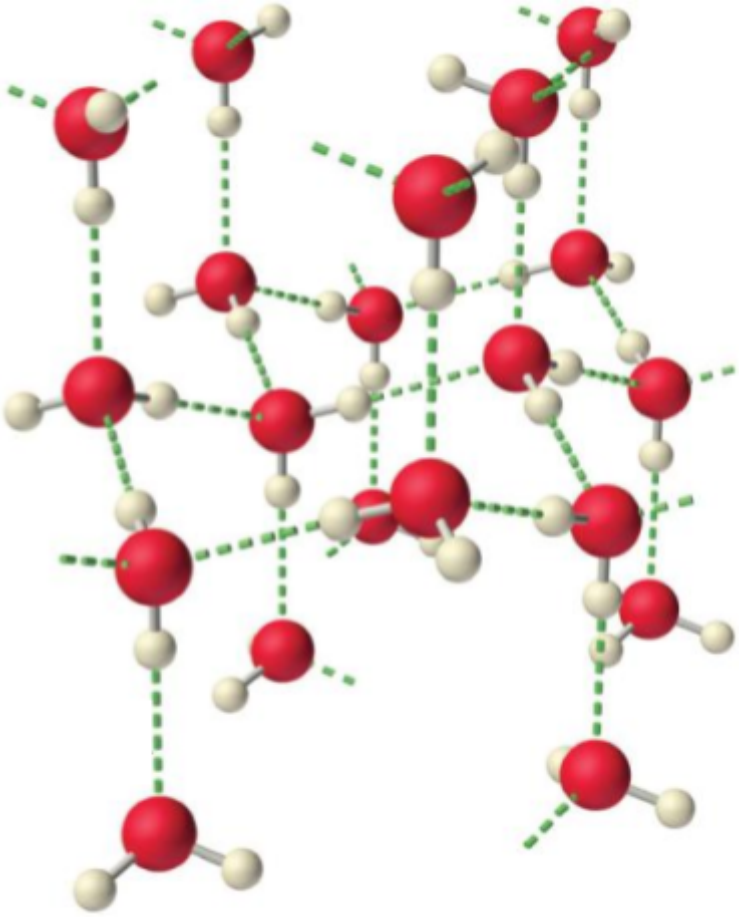


Figure 26.5.1 - Hydrogen Bonding in Ice

In fact, in ice, the water molecules are in an **orderly arrangement** held together by **hydrogen bonds**, which prevent the water molecules from moving closer. This means that there is more empty space in the structure of ice and hence ice is less dense.

Like in *Figure 26.5.2* below, each oxygen molecule is covalently bonded to 2 hydrogens and bonded by hydrogen bonds to 2 hydrogens. On the other hand, each hydrogen molecule is covalently bonded to 1 oxygen and is bonded by hydrogen bond to 1 other oxygen.

This gives ice an **tetrahedral** geometry.

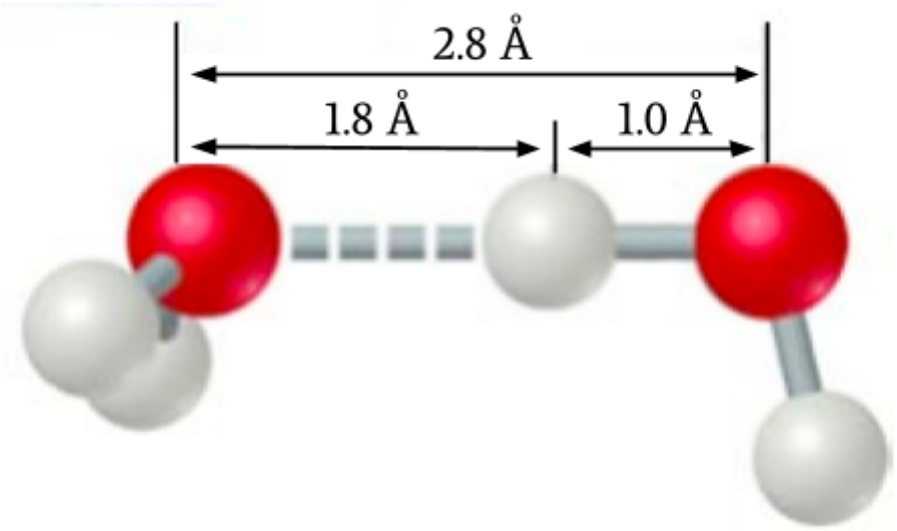


Figure 26.5.2 - Hydrogen Bond and Covalent Bond in Water Molecules

When ice melts, the structure **collapses** and the hydrogen bonds are **broken**. This allows the water molecules to move closer to one another and density increases. Hence, water is more dense than ice.

26.5.1 Factors Affecting Strength of Hydrogen Bonds

There are two main factors affecting strength of **hydrogen bonds** - **Extensiveness of Hydrogen Bonds** and **Polarity**.

Extensiveness of Hydrogen Bonds refer to the average number of hydrogen bonds each molecule will form with its neighbour. For example, in water, each molecule will form 2 hydrogen bonds, because there are 2 $H - O$ bonds and each O has 2 lone pairs.

Polarity refers to the polarity of the $H - X$ bond, where X is F, O or N. Since $H - F$ has the highest electronegativity difference, followed by $H - O$ and then $H - N$, the strength of the hydrogen bonds caused by these bonds will follow a similar order. That is, a hydrogen bond in a molecule with a $H - F$ bond is stronger than one with a $H - O$ bond, which is in turn stronger than one with a $H - N$ bond.

Polarity affects Hydrogen Bonding because a Hydrogen atom bonded to the electronegative atom will lead to the Hydrogen being **more exposed**, and hence the **attraction** between it and the **lone pair**(s) on the electronegative atom are stronger, making a stronger bond.

26.6 Miscibility: Like Dissolves Like

When two substances are held together by the same intermolecular force of attraction, they are **miscible** (i.e. mix to form a homogenous mixture) with each other in general.

This means, **polar covalent** liquids are miscible with each other in general, because the main intermolecular force of attraction is **Dipole-Dipole Forces** (or **Hydrogen Bonding**, for certain polar molecules).

Similarly, **non-polar covalent** liquids are miscible with each other in general, because the main intermolecular force of attraction is **London Dispersion Forces**.

.26.6.1 Examples of Exceptions: Solubility of Alcohol in Water and Hexane

Figure 26.6.1.1 below shows a molecule of **ethanol**.

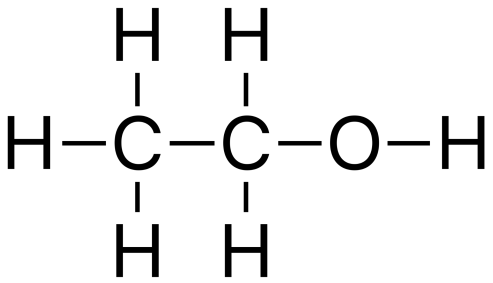


Figure 26.6.1.1 - Ethanol

We see that **ethanol** is a **polar** molecule due to the presence of the $O - H$ bond, and because of the **high electronegativity** of O, the main intermolecular force that holds molecules of ethanol together is **Hydrogen Bonding**.

As we know, **water** is also a **polar molecule** and also **experiences Hydrogen Bonding**. As such, we predict that ethanol and water is miscible. This is in fact true! (Think about the fractional distillation teacher demonstration in Year 1, fractional distillation is used for miscible liquids)

However, consider a molecule of **hexane**, as shown in *Figure 26.6.1.2* below.

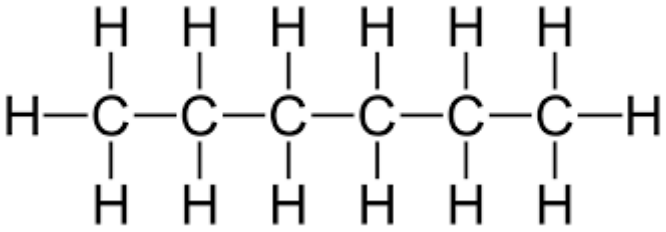


Figure 26.6.1.2 - Hexane

Hexane is a **non-polar** molecule due to its symmetry, and **low electronegativity difference** across bonds. Hence, we would predict that ethanol and hexane are immiscible. However, this is **NOT true!**

This is due to the **London Dispersion Forces** that are present on both hexane and ethanol on the **hydrocarbon chains**, which allows hexane and ethanol to be **miscible** with each other.

This is a particularly interesting special case as **ethanol** is miscible with both **hexane** and **water**, yet **water** and **hexane** are not miscible with one another!

*(FYI: In fact, mathematically, this means that the "is miscible" operator, denoted ~, is not **transitive**, that is, in logic notation, $A \sim B$ and $B \sim C$ does not imply $A \sim C$.)*

In fact, we see a trend in *Table 26.6.1.3* below.

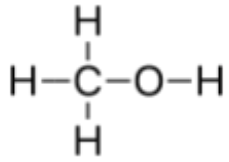
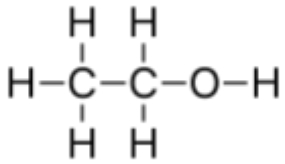
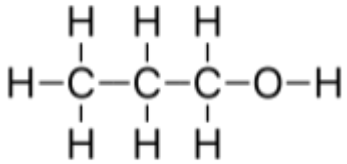
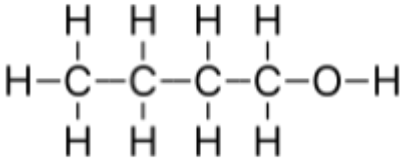
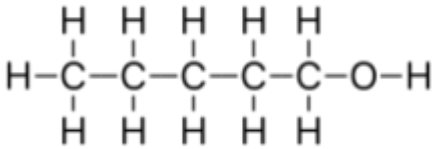
Structure	Name	Solubility in Water	Solubility in Hexane
	Methanol	Miscible	$0.592 g/cm^3$
	Ethanol	Miscible	Miscible
	Propanol	Miscible	Miscible
	Butanol	$0.073 g/cm^3$	Miscible
	Pentanol	$0.022 g/cm^3$	Miscible

Table 26.6.1.3 - Solubilities of Alcohols in Water and Ethanol

In general, as the length of the **alcohol increases**, the **solubility in water decreases**, while the **solubility in hexane increases**.

Here, we can think of "Miscible" as having infinite solubility.

This is because, as the length of alcohol increases, the $O - H$ bond is a **smaller part** of the molecule and the alcohol will behave like a more non-polar molecule, and the interaction with water decreases, so the solubility of alcohol in water decreases.

On the other hand, as the length of alcohol increases, its **polarisability increases** due to a larger and heavier molecule. Hence, the strength of **London Dispersion Forces** on the alcohol become stronger, and the interaction with hexane increases, so the solubility in hexane increases.

26.7 Ion-Dipole Forces (IDF)

Ion-Dipole Forces are the **strongest** intermolecular forces of attraction, but are still weaker than ionic and covalent bonds. They exist between **ions** and **ends of polar molecules**.

Consider the scenario of dissolving some **table salt** ($NaCl$) in **water**. Water is a **polar molecule**, and has a δ+ end on the oxygen and a δ- end on the hydrogens.

The δ- end of the water molecules are attracted to the Na^+ ions because of **Ion-Dipole Forces**,

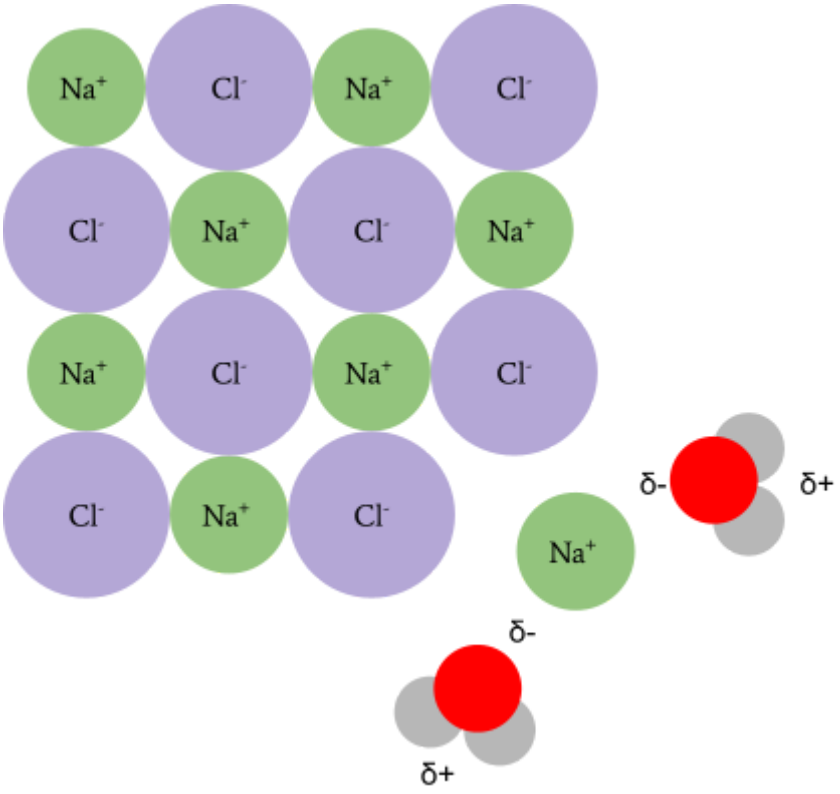


Figure 26.7.1 - Ion-Dipole Forces

And this causes the Na^+ ions to be freed from the lattice (i.e. be dissolved). A similar mechanism can be spotted between the δ+ and the Cl^- ions in the lattice, as shown in *Figure 26.7.2* below.

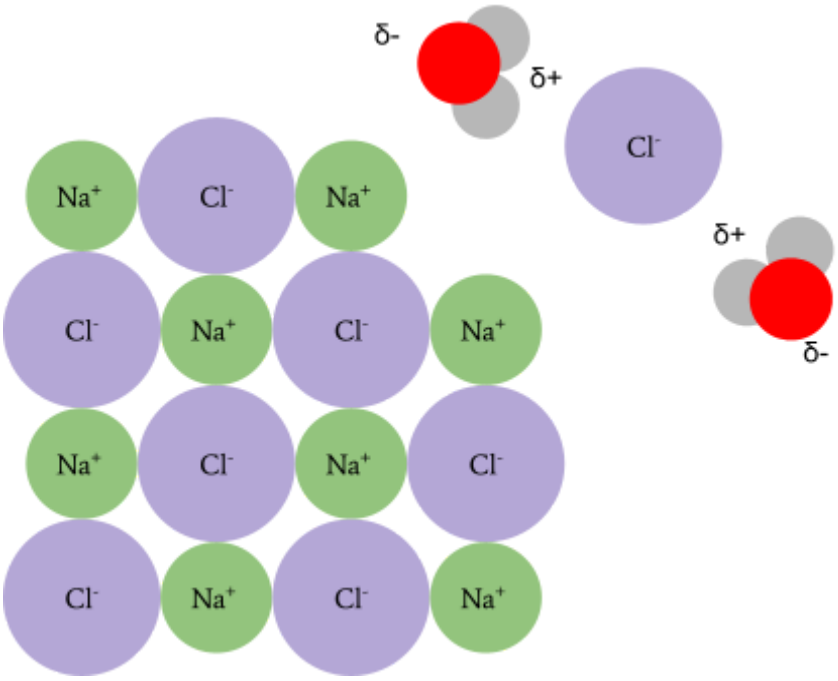


Figure 26.7.2 - Ion-Dipole Forces

26.7.1 Factors Affecting Strength of Ion-Dipole Forces

The two main factors affecting strength of **Ion-Dipole Forces** are **Charge of Ion** and **Size of Ion**:

- The **higher the charge on the ion**, the **stronger the Ion-Dipole Forces**.
- The **smaller the size of the ion** (ionic radius), the **stronger the Ion-Dipole Forces.**
- The **higher the dipole moment**, the **stronger the Ion-Dipole Forces**.

FYI: This is directly due to Coulomb's Law, which states $F = \frac{kq_1 \times q_2}{r^2}$. Since k is a constant (Coulomb's Constant), we can only increase q_1, q_2 or decrease r to increase F. This makes Ion-Dipole Forces very similar to **Ionic Bonds.**

26.8 Dipole-Induced Dipole Forces (DIDF)

As seen in Section 26.3: London Dispersion Forces, the presence of an instantaneous dipole in one molecule is able to induce an instantaneous dipole in a neighbouring molecule.

Similarly, the presence of a **dipole** in a polar molecule is able to **induce a dipole in a neighbouring non-polar molecule**. There is then a force of attraction between the δ+ and δ- ends of these dipoles, which is known as **Dipole-Induced Dipole Forces**, or **DIDF.**

26.8.1 Factors Affecting Strength of Dipole-Induced Dipole Forces

The two main factors affecting strength of Dipole-Induced Dipole Forces are the **dipole moment of the polar molecule** and the **polarizability of the non-polar molecule**:

- The **stronger the dipole moment of the polar molecule**, the stronger the resulting **Dipole-Induced Dipole Forces.**
- The **higher polarizability of the non-polar molecule**, the stronger the **Dipole-Induced Dipole Forces**.

26.9 Forces of Attraction Ranked

In general, the strength of the forces of attraction are (in decreasing order of strength):

Force of Attraction	Type
Covalent Bond	**Intra**molecular Force
Ionic Bond	
Metallic Bond	
Ion-Dipole Forces	**Inter**molecular Force
Hydrogen Bond	
Dipole-Dipole Forces	
Dipole-Induced Dipole Forces	
London Dispersion Forces	

Table 26.9.1 - Forces of Attraction Ranked

Topic 27: Hydrocarbons

27.1 Alkanes

Alkanes are one the simplest classes of organic compounds. Alkanes contain only carbon and hydrogen, and are hence called **hydrocarbons**. They also have only single covalent bonds between the carbon atoms, and for this reason we call them **saturated hydrocarbons**.

In a straight chain alkane, each carbon is connected to exactly one carbon and three hydrogens if they are on the extreme ends of the chain, or exactly two carbons and two hydrogens otherwise. For example, consider the expanded formula of hexane:

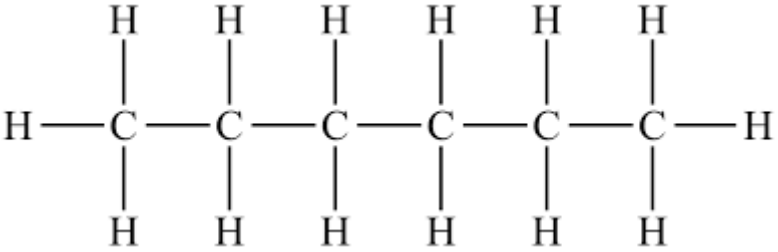


Figure 27.1.1 - Hexane

However, in fact, VSEPR Theory tells us that the carbons all have a tetrahedral geometry, since it forms 4 covalent bonds with no lone pairs. While the above representation in *Figure 27.1.1* is good enough and is the largely accepted version, technically *Figure 27.1.2* is more accurate.

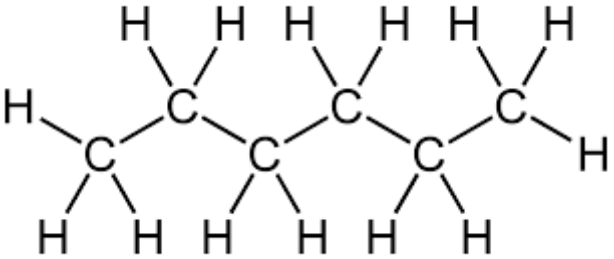


Figure 27.1.2 - Hexane

Convince yourself that we have simply just rotated some bonds slightly and the molecular structure is unchanged! Of course, this is very time-consuming, so it is **NOT necessary** to draw *Figure 27.1.2* during exams, and we can just draw *Figure 27.1.1* when asked for the structure.

However, how time-consuming *Figure 27.1.2* is serves as a motivation for something else - the **Skeletal Structure**. In short, we don't display the hydrogens that are bonded to carbon, because we can simply deduce this number from the number of bonds carbon forms, since each carbon forms a total of 4 bonds.

For example, the leftmost carbon must have 3 hydrogens bonded to it because it has only 1 bond with another carbon. Furthermore, we don't actually need to draw the carbons either as we can denote them as a bend in our drawing. Hence, the **Skeletal Structure** of hexane is:

Figure 27.1.3 - Hexane

Much easier to draw, right? Also, it is standard practice to (at least attempt to) draw the bond angles at **120°**, instead of the 109.5° suggested by VSEPR Theory, for convenience.

Of course, alkanes don't have to be just one straight chain. *Figure 6.1.4* below, for example, is also an alkane, since it only consists of single bonds between carbons.

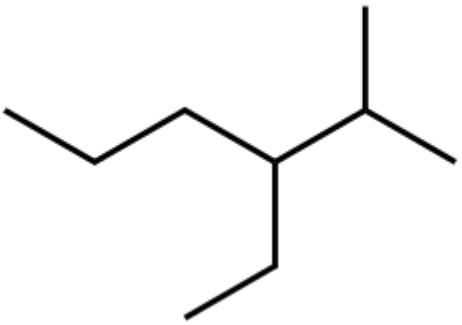


Figure 27.1.4 - Mystery Alkane

We can also draw out its full structure:

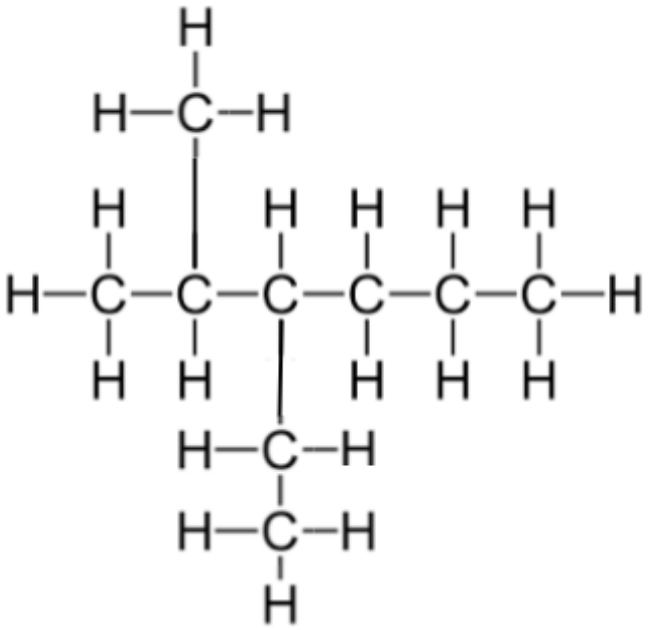


Figure 27.1.5 - Mystery Alkane but Full Structure! (This took way too long to draw)

How should we name something like this? Well, we have a set of naming rules, as will be covered later. But let's start with the simpler examples first!

27.1.1 Naming of Alkanes

To find the IUPAC (International Union of Pure and Applied Chemistry) name of an alkane, which is the widely accepted name, we will first identify the **longest carbon chain**. For example, let's see why hexane is called hexane! In hexane, the longest carbon chain is here:

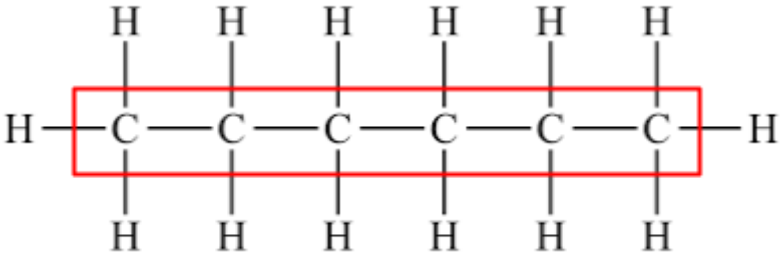


Figure 27.1.1.1 - Hexane Again

We see that there are 6 carbons in this chain! We then refer to *Table 6.1.1.2* below to find the corresponding prefix based on the number of carbons, and add "**ane**" to the end, to signify that it is an alk**ane**.

Number of Carbons	Prefix	Corresponding Alkane	Condensed Formula
1	meth-	methane	CH_4
2	eth-	ethane	CH_3CH_3
3	prop-	propane	$CH_3CH_2CH_3$
4	but-	butane	$CH_3(CH_2)_2CH_3$
5	pent-	pentane	$CH_3(CH_2)_3CH_3$
6	hex-	hexane	$CH_3(CH_2)_4CH_3$
7	hep-	heptane	$CH_3(CH_2)_5CH_3$
8	oct-	octane	$CH_3(CH_2)_6CH_3$
9	non-	nonane	$CH_3(CH_2)_7CH_3$
10	dec-	decane	$CH_3(CH_2)_8CH_3$

Table 27.1.1.2 - Prefix and Alkane Names

This is why hexane is called hexane!

Let's try another alkane. Consider *Figure 27.1.1.3* below.

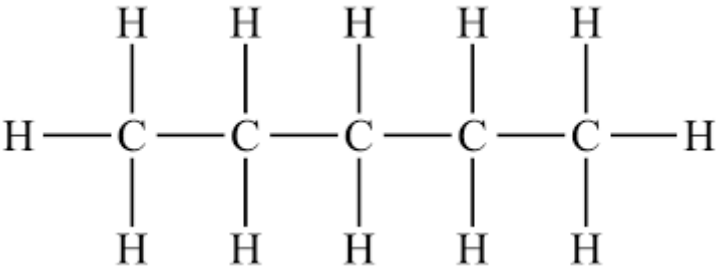


Figure 27.1.1.3: Another Mystery Alkane

Here, we count that the longest carbon chain has 5 carbons, so we use the prefix "pent-", and this is **pentane**!

Now, what if there were multiple chains instead of just one, like we saw before?

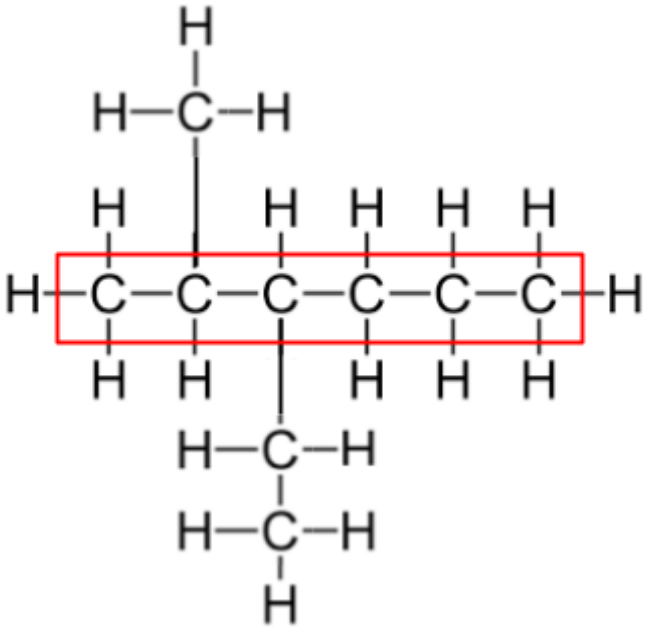


Figure 27.1.1.4 - Original Mystery Alkane is Back!

Here, we see the longest carbon chain is the one highlighted in red. If we count the number of carbons, we see that there is 6. Hence, we have the **root word** "**hexane"**.

The remaining carbons are called **side chains** and are added as prefixes to the root word.

To name a side chain, we again count the number of carbons and use the same prefixes from *Table 27.1.1.2* from before. However, instead of adding **-ane**, we add **-yl** instead.

These side chains are called **alkyls**, and they are the same as **alkanes**, but they are missing one hydrogen. This place where a hydrogen is missing is where the **alkyl** is able to bond to an **alkane**.

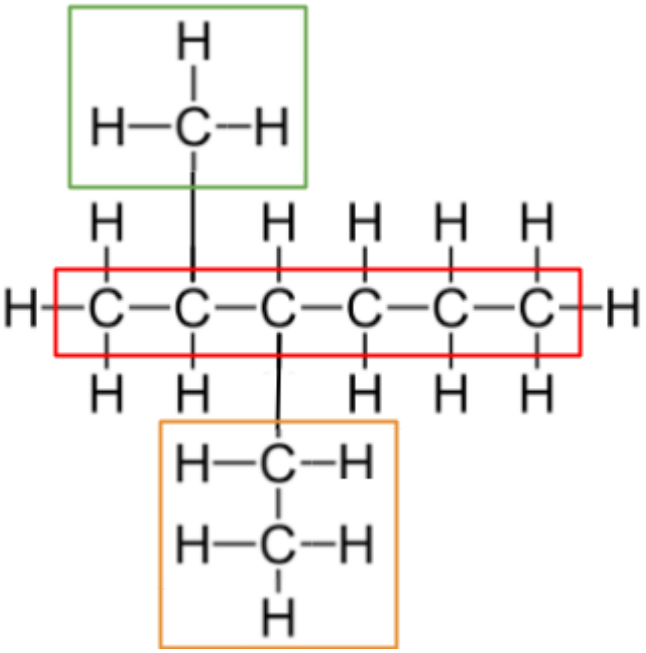


Figure 27.1.1.5 - Side Chains and Main Chain

Hence, for the side chain in **green**, we see that it has one carbon, so we use the prefix "meth-", and attach "-yl" to obtain "methyl".

Similarly, for the side chain in **orange**, we see that it has two carbons, so we use the prefix "eth-" and attach "-yl" to obtain "ethyl".

Now, we number the carbons on the **main chain** (parent chain) from left to right.

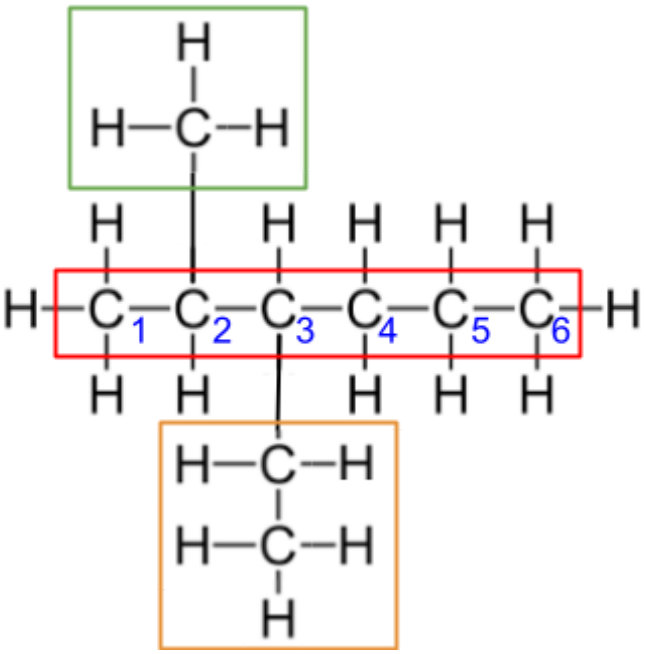


Figure 27.1.1.6 - Numbered Parent Chain

Now, the prefixes we attach come in the form "[position]-[name]". For our case, we have "2-methyl" and "3-ethyl", because the side chains are at the 2nd and 3rd carbons respectively.

Now, because “ethyl” comes first before “methyl” **alphabetically**, we write the ethyl first to get the full name “3-ethyl-2-methylhexane”.

But… not so fast! What if we numbered the carbons the other way, from right to left?

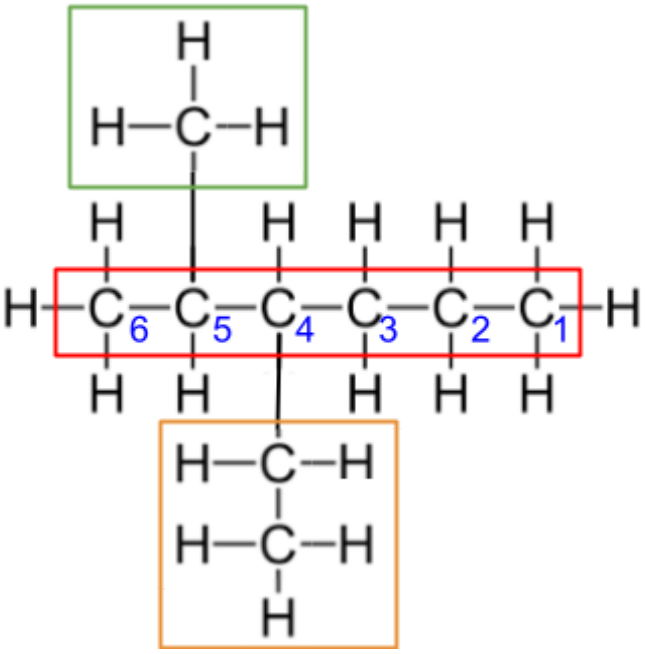


Figure 27.1.1.7 - Numbering Parent Chain in Reverse

Now, the name we get is “4-ethyl-5-methylhexane”. Which one is correct?

The rule is to prefer the one with the **smaller substituent numbers,** when the first difference is seen. Here 3 < 4, so the correct name is “3-ethyl-2-methylhexane”.

And we are done!

27.1.2 Multiple of a Substituent

Let’s consider the funny looking alkane in *Figure 27.1.2.1* below.

Figure 27.1.2.1 - Even More Mystery Alkane!

Yes, this is a real alkane. In fact, if we draw out the full structure, it would look like:

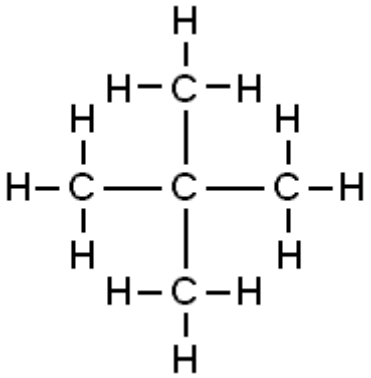


Figure 27.1.2.2 - Full Structure

We can see that the longest carbon chain can be marked as so:

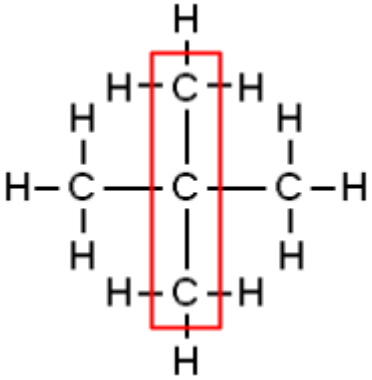


Figure 27.1.2.3 - Longest Carbon Chain

There are three carbons in this chain, so we have the root word **"propane"**.

Now, we move on to the substituents.

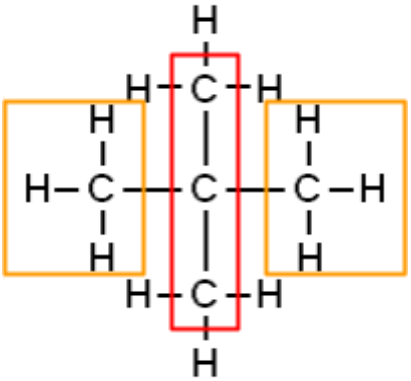


Figure 27.1.2.4 - Side Chains

Both side chains have one carbon, so they are called "methyl". We also note that they are both placed on the **second carbon**. But because they are both methyl groups, instead of writing "2-methyl-2-methylpropane", we would write "2,2-dimethylpropane".

FYI: This substance is also called **neopentane**!

This prefix before “methyl” are the same as the prefixes used for covalent substance naming: “di-” for 2, “tri-” for 3, “tetra-” for 4, “penta-” for 5, “hexa-” for 6, “hepta-” for 7 and so on.

Also, this prefix does not affect the alphabetical order in which the substituents are listed. That is, “dimethyl-” **does NOT** come before “ethyl-”.

27.1.3 Haloalkanes

Adding a **halogen** to an **alkane** is much like adding a side chain (called an **alkyl**) to an **alkane**.

Halogen	Prefix
Fluorine	Fluoro-
Chlorine	Chloro-
Bromine	Bromo-
Iodine	Iodo-

Table 27.1.3.1 - Halogens as Substituents

Consider this mystery haloalkane!

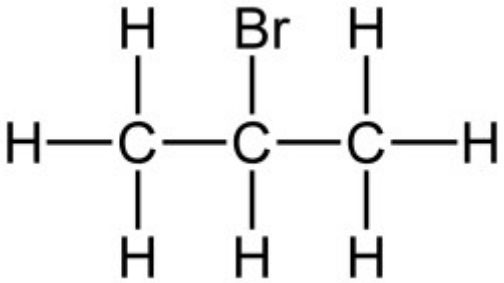


Figure 27.1.3.2 - Mystery Haloalkane

First, we find the root word. We locate the longest carbon chain, and we see that it has 3 carbons. Hence, the root word is **“propane”**.

We see that we have introduced **bromine** as a substituent onto propane, on the **second carbon**. Hence, the prefix we attach is “2-bromo”.

The name of this haloalkane is then “2-bromopropane”.

Let’s consider this other mystery haloalkane in *Figure 27.1.3.3* below.

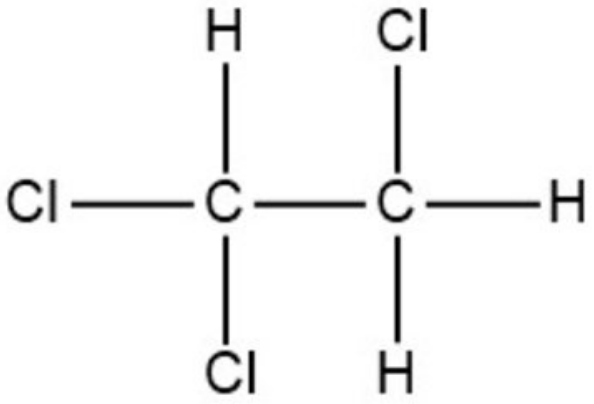


Figure 27.1.3.3 - Mystery Haloalkane No. 2

Here, we see the longest carbon chain has 2 carbons, so we have the root word **ethane**.

Now, there are three chlorine substituents. Let's try numbering the carbons from right to left.

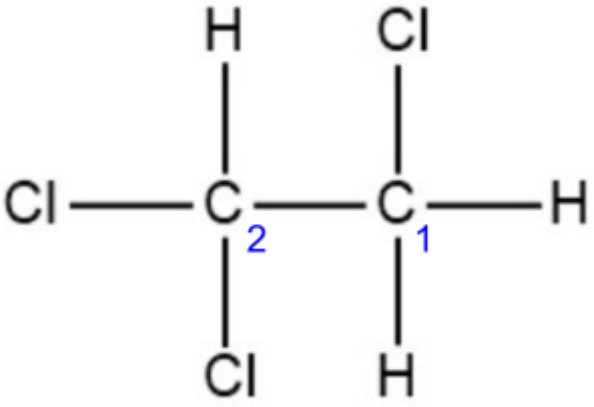


Figure 27.1.3.4 - Number from Right to Left

Here, we see that there will be 1 chlorine atom on the first carbon, and 2 chlorine atoms on the second carbon. Hence, the name in this case would be "1,2,2-trichloroethane".

However, if we number the carbons from left to right:

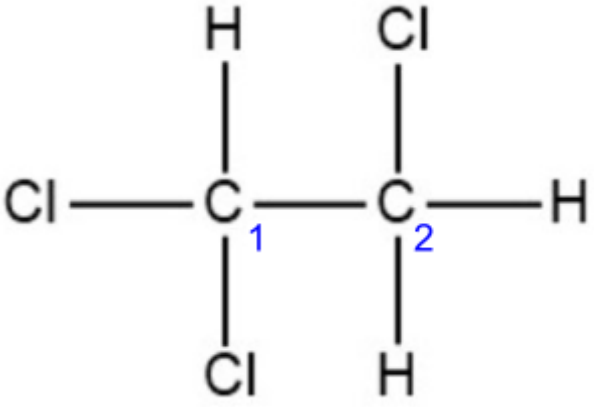


Figure 27.1.3.4 - Number from Left to Right

We would get the name "1,1,2-trichloroethane".

We check the substituent numbers. Since for the first number, 1 = 1, but for the second, 1 < 2, the correct name is actually the latter: "1,1,2-trichloroethane".

27.1.4 Structural Isomerism

Two substances are said to be **structural isomers** of one another if they have the **same molecular formula**, but they have different structures (arrangements) of atoms.

An example would be **pentane** (or n-Pentane) and **2,2-dimethylpropane** (or neopentane).

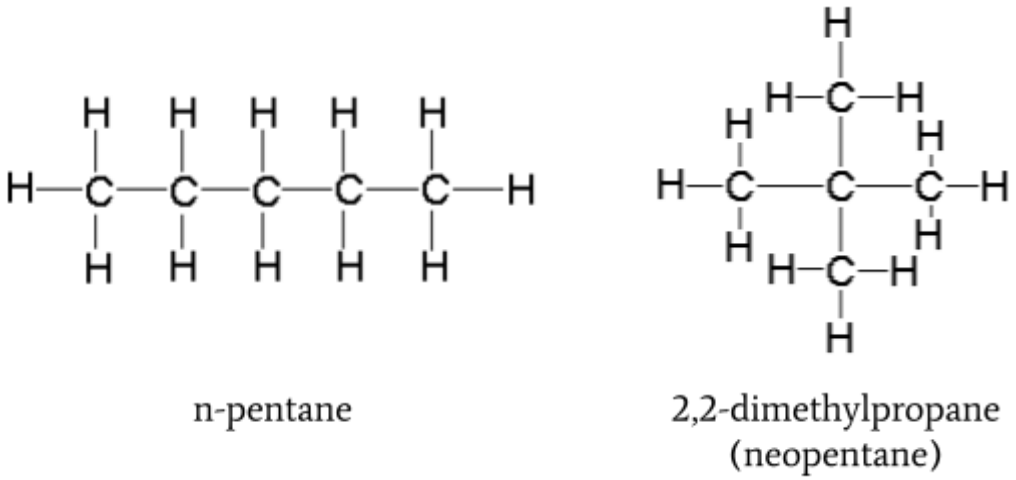


Figure 27.1.4.1 - Structural Isomers

Here, the "n-" in "n-pentane" refers to **normal**, which means that the pentane is an **unbranched chain** (i.e. straight chain).

Both n-pentane and 2,2-dimethylpropane have the molecular formula C_5H_{12}, but as we see from *Figure 27.1.4.1* above, they have different structures.

Hence, we say they are **structural isomers**.

Structural Isomers have different **IUPAC names,** but the same **molecular formula**.

27.1.5 Physical State of Alkanes

As the length of the alkane increases, the melting and boiling point of the alkane also increases.

In fact, alkanes with 1-4 carbons are **gases** at room temperature, alkanes with 5-17 carbons are **liquids** at room temperature, and alkanes with 18 or more carbons are **solids** at room temperature.

Pause and Ponder 27.1.5

Explain why as the length of alkane increases, its melting and boiling point also increases.

[Click to Go to Solution]

27.1.6 Combustion of Alkanes

Alkanes can undergo combustion by the following equation:

$$C_nH_{2n+2} + \frac{3n+1}{2} O_2 \rightarrow n\,CO_2 + (n+1)\,H_2O$$

where the entire equation is multiplied by 2 if n is even to eliminate the fractional O_2.

Through this reaction, energy is released as **heat** as a byproduct.

For example, **propane** has the combustion reaction: $C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$.
On the other hand, **ethane** has the reaction $C_2H_6 + \frac{7}{2} O_2 \rightarrow 2\,CO_2 + 3\,H_2O$. But we have to multiply by 2 to eliminate the fraction, so it is actually $2\,C_2H_6 + 7\,O_2 \rightarrow 4\,CO_2 + 6\,H_2O$.

When there is **sufficient oxygen**, the combustion is said to be a **complete combustion** and will have a **blue flame**, forming CO_2 and H_2O **only**.

When there is **insufficient oxygen**, the combustion is said to be an **incomplete combustion** and will have an **orange flame**, where CO and soot (C) are also released as products, along with CO_2 and H_2O.

27.1.7 Halogenation of Alkanes

Alkanes can undergo halogenation by the following equation:

$$C_nH_{2n+2} + X_2 \rightarrow C_nH_{2n+1}X + HX$$

and this can potentially be repeated:

$$C_nH_{2n+1}X + X_2 \rightarrow C_nH_{2n}X_2 + HX$$

and so on, where X is a halogen.

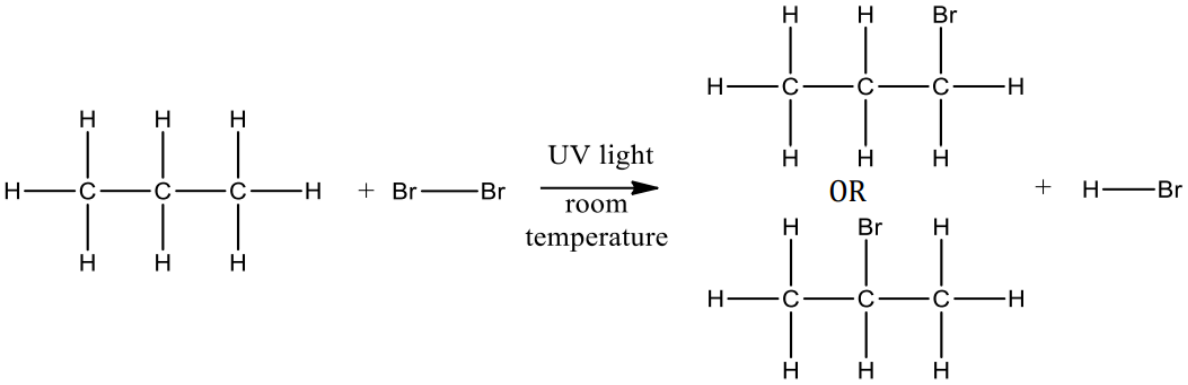
This reaction is catalysed by the presence of **UV Light**, and will not take place without UV Light under normal conditions.

For example, CH_4 can be halogenated in the following reactions:

- $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$
- $CH_3Cl + Cl_2 \rightarrow CH_2Cl_2 + HCl$
- $CH_2Cl_2 + Cl_2 \rightarrow CHCl_3 + HCl$
- $CHCl_3 + Cl_2 \rightarrow CCl_4 + HCl$

where UV Light is present as a catalyst.

Consider the **bromination** of **propane**. There are two possible haloalkane products:



This haloalkane product is either **1-bromopropane** or **2-bromopropane**.

Note that 3-bromopropane does not exist as we would simply reverse the numbering of the carbons on the propane chain and we 3-bromopropane is actually just **1-bromopropane**.

27.1.8 Cycloalkanes

A **cycloalkane** is an **alkane** that is also **cyclic**, that is, it has carbon atoms which form a **ring**.

To get a cycloalkane, you can think of starting with an alkane, and replacing a hydrogen on the leftmost carbon and one on the rightmost carbon with a covalent bond joining the two carbons.

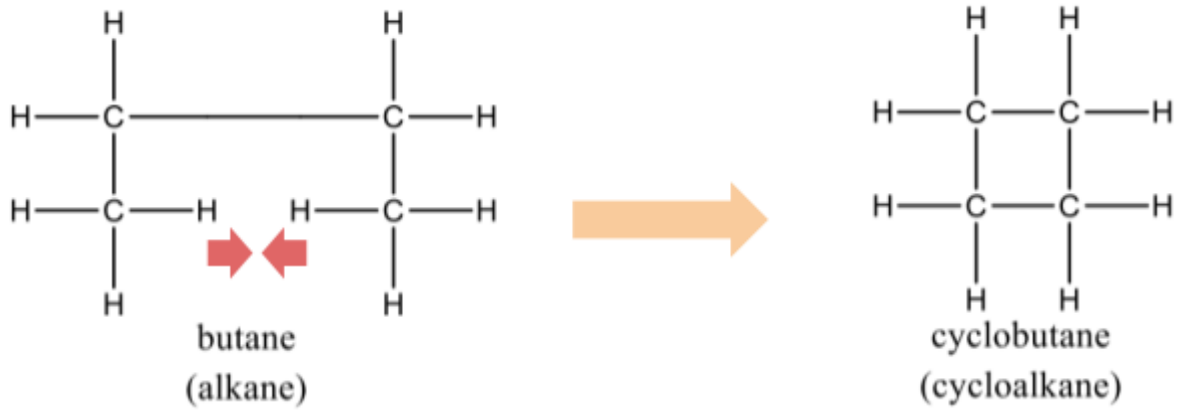


Figure 27.1.8.1 - Butane and Cyclobutane

To name cycloalkanes, we simply take the name of the corresponding alkane, and add the prefix **"cyclo-"** in front, as we see from butane and **cyclo**butane.

In fact, since a cycloalkane has the same number of carbons as the corresponding alkane, but two less hydrogens, we can find its general formula.

Since alkanes have the general formula C_nH_{2n+2}, cycloalkanes have the general formula C_nH_{2n}.

This makes a **cycloalkane** an isomer of the **alkene** with the **same number** of **carbons**!

Furthermore, to draw the **skeletal structure** of a cycloalkane, we simply draw a regular polygon with the number of vertices (or equivalently, sides) equal to the number of carbons.

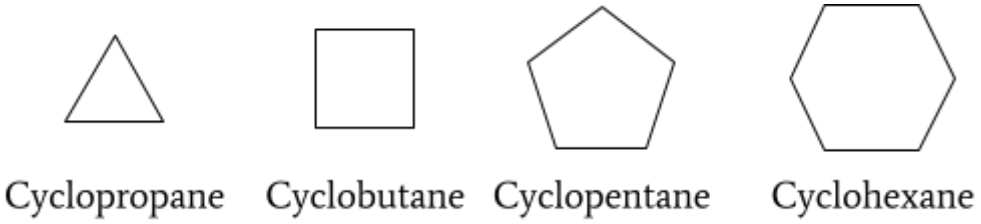
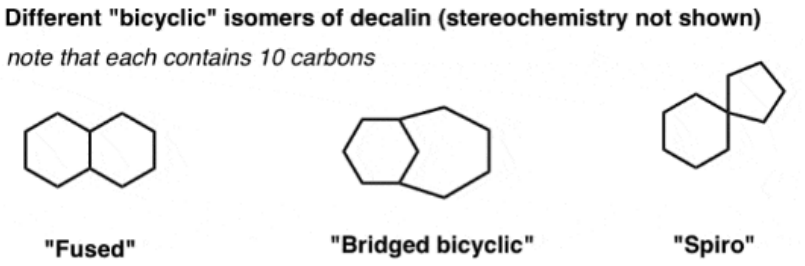


Figure 27.1.8.1 - First Four Cycloalkanes

Note that cyclomethane and cycloethane do **NOT** exist.

Extension 2.7.1.8.1: Fused Cyclic Compounds

Now, what happens if we take two rings and fuse them? We end up with 3 different possibilities of **bicycle**:



These compounds are seen in more complex organic synthesis. The IUPAC nomenclature for such compounds is as follows:

2.7.1.8.1.1 Bridged Bicycles

We start by counting the **total** number of carbon atoms in the ring system. We count the number of carbons in **each** smaller **unit ring**, excluding the bridged carbons. We then count the number of carbon atoms in the **bridge**. Then, we assign the numerical values in **descending** order:

3 carbons in "cyclohexane" segment, 4 in the "cycloheptane" segment. 1 Bridgehead carbon. Parent chain has 9 carbons so it is nonane.

Thus, we name **bicyclo-[4.3.1]-nonane**.

2.7.1.8.1.1.1 Fused Bicycles

Fused heterocycles are simply a subset of bridged heterocycles with the number of bridgehead carbons set as 0. For example, decalin (shown above) has the IUPAC name **bicyclo-[4.4]-decane**.

2.7.1.8.1.2 Spiro Bicycles

Similarly to the above, we name the compound but with perfix "spiro-". We have the above example as **spiro-[5.4]-decane**.

27.2 Alkenes and Alkynes

Recall that an alkene is just an alkane but with some $C - H$ hydrogens replaced to form $C = C$ **double bond**(s). The double bond doesn't have to be at the ends of the molecule, and there may be more than one double bond.

Here is an example of an alkene, **hex-3-ene**, or **3-hexene**:

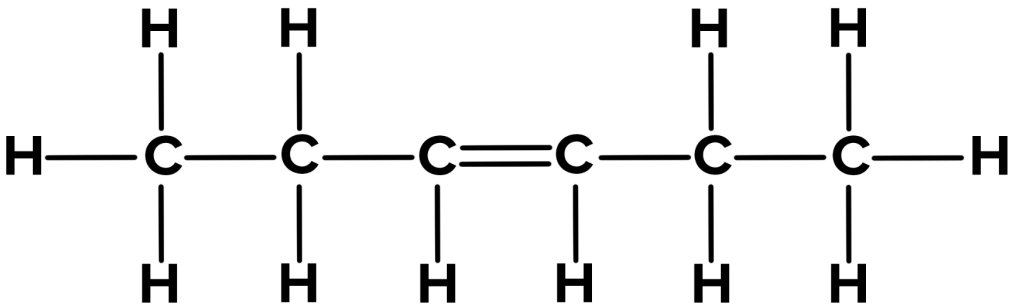


Figure 27.2.1 - Hex-3-ene

Furthermore, an alkyne is an alkane but with some $C - H$ hydrogens replaced to form $C \equiv C$ **triple bond**(s) .The triple bond doesn't have to be at the ends of the molecule, and there may be more than one triple bond.

This is an example of an alkyne, **hex-3-yne** or **3-hexyne**:

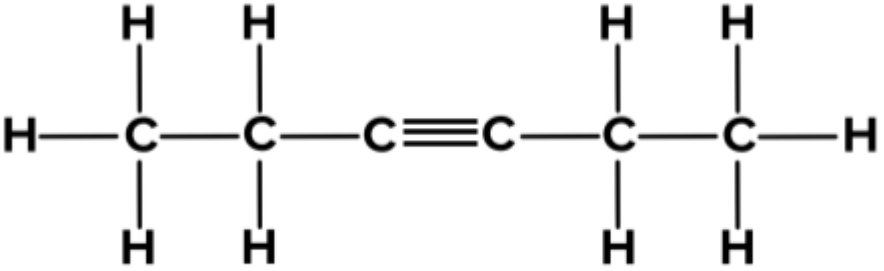


Figure 27.2.2 - Hex-3-yne

Let's see how these names are derived.

27.2.1 Naming Simple Alkenes and Alkynes

We first identify the **longest carbon chain** that contains the **double/triple bond**(s), and count the number of carbons. We use the same prefix as we did for alkanes, and attach "-ene" for alkenes and "-yne" for alkynes.

Number of Carbons	Prefix	Alk**ane** (-ane)	Alk**ene** (-ene)	Alk**yne** (-yne)
1	meth-	meth**ane**	-	-
2	eth-	eth**ane**	eth**ene**	eth**yne**
3	prop-	prop**ane**	prop**ene**	prop**yne**
4	but-	but**ane**	but**ene**	but**yne**
5	pent-	pent**ane**	pent**ene**	pent**yne**
6	hex-	hex**ane**	hex**ene**	hex**yne**
7	hep-	hept**ane**	hept**ene**	hept**yne**
8	oct-	oct**ane**	oct**ene**	oct**yne**
9	non-	non**ane**	non**ene**	non**yne**
10	dec-	dec**ane**	dec**ene**	dec**yne**

Table 27.2.1.1 - Prefixes, Alkane, Alkene and Alkyne Names

If there is only one double/triple bond, number the carbon chains from the end **nearest to the double or triple bond**.

If the bond goes from carbon n to carbon $(n + 1)$, we say its position is n.

Now, we add this position either as a **prefix** (3-hexene) or right **before -ene/-yne** (hex-3-ene). Both are acceptable naming conventions.

For alkenes and alkynes with **up to 3 carbons** (ethene, ethyne, propene, propyne), we **don't actually need to** state the **position** of the double/triple bond, because there is only one possible position (for propene/propyne, we simply number the carbons the other way).

FYI: If there is more than one double/triple bond, number the carbons such that the first difference in substituent positions is minimised.

27.2.2 Skeletal Structures of Alkenes and Alkynes

To draw the skeletal structure of an alkene or alkyne, we simply add a line to denote a double bond, or add two lines to denote a triple bond. For example, this is **2-pentene** or **pent-2-ene**:

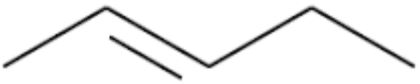


Figure 27.2.2.1 - Pent-2-ene

For a triple bond, however, we no longer create a bend and continue straight, to more accurately depict its linear geometry. This is **2-pentyne** or **pent-2-yne:**

Figure 27.2.2.2 - Pent-2-yne

27.2.3 Cycloalkenes

For a simple cycloalkene, we can just take the **alkene name**, and add the prefix “cyclo-”.

For example, this is **1-hexene** or **hex-1-ene**:

Figure 27.2.3.1 - Hex-1-ene

We can imagine curling this structure up into a ball and adding one line (the extra covalent bond that forms when a cyclic compound is formed) as in *Figure 27.2.3.2* below.

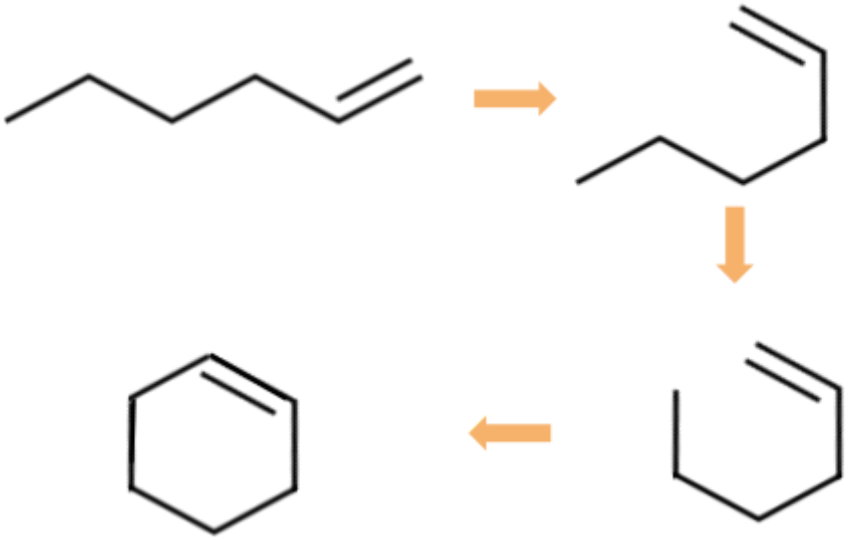


Figure 27.2.3.2 - Curling Alkene into a Ball! (NOT a real synthesis)

This gives us the skeletal structure of **cyclohexene**:

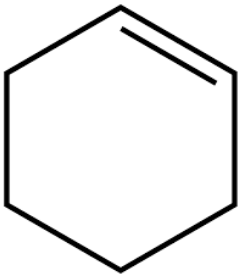


Figure 27.2.3.3 - Cyclohexene

Due to their rotational symmetries, we need not emphasise the position of the bond for cycloalkenes with one double bond.

Extension 27.2.3.1 - Cycloalkynes (FYI)

There is little evidence for the existence of smaller cycloalkynes like cyclobutyne.

Meanwhile, cyclopropyne, and cyclopentyne, cyclohexyne and cycloheptyne are too unstable to be isolated and only exist as intermediates to a reaction (Adolf *et al.* 1983).

On the other hand, cyclooctyne and cyclononyne have been isolated, but are expectedly very reactive (Adolf *et al.* 1983).

This is due to something called **angle strain**. In intuitive terms, VSEPR Theory tells us that $C \equiv C$ triple bonds have a **linear geometry**, and hence the $C - C \equiv C$ bond angle in a cycloalkyne actually wants to be as close to 180° as possible.

For smaller cycloalkynes, this is simply not possible due to the sum of interior angles of polygons being fixed in Euclidean Geometry. That is, a triangle will have smaller interior angles on average than a polygon with more sides.

These conflicting factors cause **angle strain** which build up within the cycloalkyne to form **ring strain**, which causes the ring to be unstable and hence causes smaller cycloalkynes to not exist, or to be very unstable.

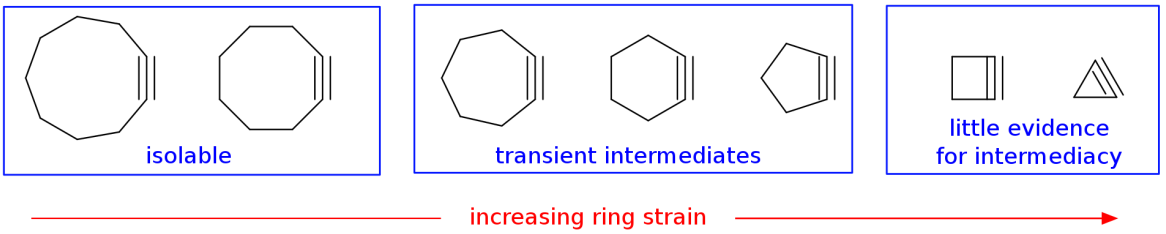


Figure 27.2.3.1.1 - Cycloalkynes in Theory

In fact, from *Figure 27.2.3.1.1* we can see very clearly the smaller bond angles for the smaller cycloalkynes, which causes the angle strain.

27.2.4 Physical Properties of Alkenes and Alkynes

Much like alkanes, alkenes and alkynes are **non-polar** and hence are soluble in non-polar solvents and insoluble in water in general.

They have a **low density**, **low boiling point** and **low melting point**.

Furthermore, many alkenes with **multiple double bonds** have **colours** associated with them.

27.2.5 Combustion of Alkenes and Alkynes

Recall that combustion of a hydrocarbon is the reaction of a hydrocarbon with oxygen in the presence of heat to form CO_2 and H_2O **only**.

However, since **alkenes** and **alkynes** have a higher Carbon : Hydrogen Ratio, most of the time the combustion would be **incomplete** due to a lack of oxygen. This means that the combustion of alkenes and alkynes will produce more **carbon monoxide** and **soot** as byproducts.

Pause and Ponder 27.2.5

Derive the general equation of the combustion of a hydrocarbon C_xH_y.

[Click to Go to Solution]

27.2.6 Addition Reactions

An addition reaction is one where a $C = C$ double bond (or $C \equiv C$ triple bond) is broken and atoms are added to form an alkane with these extra atoms.

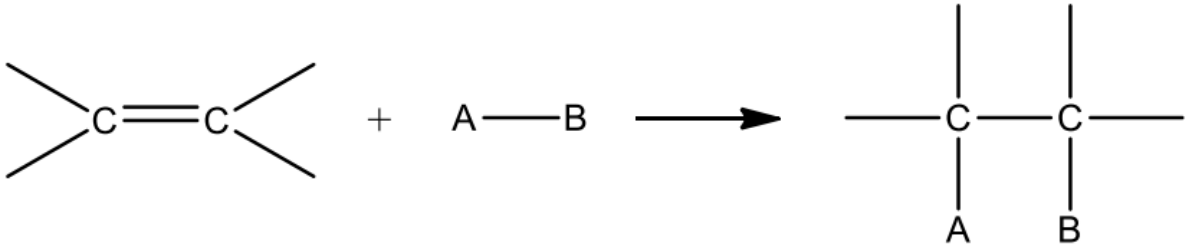


Figure 27.2.6.1 - Addition Reaction of Alkene

Certain Addition Reactions require the presence of a **catalyst**. Here are two main types of Addition Reactions: **Hydrogenation** and **Halogenation**. There is also the **Hydration of Alkene** reaction (see **Section 29.1.5: Hydration of Alkene and Dehydration of Alcohol**).

27.2.6.1 Hydrogenation of Alkene

Hydrogenation of an **alkene** involves the addition of **Hydrogen**, H_2 to an alkene, with the presence of high temperatures, high pressure and a **palladium, platinum** or **nickel** catalyst, to form an alkane.

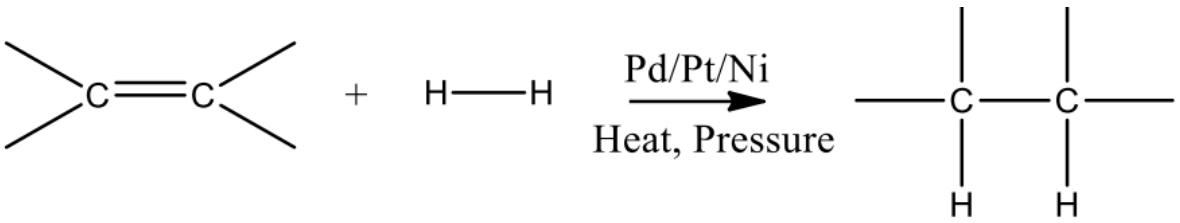


Figure 27.2.6.1.1 - Hydrogenation of Alkene

In fact, this reaction is **redox**, and can also be called **reduction of alkene**. Consider an example like the hydrogenation of ethene:

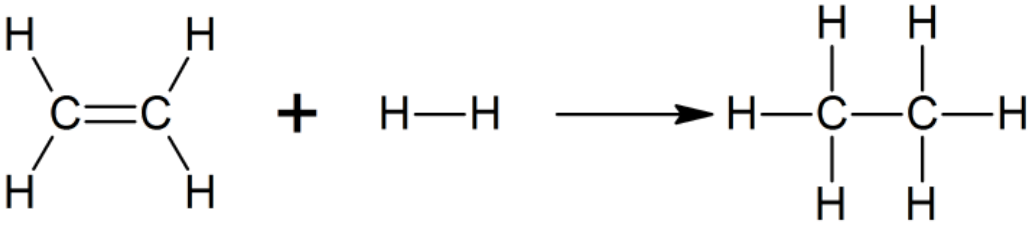


Figure 27.2.6.1.2 - Hydrogenation of Ethene

This is because both Reduction and Oxidation occur **(below is FYI)**:

- The oxidation number of H increases from 0 in H_2 to +1 in Ethane (Oxidation)
- The oxidation number of C decreases from -2 in Ethene to -3 in Ethane (Reduction)

Note: The assigning of oxidation numbers in large organic compounds is **NOT** tested.

Extension 27.2.6.1.1 - Can Benzene Be Reduced?
The short answer is yes. This also introduces the selectivity of the inert metal catalyst: When do I use Palladium-Charcoal (Pd/C), Adam's Catalyst (PtO_2) or Raney Nickel (RaNi).

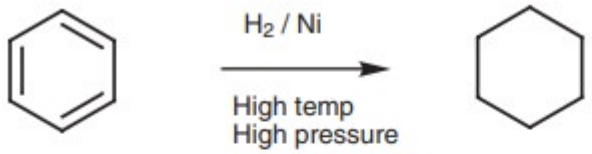


Fig. 2. Reduction of benzene to cyclohexane.

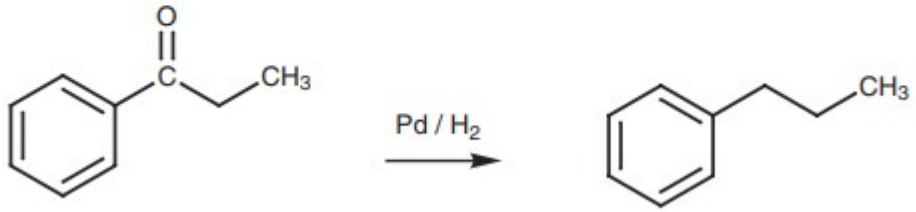
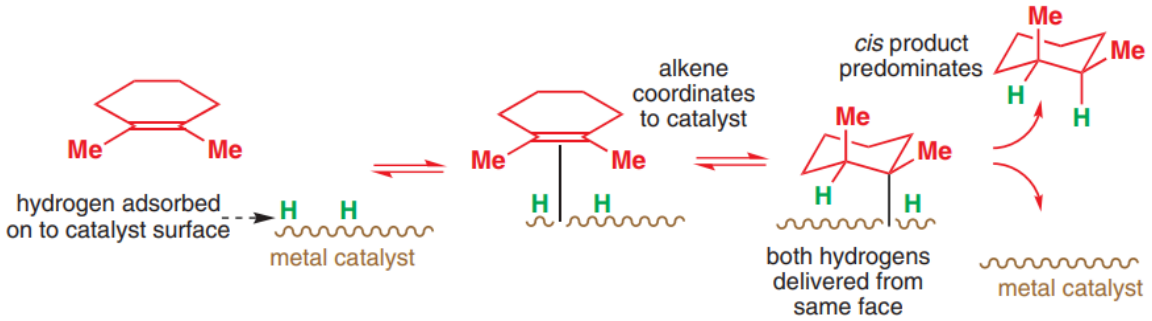


Fig. 3. Reduction of an aromatic ketone.

Benzene may be reduced selectively by Raney Nickel or PtO_2.



Palladium on Charcoal is best suited to reduce alkenes *syn*.

27.2.6.2 Halogenation of Alkene

Halogenation of **alkene** involves the addition of a halogen to an alkene to form a **dihaloalkane**.

Here, F_2 is too reactive and is unsafe in this reaction, while I_2 is too unreactive and this reaction does not take place. Hence, usually, the halogenation of an alkene refer to the addition of Cl_2 (**chlorination**) or Br_2 (**bromination**) to an alkene.

In either case, the reaction is:

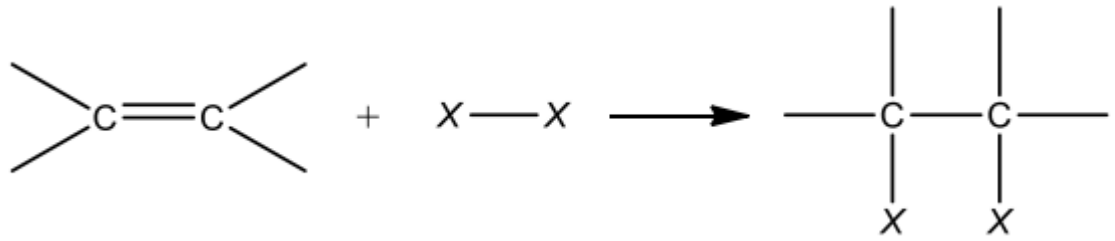


Figure 27.2.6.2.1 - Halogenation of Alkene

where X denotes either Cl or Br. No catalyst is required in this reaction. The product proceeds through a brominium, chlorinium or iodinium intermediate (CM5131)

For example, *Figure 27.2.6.2.2* below shows the **chlorination** of **ethene**.

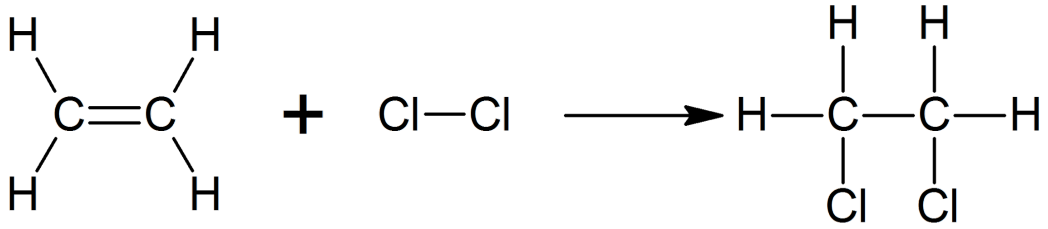


Figure 27.2.6.2.2 - Chlorination of Ethene

This reaction is also **redox**, and can be called **oxidation of alkene**, because **(below is FYI)**:

- Cl is **reduced** from oxidation number 0 in Cl_2 to -1 in 1,2-dichloroethane
- C is **oxidised** from oxidation number -2 in Ethene to -1 in 1,2-dichloroethane

Note: The assigning of oxidation numbers in large organic compounds is **NOT** tested.

Applications of halogenation, such as **bromination** can be seen in **Section 27.4: Test to Differentiate Alkane and Alkene** below.

27.2.6.3 Hydrogenation of Alkyne

Hydrogenation of **alkyne** involves the addition of 2 H_2 to obtain an **alkane**, under high temperatures, high pressures, and in the presence of a **Palladium**, **Platinum** or **Nickel** catalyst.

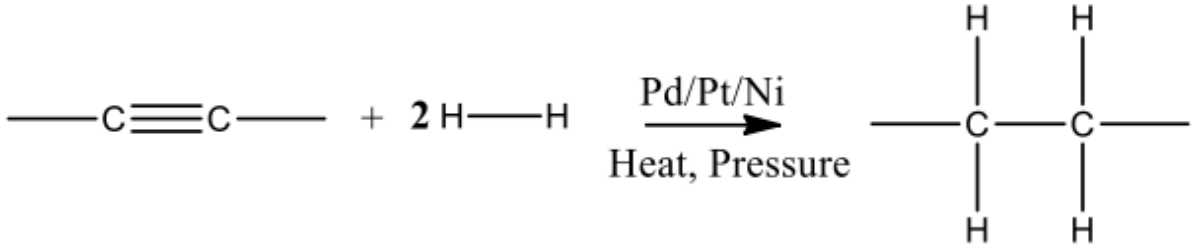


Figure 27.2.6.3.1 - Hydrogenation of Alkyne

For example, for **2-butyne**, *Figure 27.2.6.3.2* below is its hydrogenation reaction:

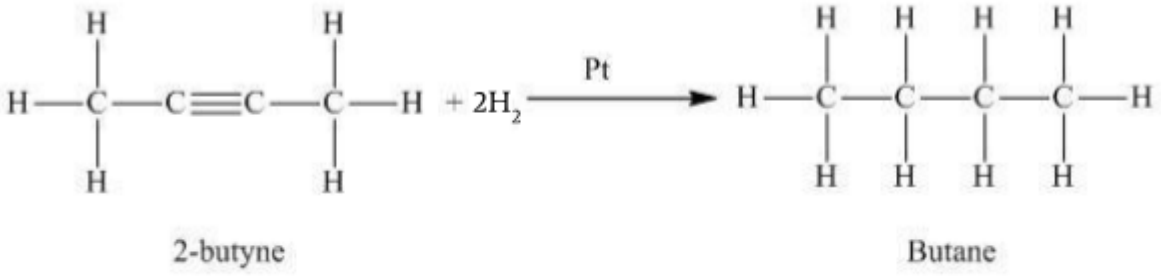


Figure 27.2.6.3.2 - Hydrogenation of 2-Butyne

Similar to **Hydrogenation** of **alkene**, this reaction is also **redox**, where the **alkyne** is reduced.

27.2.6.4 Halogenation of Alkyne

Halogenation of an **alkyne** involves the addition of two halogen molecules to form a **tetrahaloalkane**, in the following reaction:

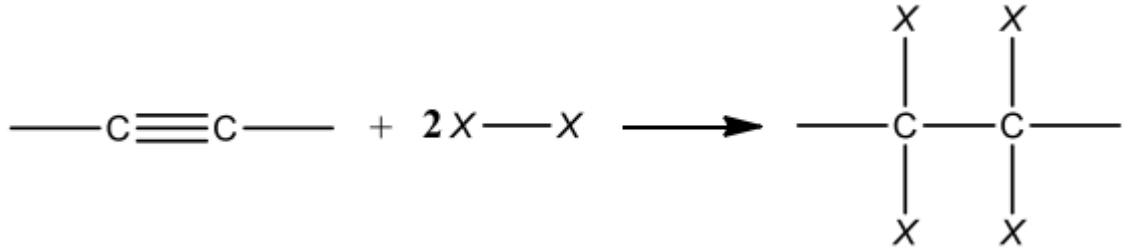


Figure 27.2.6.4.1 - Halogenation of Alkyne

Just like the **halogenation** of **alkene**, X should be Cl or Br because F_2 is too reactive, while I_2 is too unreactive.

For example, for **2-butyne**, *Figure 27.6.4.2* below is its chlorination reaction:

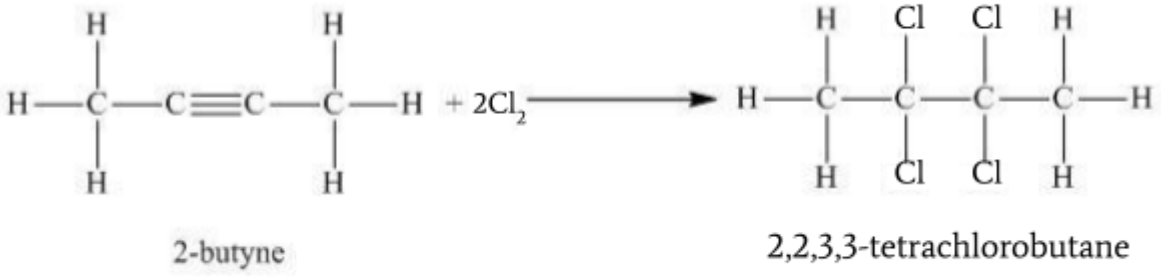


Figure 27.2.6.4.2 - Halogenation of 2-Butyne

This reaction is also **redox**, and where the **alkyne** is **oxidised**.

Extension 27.2.6.5 - Hydrogenation of Alkynes (FYI)

In fact, reactions like the hydrogenation of alkynes will have different products depending on the catalysts used.

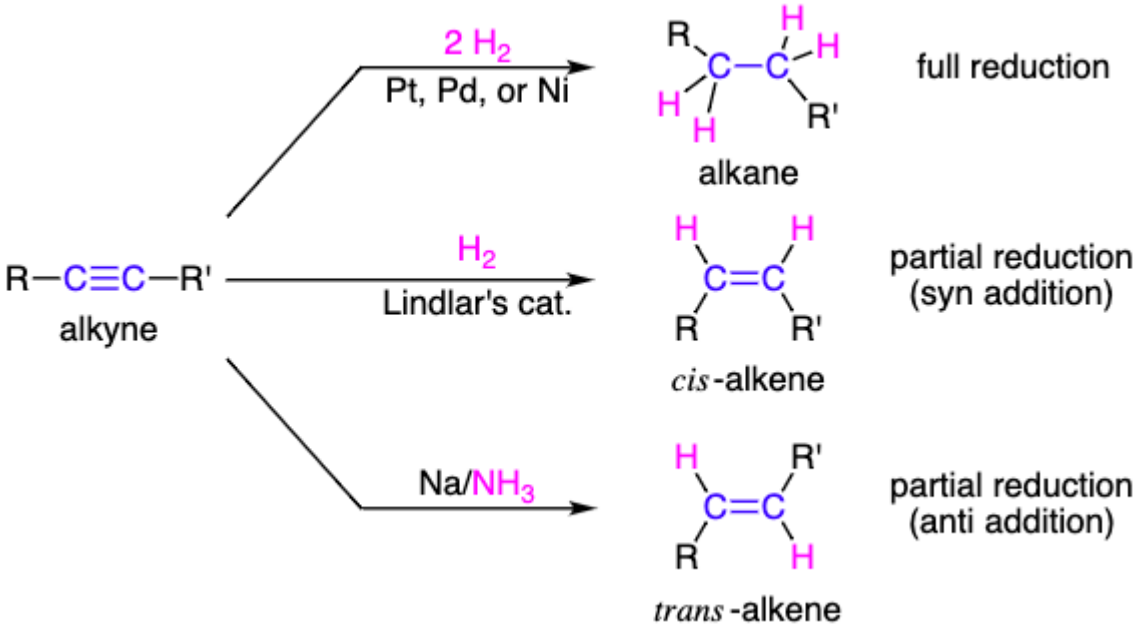


Figure 27.2.6.5.1 - Reaction Pathways

In the first reaction, a **full reduction** of the alkyne occurs, where an alkene is not isolated but is only present as an intermediate reactant, because of how effective $Pt/Pd/Ni$ is at catalysing this reaction.

In the second reaction, a **partial reduction** of the alkyne occurs and a **cis-alkene** is formed.

Here, **cis** means that the two functional groups (the alkyl groups R and R') present are on **the same** side of the molecule.

The Lindlar's Catalyst is essentially made by "poisoning", or reducing the efficiency of a Palladium catalyst, and is a mixture of three substances, as in *Figure 27.2.6.5.2* below.

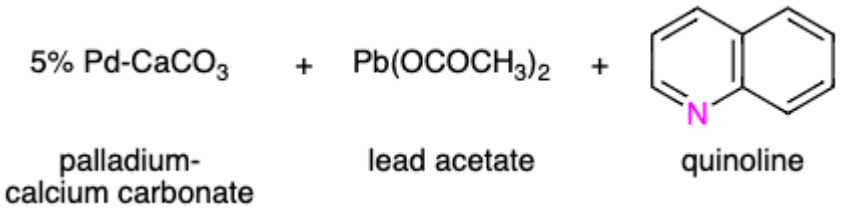


Figure 27.2.6.5.2 - Lindlar's Catalyst

In the last reaction, another **partial reduction** takes place and a **trans-alkene** is formed.

The **trans** means that the functional groups present (the alkyl groups R and R') are on **opposite** sides of the molecule.

Note that Na/NH_3 here means the addition of both Na and NH_3. Here, the Na serves as a oxidising agent due to its low 1st Ionisation Energy and donates its valence $3s^1$ electron to the alkyne to form Na^+. NH_3 also donates a proton (H^+) to the alkyne and forms NH_2^-. This method is known in industry as the Birch Reduction.

The full reaction scheme is as such:

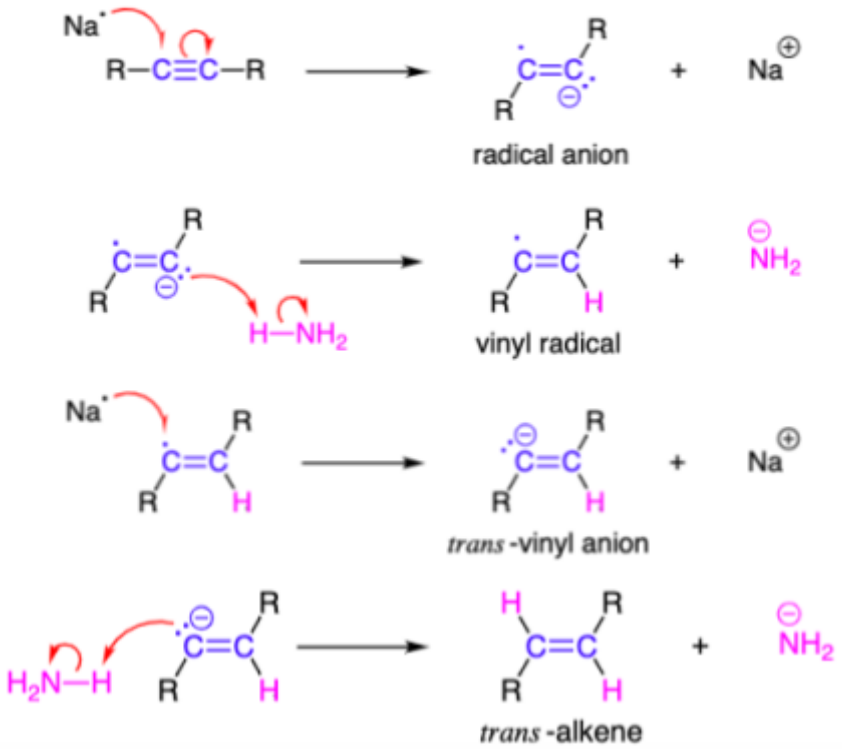


Figure 27.2.6.5.3 - Anti-Addition Pathway

27.3 Cracking Process

Cracking is a process where a **big alkane molecule** is broken down into a **smaller alkane** and a **smaller alkene**, either through high temperatures and pressures, or with a catalyst.

For example,

$$\underset{\text{(octadecane)}}{C_{18}H_{38}} \rightarrow \underset{\text{(octane)}}{C_8H_{18}} + \underset{\text{(decene)}}{C_{10}H_{20}}$$

Octadecane and Octane are **alkanes**, while Decene is an **alkene**.

Cracking is a very important industrial process because it breaks down long alkanes, which are the major components of **crude oil**, into shorter alkanes which are higher in demand, and producing alkenes at the same time.

27.4 Test to Differentiate Alkane and Alkene

The **Bromination** of **Alkene** reaction can be used to differentiate between **Alkanes** and **Alkenes**. Suppose we had two unlabelled beakers, but we know one has ethane and the other has ethene.

We can add **liquid bromine** (i.e. bromine dissolved in an **organic solvent)**, which is reddish-brown in colour, into each of the beakers.

- The bromine in the beaker with **ethene** will **change** from **reddish-brown** to **colourless**.
- The bromine in the beaker with **ethane** will remain **reddish-brown**.

This is because **Halogenation of Alkene** does not require any catalyst, while **Halogenation of Alkane** requires UV light. Hence, bromine only reacted with the alkene, which is **ethene** in this case. The resulting product is a **dihaloalkane**, 1,2-dibromoethane, which is colourless.

This test can also test for **Saturation of Hydrocarbons**.
Recall that saturation simply means **absence of double** $C = C$ **and triple** $C \equiv C$ **bonds**.

When we add bacon, which has unsaturated fats, to a flask with liquid bromine, we will observe the gradual change of colour of the liquid in the flask from **reddish-brown** to **colourless**, as the unsaturated fats undergo **halogenation**.

27.5 Addition Polymerisation

A **polymer** is a large molecule made of many **smaller repeated units**, called **monomers**.

Figure 27.5.1 below shows how an alkane can undergo addition polymerisation.

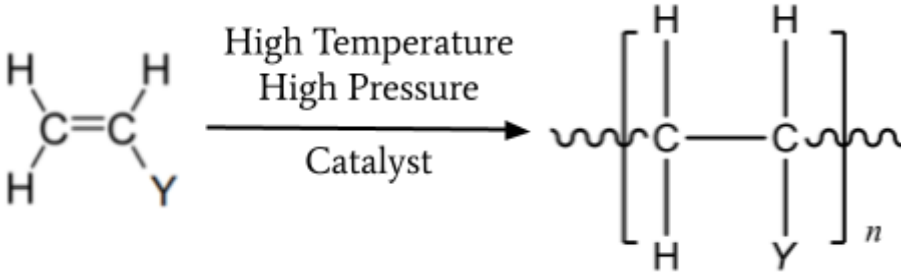


Figure 27.5.1 - Addition Polymerisation

Here, the optional wavy lines and the brackets with subscript n simply denotes the repeating structure, where n can be a number in the thousands. It reflects the structure:

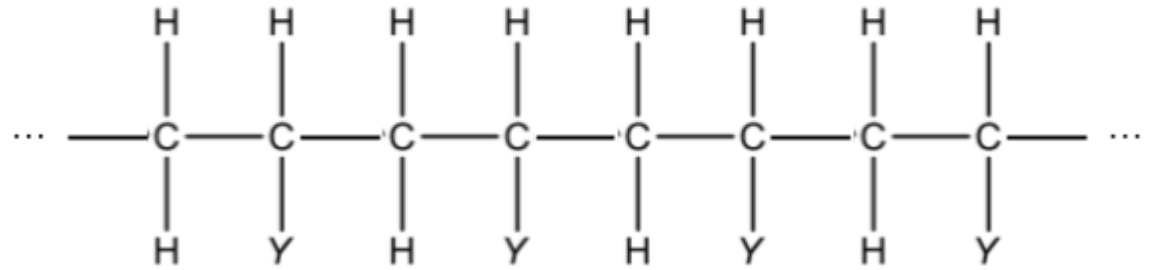


Figure 27.5.2 - Polymer

These resulting polymers are often **inert** (i.e. very unreactive) to most substances.

For example, this shows ethene forming an addition polymer **polyethylene**.

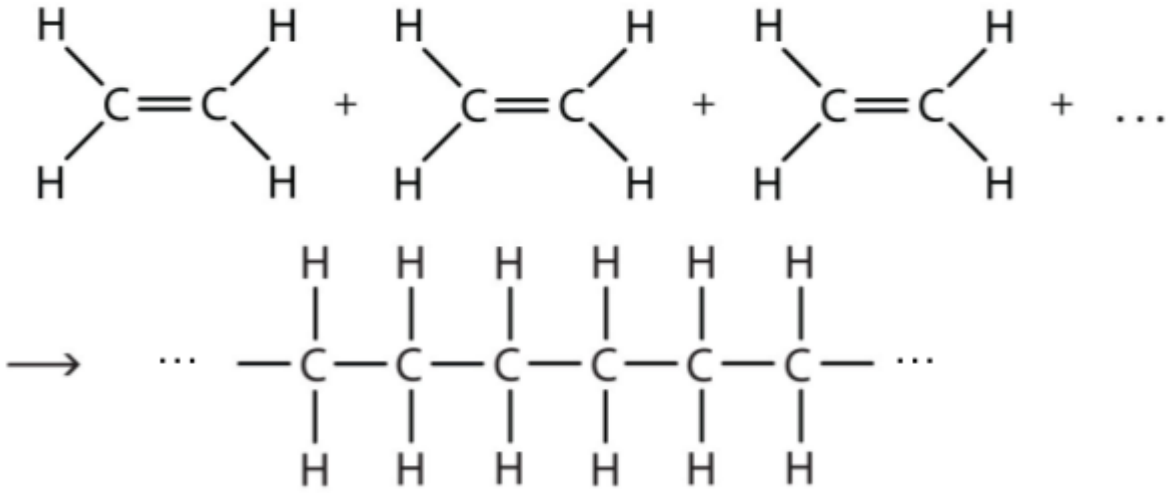


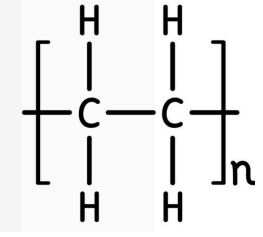
Figure 27.5.3 - Ethene Addition Polymerisation to form Polyethylene

Ethylene is actually the name of the $-CH_2-CH_2-$ group, so polyethylene is derived from this.

The actual **IUPAC name** for **polyethylene,** however, is **polyethene**.

27.5.1 Common Polymers

Table 27.5.1.1 below lists some common polymers.

Name of Polymer	Examples of Uses	Structure
Common Name: Polyethylene **IUPAC Name:** Poly(ethene)	Packaging Bottles Plastic Bags	

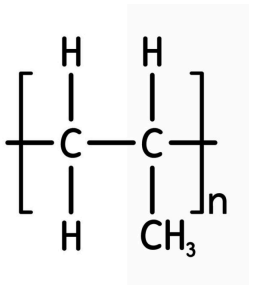
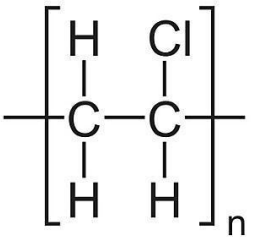
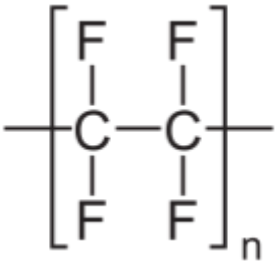
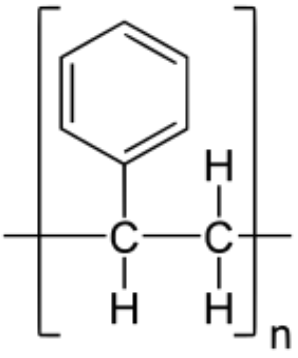
Common Name: Polypropylene **IUPAC Name:** Poly(propene)	Pails Medical Tubings	
Common Name: Polyvinyl Chloride (PVC) **IUPAC Name:** Poly(chloroethene)	Pipes Shower Curtains Credit Cards	
Common Name: Teflon **IUPAC Name:** Poly(1,1,2,2-tetrafluoroethene)	Non-stick Coating in Cooking Utensils Fire-resistant Fabrics	
Common Name: Polystyrene **IUPAC Name:** Poly(1-phenylethene)	Styrofoam Plates Styrofoam Cups	

Table 27.5.1.1 - Common Polymers

27.6 Aromatic Compounds

Aromatic compounds substances are known to have a **pleasant smell**, but they are defined by other important common characteristics - they are very **chemically stable** and they are **cyclic**.

FYI: Hückel's rule predicts that a molecule is aromatic if it has $(4n + 2)$ π electrons, where n is a non-negative integer, is planar, cyclic, and has a continuous ring of p orbitals.

Non-aromatic compounds are said to be **aliphatic** (fat-like).

27.6.1 Conjugated Double Bonds and Stability

Conjugated double bonds are **alternating double bonds** (double bonds linked together by a single bond), such as in **penta-1,3-diene** below. **Conjugated double bonds** increase the stability of a molecule.

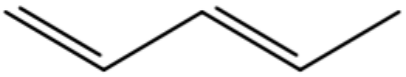


Figure 27.6.1.1 - Penta-1,3-diene

Note that this molecule is called “penta-1,3-diene” because we number the carbons as such:

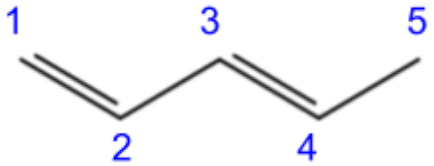


Figure 27.6.1.2 - Numbering Carbons from Left to Right

And among two double bonds we take the two smaller numbers, 1 and 3. Since there are two double bonds, we add the prefix di- before -ene. This gives us Penta-1,3-diene.

On the other hand, consider **penta-1,4-diene**.

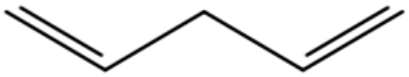


Figure 27.6.1.3 - Penta-1,4-diene

Here, the double bonds are not **conjugated**.

27.6.2 Conjugation in Rings

Conjugation can also appear in rings. Consider the mystery 6-carbon alkene molecule in *Figure 27.6.2.1* below:

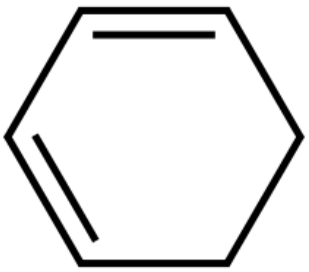


Figure 27.6.2.1 - Mystery Cyclic Molecule

IUPAC rules tell us to number the carbons such that one of the double bonds land on 1 and 2:

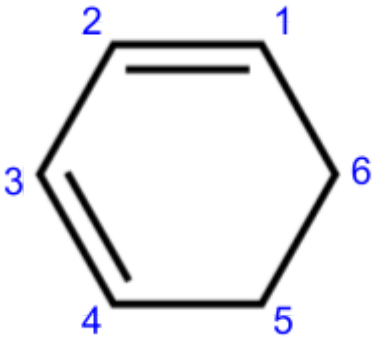


Figure 27.6.2.2 - Cyclic Compound Carbon Numbering

Now, we see that the two double bonds are on positions 1 and 3 respectively. Hence, the alkene name for this molecule would be **hexa-1,3-diene**, or **1,3-hexadiene**. But since the compound is cyclic, we add the prefix "cyclo-".

The name of this molecule is **cyclohexa-1,3-diene** or **cyclo-1,3-hexadiene**.

In this molecule, we see that the double bonds are in fact, **conjugated**. Now, what about:

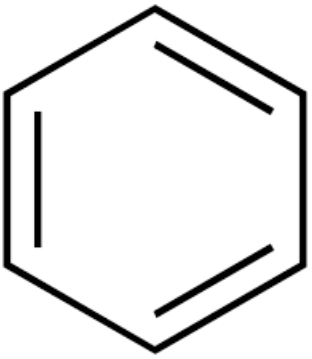


Figure 27.6.2.3 - Another Cyclic Molecule!

By similar reasoning, we might arrive at the name **cyclohexa-1,3,5-triene**. However, this is not exactly right, because this actually **benzene**, or rather, one of the **resonance structures** of **benzene**. It is unique in that it is an **aromatic compound.**

Why is "cyclohexa-1,3,5-triene" inaccurate? Since double bonds in an alkene are shorter than single bonds, we would expect something like *Figure 27.6.2.4* below (exaggerated).

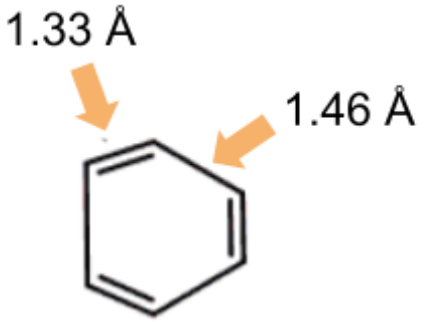


Figure 27.6.2.4 "cyclohexa-1,3,5-triene"

However, if we analyse the structure of benzene, we notice that the bond lengths are all equal at 1.39 Å. Hence, this is not an accurate representation.

Instead, the actual structure of benzene comes from two resonance structures:

Figure 27.6.3 - Benzene

For convenience, it is standard notation to represent these two structures by the following:

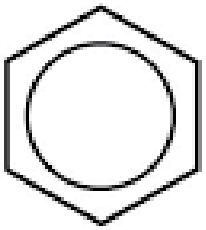


Figure 27.6.4 - Benzene

27.6.3 Benzene

Benzene is an **aromatic compound** with molecular formula C_6H_6.

Because of how all the bonds are conjugated with one another, this gives **benzene** a very high stability. As such, it loses some of the properties of alkenes.

Benzene can only go through **aromatic substitution reactions** in the presence of a catalyst where the reaction $C_6H_6 + Br_2 \rightarrow C_6H_5Br + HBr$ can take place. (See **Section 27.6.6.1: Halogenation of Benzene**)

However, it does **NOT** go through **addition reactions**, and something like $C_6H_6 + Br_2$ will not have any visible reaction, and the bromine remains reddish-brown.

This suggested that **benzene** is not like other **alkenes** and "cyclohexa-1,3,5-triene" cannot be used to describe it. It is a **fundamental structural unit**.

27.6.4 Naming Benzene-Containing Compounds

Monosubstituted containing benzene simply take on "-benzene" as the suffix.

For example, consider *Figure 27.6.4.1* below. It has a single chlorine added onto a benzene ring.

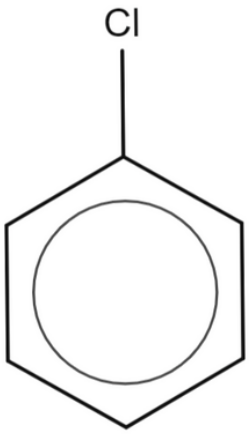


Figure 27.6.4.1 - Chlorobenzene

Hence, this substance would be called **chlorobenzene**.

However, for certain monosubstituted benzene containing compounds, there are special names assigned. For example, *Figure 27.6.4.2* below is **methylbenzene**, but together they form a new functional group by the name of **toluene**. Hence, both names are accepted.

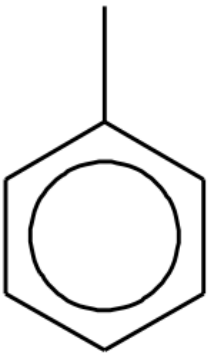


Figure 27.6.4.2 - Toluene

Some other examples include **(FYI** for now**)**:

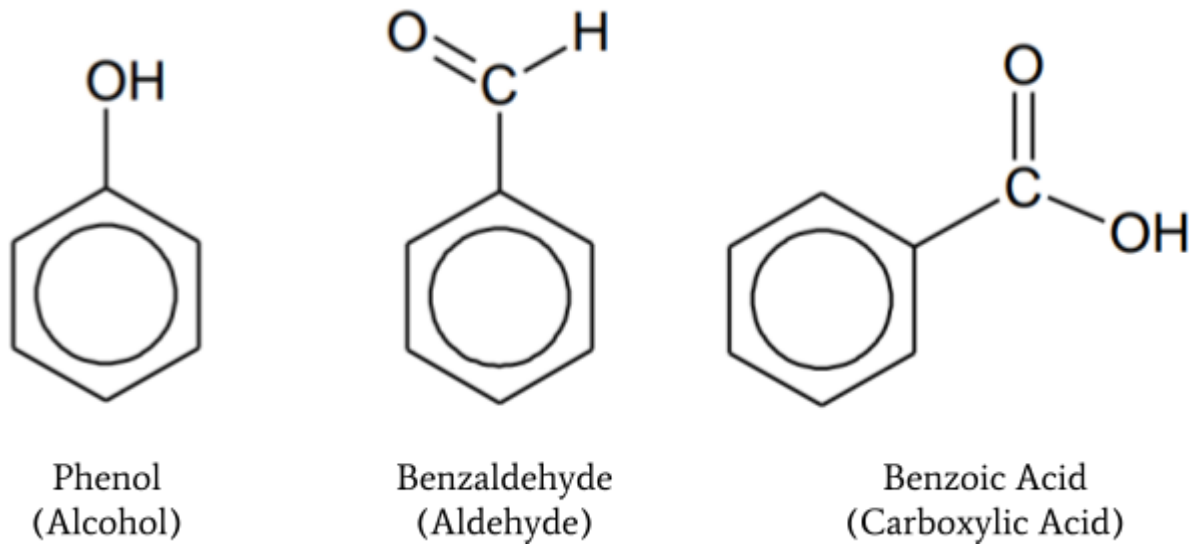


Figure 27.6.4.3 - Other Benzene-Containing Functional Groups

For alkyl groups attached to a benzene, we count the length of the alkyl.

If the alkyl has less than 6 carbons, we treat the **parent chain** as the **benzene** and hence name the compound by attaching "-benzene".

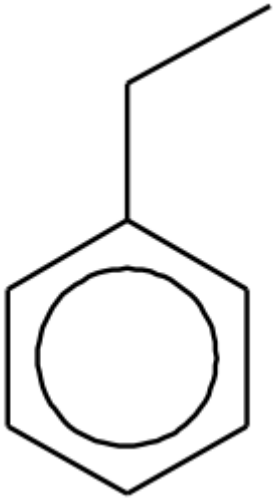


Figure 27.6.4.4 - Ethylbenzene

For example, in *Figure 27.6.4.4* above, since the alkyl has only 2 carbons, this compound is named with the suffix "-benzene". It is **ethylbenzene**.

If the alkyl has 6 or more carbons, we treat the **parent chain** as the **alkane** and hence name the compound by writing "phenyl-" as a substituent.

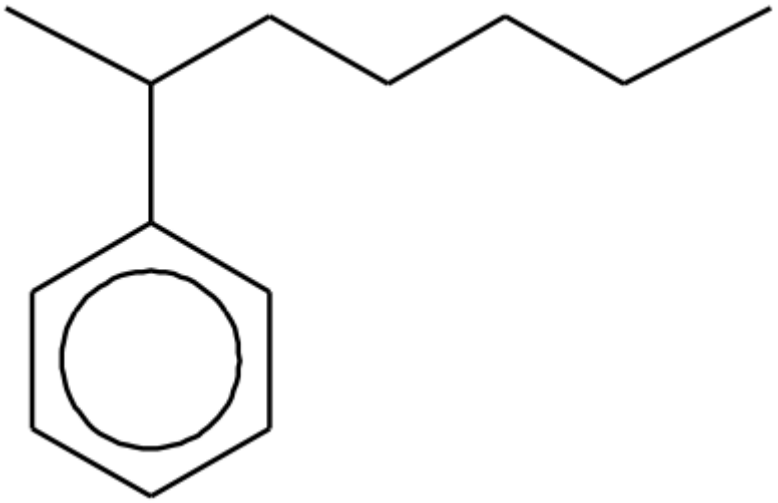


Figure 27.6.4.5 - 2-Phenylheptane

For example, in *Figure 27.6.4.5* above, the alkane has 7 carbons, so we name it as **heptane**, and add "phenyl-" as a prefix. The phenyl group is on the second carbon, so this is **2-phenylheptane**.

Phenyl refers to a **benzene ring** with a **missing hydrogen** which is where it can form a covalent bond with other atoms. **Phenyl** to **Benzene** is like an **Alkyl** to its **Alkane**. You might see the phenyl group depicted as such:

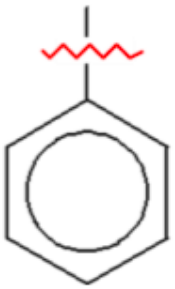


Figure 27.6.4.4 - Phenyl

Here, the red line just means that the carbon on top is **not** part of the **phenyl group**.

27.6.5 Properties of Aromatic Compounds

Aromatic compounds in general are:

- Non-polar
- Insoluble in water
- Volatile
- Flammable
- Chemically Stable

And some are **toxic** and **carcinogenic**! (oh no)

27.6.6 Aromatic Substitution

Since **benzene** is quite **inert** (i.e. chemically unreactive), it doesn't react in addition reactions, and can only be slightly modified through **aromatic substitution**. There are two main types: **Halogenation** and **Alkylation**.

27.6.6.1 Halogenation of Benzene

Halogenation of **Benzene** follows the following reaction:

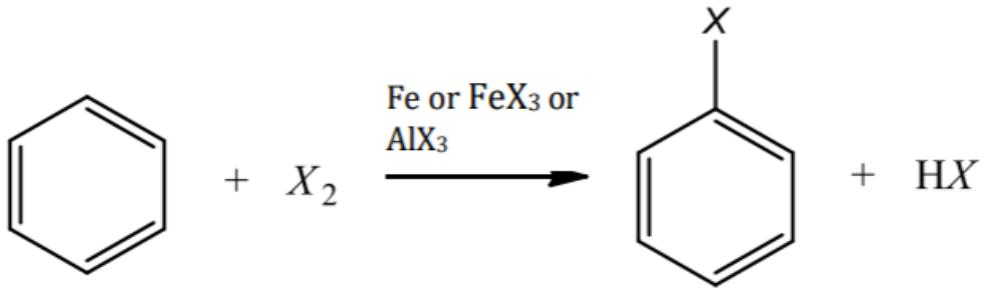


Figure 27.6.6.1.1 - Halogenation of Benzene

Like the **Halogenation of Alkene and Alkyne** reactions, X should be Cl or Br. A catalyst, is needed, namely Fe, FeX_3 or AlX_3. For example, *Figure 27.6.6.1.2* below is the reaction for **Bromination of Benzene:**

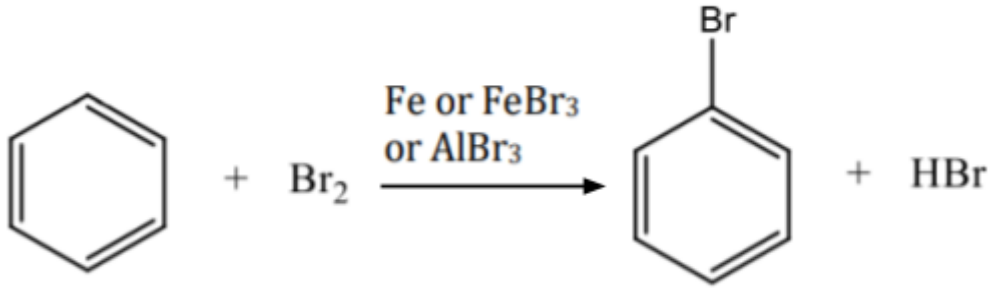


Figure 27.6.6.1.2 - Bromination of Benzene

27.6.6.2 Alkylation

Alkylation of Benzene (also called Friedel-Crafts Alkylation) follows the following reaction:

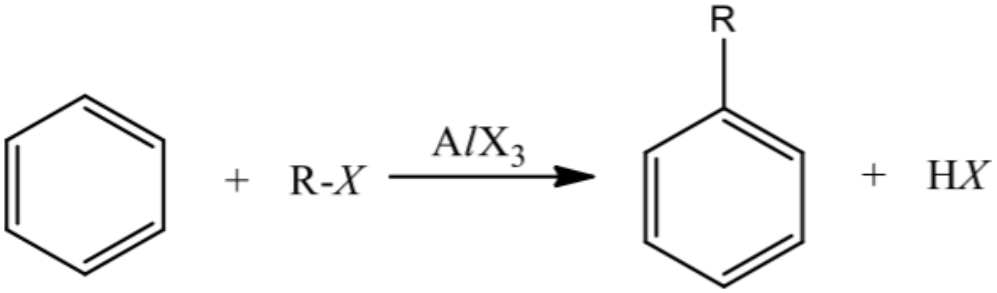


Figure 27.6.6.2.1 - Alkylation of Benzene

Here, X is almost always taken to be Cl, and R denotes an alkyl group.

27.6.7 Heterocyclic Aromatic Compounds

Heterocyclic Aromatic Compounds have **one** or **more non-carbon atoms** as part of the **aromatic ring**. This is usually **Nitrogen** but can also be other atoms. (Sulfur, Oxygen, etc.)

Pyrimidine, Purine and Pyrrole are three such compounds that contain nitrogen as shown in *Figure 27.6.7.1* below:

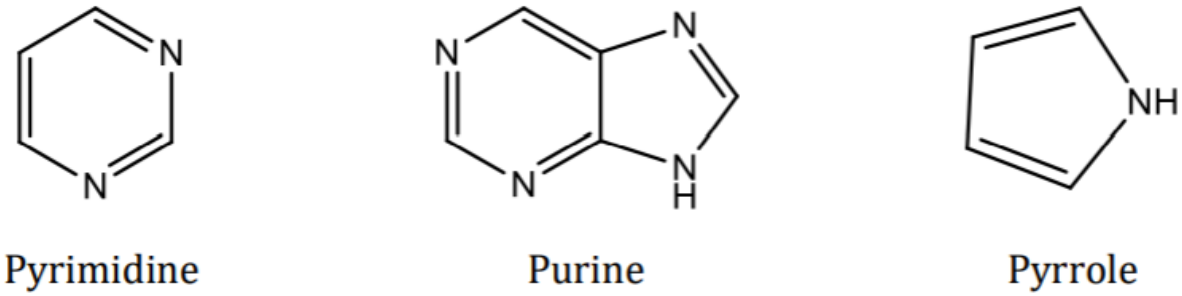


Figure 27.6.7.1 - Nitrogen-Containing Heterocyclic Aromatic Compounds

Topic 28: Air and Atmosphere

Nonexistent for now :(

Topic 29: Organic Compounds Containing Oxygen

29.1 Alcohol

(i love drinking alcohol glugglugglug)

Alcohols are molecules with a **hydroxyl** functional group ($-OH$), typically attached to an alkyl. Ethanol is one such molecule, because it consists of an alkyl group (in green), and a hydroxyl group (in red).

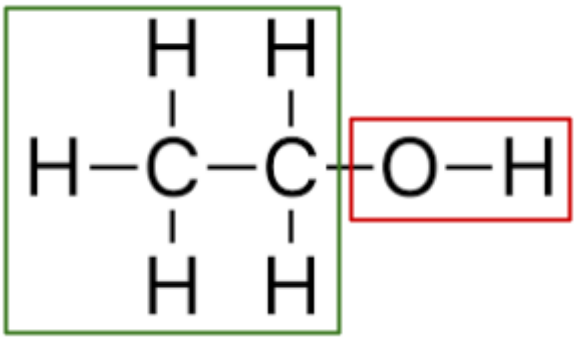


Figure 29.1.1 - Ethanol

29.1.1 Naming Alcohols

To find the IUPAC name for an alcohol, we first look at the longest carbon chain, count the number of carbons, and use the same prefix as we do for hydrocarbons (e.g. 1: meth-, 2: eth-, 3: prop- and so on), and we name the corresponding alkane.

Lastly, we replace the ending "-e" with "**-ol**" to denote alcoh**ol**, and we indicate its position either at the **very start**, or right **before** "**-ol**".

For example, consider this alcohol:

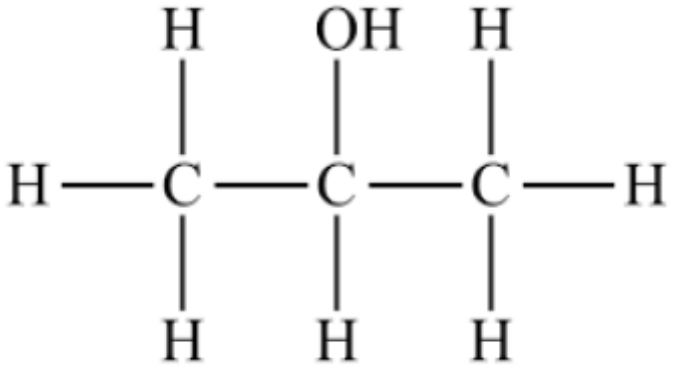


Figure 29.1.1.1: Mystery Alcohol

We first count that the longest carbon chain has 3 carbons, and hence we have the root word **"propane"**. Because it is an alcohol ($-OH$ is present), we change the "-e" to "-ol", and we have "**propanol**". Finally, because the $-OH$ group is on the second carbon, we will write 2.

Hence, the name is either **2-propanol**, or **propan-2-ol**.

FYI: Actually, this mystery alcohol in *Figure 29.1.1.1* is also known as **isopropanol**, or **isopropyl alcohol**, and is the ingredient found in **rubbing alcohol**.

If there is more than one $-OH$, we simply add the same prefix as we would add for other substituents like chloro- (e.g. 2: di-, 3: tri-, 4: tetra-, 5: penta-, and so on).

For example, *Figure 29.1.1.2* is a molecule of glycerol. What is its IUPAC name?

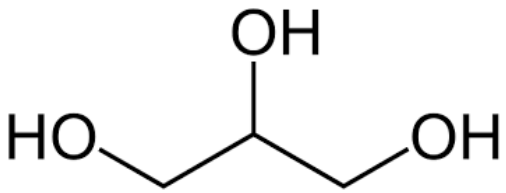


Figure 29.1.1.2 - Glycerol

Well, the main carbon chain has length 3, so the root word is **propane**. Note that there is no carbon at the two ends because there is already an oxygen present.

There are 3 $-OH$ groups present, one on each of the carbons, so we have the suffix 1,2,3-triol.

Hence, the IUPAC name of **glycerol** is **propan-1,2,3-triol** (or **1,2,3-propanetriol)**.

29.1.2 Classification of Alcohols

Alcohols can be classified by the number of **alkyl chains** connected to the carbon it bonds with.

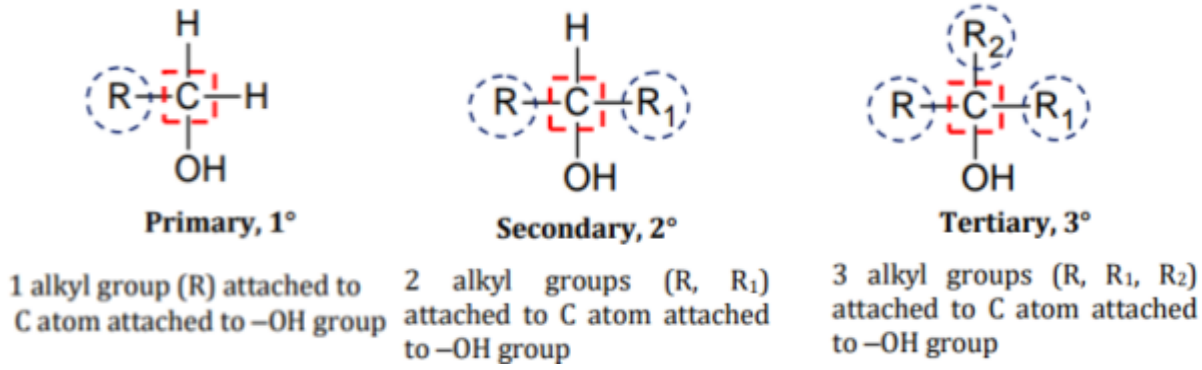


Figure 29.1.2.1 - Primary, Secondary and Tertiary Alcohols

In *Figure 29.1.2.1*, the alkyl chains are circled in blue, while the carbon the $-OH$ group bonds with is boxed in red.

29.1.3 Physical Properties of Alcohols

As the alcohols get longer, the boiling and melting points also increase, and the solubility in water decreases.

This is due to the net effect of the different **intermolecular forces of attractions**. There are two main intermolecular forces in play here:

- Hydrogen Bonding (mainly for shorter alcohols)
- London Dispersion Forces (mainly for longer alcohols)

Even though these forces don't increase and decrease at the same time, their overall effect causes the net **intermolecular forces** between alcohols to increase as the alcohols get longer.

Hence, more energy is required to overcome these forces, and the melting and boiling points increase as the alcohols get longer in general.

Alcohol	Melting Point / °C	Boiling Point / °C
1-Propanol	-124	97
1-Butanol	-89	118
1-Pentanol	-78	138
1-Hexanol	-46	157
1-Heptanol	-33	178

Table 29.1.3.1 - Melting and Boiling Points of Alcohols

Since Hydrogen Bonding is the main Intermolecular Force in shorter alcohols, and this same force is also the main Intermolecular Force in water, this allows shorter alcohols to interact with water more, and hence these shorter alcohols are **soluble** in water.

For longer alcohols, the main Intermolecular Force becomes **London Dispersion Forces**, so longer alcohols interact with water less, and hence longer alcohols are **less soluble/insoluble in water.**

Number of Carbons	1 - 4	5 - 7	≥ 8
Solubility in Water	Highly Soluble or Miscible	Moderately Soluble	Slightly Soluble or Insoluble

Table 29.1.3.2 - Solubility of Alcohols in Water

29.1.4 Oxidation of Alcohol

A **Primary Alcohol** can be **oxidised** to form an **Aldehyde**, which can then be **oxidised** again to form a **Carboxylic Acid**.

A **Secondary Alcohol** can be **oxidised** to form a **Ketone**.

A **Tertiary Alcohol** does not have any hydrogens and resists **oxidation**. It cannot be **oxidised**.

Figure 29.1.4.1 below shows the oxidation ladder of **1-butanol (primary alcohol)**, **2-pentanol (secondary alcohol)** and **2-methylpentan-2-ol (tertiary alcohol)**.

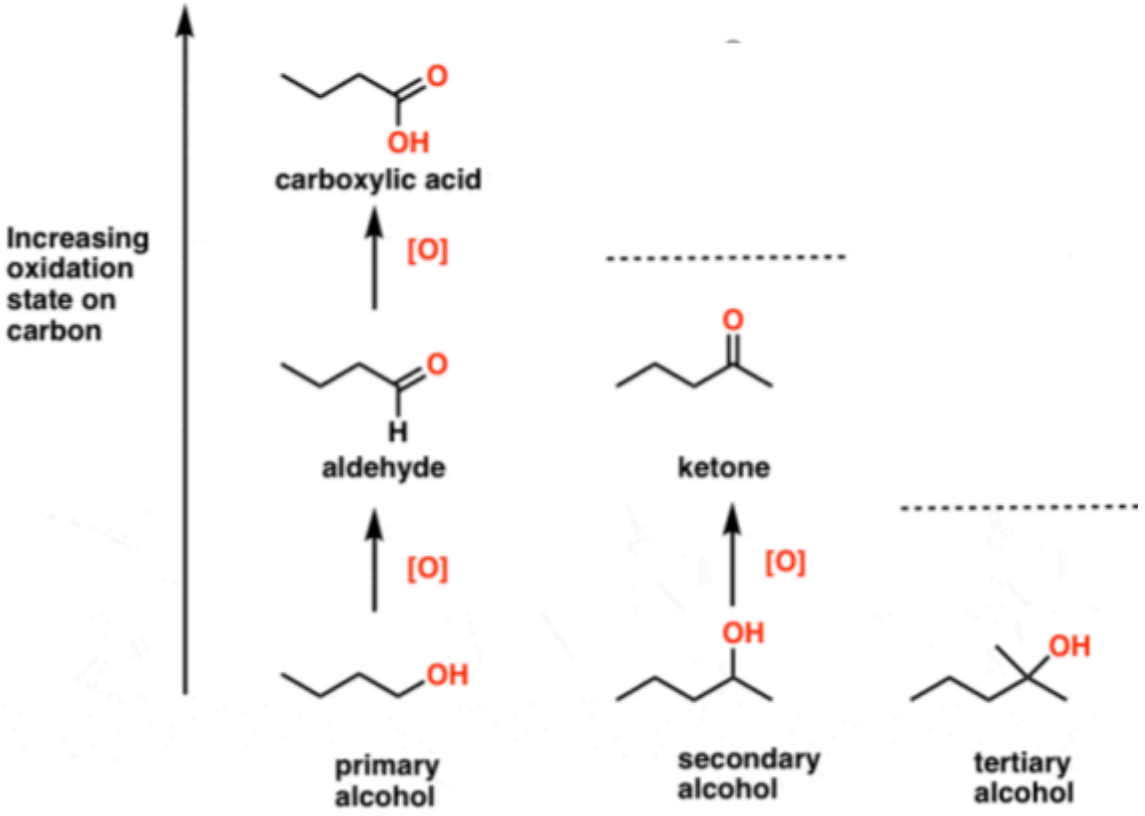


Figure 29.1.4.1 - Oxidation Ladder of Alcohols

Here [O] describes the occurrence of **oxidation**.

The compounds involved are (left to right):

- **1-butanol →1-butanal → 1-butanoic acid**
- **2-pentanol → 2-pentanone**
- **2-methylpentan-2-ol**

The naming of aldehydes and ketones are not in syllabus while the naming of carboxylic acids will be discussed later. See **Section 29.2.1: Naming of Carboxylic Acids**.

Sometimes, **strong oxidation agents** such as acidified potassium permanganate ($KMnO_4$) or acidified potassium dichromate ($K_2Cr_2O_7$) can sometimes allow the **primary alcohol** to be oxidised directly into a **carboxylic acid**.

29.1.5 Hydration of Alkene and Dehydration of Alcohol

These two reactions are actually inverses of each other. An alkene can be hydrated through an **addition reaction** called hydration.

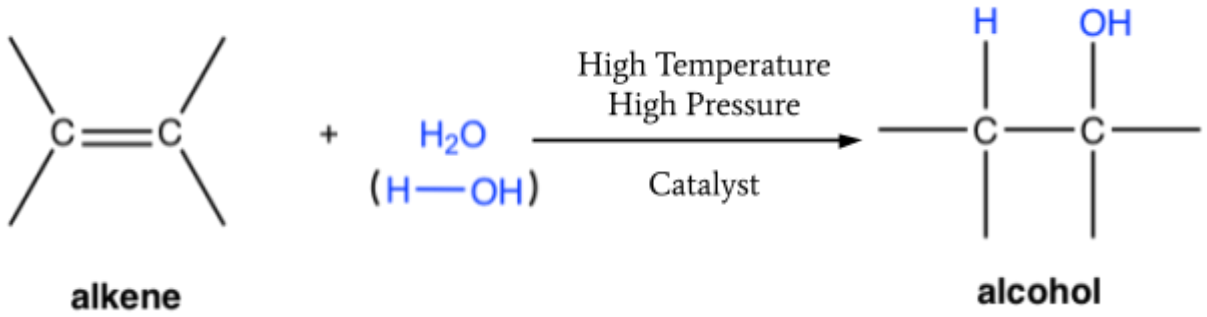


Figure 29.1.5.1 - Hydration of Alkene

For example, *Figure 29.1.5.2* below shows the **Hydration of Ethene**.

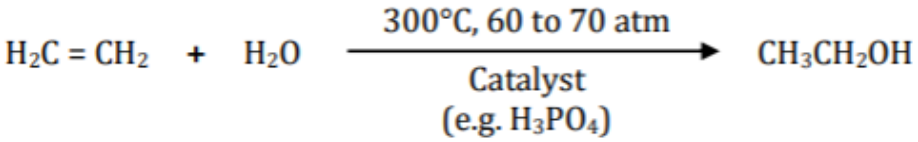


Figure 29.1.5.2 - Hydration of Ethene

Ethanol, an alcohol, is formed as the product of this reaction.

This type of reaction is reversible with heat and using excess concentrated H_2SO_4 as a catalyst.

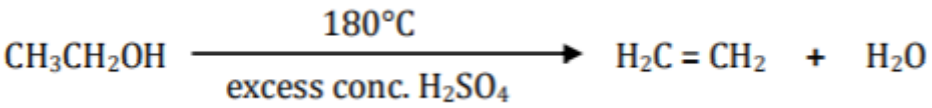


Figure 29.1.5.3 - Dehydration of Ethanol

This is called the **Dehydration of Alcohol**.

29.2 Carboxylic Acid

Carboxylic Acids contain the carboxyl ($-COOH$) functional group, which has a carbonyl group and a hydroxyl group.

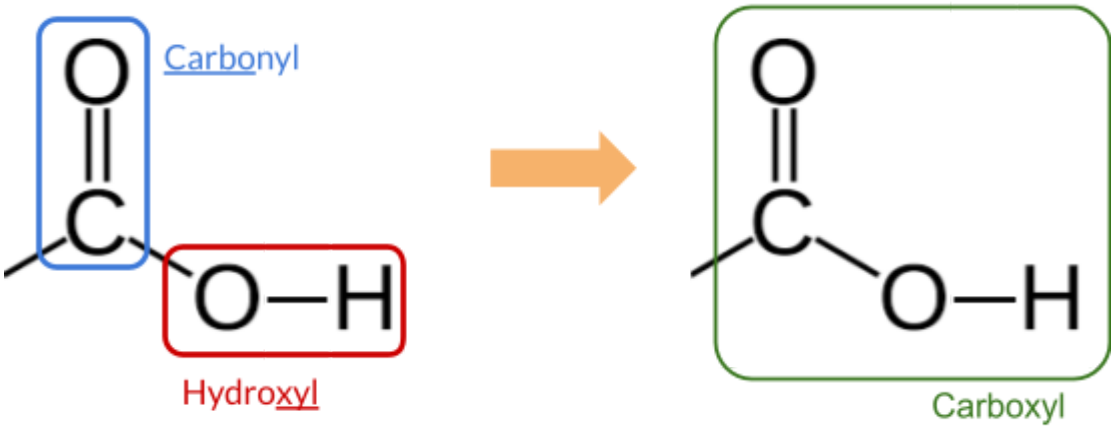


Figure 29.2.1 - Carbonyl, Hydroxyl and Carboxyl

29.2.1 Naming of Carboxylic Acids

To name a carboxylic acid, first find the longest carbon chain that contains the carboxyl group, and obtain the root word (name of alkane with the same number of carbons).

Next, replace the ending "-e" with "-oic acid" to reflect the carboxylic acid.

We then number the carbons, **starting from the carboxyl carbon as 1** and name the substituents, if any.

29.2.2 IUPAC Priority

Carboxylic Acids have the **highest priority** of any functional group. What is priority? Each functional group has a **priority** assigned to it, and the carbons in the longest carbon chain are number in the way that **minimises** the number on the **highest priority functional group**.

Only when there's a tie, do we consider the **positions of substituents**.

The IUPAC Priority is as such (from highest priority to lowest):

1. Carboxylic Acid
2. Ester
3. Amide
4. Nitrile
5. Aldehyde
6. Ketone
7. Alcohol
8. Amine
9. Alkene
10. Alkyne
11. Alkane
12. Ether
13. Alkyl Halide
14. Nitro

The general rule is, the functional group with the **highest priority** will take its **suffix form** (e.g. -ol for alcohol), whereas all other functional groups will be named as **substituents** in their **prefix forms** (e.g. hydroxy- for alcohol).

29.2.3 Properties of Carboxylic Acids

In Organic Chemistry, generally, only **Carboxylic Acids** and **Phenols** are acidic. In a **carboxylic acid,** the acidic hydrogen is the one in the **carboxyl** group.

For example, as we know, we have $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$.

29.2.4 Reactions of Carboxylic Acids

Carboxylic Acids are acidic, so they have the normal **acid reactions**, including **neutralisation**:

- Carboxylic Acid + Base → Carboxylate Salt + Water [Neutralisation]
- Carboxylic Acid + Metal → Carboxylate Salt + Hydrogen
- Carboxylic Acid + Carbonate → Carboxylate Salt + Carbon Dioxide + Water

Carboxylate just refers to the Carboxylic Acid without the **acidic hydrogen**. For example, ethanoic acid (CH_3COOH) corresponds to ethanoate (CH_3COO^-).

Furthermore, carboxylic acids can react with alcohols in a process called Esterification. See **Section 29.3.1: Esterification**.

29.3 Ester

Esters are formed from the reactions of **Carboxylic Acids** with **Alcohols**. In an ester, the $-OH$ group of the carboxylic acid is replaced by an $-OR$ group, where R is usually an alkyl.

29.3.1 Esterification

Esters have the following general structure, where R and R' are alkyls:

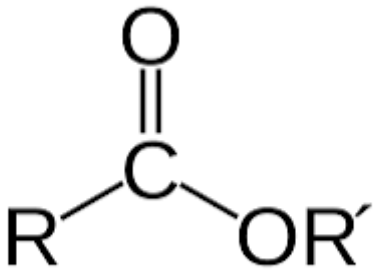


Figure 29.3.1.1 - Ester

Refer to *Figure 29.3.1.2* below.

During Esterification, the $-OH$ group from the carboxylic acid and the H from the $-OH$ in the alcohol bond to form water (boxed in black), and this water is removed as a product.

FYI: Hence, esterification is a type of **condensation** reaction.

After the water is removed, the O and C will form a single $C - O$ bond (in green).

This, along with the existing $C = O$ bond, creates an ester linkage (boxed in purple) .

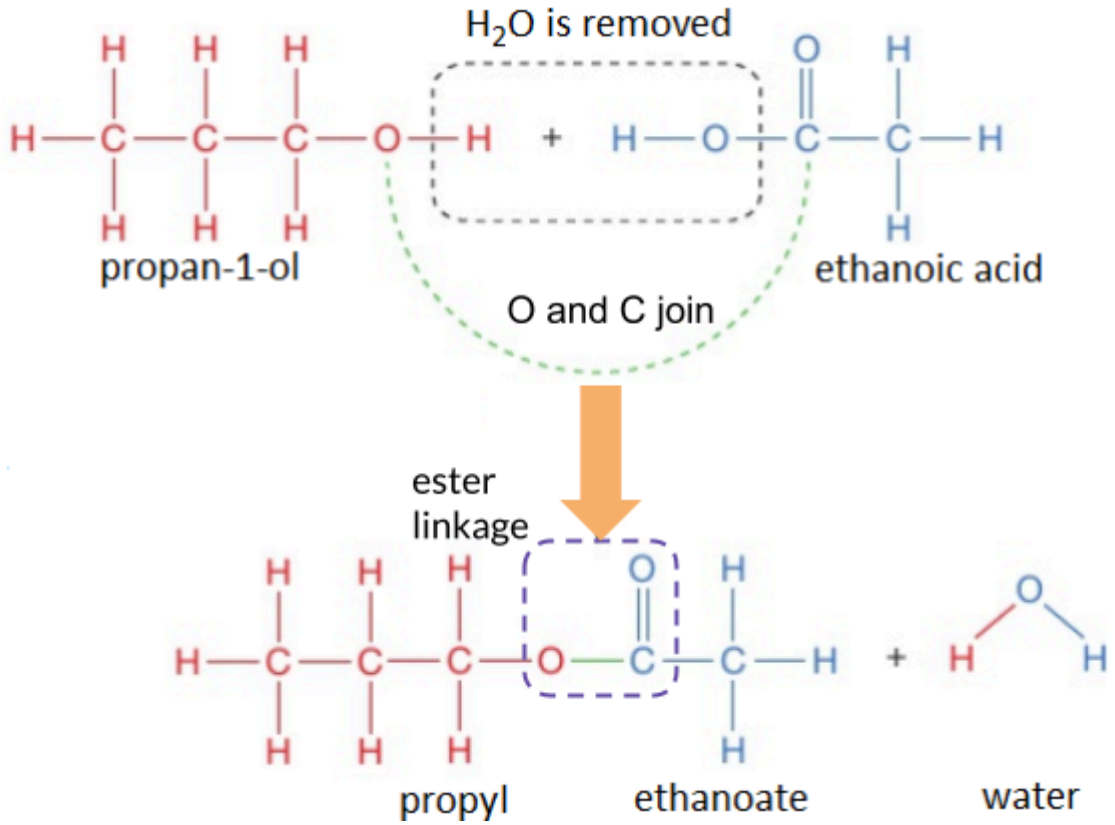


Figure 29.3.1.1 - Esterification

Esterification is catalysed by an **acid catalyst**, which is typically **concentrated** H_2SO_4.

29.3.2 Naming of Esters

To name an ester, my preferred method is to imagine cutting through the $C - O$ single bond. The side with the $C = O$ double bond is derived from the carboxylic acid, and the other side is derived from the alcohol. Now, name the alcohol and carboxylic acid it is derived from.

You can imagine an $-OH$ group where you drew the line on the carboxylic acid side, and replace the O with an OH on the alcohol side to help you name the two parts, if needed.

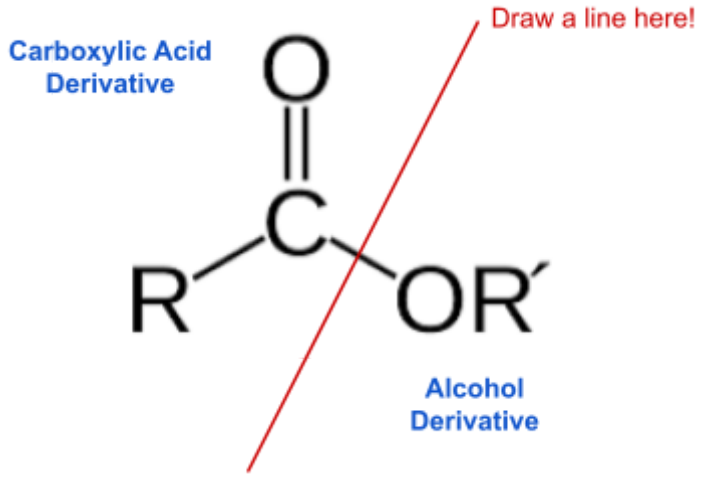


Figure 29.3.2.1 - Splitting an Ester

Take the name of the alcohol, and replace "-anol" with "-yl". In the example in *Figure 29.3.1.1* above, "propanol" becomes "propyl". This is the first part of the name of the ester.

Now take the name of the carboxylic acid, and replace "-oic acid" with "-oate". In the example in *Figure 29.3.1.1* above, "ethanoic acid" becomes "ethanoate". Add this behind the first part.

And we are done! In the example in *Figure 29.3.1.1*, the ester is called "propyl ethanoate".

29.3.3 Properties of Esters

Esters are generally **slightly polar, volatile** and have a **pleasant (fruity) odour**.

Smaller esters (i.e. with lower molecular masses) tend to be **soluble in water**.

29.3.4 Fats

Recall from BL1131 that fats are made of a glycerol and three fatty acids.

Glycerol is actually propan-1,2,3-triol, and fatty acids are long carboxylic acids. As such, fats are actually the result of the three alcohol groups on glycerol reacting with the three fatty acids, to form three ester linkages.

Hence, we say that fats are **tri-esters of glycerol**.

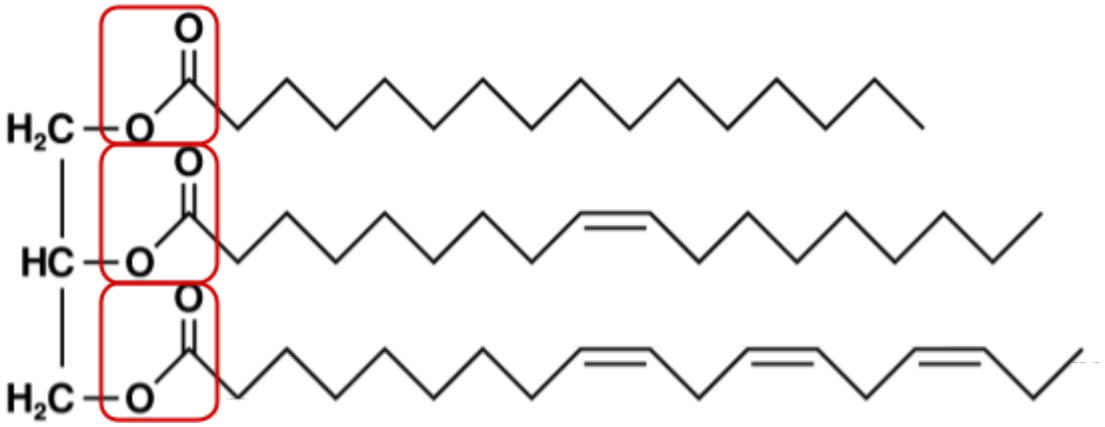


Figure 29.3.4.1 - A Fat Molecule

Figure 29.3.4.1 shows an example of a fat molecule. It has long hydrocarbon chains attached to the glycerol by three ester linkages (boxed in red).

29.3.5 Condensation Polymerisation of Polyesters

Polyesters are a group of **Polymers** which contain the **ester** functional group in every monomer. An example of a **polyester** is **terylene**.

Terylene is synthesised by the repeated **esterification** between 1,2-ethanediol and benzene-1,4-dicarboxylic acid, shown below in *Figure 29.3.5.1*.

Terylene is widely used in clothing and fabrics, to make things like curtains and sleeping bags.

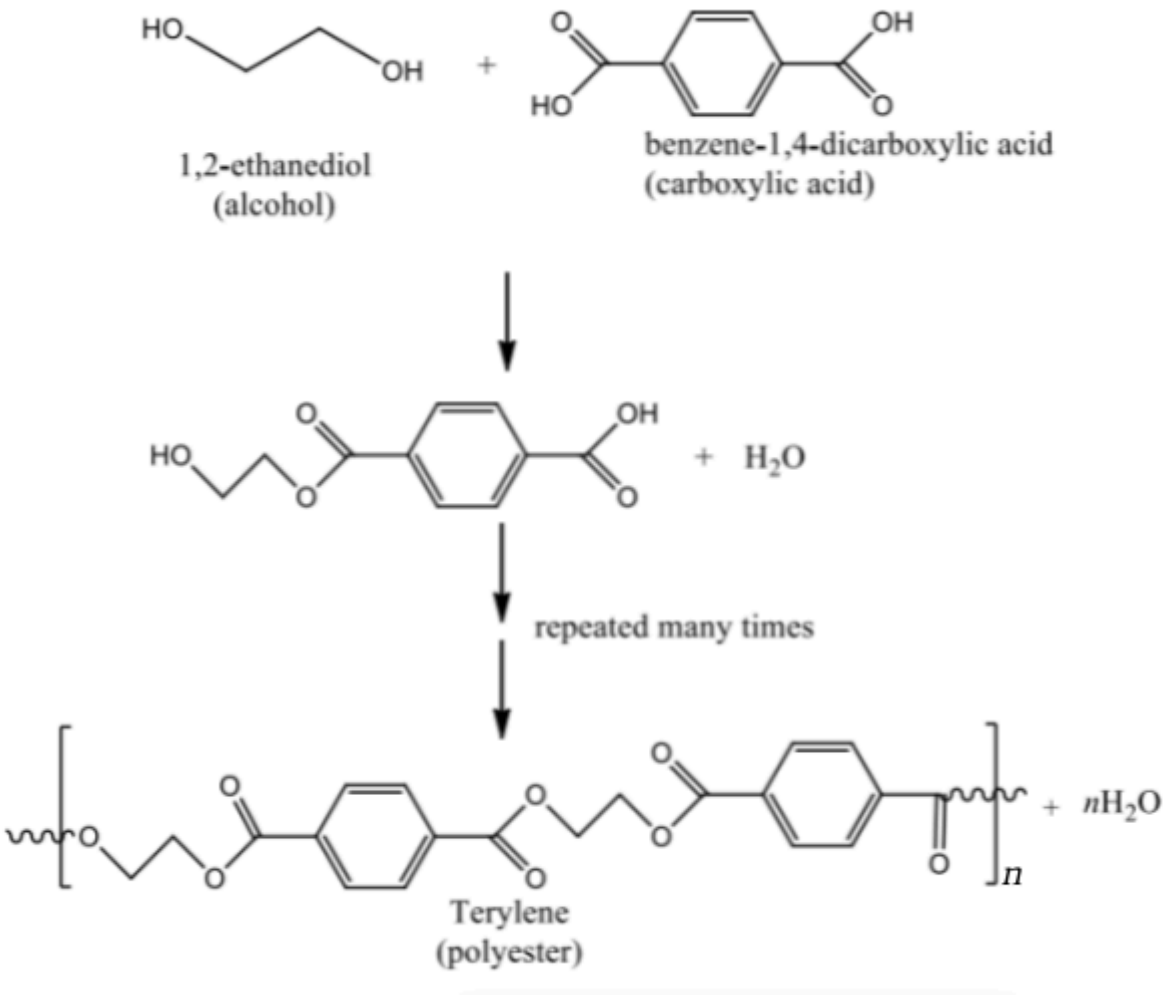


Figure 29.3.5.1 - Condensation Polymerisation of Terylene

Topic 30: Organic Compounds Containing Nitrogen

30.1 Amine

Amines are organic molecules that contain **nitrogen**. They are **organic derivatives** of **ammonia.** (i.e. you take an ammonia molecule, and substitute the hydrogens for other functional groups)

Similar to alcohols, we can classify amines into **primary, secondary** and **tertiary amines**, by the **number of alkyl chains** bonded to the **nitrogen**.

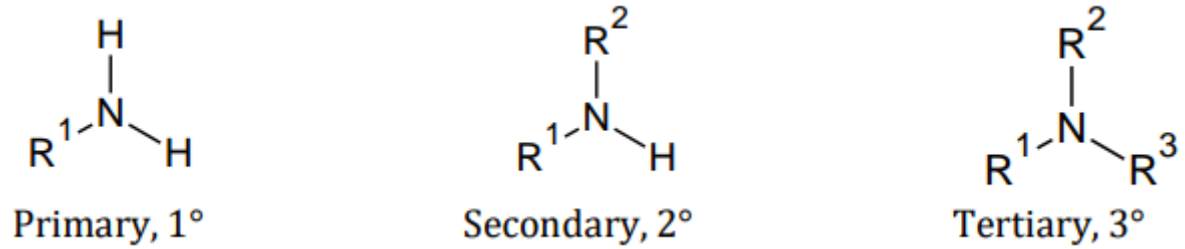


Figure 30.1.1 - Primary, Secondary and Tertiary Amines

In particular:

- Primary amines have one **alkyl chain** (R^1) bonded to the **nitrogen**.
- Secondary amines have two **alkyl chains** (R^1, R^2) bonded to the **nitrogen**.
- Tertiary amines have three **alkyl chains** (R^1, R^2, R^3) bonded to the **nitrogen**.

30.1.1 Naming Primary Amines

Amines have the **prefix form** "amino-" and the **suffix form** "-amine".

Since amines have a very low priority, (See **Section 29.2.2: IUPAC Priority**) they will usually take the prefix form as a substituent. Hence if another functional group with higher priority is present, we will name amines as **amino-**, along with the position.

Example

State the IUPAC Name of the molecule shown in *Figure 30.1.1.1* below.

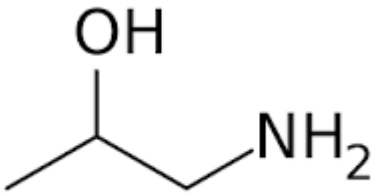


Figure 30.1.1.1 - Example Molecule

Solution

We notice that there are two functional groups, an **alcohol** and an **amine**, and the longest carbon chain has 3 carbons, as shown in *Figure 30.1.1.2* below.

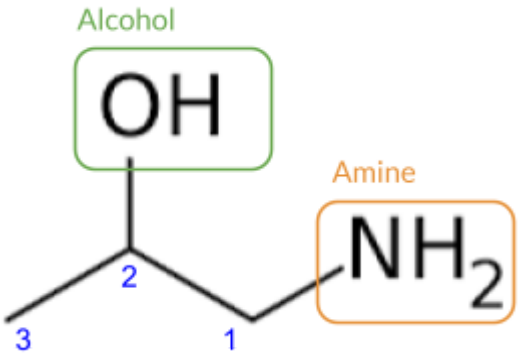


Figure 30.1.1.2 - Functional Groups and Numbers of Carbons

Alcohol has a higher priority than **Amine**, so we will use the suffix form for alcohol, "-ol". All other functional groups (i.e. Amine), will take their prefix forms. Hence, amine takes "amino-".

Also, the alcohol is on the **second** carbon, while the amide is on the **first** carbon.

Now, the root word is derived from the number of carbons. In this case, it's **"propane"**.

Putting "amino-", "propane", and "-ol" together, we get "**1-aminopropan-2-ol"**.

Note: If we had numbered the carbons from left to right instead, the group with highest priority (alcohol) would still have the same number, but the first difference in substituent numbers would increase. Hence, we would not number from left to right.

On the other hand, if there are **no other functional groups with higher priority**, we take the root word derived from the number of carbons, such as "propane", and replace the "-e" with "-amine".

Example

State the IUPAC Name of the molecule shown in *Figure 30.1.1.3* below.

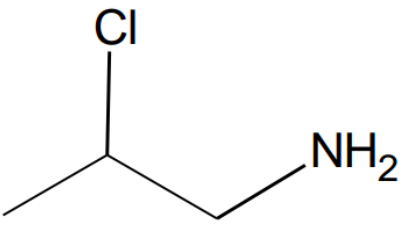


Figure 30.1.1.3 - Example Molecule

Solution

We notice that there are two functional groups, an **alkyl halide** and an **amine**, and the longest carbon chain has 3 carbons, as shown in *Figure 30.1.1.4* below.

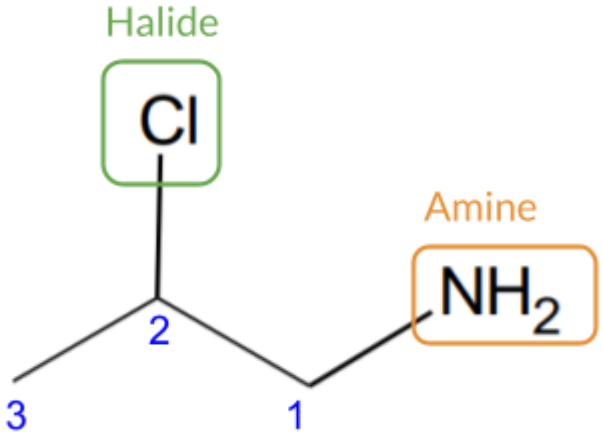


Figure 30.1.1.4 - Functional Groups and Numbers of Carbons

Amine has a higher priority than **Alkyl Halide**, so we will use the suffix form for amine, "-amine". All other functional groups (i.e. Alkyl Halide), will take their prefix forms. Hence, the chlorine takes "chloro-".

Also, the amine is on the **first** carbon, while the halide is on the **second** carbon.

Now, the root word is derived from the number of carbons. In this case, it's **"propane"**.

Putting "chloro-", "propane", and "-amine" together, we get "**2-chloropropan-1-amine"**.

Even though this example looks very similar to the previous example, the naming is completely different due to the difference in priority!

30.1.2 Physical Properties of Amines

Here are a few physical properties of amines.

30.1.2.1 Melting and Boiling Point

Amines are **polar molecules**, due to the presence of a lone pair on the nitrogen, causing the molecule to have a **trigonal pyramidal** geometry around nitrogen.

This implies that they have slightly **higher melting and boiling points** compared to similar hydrocarbons.

For example, refer to *Figure 30.1.2.1.1* below.

Note: The naming of secondary and tertiary amides as in *Figure 30.1.2.1.1* below is **NOT** tested.

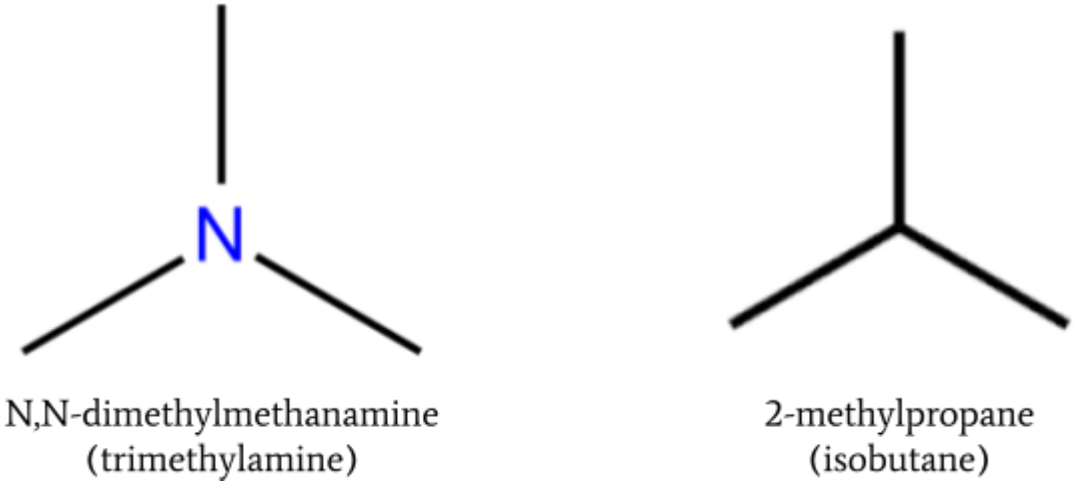


Figure 30.1.2.1.1 - Amine and Alkane

For example, **N,N-dimethylmethanamine** has a melting point of **-117.20 °C**, which is significantly higher than the melting point of **2-methylpropane**, which has a melting point of **-159.6 °C**, despite the similar structure.

This is due to the **Dipole-Dipole Forces** that are present in N,N-dimethylmethanamine (which is polar) but not in 2-methylpropane (which is non-polar).

London Dispersion Forces are also present, but the relative strengths of these forces should be about the same for both molecules since they have similar molar masses (59.112 g/mol vs 58.124 g/mol).

30.1.2.2 Solubility in Water

Amines with longer hydrocarbon chains attached will be less soluble in water. This is because amines have a highly polar $H - N$ bond, which allows it to interact by **hydrogen bonding**. Since water also interacts by **hydrogen bonding**, shorter chain amines and water will interact with each other more and hence these shorter chain amines are **soluble in water.**

On the other hand, as the carbon chains on the amine gets longer, the polar $H - N$ bond becomes a smaller part of the molecule and the main intermolecular force of attraction becomes **London Dispersion Forces**. Since these longer amines and water now interact mainly by different intermolecular forces of attraction, these longer amines are **less soluble in water**.

In particular, amines with **less than 6 carbons** attached will be **soluble** in water, while amines with **6 or more carbons** will be **insoluble** in water.

30.1.2.3 Odour

Amines have **sharp, penetrating odours** that gives **decomposing animal tissues** their smell.

30.1.3 Reactions of Amines

Amines have two main reactions: **Neutralisation** and **Formation of Amide**.

30.1.3.1 Neutralisation

Recall that NH_3 is considered a base because it has the reaction $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$, where a hydroxide is produced.

Similarly, an amine can produce OH^- ions when dissolved in water as such:

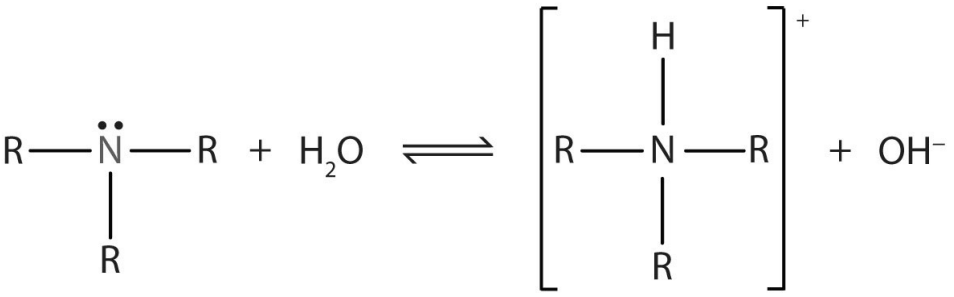


Figure 30.1.3.1.1 - Amine Producing Hydroxide Ion

This implies that amines are **basic** in nature and can undergo **neutralisation** with an acid.

For example, consider methylamine, CH_3NH_2, which is a primary amine.

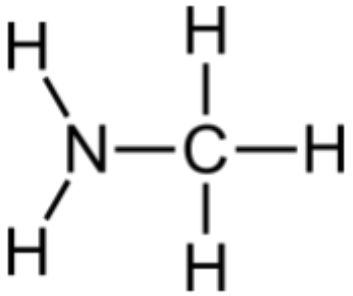


Figure 30.1.3.1.2 - Methylamine

Methylamine will dissociate to form ions in water like such:

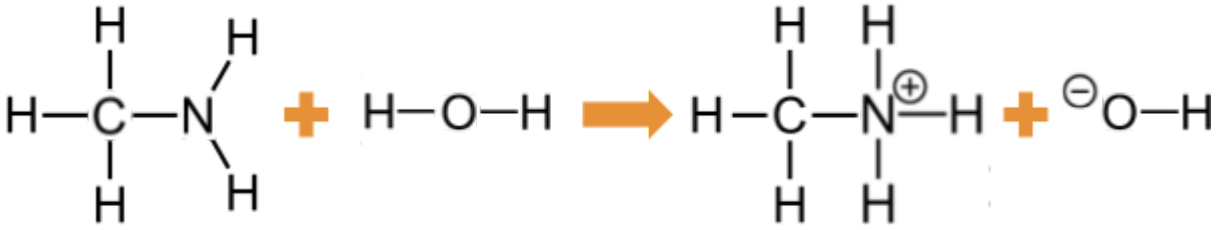


Figure 30.1.3.1.3 - Dissociation of Methylamine

Hence, if methylamine was dissolved in water, it can react with acids like hydrochloric acid like so in *Figure 30.1.3.1.4*:

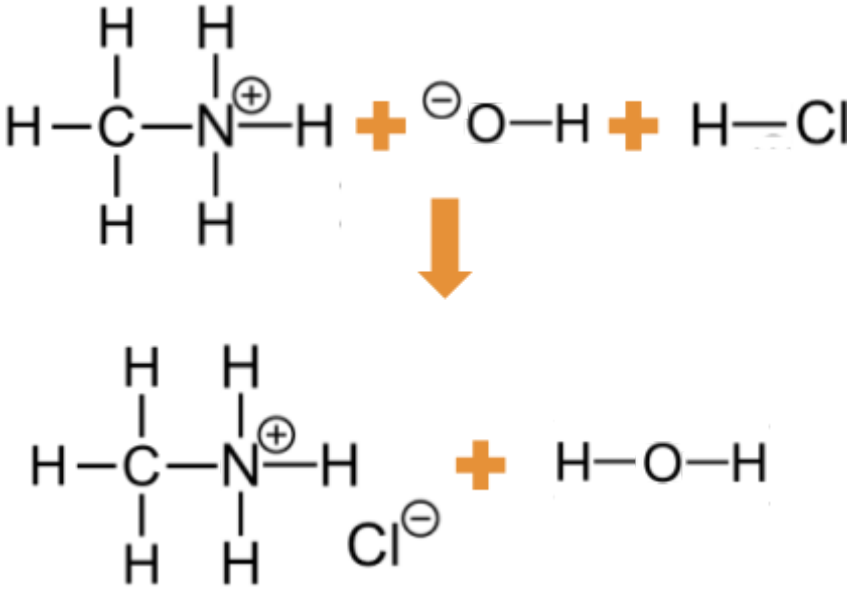


Figure 30.1.3.1.4 - Reaction Between Methylamine Solution and Hydrochloric Acid

More concisely without showing the presence of water:

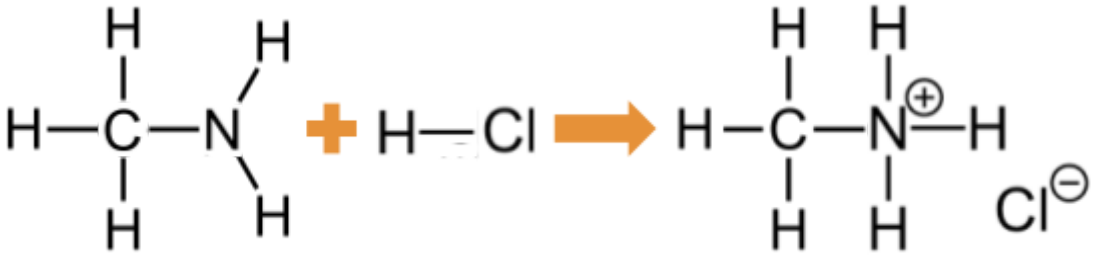


Figure 30.1.3.1.5 - Neutralisation of Amine

Other acids (e.g. sulfuric, nitric, phosphoric, carboxylic acids etc.) can also react with amines.

30.1.3.2 Formation of Amide

Recall that an amide is a molecule containing the $-CON-$ group as highlighted in red in *Figure 30.1.3.2.1* below:

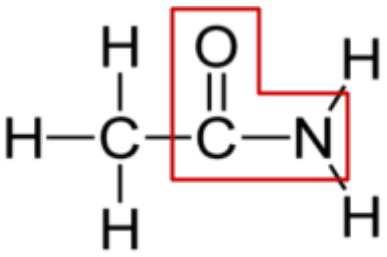


Figure 30.1.3.2.1 - Acetamide

Amides are formed by a special type of neutralisation reaction between **primary or secondary amines** and **carboxylic acids**. (See **Section 30.2: Amide**)

This is the reaction between a **primary amine** and a **carboxylic acid**:

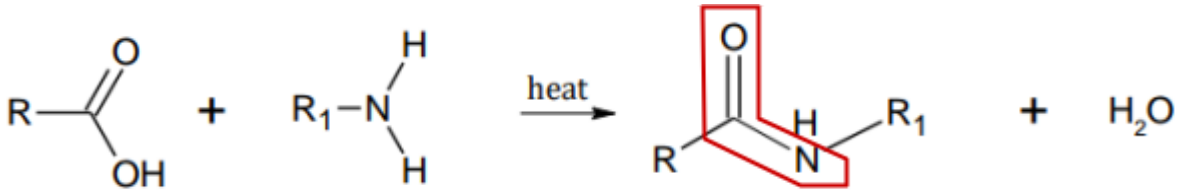


Figure 30.1.3.2.2 - Formation of Amide from Primary Amine and Carboxylic Acid

This is the reaction between a **secondary amine** and a **carboxylic acid**:

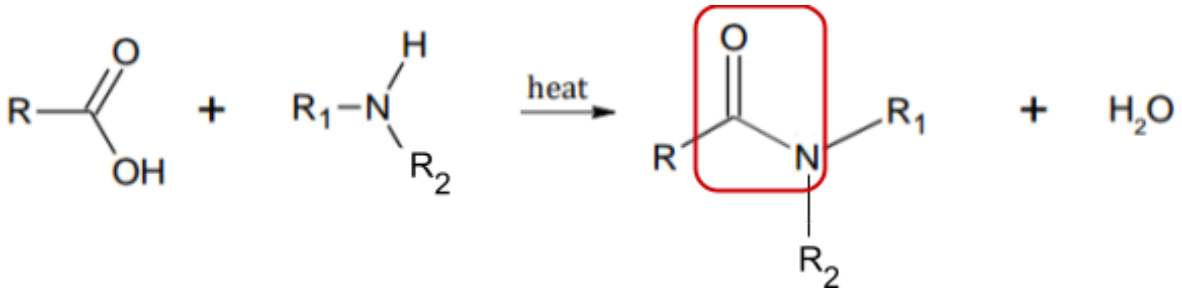


Figure 30.1.3.2.3 - Formation of Amide from Primary Amine and Carboxylic Acid

In both *Figure 30.1.3.2.2* and *Figure 30.1.3.2.3* above, the **amide group** is highlighted in red.

Since water is released as a product during the formation of an amide, this reaction is known as a type of **condensation reaction**.

30.1.4 Separation of Amine from Mixture of Organic Compounds

A technique to separate amines from a mixture of organic compounds that is dissolved in an organic solvent such as an **ether**, is by using **separation funnels**.

Since most organic compounds are **insoluble** or **only very slightly soluble** in water, but amines are an exception because they react with acids to form salts that are generally **very soluble** in water, the idea is to differentiate between these compounds by introducing an acid and water.

Compounds tend to move from one solvent to another based on its solubility in the solvent. It will tend to go to the solvent in which it is **more soluble in**.

For example, suppose our amine was RNH_2. We add the ether mixture into a **separatory funnel** and add an acid like HCl and water. This causes the amine to react with the acid in a neutralisation reaction: $RNH_2\ (in\ ether)\ +\ HCl\ (aq) \rightarrow RNH_3Cl\ (aq)$.

This causes the amine to shift to the **aqueous phase** (i.e. into the water), due to the ionic nature of the amine salt, while the other **organic compounds**, which are not soluble in water, will tend to stay in the **non-aqueous phase** (i.e. in the ether), as seen in *Figure 30.1.4.1* below.

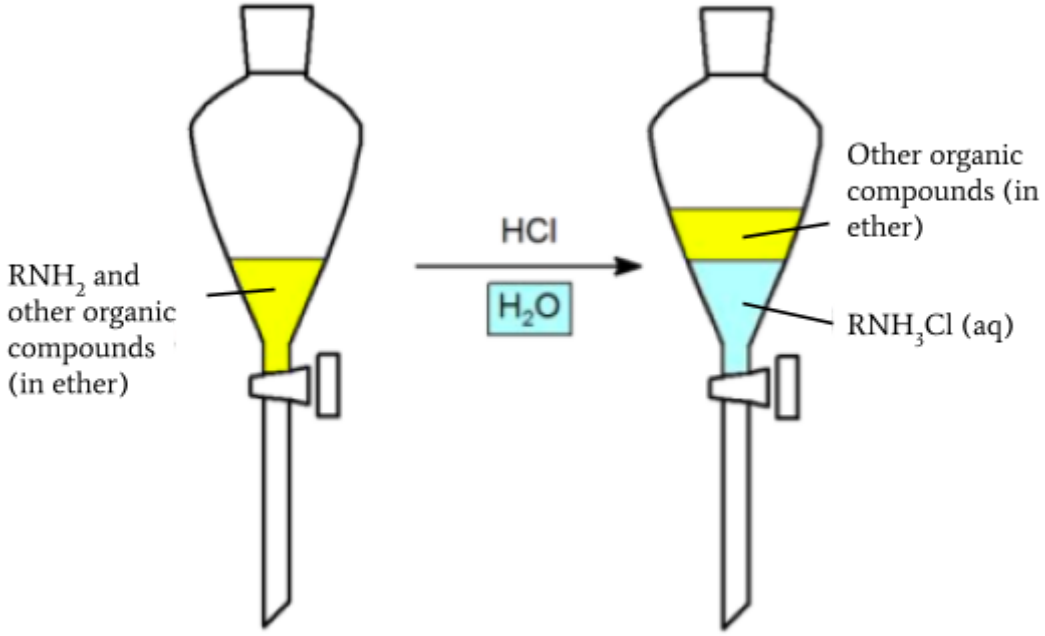


Figure 30.1.4.1 - Adding HCl and H_2O

Now, we can just collect the aqueous phase from the **separatory funnel**, obtaining RNH_3Cl solution dissolved in water. We now convert the amine salt back to an amine by introducing $NaOH\ (aq)$ and an ether, for the amine to dissolve back into. This encourages the reaction:

$$RNH_3Cl\ (aq)\ +\ NaOH\ (aq) \rightarrow NaCl\ (aq)\ +\ RNH_2\ (in\ ether)\ +\ H_2O\ (l)$$

This causes the amine to dissolve back into the **non-aqueous phase** (ether layer).

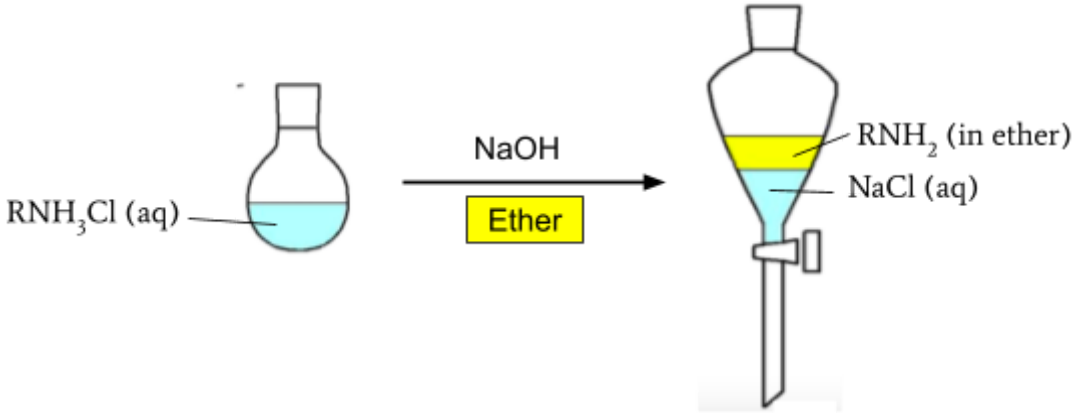


Figure 30.1.4.2 - Converting Amine Salt Back to Amine

Finally, we discard the **aqueous phase** (water layer) and collect the RNH_2 in ether.

30.2 Amide

Amides are **carboxylic acid derivatives** of **amines**, and are formed from a condensation reaction between **ammonia**, or a **primary or secondary amine**, and a **carboxylic acid**.

They contain the **amide linkage** ($-CON-$) as shown in *Figure 30.2.1* below, and the amide linkage itself contains a **carbonyl group** (the $C = O$ bond).

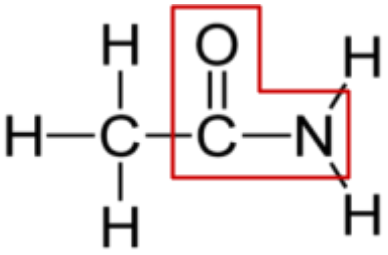


Figure 30.2.1 - Amide Linkage in Acetamide

30.2.1 Classification of Amides

We classify Amines as **primary, secondary** or **tertiary** based on the number of **carbons** bonded to the **nitrogen**.

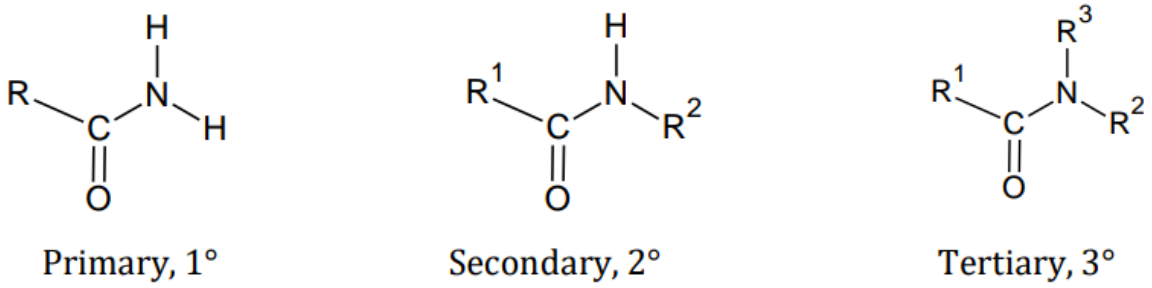


Figure 30.2.1.1 - Primary, Secondary and Tertiary Amides

Primary amides have the **carbonyl carbon** bonded to the **nitrogen**.
Secondary amides have the **carbonyl carbon** and one **alkyl chain** (R^2) bonded to the **nitrogen**.
Tertiary amides have the **carbonyl carbon** and two **alkyl chains** (R^2, R^3) bonded to the **nitrogen**.

Note that the alkyl chain R or R^1 do not count to this number of carbons because it this chain is not **directly bonded** to the nitrogen.

30.2.2 Naming of Primary Amides

Amides take the suffix form "-amide". To name an amide, we will first have to name the carboxylic acid it is derived from.

To find the carboxylic acid, the trick I prefer is to draw a line through the $C - N$ single bond, and the sides that contains the $C = O$ double bond is the part that is derived from the carboxylic acid, as in *Figure 30.2.2.1 below*.

For the carboxylic acid side, just imagine an $-OH$ group where you drew the line.

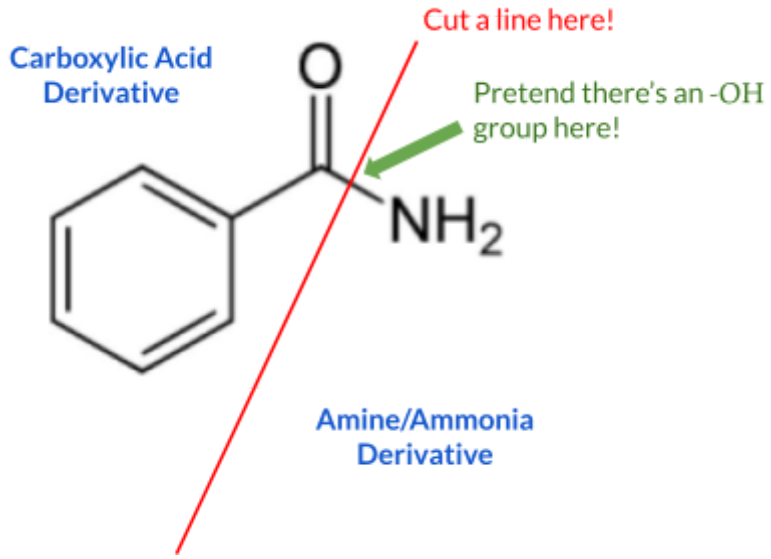


Figure 30.2.2.1 - Separating an Amide

In this example, the carboxylic acid is **Benzoic Acid**.

After we name the carboxylic acid, we can change the ending "-oic acid" to "-amide". In this example, we get "**Benzamide**".

If there are any substituents, number the **carbonyl carbon** (the one in the $C = O$ double bond) as the **first** carbon, and name the substituents accordingly.

30.2.3 Properties of Amides

Like amines, the longer the carbon chains attached to an amide, the **less soluble in water** it becomes. Also similar to amines, when there are less than 6 carbons attached to an amide, it is **soluble in water** in general, and when there are more than 6 carbons, it is **insoluble in water**.

Unlike amines, however, **amides are not basic**.

30.2.4 Amino Acids

An amino acid is an **amine derivative** of a **carboxylic acid**. That is, it contains both an **amine group** and a **carboxylic acid group**. Many amino acids have the amine group and the carboxylic acid group bonded to the same carbon, like in *Figure 30.2.4.1* below.

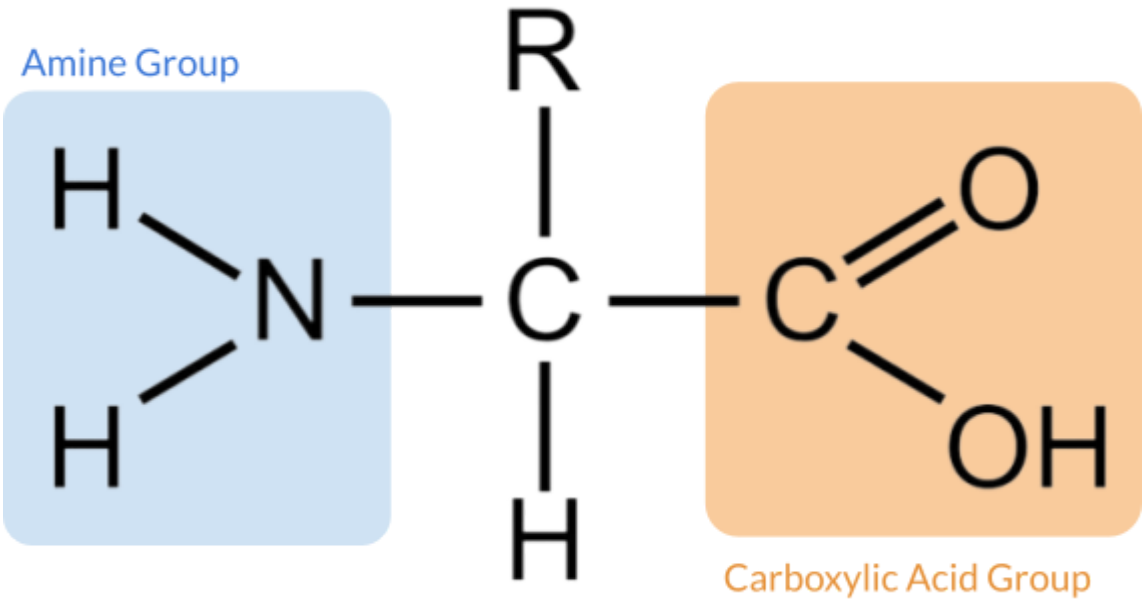


Figure 30.2.4.1 - Amino Acid

where R denotes a side chain. Because of the presence of both an **amine** and a **carboxylic acid** in amino acids, they can react with other **amino acids** to form **amide linkages**, also referred to in this case as **peptide bonds**, as shown in *Figure 30.2.4.2* below.

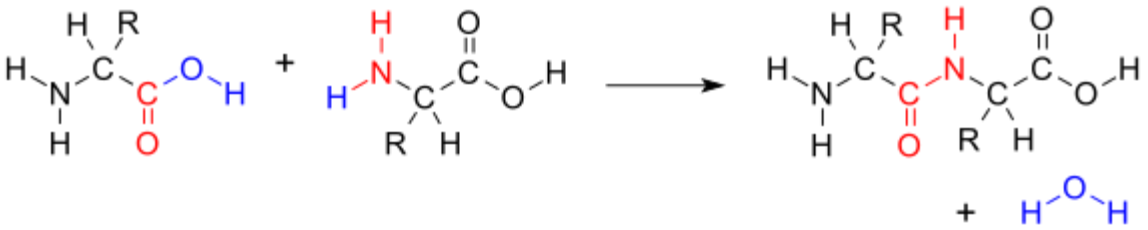


Figure 30.2.4.2 - Formation of Peptide Bond

As such, these **amino acids** are able to bond together in order to form long chains called **polypeptides,** which you might recall from BL1131 are the **proteins** that can be found in all living organisms, for example, in our DNA.

30.2.5 Condensation Polymerisation of Polyamides

Condensation Polymerisation can be used to synthesise **polyamides** like **nylon**.

This can be from **amino acids**, such as **6-aminohexanoic acid**, which forms **nylon-6**, a polyamide, as shown in *Figure 30.2.5.1* below:

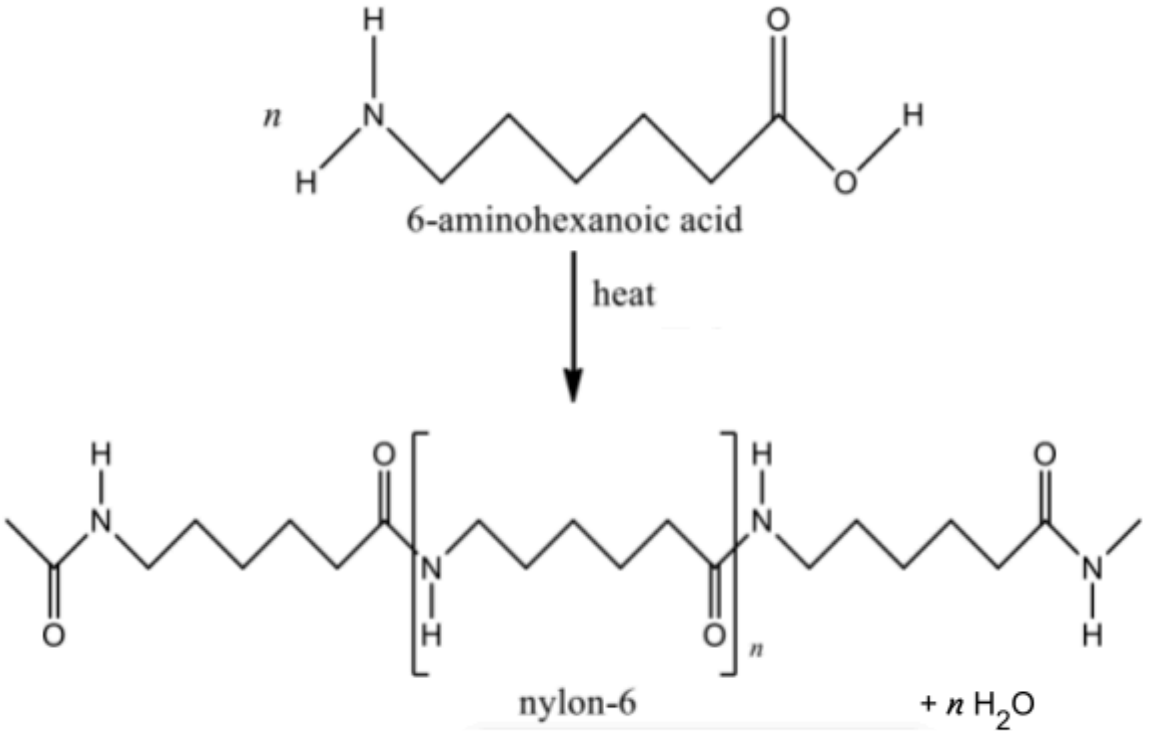


Figure 30.2.5.1 - Formation of Nylon-6

Furthermore, the **amide linkage** could be formed from two different compounds, with one containing an **amine** and one containing a **carboxylic acid**, such as the formation of **nylon-6,6** from **1,6-hexanediamine** (contains amine) and **1,6-hexanedioic acid** (contains carboxylic acid), as in *Figure 30.2.5.2* below.

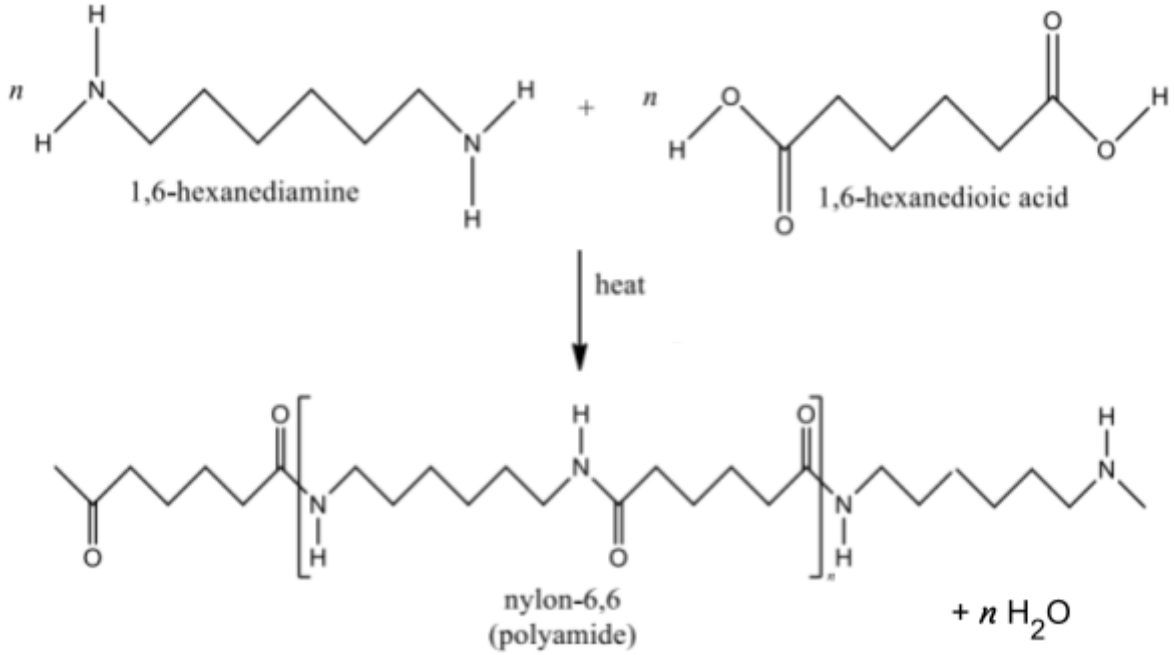


Figure 30.2.5.2 - Formation of Nylon-6,6

Furthermore, we can also use a carboxylic acid derivative such as **1,10-decanedioyl dichloride** (sebacoyl chloride) with an amine like **1,6-hexanediamine** (hexamethylene diamine), to form Nylon-6,10 as shown in *Figure 30.2.5.3*.

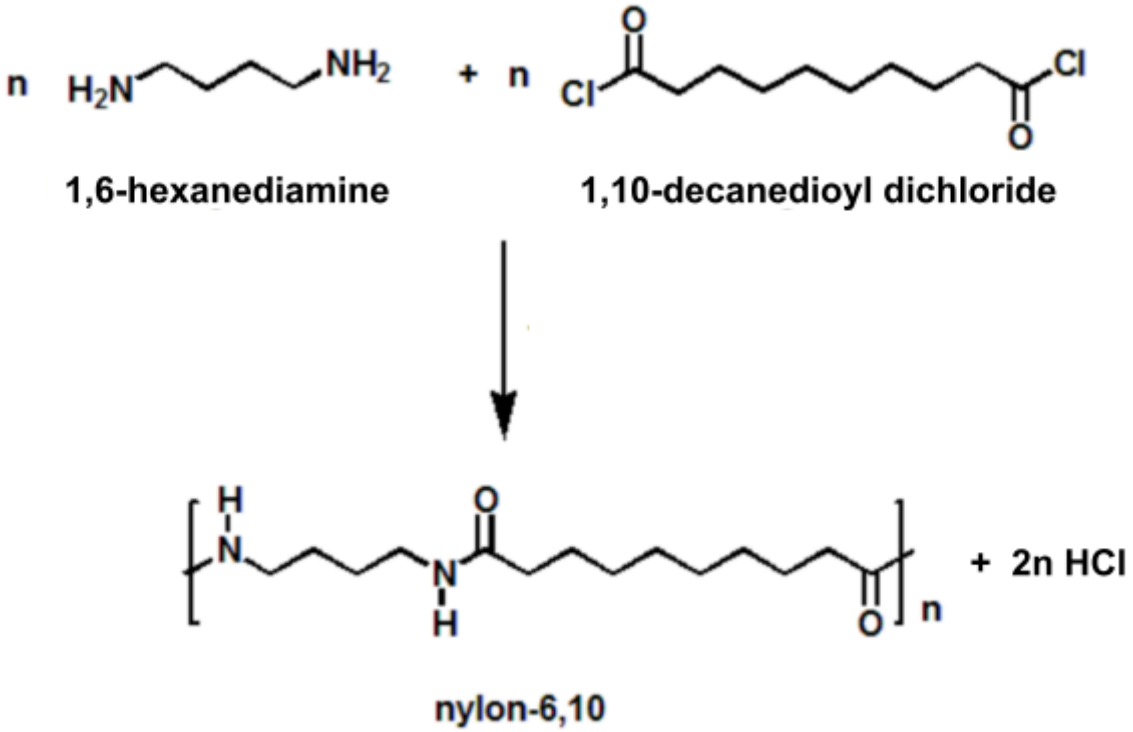


Figure 30.2.5.3 - Formation of Nylon-6,10

Topic 31: Chemical Energetics

31.1 Exothermic and Endothermic Reactions

When the energy **released** in a reaction is **more** than the energy **absorbed**, there is a net release of energy to the surroundings as **heat**. As a result, the **temperature** of the surroundings will **increase** and the reaction mixture will feel hot. This is called an **exothermic** reaction.

When the energy **released** in a reaction is **less** than the energy **absorbed**, there is a net absorption of energy from the surroundings. As a result, the **temperature** of the surroundings will **drop** and the reaction mixture will feel cold. This is called an **endothermic** reaction.

Exothermic Reactions	Endothermic Reactions
Bonds Formed stronger than Bonds Broken	Bonds Broken stronger than Bonds Formed
More energy Released than Absorbed	More energy Absorbed than Released
Net Release of Energy to surroundings	Net Absorption of Energy from surroundings
Temperature of surroundings increases	Temperature of surroundings decreases
Reaction Mixture feels warm	Reaction Mixture feels cold

Table 31.1.1 - Exothermic vs Endothermic Reactions

Here are some examples of **exothermic** reactions:

- Combustion
- Respiration
- Neutralisation
- Dissolving Acids
- Dissolving Alkalis
- Oxidation of Metals (e.g. Rusting)
- Nuclear Reactions

Here are some examples of **endothermic** reactions:

- Thermal Decomposition
- Photosynthesis
- Dissolving Certain Ionic Salts
- Photolytic Decomposition (e.g. $2\ AgBr\ (s) \rightarrow 2Ag\ (s)\ +\ Br_2\ (g)$ under sunlight)

31.2 Enthalpy Change in Reaction

The **change in enthalpy,** denoted ΔH, of a reaction refers to the difference between the energy content of the products and reactants. It measures the **amount of heat flowing into or out of a reaction system** at constant pressure, where ΔH is given by:

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

Where $H_{products}$ and $H_{reactants}$ refer to the amount of energy stored in one mole of the product and reactant respectively, and are typically measured in kJ/mol. ΔH is also measured in kJ/mol.

31.2.1 Energy Level Diagrams

Energy Level Diagrams can be used to represent the change in energy of the **reaction mixture**.

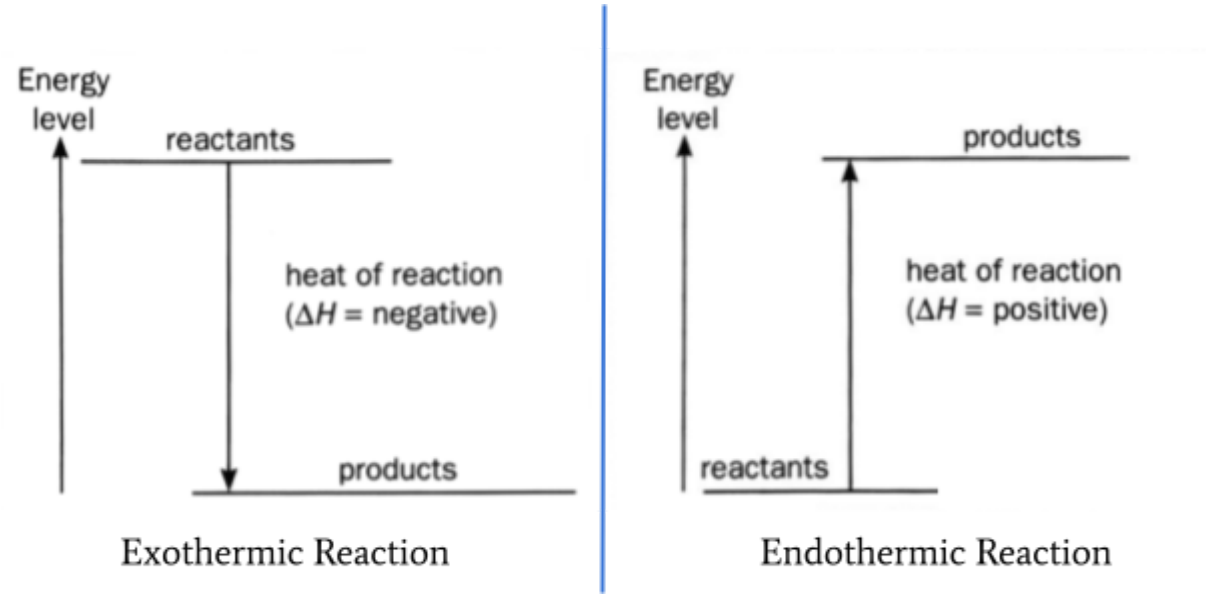


Figure 31.2.1.1 - Energy Diagram for Exothermic and Endothermic Reaction

Note: The energy diagram represents the change in energy of the **reaction mixture**, so for an Endothermic reaction, while the change in energy of the **environment** is **negative**, the change in energy of the **reaction mixture** is actually **positive**, and vice versa for Exothermic reactions.

For example, we can consider the reaction $2\ H_2\ (g)\ +\ O_2\ (g) \rightarrow 2\ H_2O\ (l)$. Its reaction enthalpy is given by $\Delta H\ = -484 kJ/mol$.

Here, the reaction is **exothermic**, because the energy is lost during the reaction, implying that energy was transferred to the environment.

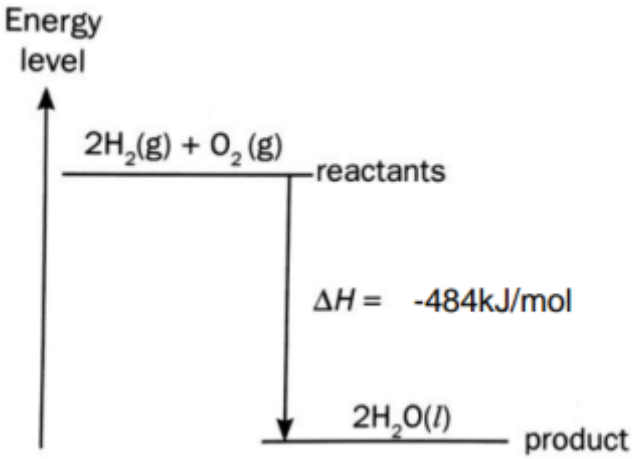


Figure 31.2.1.2 - Example Energy Diagram

31.2.2 Energy Profile Diagrams

Consider the above exothermic reaction, $2\ H_2\ (g)\ +\ O_2\ (g) \rightarrow 2\ H_2O\ (l)$ again.

To form H_2O, energy is **must first be taken** in to break the $H\ -\ H$ bonds and the $O\ =\ O$ bonds present in H_2 and O_2 must first be broken before the $H\ -\ O$ bonds in H_2O can be formed which then releases back **more energy than was taken in**, making the reaction **exothermic**.

The energy that must be taken in to break the bonds is known as the reaction's **Activation Energy**, denoted E_a, and this energy barrier must be overcome before the reaction can occur, and comes from the **kinetic energy** possessed by the reactants before colliding.

This means that a more accurate representation of the energy in an exothermic reaction should first show the amount of energy increasing, then decreasing to below the original level, like in an **Energy Profile Diagram**, shown in *Figure 31.2.2.1* below.

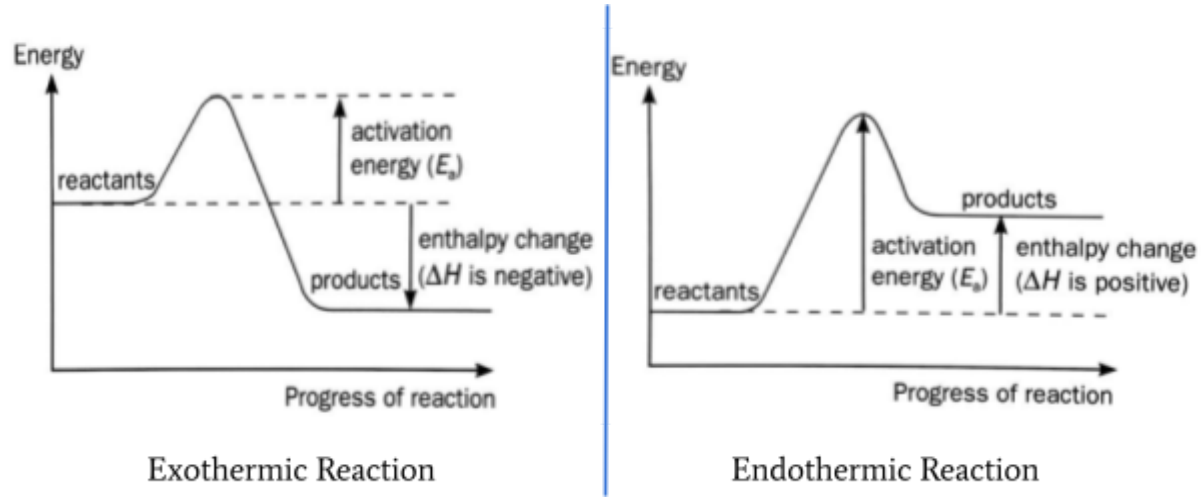


Figure 31.2.2.1 - Energy Profile Diagram of Exothermic and Endothermic Reaction

31.3 Bond Energy and Enthalpy Change

The **Bond Energy** of a specific type of **bond** is the amount of energy needed to break **one mole** of a given bond. For example, we are talking about a specific type of bond like a $O - H$ bond.

Note: This is different from **Section 22.3: Covalent Bond and Bond Energy**, where we talk about the Bond Energy of a **molecule**, not a specific bond in a molecule.

We can calculate the **Enthalpy Change** of a reaction equivalently by:

$$\Delta H = (sum\ of\ Bond\ Energies\ of\ Bonds\ Broken) - (sum\ of\ Bond\ Energies\ of\ Bonds\ Formed)$$

For example, consider the reaction $N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$, assuming we are given the following table of values, where we wanted to check if it is **endothermic** or **exothermic.**

Bond	Average Bond Energy at r.t.p / kJ mol^{-1}
$N \equiv N$	941
$N - H$	391
$H - H$	436

Table 31.3.1 - Bond Energies of Some Bonds

We would calculate the ΔH value. Since one $N \equiv N$ bond and 3 $H - H$ bonds were broken while 6 $N - H$ bonds were formed, we deduce that:

$$\begin{aligned}\Delta H &= 1(BE_{N \equiv N}) + 3(BE_{H-H}) - 6(BE_{N-H}) \\ &= 1(941\ kJ/mol) + 3(436\ kJ/mol) - 6(391\ kJ/mol) \\ &= -97\ kJ/mol\end{aligned}$$

Since ΔH is negative, this implies that energy is **released**, so this reaction is **exothermic**.

Now suppose we wanted to draw a Energy Profile Diagram for this reaction. The Activation Energy is the energy required to break all the bonds, and this would amount to:

$$\begin{aligned}E_a &= 1(BE_{N \equiv N}) + 3(BE_{H-H}) \\ &= 1(941\ kJ/mol) + 3(436\ kJ/mol) \\ &= 2249\ kJ/mol\end{aligned}$$

We can then draw the energy profile diagram as such:

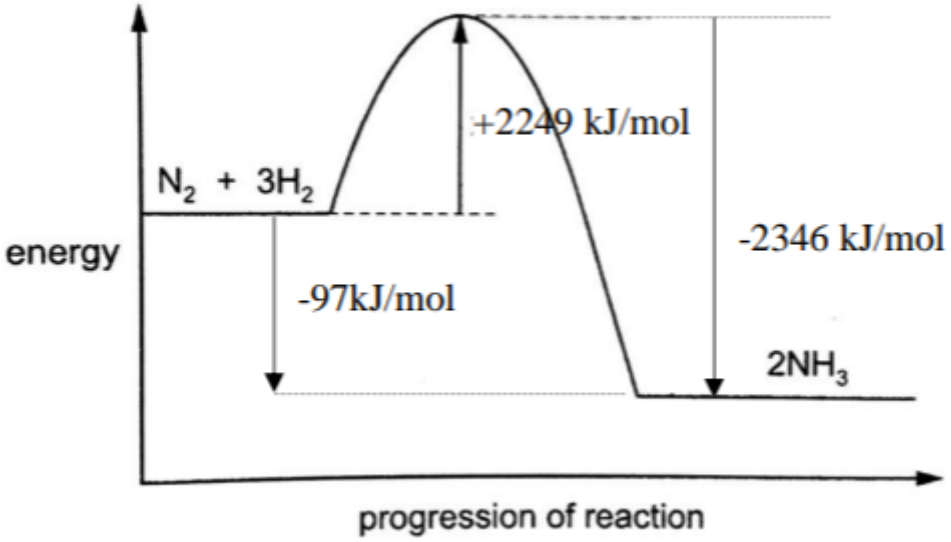


Figure 31.3.2 - Energy Profile Diagram of Reaction

Topic 32: Electrolysis

32.1. Components of an Electrolysis Setup

Electrolysis contains three main components: A **power source**, **electrodes** and an **electrolyte**.

The **power source** (which is usually a **battery** but can be any other DC power source), acts as an **electron pump** that "pushes" electrons along the circuit, transferring electrons from the anode to the cathode when the circuit is closed.

The **electrodes** (**cathode** and **anode**) are usually made of **graphite** or a **metal** and the cathode is attached to the **negative** terminal of the power source, while the anode is attached to the **positive** terminal.

The **electrolyte** is usually an ionic compound in its **aqueous** or **molten** form, such that it will dissociate into **cations** and **anions**, and there are mobile ions that will complete the circuit by conducting electricity.

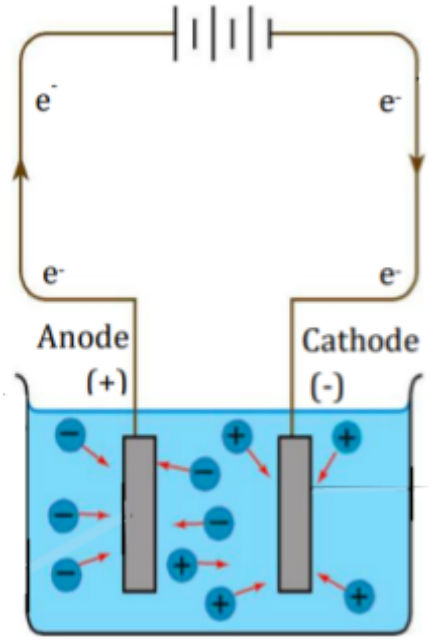


Figure 32.1.1 - Electrolysis Setup

During electrolysis, the **anions** of the electrolyte move towards the **anode**, where they **lose electrons** to the anode and get **oxidised**.

These electrons are transported from the **anode** to the **cathode** by the battery.

On the other hand, the **cations** of the electrolyte move towards the **cathode**, where they **gain electrons** from the cathode and get **reduced**.

Do note that the power source is denoted by two lines, where the longer line denotes the **positive terminal** and the shorter line denotes the **negative terminal**.

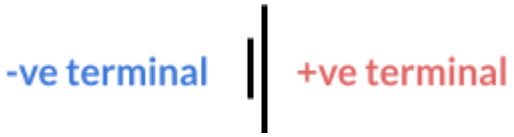


Figure 32.1.2 - Power Source

32.1.1 Precautions

Here are three things to take note:

- The **electrodes** must be made of a **conductive material**, such as **metal** or **graphite**.
- The **electrodes** should be connected to the **battery** by **conductive electrical wires**.
- The **electrodes** must be immersed in the **electrolyte** (not necessarily fully) and connected to the **power supply**, and the **cathode and anode** must **NOT** touch.

32.2 Oxidation and Reduction at the Electrodes

At the **cathode**, the cation, say A^+, will gain electron(s) and become reduced: $A^+ + e^- \rightarrow A$.
If an **aqueous electrolyte** is used, water can also be reduced: $2H_2O + 2e^- \rightarrow H_2 + 2\,OH^-$.

At the **anode**, the anion, say B^-, will lose electron(s) and become oxidised: $B^- \rightarrow B + e^-$.
If an **aqueous electrolyte** is used, water can also be oxidised: $2H_2O \rightarrow O_2 + 4H^+ + 2e^-$.

32.2.1 Standard Reduction Potential, $E°_{red}$

The **Standard Reduction Potential** of a substance, denoted E^0_{red}, measures the **tendency for a substance to be reduced**, and is measured in volts (V). The more **positive** the E^0_{red} value, the **more easily reduced** a substance is.

This means, the more positive E^0_{red} is, the more **favourable** the **reduction process**.

Since **oxidation** is the **opposite** of **reduction**, the more **negative** the E^0_{red} value, the **more easily oxidised** a substance is.

This means, the more negative E^0_{red} is, the more **favourable** the **oxidation process**.

FYI: In fact, **Standard Oxidation Potential,** E^0_{ox}, is equal to $-E^0_{red}$.

Here are some common E^0_{red} values: (For oxidation half-equations, simply flip the arrow)

Reduction Half-equation	E^0_{red} / V
$F_2 + 2e^- \rightarrow 2F^-$	+2.87
$Au^{3+} + 3e^- \rightarrow Au$	+1.50
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.23
$Br_2 + 2e^- \rightarrow 2Br^-$	+1.07
$Ag^+ + e^- \rightarrow Ag$	+0.80
$I_2 + 2e^- \rightarrow 2\,I^-$	+0.53
$2H_2O + O_2 + 4e^- \rightarrow 4OH^-$	+0.40
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
$2H^+ + 2e^- \rightarrow H_2$	0.00
$Pb^{2+} + 2e^- \rightarrow Pb$	-0.13
$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	-0.83
$Al^{3+} + 3e^- \rightarrow Al$	-1.66
$Mg^{2+} + 2e^- \rightarrow Mg$	-2.37
$Na^+ + e^- \rightarrow Na$	-2.71
$Ca^{2+} + 2e^- \rightarrow Ca$	-2.87
$K^+ + e^- \rightarrow K$	-2.92

Table 32.2.1.1 - Some Standard Reduction Potentials

32.2.2 Predicting the Net Reaction of Electrolysis with Molten Ionic Electrolytes

When a Molten Ionic Electrolyte is used, the **anion** is always the one that is oxidised, since water is not present. Similarly, the **cation** is always the one that is reduced.

Suppose the electrolyte used was molten $NaCl$, where $NaCl\ (l) \rightarrow Na^{+}\ (l)\ +\ Cl^{-}\ (l)$.

We can then write the half-equations:

- $Na^{+}\ (l)\ + e^{-} \rightarrow Na\ (l)$ (Reduction At Cathode)
- $2Cl^{-}\ (l) \rightarrow Cl_{2}\ (g)\ + 2e^{-}$ (Oxidation At Anode)

We can now balance the electrons and combine these half-equations to obtain the overall electrolysis net reaction. In this case, multiply the first equation by 2, then combine.

$$2NaCl\ (l) \rightarrow 2Na\ (l)\ +\ Cl_{2}\ (g)$$

We can now predict what will happen:

- Silvery liquid (molten sodium) will form near the cathode.
- Yellowish-green gas (chlorine gas) will form near the anode.

Note: Because $NaCl$ has to be kept in a molten form, the temperature is consistently very high, so the state symbol for Na is $Na\ (l)$.

32.2.3 Predicting the Net Reaction of Electrolysis with Aqueous Electrolytes

To predict what happens during electrolysis with an aqueous electrolyte is quite a bit more complex. This is because, now with the addition of water, water could be the one being **oxidised/reduced** instead of the **anion/cation** respectively.

Hence, we have to consider ΔH^{0}_{red} values.

Note: Larger polyatomic ions like SO_{4}^{2-} and NO_{3}^{-} will **not** be oxidised and discharged. If one of these ions is the anion, **water** will be the one **oxidised** instead.

Furthermore, if the ΔH^{0}_{red} value of the cation/anion makes it less preferable than water, but the cation/anion is very **concentrated**, and the difference in ΔH^{0}_{red} is relatively small, then sometimes the cation/anion can be **reduced/oxidised instead of water**.

Let's consider some examples.

32.2.3.1 Dilute Aqueous Sodium Chloride with Inert Electrodes

For example, let's consider the electrolysis of dilute $NaCl\ (aq)$.

Since both Na^+ and H_2O are present at the **cathode**, we should figure out which particle is the one being **reduced**.

The two reduction half-equations are:

- $Na^+ + e^- \rightarrow Na$ $\quad (\Delta H^0_{red} = -2.71\ V)$
- $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $\quad (\Delta H^0_{red} = -0.83\ V)$

The more positive the ΔH^0_{red} value, the more favourable the reduction of that substance. Since H_2O has a more positive ΔH^0_{red}, this means that **water** is the one that is **reduced** at the **cathode**.

We do a similar analysis at the **anode**, with Cl^- and H_2O. The two reduction half-equations are:

- $Cl_2 + 2e^- \rightarrow 2\ Cl^-$ $\quad (\Delta H^0_{red} = +1.36\ V)$
- $4H^+ + O_2 + 4e^- \rightarrow 2H_2O$ $\quad (\Delta H^0_{red} = +1.23\ V)$

Note: Alternatively, you can also write the half-equation in the other direction to better reflect oxidation, but it's written this way because we are using the standard **reduction** potentials.

The more negative the ΔH^0_{red} value, the more favourable the oxidation of that substance. Since H_2O has a more negative ΔH^0_{red} value, this means that **water** is the one **oxidised** at the **anode.**

Hence, the two half-equations are:

- $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $\quad$ (Reduction at Cathode)
- $2H_2O \rightarrow 4H^+ + O_2 + 4e^-$ $\quad$ (Oxidation at Anode)

We can now balance the electrons by multiplying the first equation by 2 and combining (where H^+ and OH^- are combined to form H_2O):

$$6H_2O\ (l) \rightarrow 2H_2\ (g) + 4H_2O\ (l) + O_2\ (g)$$

Subtracting $4\ H_2O$ from both sides:

$$2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)$$

This is the **net reaction** of this electrolysis reaction.

32.2.3.2 Concentrated Aqueous Sodium Chloride with Inert Electrodes

There is a slight difference in the electrolysis of concentrated $NaCl\ (aq)$.

Because the **standard reduction potential**, ΔH^0_{red}, of the two reactions at the anode are so similar, as seen below:

- $Cl_2 + 2e^- \rightarrow 2Cl^-$ $(\Delta H^0_{red} = +1.36\ V)$
- $4H^+ + O_2 + 4e^- \rightarrow 2H_2O$ $(\Delta H^0_{red} = +1.23\ V)$

By using **concentrated sodium chloride** solution, we can increase the concentration of Cl^- significantly and force Cl^- to be the one **oxidised** instead of water, despite having a less negative ΔH^0_{red}. Note that this can only happen because the ΔH^0_{red} values are relatively close.

This is the **concentration effect**.

Just like in the dilute case, the reduction that occurs is still $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ occurring at the **cathode**.

Hence, the two half-equations that take place are:

- $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ (Reduction at Cathode)
- $2Cl^- \rightarrow Cl_2 + e^-$ (Oxidation at Anode)

Multiplying the second equation by 2 and combining with the first equation:

$$2H_2O\ (l) + 2Cl^-\ (aq) \rightarrow H_2\ (g) + Cl_2\ (g) + 2OH^-\ (aq)$$

This is the **net reaction** of this electrolysis reaction.

We can predict some observations:

- The solution will turn alkaline due to OH^- being released at the cathode.
- An effervescence of a yellowish-green gas (Cl_2) will be formed near the anode.

32.2.3.3 Copper (II) Sulfate with Active Copper Anode

In this case, the **cathode** is still made of **graphite**, an **inert** material, but the **anode** is replaced with a **copper anode**.

Now, the particles present at the cathode are Cu^{2+} (from $CuSO_4$) and H_2O. The ΔH^0_{red} values of these two particles are:

- $Cu^{2+} + 2e^- \rightarrow Cu$ $\quad (\Delta H^0_{red} = +0.34V)$
- $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $\quad (\Delta H^0_{red} = -0.83V)$

Since Cu^{2+} has the most **positive** ΔH^0_{red} value, it is the one that is reduced.

Things are a little different at the **anode**. Now, there are three particles which are present at the anode. SO_4^{2-}, H_2O and Cu (from the anode itself). Firstly, SO_4^{2-} will **not** be discharged, so we can disregard it, and we are left with H_2O and Cu.

The ΔH^0_{red} values of the half-equations involving these two particles are:

- $Cu^{2+} + 2e^- \rightarrow Cu$ $\quad (\Delta H^0_{red} = +0.34V)$
- $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ $\quad (\Delta H^0_{red} = +1.23V)$

Since Cu has the most **negative** ΔH^0_{red} value, it is the one that is oxidised.

Hence, the two half-equations that occur are:

- $Cu^{2+} + 2e^- \rightarrow Cu$ (Reduction at Cathode)
- $Cu \rightarrow Cu^{2+} + 2e^-$ (Oxidation at Anode)

Combining, we get the net equation of $Cu\,(s)$ *(anode)* $+ Cu^{2+}\,(aq) \rightarrow Cu^{2+}\,(aq)$ *(cathode)* $+ Cu\,(s)$.

It might seem like there is no change! However, we see that the $Cu\,(s)$ is being reacted away from the anode, and additional $Cu\,(s)$ is being formed at the cathode!

This means, we will actually observe the **size of the anode** slowly decreasing as the copper is oxidised, while a **reddish-brown** solid (Cu) is being deposited on the cathode.

32.3 Electroplating

Electroplating is a process where a metal is deposited onto an item (that is conductive).

The **electroplated item** is placed as the **cathode**. For example, we could electroplate an item like a key, as shown in *Figure 32.3.1* below.

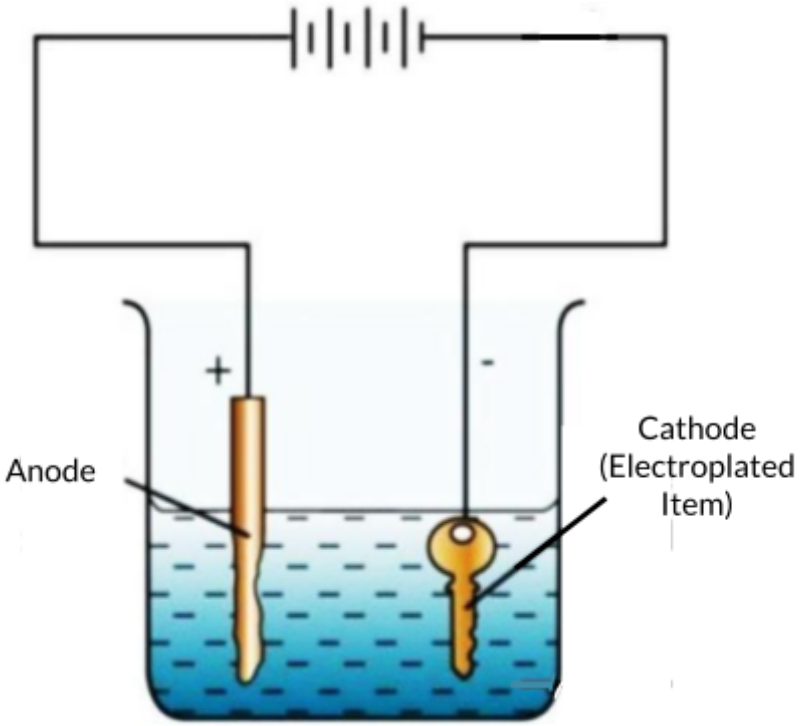


Figure 32.3.1 - Electroplating

For example, if we wanted to electroplate the key with **copper**, we would use a **electrolyte** like **Copper (II) Sulfate,** so that the copper ions in the solution can be reduced at the **cathode** to form **copper metal**, which is then deposited onto the item.

The item has been **electroplated**.

The longer the electroplating setup is allowed to run, the thicker the layer of metal deposited.

32.4 Electrolytic Refining of Metals

See **Section 33.2.1: Electrolysis of Aluminium Oxide** for more information.

Electrolytic Refining is a technique used to remove impurities from a block of impure metal, such as refining copper, as shown in *Figure 32.3.1* below.

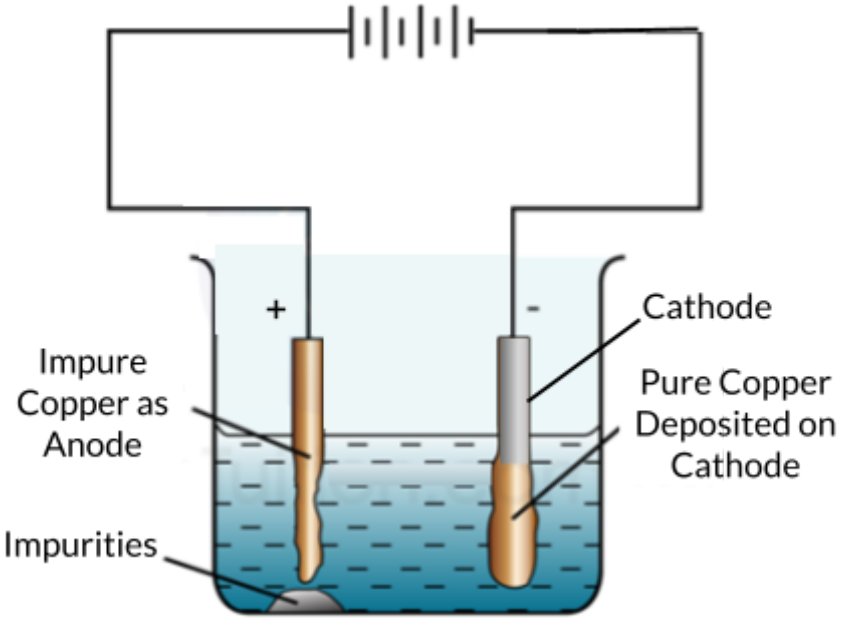


Figure 32.4.1 - Electrolytic Refining of Copper

We can place the impure copper as the **anode**. The copper in the anode will be oxidised to form copper ions which become part of the electrolyte. As the amount of copper decreases, the impurities will sink to the bottom.

As electrons are pushed out of the **cathode**, the copper ions are reduced to form **pure copper metal**, and are deposited on the cathode.

The impure copper has been **refined**.

Topic 33: Extraction of Metals and Rust Prevention

33.1 Stability and Reactivity

In general:

- The **more reactive** a metal is, the **more stable** its ionic compounds are.
- The **less reactive** a metal is, the **less stable** its ionic compounds are.

For example, since Na and K are **very reactive metals**, compounds containing Na^+ and K^+ are **more stable**, and hence Na_2CO_3 and K_2CO_3 are **resistant to thermal decomposition**.

33.2 Extraction of Metals

Most metals are not found in their elemental form in nature, and have to be **extracted** by chemical reactions. In fact, most metals react with oxygen in the air and are found as **oxides**.

Reactivity Decreases ↓

Metal	Category	Method of Extraction
Potassium (K)	Very Reactive Metals	Electrolysis
Sodium (Na)		
Calcium (Ca)		
Magnesium (Mg)	Reactive Metals	
Aluminium (Al)		
Carbon (C)		
Zinc (Zn)	Reactive Metals	Displacement by Carbon in Redox
Iron (Fe)		
Tin (Sn)		
Lead (Pb)		
Hydrogen (H)		
Copper (Cu)	Unreactive Metals	Found Naturally in Elemental Form
Silver (Ag)		
Gold (Au)		
Platinum (Pt)		

Table 33.2.1 - Methods of Extraction of Metals

As shown in *Table 33.2.1* above, the most reactive metals (K to Al), which are those placed above **Carbon** on the **Reactivity Series**, form relatively **stable** compounds, and hence the metal has to be extracted by **electrolysis**.

Furthermore, the metals below **Carbon** and above **Hydrogen** are those that can be extracted by a **redox displacement reaction** with carbon.

Lastly, the metals below **Hydrogen** are very **unreactive** and are found as **free elements** in nature, due to their low reactivity.

33.2.1 Electrolysis of Aluminium Oxide

An example of when electrolysis is used is when we extract Aluminium, from Aluminium Ore, which is in the form of Aluminium Oxide, Al_2O_3.

Using molten Al_2O_3 would be too inconvenient due to the high melting point of Al_2O_3, which is around 2072 °C. Instead we use **cryolite**, which has a melting point of 812 °C, as a solvent.

Aluminium oxide is **dissolved** in molten **cryolite**, Na_3AlF_6. This acts as the **electrolyte**.
In this electrolysis reaction, both electrodes are made of **graphite**.

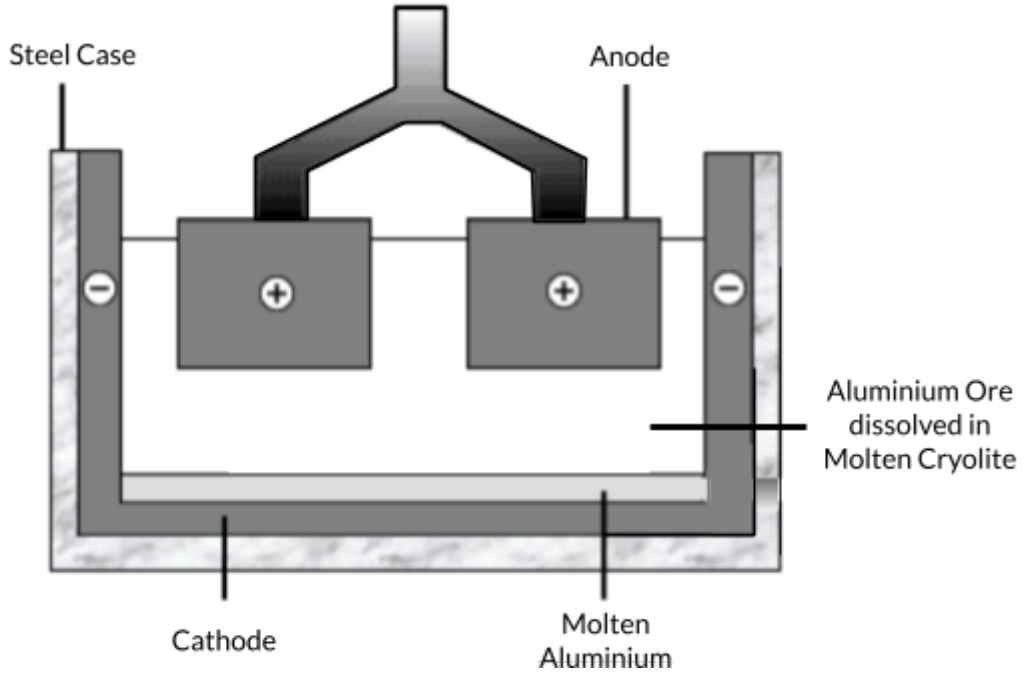


Figure 32.2.1.1 - Electrolysis Setup for Extraction of Aluminium

The Al^{3+} ions will gain electrons from the **cathode**, which surrounds the entire tank, and undergo the reduction $Al^{3+}(l) + 3e^- \rightarrow Al(l)$ and is deposited on the cathode at the bottom, and is removed by an opening in the tank periodically.

On the other hand, the oxidation reaction at the **anodes** are $2O^{2-}\ (l) \rightarrow O_2\ (g)\ +\ 4e^-$. But recall that the **anodes** are made of graphite (C).

This means that the **anodes** will react with the oxygen formed: $C\ (s)\ +\ O_2\ (g) \rightarrow CO_2\ (g)$ and have to be replaced constantly.

33.2.2 Carbon Displacement of Iron (III) Oxide

Iron Ore is usually found in the form of Fe_2O_3, sometimes called **haematite**, where "**haema**", greek for "blood", is a reference to its **reddish-brown** colour.

Figure 32.2.2.1 - Haematite

Iron is **lower** in the **Reactivity Series** than **Carbon**, so carbon is able to displace **Iron** from its oxide. In this case, this will give us **Elemental Iron**.

Coke (a form of carbon, shown in *Figure 32.2.2.2* below), is used both as the **reducing agent** for the reaction, and also as a **heat source**, because it has an **exothermic combustion reaction**.

Figure 33.2.2.2 - Coke

Lastly, **limestone**, or calcium carbonate, is also added, to remove impurities.

All of these ingredients are added into a **blast furnace**, where temperatures can reach up to 1500 °C such that the reactions can occur.

The combustion of coke in oxygen forms carbon dioxide, then carbon monoxide:

$$C\ (s)\ +\ O_2\ (g) \rightarrow CO_2\ (g)$$
$$C\ (s)\ +\ CO_2\ (g) \rightarrow 2\ CO\ (g)$$

Now, carbon monoxide serves as a **reducing agent** and the carbon **displaces** iron from its oxide, forming **molten iron** (due to the high temperatures):

$$3\ CO\ (g)\ +\ Fe_2O_3\ (s) \rightarrow 2\ Fe\ (l)\ +\ 3\ CO_2\ (g)$$

However, the iron ore likely contains impurities like sand, SiO_2, which is an **acidic oxide**. Hence, we want to introduce a **basic oxide** to neutralise it. This is why we add $CaCO_3$, limestone.

$$CaCO_3\ (s) \rightarrow CaO\ (l)\ +\ CO_2\ (g)$$
$$CaO\ (l)\ +\ SiO_2\ (l) \rightarrow CaSiO_3\ (l)$$

The product here, $CaSiO_3\ (l)$, which is called **slag**, or **calcium silicate**, will float on top of the molten iron and can be removed.

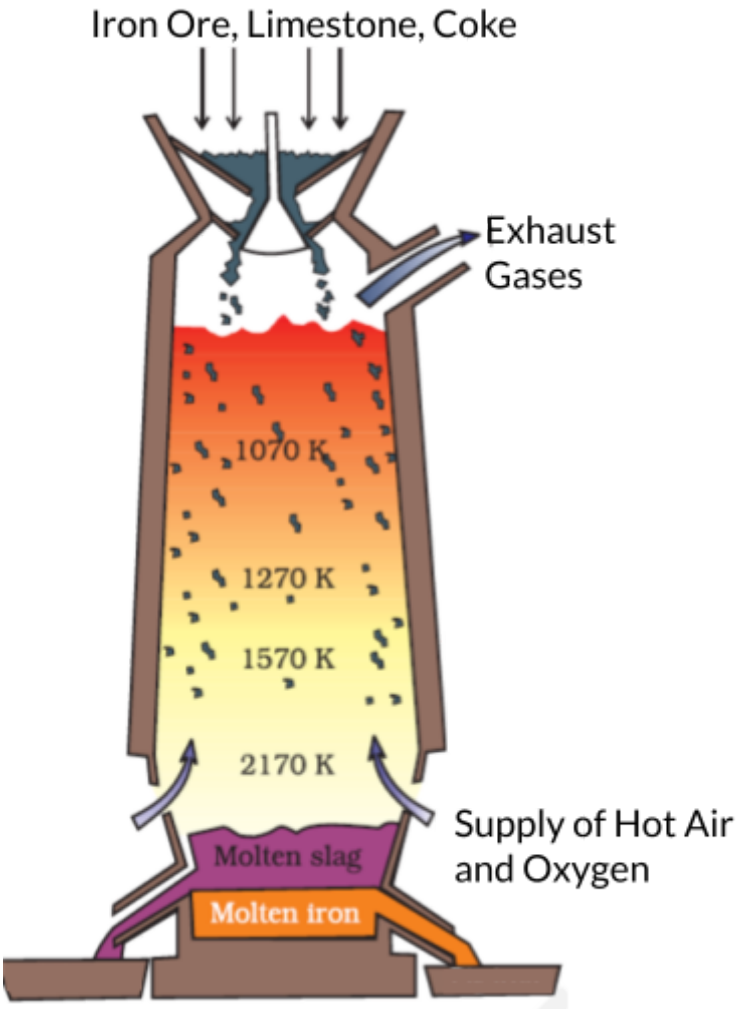


Figure 33.2.2.3 - Blast Furnace

33.3 Rusting of Iron

Rust is a **hydrated form** of **Iron (III) Oxide**, although the number of water molecules in its structure can vary. Hence, the formula for rusting is:

$$4\,Fe\,(s)\,+\,3\,O_2\,(g)\,+\,n\,H_2O\,(l) \rightarrow 2\,Fe_2O_3\cdot n\,H_2O\,(s)$$

Pause and Ponder 33.3

Rusting makes Iron become **flaky** and **brittle**. Why is this so?

[Click to Go to Solution]

33.3.1 Sacrificial Protection from Rust

To prevent rusting, we can attach a more reactive metal such as **zinc** or **magnesium**, called the **sacrificial metal**. This **sacrificial metal** will oxidise first before iron, serving as a temporary protection for the iron, hence the term **sacrificial.**

However, we should not use a metal like **potassium** or **sodium** since they are **overly reactive** as this could both be dangerous due to the **highly exothermic nature** of their oxidisation, and wasteful because they will be oxidised very quickly and the protection will not last long.

For example, in underground steel (which contains iron) pipelines, blocks of **magnesium** are connected to the steel by an insulated wire.

Figure 33.3.1.1 - Sacrificial Protection

This makes the **magnesium** corrode first, before the **iron** does.

FYI: In fact, this is a type of electrolysis reaction!

33.3.2 Barrier Method of Protection from Rust

BL2131 Presentation? hMmMmm

A **barrier**, which can be either **metallic** or **non-metallic**, can be coated on the iron surface to prevent it from coming into direct contact with the moisture and oxygen in the air.

This will prevent the **rusting reaction** from occurring.

Here are some common methods used to create a barrier to prevent the rusting of iron:

Method	Examples of Applications	Pros/Cons
Painting	Cars, Ships, Bridges, etc.	If the paint scratches, rusting will occur under the paint surface.
Oiling or Greasing	Tools and Machine Parts	The protective layer gathers dust and, and must often be replaced.
Galvanising (Zinc Plating)	Baskets, Dustbins, Kitchen Sinks, etc.	The iron doesn't rust even if the zinc is damaged due to the reactivity of zinc.
Tin Plating	Food Cans	If the tin is damaged or scratched, the iron below will rust.
Chrome Plating (Chromium Plating)	Taps, Kettles, Bicycle Handle Bras, etc.	Gives the surface a bright, shiny look.
Stainless Steel	Cutlery, Surgical Instruments, etc.	Stainless Steel contains chromium and nickel, which shield the iron from rust.

Table 33.3.2.1 - Methods of Rust Prevention by Barrier Method

Joel’s Area

He cooked. (kinda)

If it’s not clear, these are extensions.
You do **NOT** need to understand them.

Sorry i was smoking smth pls dont look at this
- Joel

Extension 23.5.1 - Molecular Orbital Theory (The more cursed MO)

Let us consider oxygen. It is a funny blue liquid that is paramagnetic (indicative of unbonded electrons that behaves like a diradical (has two singlet electrons).

Image 23.5.1 - Oxygen (Blue) between 2 magnets

Now, let us inspect oxygen's lewis structure.

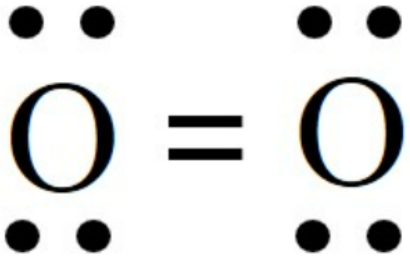


Image 23.5.2 - Dioxygen's Lewis Structure

All the electrons are bonded. So, how does oxygen act as a diradical? Well, Lewis Structures are a lie. Oxygen does not really exist like this. Instead, it is represented by a molecular orbital diagram as follows, which indicates oxygen's true bonding.

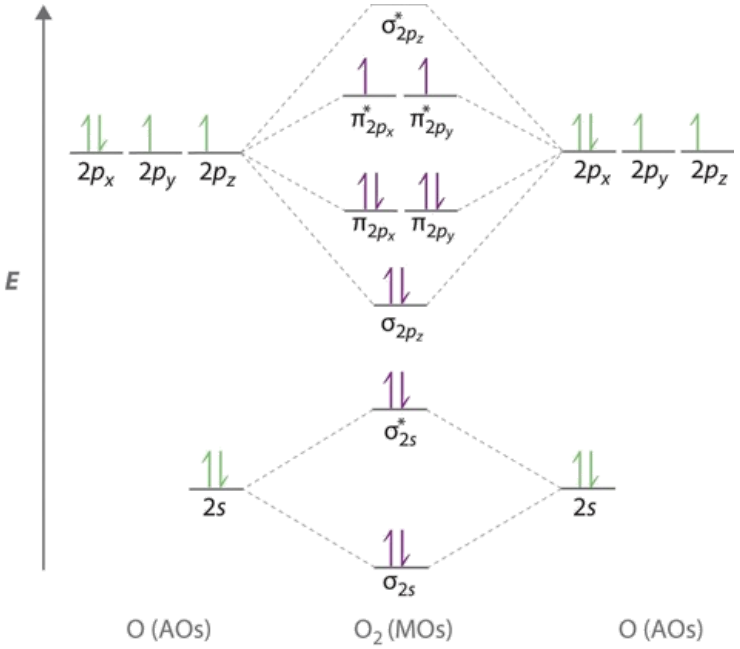


Image 23.5.3: THe MO diagram of dioxygen

Extension 24.3.1 - Bent's Rule and the n>6 Domain Geometry

As is so common in foundational chemistry, VSEPR is a lie. Molecules do not always obey the most ideal structure - not in the way one expects. Take water as an example: from the textbook, you may say with certainty that the bond angle is different due to electron pairs. But this is not exactly true. Let's say we want to explain bond length. How do we do that?

Bent's Rule:
Bent's Rule basically defines any "hybrid" orbital to have greater *s* character when directed at an electropositive substituent (electrons are considered most electropositive). The s orbital has greater electron density; this favours the electropositive atoms. Thus, this would result in the following:

- Bond angle increases between 2 electropositive atoms
- Bond length decreases with greater s character

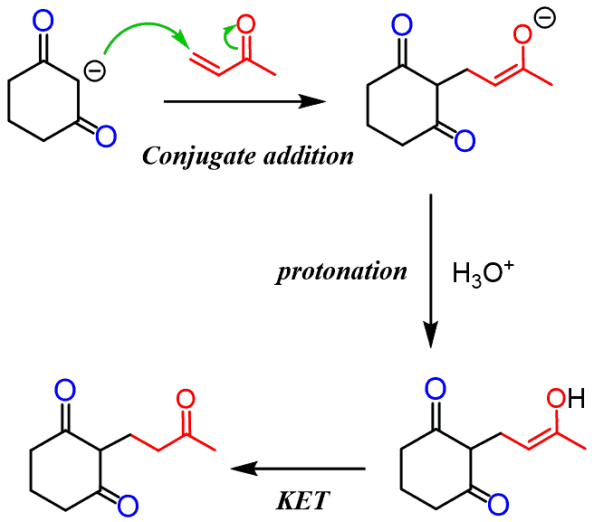
These statements summarise the overall applications of the rule and is not fully detailed.

n>6 Domain Geometry:

7	7	0	AB_7	Pentagonal bipyramidal	72°, 90°	IF_7
				Capped octahedral	---	MoF_7^{-1}
				Capped trigonal prismatic	---	TaF_7^{-2}
	6	1	AB_6E	Pentagonal pyramidal	72°, 90°	IOF_5^{-2}
				Distorted octahedral	---	IF_6^{-1}
8	8	0	AB_8	Square antiprismatic	70.5° 99.6° 109.5°	IF_8^{-1}
				Dodecahedral	---	$Mo(CN)_8^{-4}$
				Bicapped trigonal prismatic	---	ZrF_8^{-4}
9	9	0	AB_9	Tricapped trigonal prismatic	---	TcH_9^{-2}

Extension 29.3.5.1 - Michael Addition
If an alkene and a ketone are present in an α,β-unsaturated compound, nucleophilic addition may not occur at the carbonyl carbon. Instead, it may undergo conjugate addition to the alkene through an enolate intermediate.

Michael Addition Mechanism



But, what is an enol? An enol is a sp^2 carbon with an alcohol attached to it. It is a tautomer of a ketone.

"But, Mr. Joel, can aldehydes enolise too?"
Well, yes, but not really, The aldehyde enol is too unstable to be isolatable as with most enols. Enols have the electrophilic alkene region and a nucleophilic oxygen next to each other, and the polarity is reduced.

"But if enols are so unstable, how can we know they exist?"
You encounter an enol often in organic chemistry able to be bought *en masse*: Phenol.

Pause and Ponder Solutions

Pause and Ponder 5.4 Solution

As covered under **Section 5.3.1: Proton Number**, the proton number defines the identity of an element. Since Isotopes are atoms with the same number of protons, then by definition, it must be the same element.

[Click to go back]

Pause and Ponder 5.4.1 Solution

Since Neon-20 has a relative atomic mass of 20, we take 20 × 90.48%.
Similarly, we repeat for Neon-21 and Neon-22, obtaining 21 × 0.27% and 22 × 9.25% respectively.

Then we sum the numbers and divide by 100%, obtaining:

(20 × 90.48% + 21 × 0.27% + 22 × 9.25%) ÷ 100% = **20.19** (4sf)

Note: We have to divide by 100% and we always round to 4sf, unless otherwise stated in the question.

[Click to go back]

Pause and Ponder 5.5.4a Solution

Aluminium has the electronic configuration 2.8.3. It loses 3 electrons to obtain the noble gas valence shell electronic structure, so it forms the 3+ ion Al^{3+}.

Hence the charge on an Aluminium Ion is **+3**.

In general, atoms in group 1 have 1 electron in their valence shell, so it forms a 1+ ion.
atoms in group 2 have 2 electrons in their valence shell, so it forms a 2+ ion.
atoms in group 13 have 3 electrons in their valence shell, so it forms a 3+ ion.

[Click to go back]

Pause and Ponder 5.5.4b Solution

Chlorine has the electronic configuration 2.8.7. It gains 1 electron to obtain the noble gas valence shell electronic structure, so it forms the 1- ion Cl^-.

Hence the charge on an Chloride Ion is **-1**.

In general, atoms in group 15 have 5 electrons in its valence shell, so it forms a 3- ion.
atoms in group 16 have 6 electrons in its valence shell, so it forms a 2- ion.
atoms in group 17 have 7 electrons in its valence shell, so it forms a 1- ion.

[Click to Go Back]

Pause and Ponder 6.6 Solution

In the liquid state, the ions are free to slide over one another, essentially meaning that they are no longer in a rigid lattice structure.

The next part on aqueous state is not tested in Year 1, and is part of **CM3131 Foundations in Chemistry III**. See Section 26.7: Ion-Dipole Forces.

In the aqueous state, since water is a polar molecule, it has a slightly positively charged end and a slightly negatively charged end, denoted δ+ and δ- respectively. Refer to *Figure 6.6.2.1* below.

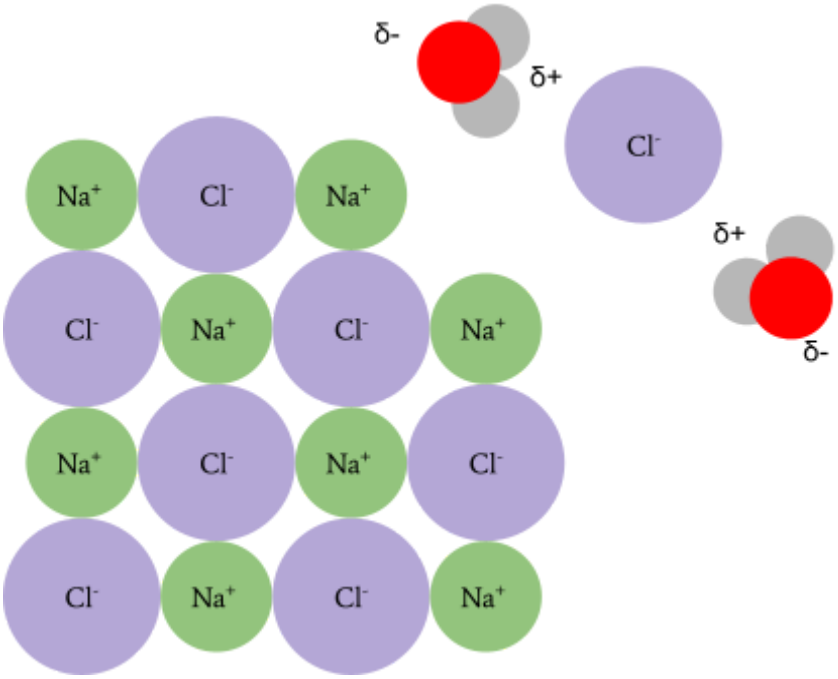


Figure 6.6.2.1: Water dissolving NaCl

As shown above, the water molecules have a slight charge that allows it to attract Chloride ions and Sodium ions away from the lattice, freeing them from the lattice.

[Click to Go Back]

Pause and Ponder 10.5.1 Solution

Let's first write this as a chemical equation.

The reactants are $NaOH\ (aq)$ and $CuCl_2\ (aq)$.

After balancing the equation, we have:

$$2NaOH\ (aq)\ +\ CuCl_2\ (aq) \rightarrow Cu(OH)_2\ (s)\ +\ 2NaCl\ (aq)$$

Let's now break all the (aq) chemicals into its ions.

$$2Na^+(aq)\ +\ 2OH^-\ (aq)\ +\ Cu^{2+}(aq)\ +\ 2Cl^-\ (aq) \rightarrow Cu(OH)_2\ (s)\ +\ 2Na^+(aq)\ +\ 2Cl^-\ (aq)$$

Note: We didn't break $Cu(OH)_2\ (s)$ into Cu^{2+} and $2OH^-$ because it has the state symbol (s) instead of (aq), signifying that dissociation would not occur since it is insoluble.

We can now cancel the spectator ions, which in this case, is $2Na^+$ and $2Cl^-$, and we arrive at the **net ionic equation**.

$$2OH^-\ (aq)\ +\ Cu^{2+}(aq) \rightarrow Cu(OH)_2\ (s)$$

[Click to Go Back]

Pause and Ponder 10.6.2 Solution

a) $Li_2CO_3\ (s) \rightarrow 2Li^+(aq)\ +\ CO_3^{2-}\ (aq)$

$CO_3^{2-}\ (aq)\ +\ H_2O\ (l) \rightarrow HCO_3^-\ (aq)\ +\ OH^-\ (aq)$

b) $NaHCO_3\ (s) \rightarrow Na^+(aq)\ +\ HCO_3^-\ (aq)$

$HCO_3^-\ (aq)\ +\ H_2O\ (l) \rightarrow H_2CO_3\ (aq)\ +\ OH^-\ (aq)$

[Click to Go Back]

Pause and Ponder 11.3 Solution

First, we recall the formula for oxygen gas is O_2, which means there are 2 atoms of oxygen in each oxygen gas particle.

Hence, $No.\ of\ Oxygen\ Atoms = 2 \times No.\ of\ Oxygen\ Gas\ Molecules$
$= 2 \times 5 \times (6.02 \times 10^{23})$
$= 6.02 \times 10^{24}$
$= 6 \times 10^{24}\ (1\ sf)$

[Click to Go Back]

Pause and Ponder 11.6 Solution

The M_r is defined as the sum of A_r's of the component atoms, and the Molar Mass of a compound is the sum of the Molar Masses of the component atoms.

Since the A_r of a component atom is the same as its Molar Mass numerically, the sum of A_r's must be equal to the sum of the Molar Masses, numerically.

Hence, we reach the conclusion.

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Pause and Ponder 11.7 Solution

By definition, 1 mole of oxygen **atoms** contains 6.02×10^{23} oxygen atoms.
$Mass\ of\ 1\ mole\ of\ oxygen\ atoms = 1\ mol \times 16.00\ g/mol$
$= 16.00\ g$

Hence $Mass\ of\ 1\ oxygen\ atom = 16.00\ g \div (6.02 \times 10^{23})$
$= 2.66 \times 10^{-23} g\ (3\ sf)$

Since oxygen gas has formula O_2, the mass of 1 molecule of oxygen gas is the mass of 2 oxygen atoms.

Hence the answer is $2.66 \times 10^{-23} g \times 2 = 5.32 \times 10^{-23} g$.

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Pause and Ponder 11.9 Solution

As usual, we first consider the mass of each element in 100g of Albite.

Element	Percentage Composition	Mass in 100g of Compound (g)
Na	8.7668%	8.7668
Al	10.288%	10.288
Si	32.134%	32.134
O	48.810%	48.810

Table 11.9.5

We then consider their molar masses and find the number of moles in 100g of Albite.

Element	Mass in 100g of Compound (g)	Molar Mass (g/mol)	Number of moles in 100g of compound (mol)
Na	8.7668	22.99	0.3813
Al	10.288	26.98	0.3813
Si	32.134	28.09	1.144
O	48.810	16.00	3.051

Table 11.9.6

Then, we consider the ratio of the number of moles.

Na	Al	Si	O
0.3813	0.3813	1.144	3.051

1	1	3.0002	8.0016	*(divide by 0.3813)*
1	1	3	8	*(round)*

Table 11.9.7

Hence, we can determine the empirical formula of Albite to be $NaAlSi_3O_8$.

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Pause and Ponder 13.2.2 Solution

We know that:

$$Mass\ Concentration = \frac{Mass\ of\ Solute}{Volume\ of\ Solution}$$

$$Molarity = \frac{No.\ of\ moles\ of\ Solute}{Volume\ of\ Solution}$$

$$Molar\ Mass = \frac{Mass\ of\ Solute}{No.\ of\ moles\ of\ Solute}$$

Hence, $Molarity = \frac{Mass\ Concentration}{Molar\ Mass}$.

[Click to Go Back]

Pause and Ponder 13.3 Solution

Firstly, we calculate the amount of sucrose present.

$$\begin{aligned} No.\ of\ moles\ of\ sucrose\ in\ 6.9M\ solution &= Molarity \times Volume \\ &= 6.9\ mol/L \times 0.420\ L \\ &= 2.89\ mol \end{aligned}$$

Since dilution does not affect the number of moles of sucrose, we can determine that the number of moles of sucrose in the $2.0M$ solution is 2.89 mol.

$$\begin{aligned} Volume\ of\ 2M\ solution &= \frac{No.\ of\ Moles}{Molarity} \\ &= 2.89\ mol \div 2.0\ mol/L \\ &= 1.44L \end{aligned}$$

$$Volume\ of\ water\ added = 1440ml - 420ml$$

$$= 1020ml$$

$$No.\ of\ moles\ of\ water = \frac{Volume}{Molar\ Volume}$$
$$= \frac{1020\ cm^3}{18.0\ cm^3/mol}$$
$$= 56.7\ mol\ (3\ sf)$$

[Click to Go Back]

Pause and Ponder 15.5.2 Solution

This is because of the methyl orange indicator that we added. We don't know if it will cause the crystals to come out as impure after crystallisation, and we are not entirely sure how it will affect the crystallisation process.

Hence, we do not "take the risk" and perform crystallisation straight away.

Instead, we use a new setup.

[Click to Go Back]

Pause and Ponder 17.5 Solution

Silicon Dioxide is a **poor conductor of electricity** as all the valence electrons are used for covalent bonding, and hence there are no delocalised electrons and hence mobile charge carriers.

Silicon Dioxide has a **high melting/boiling point** as the strong covalent bonds between the Silicon and Oxygen atoms require a lot of energy to overcome in its **Giant Covalent Structure**.

[Click to Go Back]

Pause and Ponder 18.3.3 Solution

Since F is more electronegative (4.0) than O (3.5), the electrons are pulled closer to F.

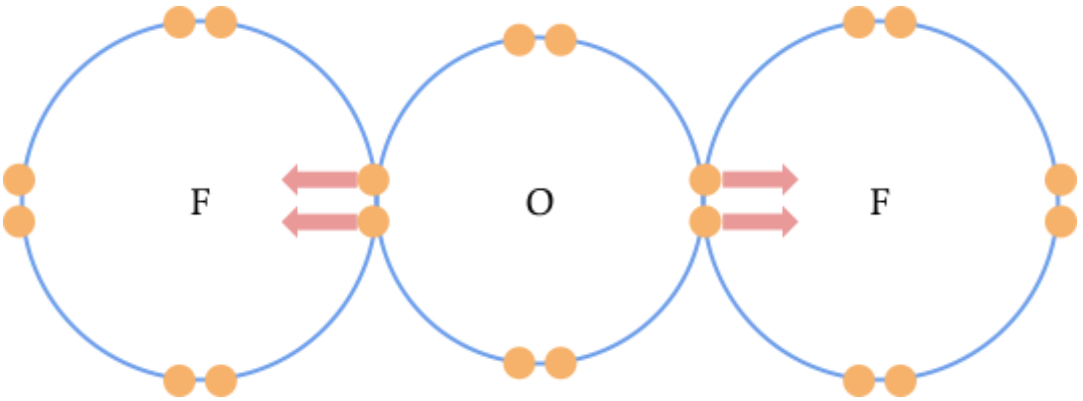

Figure 18.3.3.2

The 4 electrons shared between O and F_2 are pulled closer to the F atoms, and of those 4 electrons, 2 originally belonged to O.

Hence, we say O lost control of 2 electrons, so it has an oxidation number of **+2**.

However, in general, since O has a relatively high electronegativity number, it normally is the one that takes control of the electrons, so normally it has an oxidation number **-2**, it is only when a more electronegative element (F) is present does it have an oxidation number **+2**.

[Click to Go Back]

Pause and Ponder 20.6 Solution

The equation is: $Ca(OH)_2\ (aq)\ +\ CO_2\ (g) \rightarrow CaCO_3\ (s)\ +\ H_2O\ (l)$, so the precipitate formed is $CaCO_3\ (s)$.

[Click to Go Back]

Pause and Ponder 20.7 Solution

If there are 12 H_2O molecules in 2 formula units of hydrated $CoCl_2$, there is then 6 H_2O molecules in each formula unit, making the formula for hydrated $CoCl_2$ as $CoCl_2.6H_2O$.

Hence, the reaction is:

$$CoCl_2 + 6H_2O \rightarrow CoCl_2.6H_2O$$

(blue) *(pink)*

Now wouldn't that make for a great gender reveal? Maybe that's the real thing to Pause and Ponder about…

[Click to Go Back]

Pause and Ponder 21.4 Solution

Since we know the oxidation number of manganese is +4 from the (IV) after it, we can use the rule that the sum of oxidation numbers in a compound must be 0 to figure out the number of oxygen atoms per formula unit.

We know that oxygen has an oxidation number -2 when not combined with fluorine. Hence, we let the number of oxygen atoms for every 1 atom of manganese be x, and we have:

$$\begin{aligned} 4 + x(-2) &= 0 \\ x &= 2 \end{aligned}$$

Hence, there are 2 oxygen atoms for every 1 manganese atom in Manganese (IV) Oxide, and hence the empirical formula of Manganese (IV) Oxide must be MnO_2.

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Pause and Ponder 22.2 Solution

Let's consider the formula $\Delta H_{LE} \propto \frac{q_1 q_2}{r}$.

As we know from MA1132, since this is in direct proportion, we can write it as $\Delta H_{LE} = \frac{kq_1 q_2}{r}$ for some value of k.

Then, we know since r measures a distance, it must be strictly positive.

On the other hand, since this is an ionic substance, the charge of one ion should be strictly positive, while the other is strictly negative.

Hence, $\frac{q_1 q_2}{r}$ is **strictly negative**.

And since the ΔH_{LE} value of $NaCl$ is negative, we deduce that the value of k must be **strictly positive**.

Hence, using $\Delta H_{LE} = \frac{kq_1q_2}{r}$, we get that ΔH_{LE} is **strictly negative**.

FYI: This value of k is known as the Coulomb Constant, and is denoted with k_e. It has the value $k_e = 8.99 \times 10^9\ kg\ m^3\ s^{-4}\ A^{-2}$ in Base SI Units.

[Click to Go Back]

Pause and Ponder 23.4.1 Solution

The nitrosonium ion, NO^+, has two resonance structures.

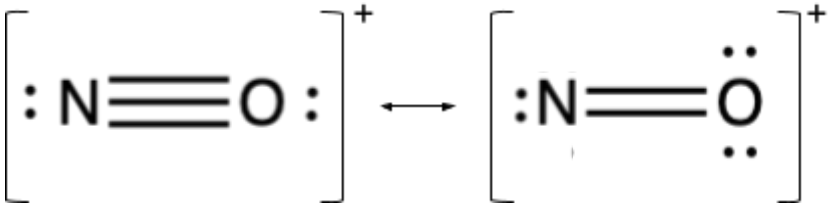

Figure 23.4.1.2

We see in the left structure, N has formal charge 0, while O has formal charge +1;
in the right structure, N has formal charge +1, while O has formal charge 0.

We may verify that in both cases, the sum of the formal charge is +1, matching the +1 charge of the Nitrosonium Ion, and is the minimal possible. Theoretically, the right resonance structure is more optimal, since O is more electronegative. Experimentally, however, both resonance structures are valid. Specifically, in the left resonance structure, one of the bonds is a dative bond with oxygen donating one of its lone pairs.

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Pause and Ponder 24.6 Solution

a) Its chemical formula is BF_3.

b) The difference in electronegativity is 3.98 - 2.04 = 1.94. This implies polar covalent bonding.

c) The Lewis Structure of BF_3 is shown in *Figure 24.6.4* below, with the dipole moments marked by red arrows.

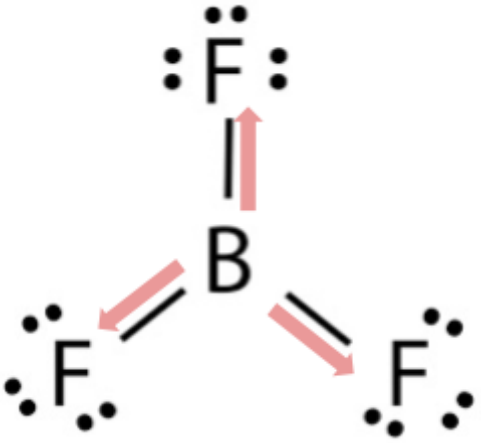

Figure 24.6.4 - Boron Trifluoride Lewis Structure

Despite having 3 polar bonds, the partial charges are distributed **symmetrically** and the dipole moments induced by these bonds **cancel each other out**, seeing as the vectors will sum to 0, by vector addition. Hence, it has no net dipole moment.

d) It should have high reactivity with ammonia. Boron in BF_3 has an incomplete octet, with only 6 valence electrons, and will readily accept a lone pair. On the other hand, Nitrogen in NH_3 has a lone pair, and NH_3 will tend to donate a lone pair to BF_3 and form a dative bond, suggesting high reactivity between BF_3 and NH_3.

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Pause and Ponder 27.1.5 Solution

Alkanes are **non-polar covalent** molecules, because of their symmetric structure and the small electronegativity differences between atoms leading to non-polar bonds. Hence, the only intermolecular force of attraction in an alkane is the **London Dispersion Forces**.

As an alkane gets longer, its **polarisability increases**, and its **surface area increases**. Hence, the strength of **London Dispersion Forces** between alkane molecules increase, and more energy is required to overcome these intermolecular forces of attraction, hence its boiling and melting points both increase.

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Pause and Ponder 27.2.5 Solution

We let the general equation be $C_xH_y + p\,O_2 \rightarrow q\,CO_2 + r\,H_2O$.

Because we have 3 unknowns ($p,\ q,\ r$) and we have 3 equations (conservation of the number of atoms of $C,\ H,\ O$), we can solve this simply with simultaneous equations!

Comparing the number of atoms of:

- C: $x = q$
- H: $y = 2r$
- O: $2p = 2q + r$ ---------- (1)

Now, since $y = 2r, r = \frac{y}{2}$. Plugging this and $x = q$ into (1) we get: $2p = 2x + \frac{y}{2}$.

Rearranging, $p = x + \frac{y}{4}$.

Now, we can substitute back into the original equation. This is our general equation:

$$C_x H_y + (x + \tfrac{y}{4})\, O_2 \rightarrow x\, CO_2 + \tfrac{y}{2}\, H_2O$$

Remember to multiply by 2 or by 4 if the coefficient of O_2 is fractional.

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Pause and Ponder 33.3 Solution

Rust, or **Iron (III) Oxide**, is an **ionic compound**. When a force is applied onto a layer of an ionic compound, the ions of the same charge are brought together like in *Figure 33.3.1* below.

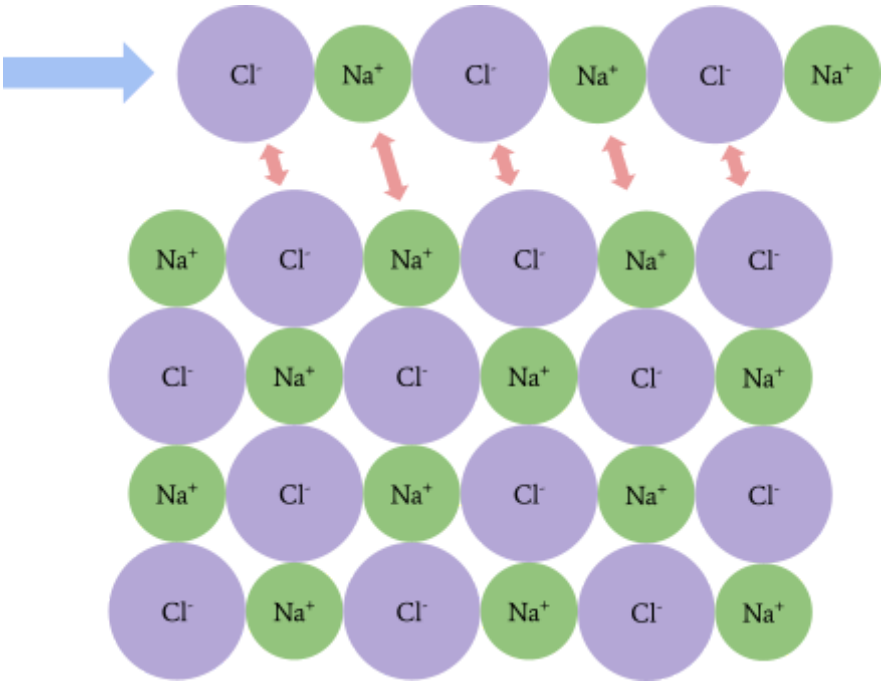

Figure 33.3.1 - Force Applied on an Ionic Compound

Similarly for Iron (III) Oxide, when a force is applied to one of its layers, it causes the Fe^{3+} ions to be pushed close together to each other, and the O^{2-} ions to be pushed closer to each other.

This creates a repulsive force between these ions with the same charge, and hence this causes the ionic lattice layer to break apart. This makes **rust** very **flaky** and **brittle**.

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Sources and Credits

Pictorial Sources

https://artsexperiments.withgoogle.com/periodic-table/ (*Figure 2.1.1.1* and *Figure 2.1.3.1*)
https://learnwithdrscott.com/diatomic-elements/ (*Figure 3.5.1*)
https://chemix.org/ (*Figure 4.5.1.2*, *Figure 4.5.1.3*, *Figure 4.6.1.1*, *Figure 4.8.1.1*, *Figure 4.10.1.1*, *Figure 4.11.1.1.1*, *Figure 4.11.2.1.1*, *Figure 4.12.1.1 and Figure 4.14.1.1*)
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http://www.nygh.sg/interdisciplinary/sec_one/the_built_environment/reactions_of_metals_ans.pdf, p. 3 (*Figure 8.6.1.1*)
https://igcse-chemistry-2017.blogspot.com/2017/07/244-describe-tests-for-these-gases.html (*Figure 10.7.1*)
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Much appreciation to all above sources for making this set of notes possible.

Thank You!

As always, thanks for making it to the end of my notes!
Good luck for your Chemistry Examinations/Assessments and...

Have a nice day! :D

Love Chemistry. Question Physics. Fear Biology.

❤️

640 Days
100 Hours of Work
19 Feb 2022 - 21 Nov 2023
Goodnight.