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Topical Notes for Cambridge IGCSE™ Chemistry (0620)

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Version 1

Chapter 1

1.1 States of matter

The 3 states of matter

Matter is a physical substance that occupies space and has mass.

There are many millions of different substances known to us and all of them can be categorized as solid, liquid or gas.

These are known as the three states of matter.

There are more states of matter other than these three but they are not in the syllabus so they will not be covered.

Property	Solid	Liquid	Gas
Volume	Fixed volume	Fixed volume	Indefinite volume
Shape	Fixed shape	Indefinite Shape Takes shape of the container stored in	Indefinite Shape Takes shape of the container stored in
Can Compress?	×	🔽 but very little	

Solids

In solids, particles are packed close together. They cannot move and can only vibrated in a fixed position.

Solids also have strong intermolecular forces due to which along with the particles being packed close together result in a fixed shape and volume.

Solids expand only a little bit when heated.

Liquids

Particles in liquids are slightly further apart. They also have the ability to move past each other. Liquids have slightly weaker intermolecular forces, allowing them to flow and take the shape of their container.

Liquids expand on heating more than solids, but less than gases.

Gases

Particles in gases are much further apart and usually don't touch each other.

Gases have really weak intermolecular forces and large inter-particle distances, due to which gases

can flow easily and can take the shape of the container.

Gases also expand the most when heated.



Fig. 1: Arrangement of particles in solids, liquids and gases

Changes of state

When a solid turns into a liquid, we call it melting. When a liquid turns into a gas, we call it boiling/evaporation.

When a gas turns into a liquid, we call it condensation When a liquid turns into a solid, we call it freezing.

The diagram below summarizes these changes of state.



Fig. 2: Changes of state

Kinetic particle theory

The kinetic particle theory states that all matter is made up of tiny particles which are always moving randomly

When these particles have higher kinetic energy, they move faster and vice versa.

Heating

Heating of the object means the particles inside the object have a higher average kinetic energy and therefore move faster, once these particles are moving fast enough, they are able to break the **intermolecular forces** holding them together, causing a change in state.

However, not all particles escape at the same time, and some may escape first while others may escape later.

Cooling

Cooling reduces the average kinetic energy of the particles in an object, this causes the **intermolecular forces** to be able to hold the particles back into a relatively fixed position.

Heating and cooling curves

Heating curves show how a substance's temperature increases as it's heated, while cooling curves show how it decreases as it cools, with flat portions on the graph indicating phase changes. (see the graphs below)



Fig. 3: Diagram showing the heating curve of water from -15°C to 110°C

Effect of pressure and temperature on gases

Temperature

Higher temperature causes gases to expand and increase in volume.

This happens because when the temperature increases, so does the average kinetic energy of the particles of the object, making them move faster. When these particles collide with the walls of the container, they exert greater pressure on the walls and start to occupy more space causing the gas to expand.

Pressure

Higher pressure causes the volume of gases to decrease as it compresses the gas particles together.

This reduces the space between particles, leading to a decrease in volume.

1.2 Diffusion

As we read earlier, particles are always moving. Diffusion is a result of this movement of particles.

Diffusion occurs because particles move randomly and collide with each other, spreading out from areas of higher concentration to lower concentration, equalizing the distribution over time.

Effect of relative molecular mass on diffusion of a gas

Relative molecular mass is the sum of the masses of the atoms in a molecule, giving the molecule's weight.

You'll read more about relative molecular mass in chapter 3.

The higher the relative mollecular mass of a gas, the slower it will diffuse and it will move slower than the gases which are lighter.

Chapter 2

2.1 Elements, compounds and mixtures

Elements

- Elements consist of only one type of atom
- They cannot be broken down into simpler substances by chemical means
- Each element has it's own symbol (for example, H for hydrogen, O for oxygen)

All elements can be classified as either metals or non-metals.

Metals usually have high melting and boiling points while non-metals have low melting and boiling points.

Compounds

- Compounds are formed when two or more different elements **chemically** combine in fixed ratios to form a new substance.
- Their properties are different than that of the compounds that they are made from
- Compounds can only be separated into their constituent elements through chemical reactions.
- Forming a compound is a chemical change
- Examples include: H_2SO_4 , HCl, CO_2 etc.

Mixtures

- Mixtures are combinations of 2 or more substances (elements or compounds) not chemically bonded together.
- Mixtures can be separated by physical methods, such as filtration or distillation, without breaking chemical bonds.
- The properties of mixtures are based on the proportions of its constituents.
- Examples include: sea water, air, alloys (brass)

Mixture	Compound
It contains two or more substances	It is only one substance
No chemical change when mixture is formed	Formation of new substances involves chemical change
The properties of the mixture are of those of the individual elements/compounds in it	The properties of the new substance are very different from either of the elements used to form the substance

In chapter one, you learned that matter is made up of very tiny particles, these particles are known as atoms. Atoms are also known as the smallest particles.

Each atom contains a nucleus and electrons around the nucleus. The nucleus holds protons and neutrons.

These electrons are placed in shells as shown by the diagram below.



Fig. 1: Atomic structure of a carbon atom

As you can tell by the diagram above, electrons are negative while protons are positive. Neutrons however are neutral and have no charge.

Particle	Symbol	Relative mass	Charge
Proton	р	1	+1

Particle	Symbol	Relative mass	Charge
Neutron	n	1	0
Electron	е	1/1837	-1

Atomic number

The atomic number (also known as the proton number) of an element is the number of protons in one atom of that element.

For example, in the diagram above (figure 1), carbon has 6 protons, which means its atomic number is 6.

The atomic number is also known as Z.

Mass number

The mass number (also known as nucleon number) of an element is the total number of protons and neutrons in the nucleus of one atom of that element.

For example, in the diagram above (figure 1), carbon has 6 protons and 6 neutrons, which means its mass number is 12.

The mass number is also known as A.

Electronic configuration

Electronic configuration is the arrangement of electrons in the electron shells of an atom which tells us about the distribution of electrons in that atom.

The first shell cannot have more than 2 electrons and the 2nd and 3rd shell cannot have more than 8 electrons.

For example, in figure 1 (yes again), carbon has 6 electrons, 2 of which are in the first shell and 4 of which in the second shell. Therefore the electronic configuration of the atom would be 2, 4.

Remember

- Group VIII elements are known as noble gases and have a full outer shell, meaning they are unreactive.
- The number of outer shell electrons is equal to the group number in Groups I to VII.
- The number of electron shells is equal to the period number

2.3 Isotopes

Isotopes are different atoms of the same element that have the **same number of protons** but **different numbers of neutrons**.

AX

Isotopes of the same element have the same chemical properties because they have the **same number of electrons** and therefore the **same electronic configuration**.

To write isotopes, we use the following notation:

where A is the mass number,

Z is the proton number and

X is the symbol of the element itself.

For example, we can write carbon as:

${}^{12}_{6}C$

Calculate relative atomic mass of an element from its isotopes

To calculate the relative atomic mass (A_r) of an element from its isotopes, we can 2 things:

- the relative mass of its isotopes
- abundance of each isotope (percentage)

Using this information, we can use the following formula and calculate the relative atomic mass of an element:

$$A_r = (ext{ mass of isotope } 1 imes rac{ ext{abundance}}{100}) + (ext{ mass of isotope } 2 imes rac{ ext{abundance}}{100})$$

2.4 lons and ionic bonds

lons

lons are atoms or molecules that have a charge (are not neutral; number of protons is not equal to number of electrons) due to the gain or loss of electrons.

Remember

Positive ions are known as cations.

They are formed when atoms lose electrons, resulting in a positive charge due to more protons than electrons.

Negative ions are known as anions.

They are formed when atoms gain electrons, resulting in a negative charge due to more electrons than protons.

Ionic Bonding

An ionic bond is a strong electrostatic attraction between oppositely charged ions.

lonic bonds are usually found in compounds that contain metals combined with non-metals.

Electrons are transferred from the metal atoms to the non-metal atoms.

This makes the atom stable because of their full outer shell.

We use a dot and cross diagram to show these bonds





In the diagram above, one electron is transferred from the sodium atom to the chlorine atom, forming sodium ion and chloride ion which together form sodium chloride, an ionic compound.

Structure of an ionic compound

lonic compounds have a repeating, closely packed structure where positively charged ions (cations) and negatively charged ions (anions) alternate, and they are held together by strong electrostatic forces.

This arrangement is known as a lattice. Many millions of ions would be arranged in this way to make up the giant ionic lattice structure.

Properties of an ionic compound

Properties of ionic compounds include:

- high melting points and boiling points
- good electrical conductivity when aqueous or molten and poor when solid
- usually solid at room temperature

High melting points and boiling points

These compounds have strong electrostatic forces of attraction between oppositely charged ions throughout the entire lattice, which require a lot of energy to break/overcome which leads to higher melting and boiling points

Good electrical conductivity when aqueous or molten, poor when solid

First we need to know what conducts electricity and what is electricity. In one simple line, electricity is the flow of electrons and therefore is conducted by either electrons or ions.

Solid ionic compounds are held firmly in place in the lattice and hence ions in them cannot move, making them a bad conductor.

Meanwhile when this ionic compound is dissolved in water or molten, the ions in the substance disassociate and are able to move, which means that they can now conduct electricity.

Forming ionic compounds

To form a stable compound, the charges on the opposing ions must always cancel out each other so that the overall charge of the compound is 0.

For example, One Na^+ ion + one Cl^- ion $\longrightarrow NaCl$ and similarly: One Mg^{2+} ion + two Cl^- ions $\longrightarrow MgCl_2$

To figure out the charge of the ion of an element, we can refer to it's oxidation number.

Transition elements have multiple oxidation numbers, and they have names like copper(ii) which means that its oxidation number is 2

To find out the oxidation number for other elements however, we use the number of valence electrons (which is simply the number of electrons in the last shell of an atom of that element).

If number of valence electrons (electrons in last shell) is more than 4 (so 5, 6, 7, 8) then we subtract the number of valence electrons by 8. (valence electrons - 8)

As we read earlier in 2.2 that the number of outer shell electrons is equal to the group number, which means that Group I elements will have an oxidation number of 1, Group II will have an oxidation number of 2 and Group VI will have a oxidation number of -2.

Usually, metals have positive oxidation numbers and non-metals have negative oxidation numbers.

2.5 Simple molecules and covalent bonds

Covalent bonds are bonds in which atoms share pairs of electrons to form stable compounds.

A simple example would be Hydrogen gas (H_2)

In hydrogen gas, 2 hydrogen atoms share electrons to become stable and form a compound. (see diagram below)



Fig. 3: Electronic configuration of hydrogen gas

Similarly, oxygen gas (O_2) also has covalent bonds along with:

- chlorine gas (Cl_2)
- water (H_2O)
- methane (CH₄)
- ammonia (NH₃)
- hydrochloric acid (HCl)
- metanol (CH_3OH)
- ethylene (C_2H_4)
- carbon dioxide (CO₂) and many more.



Fig. 4: Structures of oxygen, carbon dioxide and methanol

Properties of simple covalent compounds

Properties of simple covalent compounds include:

- Low melting points and boiling points
- Poor electrical conductivity

Low melting points and boiling points

Simple covalent compounds have weak intermolecular forces (forces that hold atoms together within a molecule) which means that less energy is required to break these forces which means that they will have low melting and boiling points.

Poor electrical conductivity

As you read earlier, to conduct electricity, we need either free ions or delocalized electrons and covalent compounds have none of these therefore they cannot conduct electricity.

2.6 Giant covalent structures

Giant covalent structures are basically "networks" of atoms bonded together by strong covalent bonds, examples include diamond, graphite and silicon(IV) oxide.

Structure of graphite

Graphite consists of **many layers** carbon atoms arranged in a **hexagonal lattice** structure where each carbon atom forms three covalent bonds with three neighboring carbon atoms in the same layer.

These layers have weak forces of attraction holding them together, allowing the layers to slide over each other.

Graphite is a non-metal, however, it does conduct electricity. As you read above, carbon atoms form only **three covalent bonds** which means that an unbonded electron is available on each carbon atom. These spare/delocalized electrons between the layers allow graphite to conduct electricity.

Graphite is used as a lubricant since it is slippery thanks to the weak forces holding the layers together.

It is also used as an electrode because of its good electrical conductivity and how cheap it is.

Structure of diamond

Diamond is composed of a network of carbon atoms, each bonded to four other carbon atoms through strong covalent bonds in a **tetrahedral** arrangement.

This network results in a very hard and rigid structure, making diamond one of the hardest substances known.

Due to its strong covalent bonds, diamond has an exceptionally high melting point

Diamond does not conduct electricity because it does not have free electrons or ions.

Diamond is used as a cutting tool (it is really hard) as well as in jewellery (it is shiny and looks nice)

Structure of silicon(IV) oxide (SiO₂)

Silicon(IV) oxide, SiO2, is also known as silica or quartz (calm your horses fellow minecraft players).

It has a giant covalent structure where each silicon atom is bonded to four oxygen atoms through strong covalent bonds in a **tetrahedral** arrangement.

These tetrahedral units form a network throughout the crystal lattice, creating a hard and rigid structure with a high melting point.

Silicon dioxide is also not a conductor of electricity as there are no free ions or electrons in its structure.

Now, the structure may seem similar to that of diamond, except for a key difference, diamond is made up of entirely carbon while silicon(IV) oxide is made up of both silicon and oxygen atoms and does not contain any carbon.

There are also many similiarities between their structures such as:

- Hardness
- High melting points
- Do not conduct electricity

2.7 Metallic bonding

Metallic bonding is the **electrostatic attraction** between the **positive ions** in a giant metallic lattice and a sea of **delocalised**, **mobile electrons**.



Fig. 5: Metallic bonding

Properties of metals

Metals have good electrical conductivity, and are ductile (drawn into wires) and malleable (hammered into thin sheets).

Good electrical conductivity

As we read earlier, to conduct electricity we need free electrons and metals do indeed have free electrons which means that they can conduct electricity.

Malleability and ductility

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Metal bonds are indeed rigid but they aren't fixed unlike bonds in diamond which means that when a force is applied, layers of metal ions can easily slide past each other without breaking the metallic bonds allowing them to be drawn into wires (ductility) or thin sheets (malleability).

3.1 Formulae

Molecular formula: The molecular formula of a compound tells you how many and what kinds of atoms are in one molecule of that substance.

Empirical formula: The empirical formula of a compound tells you the simplest ratio of atoms in one molecule of that substance.

you can figure out the compound's molecular formula by counting the atoms in it's structure using which you can figure out it's empirical formula by reducing the ratios to their simplest whole numbers.



Fig. 1: Displayed formula of ethane ()NEUROtiker, Public domain, via Wikimedia Commons

For example, we can count that there are 2 carbon atoms and 6 hydrogen atoms in the figure above, which means it's molecular formula would be C_2H_6 . Now to get the empirical formula, we can simply simplify this to get CH_3 .

Similarly with ionic compounds, you can deduce the formula by balancing the charges on the ions in a model or diagram, ensuring that the total charge is neutral (0), to find the compound's formula.

Constructing equations

Word equations and symbol equations are used to describe chemical reactions, including the reactants, products, and their physical states.

A simple example of a word equation would be:

Hydrogen gas + Oxygen gas \longrightarrow Water vapor

While the symbol equation for the same reaction would be:

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$$\mathrm{H}_{2}\left(\mathrm{g}
ight)+\mathrm{O}_{2}\left(\mathrm{g}
ight)\longrightarrow\mathrm{H}_{2}\mathrm{O}\left(\mathrm{g}
ight)$$

Notice the (g) next to the symbols? This is known as a state symbol which tells us about the physical state of the substance.

(g) means gas
 (1) means liquid
 (s) means solid
 (aq) means aqueous which means that the substance is dissolved in water

Constructing ionic equations

Now let's say we have this equation right here:

 $NaCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

which is a double displacement reaction (two compounds exchange ions to form two new compounds) in which sodium chloride and silver nitrate react to form silver chloride and sodium nitrate.

From the equation above, we can tell that the ions undergoing change are chloride ions (from NaCl) and silver ions (from $AgNO_3$).

So for this reaction, we can write that the ionic equation for this reaction is:

 $\mathrm{Cl}^{-}\left(\mathrm{aq}\right) + \mathrm{Ag}^{+}\left(\mathrm{aq}\right) \longrightarrow \mathrm{AgCl}\left(\mathrm{s}\right)$

3.2 Relative masses of atoms and molecules

There are 118 different known elements in the periodic table. All have a different mass. The masses of atoms of these elements are really, really small and hard to work with which is why we use something called **relative atomic mass** (A_r).

In this scale, an atom of carbon is given a relative atomic mass of 12.

In simple words, relative atomic mass, A_r , is the average mass of the isotopes of an element compared to 1/12th of the mass of an atom of carbon-12.

Relative molecular mass, M_r, is the sum of the relative atomic masses of all the atoms in a compound.

For example we have this equation:

$$\mathrm{C} + \mathrm{O}_2 \longrightarrow \mathrm{CO}_2$$

The relative atomic mass of C is 12 and O is 16.

Now there are 2 oxygen atoms, which means that the relative atomic mass of O_2 will be 32.

Using this, we can find the relative molecular mass of CO_2 by adding 12 and 32 which equals to 44. Therefore the mass of CO_2 is 44.

Now we can add units to this to make these into actual masses, so if 12g of carbon was burned in air (another way of saying reacted with excess oxygen; don't get confused with decomposition) and 44g of CO_2 was produced.

Using this statement, we can figure out that when 6g of carbon is burned in air, 22g of $\rm CO_2$ will be produced.

3.3 The mole and the Avogadro constant

A mole is a unit of measurement used to express the quantity of atoms, molecules, ions, or particles in a substance.

One mole contains 6.02×10^{23} particles (atoms, ions, molecules). This number is known as the **Avogadro's constant**.

Calculating moles

To calculate moles, we can use the formula:

 $moles = \frac{mass}{molar mass}$

Where moles is in mol, mass is in g and molar mass is in g/mol.

We can further multiply this number by 6.02×10^{23} to find out the number of particles in the substance.

For example, to calculate the number of particles in 24g of carbon, we need to do the following:

The mass given is 24g, and the molar mass of carbon is 12 g/mol. Now we need to use the equation above to calculate the number of moles:

moles =
$$\frac{\text{mass}}{\text{molar mass}} = \frac{24}{12} = 2$$

which comes out to 2 mol.

Now we can simply multiply this number by $6.02 imes 10^{23}$ to find the number of particles:

 $2\times 6.02\times 10^{23}=1.204\times 10^{24}$ particles

Gases and moles

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One mole of any substance takes up a volume of 24 dm³ at room temperature and pressure (r.t.p.).

Using this statement, we can deduce the following formula:

number of moles in a gas =
$$\frac{\text{volume of gas in dm}^3}{24 \text{ dm}^3}$$

Remember, $1 \text{ dm}^3 = 1000 \text{ cm}^3$.

Moles and solutions

Sometimes concentration is measured in ${\rm ~g/dm}^3$ but more often it is measured in ${\rm ~mol/dm}^3$.

concentration (in mol/dm³) =
$$\frac{\text{number of moles}}{\text{volume (in dm3)}}$$

The same equation above can be rearranged to find out the number of moles:

number of moles = concentration (in mol/dm³) \times volume (in dm³)

Tip

An easy way to remember formulas is using the units. For example the unit of concentration is mol/dm³, and the formula to calculate concentration is moles/volume. The unit of moles is mol and the unit of volume is dm³

Calculating formulae

To get the molecular formula of the compound, we use the formula below

$$(\frac{\text{relative molecular mass of compound}}{\text{mass of empirical formula}}) \times \text{ empirical formula}$$

You read about the molecular and empirical formula of ethane earlier. Ethane's molecular formula is C_2H_6 while it's empirical formula is CH_3 .

Now we can find the relative molecular mass of the compound by adding the relative atomic mass of all the atoms together which is:

 $(2 \times 12) + (6 \times 1) = 30$

And the relative molecular mass of the empirical formula would be:

$$(1\times 12)+(3\times 1)=15$$

Now we can use the formula above to find the molecular formula of ethane:

$$rac{30}{15} imes \mathrm{CH}_3 = 2 imes \mathrm{CH}_3 = \mathrm{C}_2\mathrm{H}_6$$

Moles and chemical equations

The law of conservation of mass states that mass can neither be created nor destroyed. This means that the mass on the left side of the equation has to be equal to the right side of the equation.

Percentage Yield

Chemical reactions are not 100% efficient, they do not always produce the expected amount of product. The percentage yield of the reaction is based on the amount of product that is actually produced against what should have been produced if the reaction was 100% efficient. It can be calculated using the formula below:

percentage yield =
$$(\frac{\text{actual yield}}{\text{theoretical yield}}) \times 100$$

Percentage composition

Percentage composition is used to describe the percentage by mass of each element in the compound. The formula to calculate it is:

Percentage of element =
$$(\frac{\text{Mass of element}}{\text{Total mass of compound}}) \times 100$$

Percentage purity

The purity of a substance is very important. It can be obtained using the formula below:

$$Percentage Purity = (\frac{Mass of the pure product}{Mass of the impure product obtained}) \times 100$$

The reactants that are left over at the end of reaction are known as excess reactants

4.1 Electrolysis

Electrolysis is the **decomposition** of an **ionic compound**, when **molten or in aqueous solution**, **by the passage of electricity**.

Electrolysis occurs in an electrolytic cell.

A simple electrolytic cell consists of:

- the anode, the positive electrode
- the cathode, the negative electrode
- the electrolyte as the molten or aqueous substance that undergoes electrolysis

Important

Electrodes are usually made from unreactive metals, such electrodes are known as inert electrodes. Graphite and Platinum are the commonly used inert electrodes.

Cathode (negative) attracts cations (positive) Anode (positive) attracts anions (negative)

Metals or hydrogen (exception) are formed at the cathode (because metal ions are positive/cations and cathodes are negative) and non-metals (except hydrogen) are formed at the anode.

In electrolysis, an electric current is passed through an electrolyte causing the electrolyte to decompose.

The electricity is carried through the electrolyte by ions which is why the ionic compound needs to be in a molten or aqueous state as the ions cannot move if it's in a solid state.

In this process, the transfer of charge occurs because of:

- the movement of electrons from the cathode to the anode
- the loss or gain of electrons at the electrodes
- the movement of ions in the electrolyte

Important Oxidation and reduction

Reduction is a process in which an atom, ion, or molecule gains electrons. Oxidation is the opposite of reduction, i.e. an atom, ion or molecule loses electrons.

To remember this, use this mnemonic OIL RIG:

Oxidation Ts	
Loss	
Reduction Is Gain	
Something called reducing agents and oxidising agents also exist.	

The substance being oxidised is the reducing agent while the substance being reduced is the oxidising agent.

Electrolysis of lead(II) bromide

Precaution: must be done in a fume cupboard

Binary compounds are those that contain two elements chemically combined. Lead(II) Bromide is an example of this.

To carry out the electrolysis of lead(II) bromide, we first need to melt it since it is an ionic compound We use an electrolytic cell like the one below:



Fig. 1: An electrolytic cell showing the electrolysis of lead(ii) bromide

Here is the equation of the electrolysis of Lead(II) Bromide:

 $PbBr_{2}(l) \longrightarrow Br_{2}(g) + Pb(l)$

In the process, the lead ions are **reduced** to the lead atom (gain of electrons), here's the half-ionic equation for it:

$$\mathrm{Pb}^{2+}\left(\mathrm{l}\right) + 2\,\mathrm{e}^{-} \longrightarrow \mathrm{Pb}\left(\mathrm{l}\right)$$

The lead metal is deposited at the cathode since it's a cation as it has a positive charge.

The bromide ions lose an electron at the anode to become stable and become bromine atoms

$$\mathrm{Br}^{-}\left(\mathrm{l}
ight)\longrightarrow\mathrm{Br}+\mathrm{e}^{-}$$

Then two bromine atoms combine to form a bromine molecules The bromide ions are oxidised in this case (they lose electrons).

During this process, you will notice you will see that an orange-red gas is being produced at the anode, that is bromine.

You will also notice that lead is being formed at the cathode.

Electrolysis of aluminium oxide

Bauxite (aluminium ore) is an impure form of aluminium oxide.

It is first treated with sodium hydroxide to obtain pure aluminium oxide, this is done to improve the conductivity of the molten aluminium oxide

This is then dissolved in molten cryolite (Na_3AlF_6), this is done to reduce the melting point of pure aluminium oxide (which is 2017°C) down to between 800°C and 1000°C.



Then, the Hall-Heroult cell is used for the process of electrolysis.

Fig. 2: Diagram of a Hall-Heroult cell showing the electrolysis of aluminium oxide

During this process, Oxide ions are oxidised and Aluminium ions are reduced. The half-ionic equations for both are given below.

$$2 \operatorname{O}^{2-}(\mathrm{l}) \longrightarrow \operatorname{O}_2(\mathrm{g}) + 4 \operatorname{e}^ \operatorname{Al}^{3+} + 3 \operatorname{e}^- \longrightarrow \operatorname{Al}(\mathrm{l})$$

Approximately 15 kWh of electricity is used to produce 1 kg of aluminium.

The overall equation for this is:

$$2 \operatorname{Al}_2 \operatorname{O}_3(l) \longrightarrow 4 \operatorname{Al}(l) + 3 \operatorname{O}_2(g)$$

As you may notice oxygen gas is being produced and since the electrodes are made of graphite, which is a form of carbon, the oxygen reacts with the carbon producing carbon dioxide as it has enough activation energy for the reaction to occur (due to the temperature of the cell).

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This means that eventually, the graphite anode will have to be replaced as it will burn away. (only the anode because oxygen will be formed at the anode, not the cathode).

Electrolysis of concentrated aqueous sodium chloride

Concentrated aqueous sodium chloride is a strong saltwater solution with a **high amount of salt/sodium chloride** dissolved in water.

Since sodium chloride is an ionic compound, we have to dissolve it in water so that it can conduct electricity. By doing this, the ions in sodium chloride disassociate in water.

There are 4 ions in this solution:

- ${\rm H}^+$ and ${\rm OH}^-$ from water
- Na^+ and Cl^- from sodium chloride

When the current flows, the H^+ ions and Na^+ ions are both attracted to the cathode.

But as hydrogen has stronger electrostatic forces, it accepts electrons more easily than the Na^+ ions and that is why hydrogen gas is produced at the cathode instead of sodium metal.

$$2\,\mathrm{H^{+}}\left(\mathrm{aq}
ight)+2\,\mathrm{e^{-}}\longrightarrow\mathrm{H}_{2}\left(\mathrm{g}
ight)$$

Again, both OH^- ions and Cl^- ions are attracted to the anode. The Cl^- ions lose electrons more easily as they have weaker electrostatic forces (due to more number of shells) and chlorine gas is produced at the anode.

$$2\,\mathrm{Cl^{-}}\,\mathrm{(aq)}\longrightarrow\mathrm{Cl}_{2}\,\mathrm{(g)}+2\,\mathrm{e^{-}}$$

Electrolysis of dilute sulfuric acid

Pure water is a very poor conductor of electricity because there are so few ions in it. It can be made to decompose using a Hoffman voltameter.



Fig. 3: A hoffman voltameter, the anode and cathode is made of platinum

To enable water to conduct electricity better, some dilute sulfuric acid or sodium hydroxide solution is added to it.

When the power is turned on, gases are seen to be produced at both the electrodes. Twice as much gas is produced at the cathode than at the anode.

The gas at the cathode burns with a pop, proving that it is hydrogen.

In this process, positively charged hydrogen ions moved to the cathode and is being oxidised.

 $4\,\mathrm{H^{+}}\,\mathrm{(aq)} + 4\,\mathrm{e^{-}} \longrightarrow 2\,\mathrm{H_{2}}\,\mathrm{(g)}$

The gas collected at the anode can relight a glowing splinter, proving it to be oxygen.

$$4\,\mathrm{OH^{-}}\left(\mathrm{aq}
ight) {\longrightarrow} 2\,\mathrm{H_{2}O}\left(\mathrm{l}
ight) + \mathrm{O_{2}}\left(\mathrm{g}
ight) + 4\,\mathrm{e^{-}}$$

Electrolysis guidelines

- At an inert anode, chlorine, bromine and iodine (all the halogens) are prefered over oxygen.
- At an inert cathode, hydrogen is prefered over metals unless an unreactive metal is present.

Electrolysis of aqueous copper(II) sulfate

Copper(II) sulfate solution ($CuSo_4(aq)$) may be electrolysed by using inert graphite electrodes.

During this process, copper metal and oxygen gas are formed.

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There are 4 ions present in the solution:

- H^+ and OH^- from water
- + ${\rm Cu_2}^+$ and ${\rm SO_4}^{2-}$ from sodium chloride

Copper ions are preferred over Hydrogen ions due to preferential discharge.

Preferential discharge in electrolysis means that certain ions in a solution are more likely to react and be discharged (or "stick") to the electrode compared to others, based on their chemical properties. This leads to a more efficient reaction and greater production of desired products.

The copper metal is therefore deposited at the cathode.

 $\mathrm{Cu_{2}}^{+}\left(\mathrm{aq}
ight)+2\,\mathrm{e}^{-}\longrightarrow\mathrm{Cu}\left(\mathrm{s}
ight)$

Similarly with hydroxide and sulfate ions, the hydroxide ions release electrons more easily, so oxygen gas and water are produced at the anode

$$4\,\mathrm{OH^{-}}\left(\mathrm{aq}
ight) {\longrightarrow} \mathrm{O}_{2}\left(\mathrm{g}
ight) + 2\,\mathrm{H}_{2}\mathrm{O}\left(\mathrm{l}
ight) + 4\,\mathrm{e^{-}}$$

Purification of copper

Copper is a very good conductor of electricity and heat. Even small amounts of impurities can affect this conductivity noticeably, whether in fine wires or larger cables.

To ensure purity, newly extracted copper has to be purified by electrolysis.

The impure copper is used as the anode and the cathode is made from very pure copper.

When the current flows, the copper moves from the impure anode to the pure cathode. Any impurities fall to the bottom of the cell and collect below the anode. This usually consists of precious metals such as gold and silver.

Half-ionic equations

- Anode: $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$
- Cathode: $Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$

Electroplating

Electroplating is the process in which one metal is plated/coated with another. The purpose of this is to **improve the appearence of the object** and **give a protective coating to the metal** underneath it.

For silver plating, the electrolyte is a solution of a silver salt.

The item to be plated is made the cathode in the cell so that the metal ions move to it when the current is switched on.

$$\mathrm{Ag^{+}}\left(\mathrm{aq}
ight)+\mathrm{e^{-}}\longrightarrow\mathrm{Ag}\left(\mathrm{s}
ight)$$

4.2 Hydrogen-oxygen fuel cells

A hydrogen–oxygen fuel cell uses only hydrogen and oxygen to produce electricity with water as the only product produced.

The NaOH(aq) electrolyte is used for this reaction and the electrodes are porus.

$$\mathrm{O}_{2}\left(\mathrm{g}
ight)+2\,\mathrm{H}_{2}\mathrm{O}\left(\mathrm{l}
ight)+4\,\mathrm{e}^{-}\longrightarrow4\,\mathrm{OH}^{-}\left(\mathrm{aq}
ight)$$

Then the hydroxide ions formed are removed from the fuel cell by a reaction with hydrogen.

$$\mathrm{H}_{2}\left(\mathrm{g}
ight)+2\,\mathrm{OH}^{-}\left(\mathrm{aq}
ight)\longrightarrow2\,\mathrm{H}_{2}\mathrm{O}\left(\mathrm{l}
ight)+2\,\mathrm{e}^{-}$$

Advantages and disadvantages of fuel cells

Advantages:

- Uses hydrogen and oxygen to make non-polluting water while generating electricity (compared to petrol engines which produce many pollutants)
- Similar to battery, but no external charging required
- Capable of producing electricity as long as hydrogen and oxygen are supplied

Disadvantages:

- Hydrogen is difficult to store in a car as it is a gas at room temperature
- There are no existing infrastructure (refueling stations) for this
- Electric motors and fuel cells are less durable than petrol or diesel engines
- Fuel cells are very expensive at the moment.

Chapter 5

5.1 Exothermic and endothermic reactions

Exothermic reactions

An exothermic reaction transfers thermal energy to the surroundings leading to an increase in the temperature of the surroundings.

Examples of exothermic reactions include combustion and neutralization.

Combustion

When a fuel, for example natural gas, burns in excess air, it produces a large amount of energy.

```
methane + oxygen \longrightarrow carbon dioxide + water + \Delta H
```

 ΔH means heat released.

However if not enough air is available for this reaction, it will not produce as much energy and also will produce carbon monoxide, which is poisonous.



Fig. 1: Reaction pathway diagram for an exothermic reaction

Endothermic reactions

An endothermic reaction takes in thermal energy from the surroundings leading to a decrease in the temperature of the surroundings.

Examples include: Thermal decomposition and Photosynthesis



Fig. 2: Reaction pathway diagram for an endothermic reaction

Enthalpy change

The transfer of thermal energy during a reaction is called the **enthalpy change**, ΔH , of the reaction.

 ΔH is negative for exothermic reactions and positive for endothermic reactions.

Activation energy

Activation energy, E_a , is the minimum energy that colliding particles must have to trigger a reaction.

If the particles have less energy than this amount and they collide, they will not react.

More about this in chapter 6 (6.2).

Making and breaking bonds

Bond breaking is an endothermic process (to break bonds, you need to take in energy).

While bond making is an exothermic process.

In a reaction, there is both bond breaking and bond making.

A reaction in which breaking the bonds uses more energy than making the bonds produce, is known as an endothermic reaction.

While a reaction in which breaking the bonds uses less energy than the amount of energy making the bonds produce, is known as an exothermic reaction.

Calculating enthalpy change

To calculate enthalpy change, ΔH of a reaction, we need to first list out the bonds in the reactants and products side.

For example, let's take the equation:

$$\mathrm{CH}_{4}\left(\mathrm{g}\right)+2\,\mathrm{O}_{2}\left(\mathrm{g}
ight)\longrightarrow\mathrm{CO}_{2}\left(\mathrm{g}
ight)+2\,\mathrm{H}_{2}\mathrm{O}\left(\mathrm{l}
ight)$$

from the combustion reaction we talked about above.

From this equation, we can figure out that there are:

- $4 \mathrm{C-H}$ bonds
- 2 O=O bonds

on the reactants side while there are:

- 2 C=O bonds
- $4 \,\mathrm{H}{-}\mathrm{O}$ bonds

on the products side.

Now, we'd be given a table listing the bond energies like:

Bond	Bond energy/kJ/mol
C-H	435
0=0	497
C=O	803
H-O	464
C-C	347
C-O	358

Using this table we can calculate the energies on both the sides and use that to calculate the enthalpy change:

Reactants (bond breaking)

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- + 4 m C-H bonds = 4 imes 435 = 1740 kJ
- 2 O=O bonds = 2×497 = 994 kJ

Which comes to a total of 2734 kJ of energy which is required to break the bonds in the reactants.

Products (making bonds)

- + 2 $C{=}O$ bonds = 2×803 = 1606 kJ
- + 4 $H{-}O$ bonds = 4 \times 464 = 1856 kJ

Which comes out to a total of 3462 kJ of energy given out when the bonds are made.

Calculating enthalpy change

To calculate enthalpy change, we use the following formula:

 $\Delta E =$ energy required to break bonds – energy given out when bonds are made

now subsitute the values:

$$\Delta E = 2734 - 3462 = -728$$

The negative sign shows that the reactants are losing energy to their surroundings.

The enthalpy change is negative, which means that the reaction is exothermic.

6.1 Physical and chemical changes

Physical Change	Chemical Change
Does not result in formation of a new substance	Forms a completely new substance
Only the state changes	The chemical composition changes
Usually reversible	Usually cannot be reversed

6.2 Rate of reaction

The rate of reaction is the speed at which reactants are converted into products in a chemical reaction.

Factors affecting rate of reaction

There are various factors which affect this, such as:

- changing the concentration of solutions (the more the concentration, the higher the rate of reaction)
- changing the pressure of gases (the more the pressure, the higher the rate of reaction)
- changing the surface area of solids (the larger the surface area, the higher the rate of reaction)
- changing the temperature (the higher the temperature, the higher the rate of reaction)
- adding or removing a catalyst, including enzymes (adding a catalyst increases the rate of reaction)

A catalyst is a substance that alters/changes the rate of reaction without changing the products of the reactions, it itself also is unchanged at the end of the reaction.

Collision theory

The collision theory tells us that for a reaction to occur, the particles of the reactants must collide with each other and have enough energy (activation energy) to trigger a reaction.

Collisions which have enough activation energy and result in the formation of products are known as **successful collisions**.

All the factors affecting rate of reaction can be explained by the collision theory.

Higher concentration

Higher concentration of a substance means that there are more particles of that substance, therefore more successful collisions.

Pressure of gases

A higher pressure means there are more gas molecules in a given volume of gas, which means that there will be more frequent collisions and hence more successful collisions.

Surface area

In reactions involving solids, only the exposed surface area of reactants is available for collisions. When you increase the surface area by grinding or cutting the solid into smaller pieces, more reactant particles are exposed, allowing for a greater number of collisions which means more successful collisions.

Temperature

A higher temperature means that particles will have more kinetic energy which means that they will move faster. If the particles move faster that means there will be more collisions leading to more successful collisions.

Additionally, collisions are more likely to have enough activation energy if the temperature is higher which increases the chance of a successful collision.

Catalyst

Catalysts decrease the amount of the activation energy required for a reaction to take place, which means that there is a higher chance of a successful collision.

Experiment To find out the rate of reaction

To measure the rate of reaction, we can measure:

- the volume of gas produced in a specific amount of time
- the "loss" of mass (due to gas being released) in a specific amount of time

Volume of gas produced

In such an experiment, we use a gas syringe attached to a beaker/flask/test-tube with a lid (diagram below).



Fig. 1: Diagram showing a gas syringe attached to a test-tube

By measuring the volume of gas produced in a specific amount of time, we can calculate the rate of reaction for that reaction.

Loss of mass

In such an experiment, mass is "lost" because the gas produced by the reaction is released.


Fig. 2: Diagram showing a beaker kept on a balance during a reaction

By measuring the loss of mass in a specific amount of time, we can calculate the rate of reaction for that reaction.

Now you may notice in the diagram, there's cotton wool used to cover the beaker, this is done to prevent acid spray. Now why wasn't a bung/lid used? because then the gas produced in the reaction would not have been released and would've been trapped inside.

This method is not usually preferred as gas is released into the atmosphere.

Plotting graphs

Once we are done with our observations and after noting down the loss of mass/volume of gas produced, we can now plot a rate of reaction graph by plotting the loss of mass/volume of gas on the y axis and time on the x axis.

Rate of reaction graphs

As we read earlier that the surface area of the solid reactants affect the speed of a reaction, we can see this by plotting rate of reaction graphs for the same reaction but using different sizes of solids in each of them. Note that we need to use the same mass/volume of reactants and the temperature should be kept the same.

The graph would look something like this:



Fig. 3: A rate of reaction graph showing the effect of surface area on rate of reaction

As you can see, both reactions have an equal loss in mass, however, the one with powdered solid is faster.

6.3 Reversible reactions and equilibrium

Reversible reactions are chemical reactions that can go both forward and backward unlike traditional chemical reactions which can only go forward.

We use the symbol \implies to denote a reversible reaction.

A reversible reaction in a closed system is at **equilibrium** when the rates of the **forward** and **reverse reactions** are **equal**, and the **concentrations** of reactants and products **remain constant/do not change**.

Two examples for reversible reactions are:

```
hydrated copper(II) sulfate \Longrightarrow anhydrous copper(II) sulfate + water
```

The color change in this reaction is from blue to white.

and

```
anhydrous cobalt(II) chloride + water \implies hydrated cobalt(II) chloride
```

The color change in this reaction from blue to pink.

Changing conditions for a reversible reaction

Changing the conditions can change the direction of a reversible reaction.

The effect of heat on hydrated compounds

The effect of heat on hydrated compounds, such as copper(II) sulfate or cobalt(II) chloride, can change the direction of a reversible reaction by moving it towards dehydration.

When heat is applied, the hydrated compounds lose their water molecules and become anhydrous, shifting the equilibrium from the hydrated form to the anhydrous form.

Let's take the reaction we talked about above as an example

hydrated copper(II) sulfate \implies anhydrous copper(II) sulfate + water

In this reaction, when heat is applied to hydrated copper(II) sulfate, it starts to become anhydrous and releases water, when heat is no longer being applied, the reaction starts moving in the opposite direction and hydrated copper(II) sulfate is once again formed.

The addition of water to anhydrous compounds

The addition of water to anhydrous compounds like copper(II) sulfate or cobalt(II) chloride can drive the reversible reaction in the opposite direction, causing hydration.

As water is added, the anhydrous compounds can absorb the water molecules and transform back into their hydrated forms, shifting the equilibrium towards the hydrated compounds.

Factors affecting position of equilibrium

The position of equilibrium can be affected by the following factors:

- changing temperature
- changing pressure
- changing concentration
- using a catalyst

Changing temperature

The position of equilibrium is affected by temperature changes differently depending on whether the reaction is exothermic or endothermic.

• For exothermic reactions (which release heat), increasing the temperature shifts the equilibrium towards the reactants (left) because it opposes the heat-producing direction to neutralize the temperature rise, promoting the reverse reaction.

• For endothermic reactions (which absorb heat), increasing the temperature shifts the equilibrium towards the products (right) because it favors the heat-absorbing direction to neutralize the temperature rise, promoting the forward reaction.

Basically, reversible reactions always try to maintain equilibrium.

Therefore when a reaction requires heat (endothermic reactions), they favour the direction which absorbs heat to avoid an increase in temperature and the opposite occurs for reactions which give out heat.

Changing pressure

The effect of pressure change depends on the **number of moles of gas molecules** involved in the reaction.

Increasing pressure generally shifts the equilibrium towards the side with fewer moles of gas molecules.

Decreasing pressure shifts the equilibrium towards the side with more moles of gas molecules.

Changing concentration

Increasing the concentration of reactants shifts the equilibrium towards the products, favoring the forward reaction.

Increasing the concentration of products shifts the equilibrium towards the reactants, favoring the reverse reaction.

Using a catalyst

A catalyst doesn't affect the position of equilibrium; instead, it speeds up the attainment of equilibrium. It facilitates both the forward and reverse reactions equally, making the system reach equilibrium more rapidly but with the same equilibrium position.

Haber process

The haber process is used for manufacturing ammonia. This reaction requires nitrogen and hydrogen.

Obtaining nitrogen

The nitrogen needed for this reaction is obtained from the atmosphere via **fractional distillation** of **liquid air**.

Obtaining hydrogen

The hydrogen for this reaction is obtained from a method known as **steam re-forming**.

This process involves reacting methane with steam to produce hydrogen and carbon monoxide.

$$\mathrm{CH}_{4}\left(\mathrm{g}
ight)+\mathrm{H}_{2}\mathrm{O}\left(\mathrm{g}
ight)\longrightarrow3\,\mathrm{H}_{2}\left(\mathrm{g}
ight)+\mathrm{CO}\left(\mathrm{g}
ight)$$

This process needs to be carried out at a temperature of 750°C, a pressure of 3000 kPa (30 atmospheres) with a catalyst of nickel.

The carbon monoxide produced from this reaction is then used to react with steam again to form even more hydrogen and carbon dioxide.

$$\mathrm{CO}\left(\mathrm{g}
ight)+\mathrm{H}_{2}\mathrm{O}\left(\mathrm{g}
ight)\longrightarrow\mathrm{H}_{2}\left(\mathrm{g}
ight)+\mathrm{CO}_{2}\left(\mathrm{g}
ight)$$

Making ammonia

In this reaction, for every one mole of nitrogen, 3 moles of hydrogen are needed. The ratio of nitrogen to hydrogen is 1:3.

 $\mathrm{N}_{2}\left(\mathrm{g}
ight)+3\,\mathrm{H}_{2}\left(\mathrm{g}
ight)\Longrightarrow2\,\mathrm{NH}_{3}\left(\mathrm{g}
ight)$

The conditions required for this reaction are:

- a temperature of 450°C
- a pressure of 20 000 kPa / 200 atm
- a catalyst of freshly cut and finely divided iron.

Contact process

Contact process is a way to produce sulfuric acid.

The process goes like this

Obtaining sulfur dioxide

First, sulfur dioxide is produced by burning sulfur with excess oxygen in air.

 $sulfur + oxygen \longrightarrow sulfur dioxide$

 $S(s) + O_2(g) \longrightarrow SO_2(g)$

Sulfur dioxide can also be produced by roasting sulfide ores (for example, zinc sulfide) in air.

Obtaining sulfur trioxide

Now to produce sulfur trioxide, we need to react sulfur dioxide with oxygen, which is a reversible reaction.

 $sulfur dioxide + oxygen \Longrightarrow sulfur trioxide$

 $2 \operatorname{SO}_{2}(\mathrm{g}) + \operatorname{O}_{2}(\mathrm{g}) \rightleftharpoons \operatorname{SO}_{3}(\mathrm{g})$

The conditions required for this reaction are:

- a temperature of approximately 450°C
- a catalyst of vanadium(V) oxide
- a pressure of 200 kPa / 2 atmospheres

Making sulfuric acid

Sulfuric acid is manufactured by reacting sulfur trioxide with water.

 $sulfur \ trioxide + water \longrightarrow sulfuric \ acid$

 $SO_{3}(g) + H_{2}O(l) \longrightarrow H_{2}SO_{4}(l)$

Precautions for this reaction are:

- Ensure sulfur dust explosions are prevented by wetting down the sulfur
- Carefully managing and neutralising acid sprays/leaks if any.

All the conditions for all reactions are done to maximize the yield of the products in both the processes.

6.4 Redox

Oxidation numbers

As talked about in 2.4, all elements have an oxidation number. Transition elements have multiple oxidation numbers. We use roman numeral to denote oxidation number, for example, cobalt(II) chloride tells us that the oxidation number of cobalt is 2.

Oxidation and reduction

Oxidation and reduction have been briefly talked about in Chapter 4.

Important Oxidation and reduction

Reduction is a process in which an atom, ion, or molecule gains electrons. Oxidation is the opposite of reduction, i.e. an atom, ion or molecule loses electrons.

To remember this, use this mnemonic OIL RIG:

Oxidation Is		
Loss		
Reduction		
Is		
Gain		

.....

Something called reducing agents and oxidising agents also exist. The substance being oxidised is the reducing agent while the substance being reduced is the oxidising agent.

Potassium iodide is a very common reducing agent. When it is oxidised, the iodide ion is oxidised to form iodine, resulting in a color change from colourless to yellow-brown.

Oxidation is also the gain of oxygen and Reduction is also the loss of oxygen.

When a substance is oxidised, there's an increase in its oxidation number.

When a substance is reduced, there's a decrease in its oxidation number.

It is important to note the following:

- The oxidation number of the free element is always 0, for example in metals such as zinc and copper.
- In simple monatomic ions, the oxidation number is the same as the charge on the ion. So for example, lodine has an oxidation number of 0 in I_2 (charge is 0) but an oxidation number of -1 in I^- (charge is -1)
- Compounds have no overall charge.
- The sum of the oxidation numbers in an ion is equal to the charge on the ion. This is very important and useful. So for example you have the sulfate ion, SO_4^{2-} and imagine you do not know the charge on it, you can simply calculate it by adding the oxidation numbers of the atoms together. So the oxidation number of S is +6 and the oxidation number of oxygen is -2 and there are 4 oxygen

atoms which means that the sum of the oxidation number would be: $6 + (-2 \times 4)$ which is equal to -2, which is the charge on the sulfate ion.

Redox reactions

Redox reactions are reactions in which both oxidation and reduction occur at the same time.

Chapter 7

7.1 The characteristic properties of acids and bases

The pH scale

The pH scale is a measurement system that quantifies the acidity or alkalinity (basicity) of a solution. It ranges from 0 to 14, where:

pH below 7 indicates acidity, with 0 being the most acidic.pH 7 is considered neutral. Pure water is neutral.pH above 7 indicates alkalinity, with 14 being the most alkaline (basic).

Acids

Acids are substances that release hydrogen ions (H^+) when mixed in water. Acids have a pH of less than 7. The lower the pH, the stronger the acid. The stronger the acid, the higher the concentration of hydrogen ions (H^+) .

Acids are also known as proton donors.

A strong acid is an acid that **completely dissociates** in an aqueous solution and a weak acid is an acid that **partially dissociates** in aqueous solution.

An example of strong acid is hydrochloric acid, HCl (equation: $HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$) while an example of a weak acid is ethanoic acid CH_3COOH (equation: $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$; reversible reaction indicating partial disassociation)

Bases

Base are substances that release hydroxide ions (OH^{-}) when mixed in water.

Bases have a pH of more than 7.

The higher the pH, the stronger the base.

The stronger the base, the higher the concentration of hydroxide ions (OH^{-}) .

Bases are usually oxides or hydroxides of metals. Bases which are soluble are known as **alkalis**.

Bases are also known as proton acceptors.

Amphoteric substances

Some substances behave as both, proton acceptors and donors, these are known are amphoteric substances.

Examples include, aluminium oxide (Al_2O_3) and zinc oxide (ZnO).

Indicators

Indicators are substances that change color to show if a solution is acidic or basic. They can be used to identify acids and bases.

Indicator	Colour in Acid	Colour in Base		
Litmus	Red	Blue		
Methyl Orange	Pink	Yellow		
Thymolphthalein	Colourless	Blue		

Another really popular indicator is the universal indicator.



Fig. 1: pH scale for the universal indicator solution

Acid reactions

Acids usually react with:

- metals
- bases
- carbonates

Acid + metal

Acids react with metals to produce salt and hydrogen gas.

Precaution: This can only be done with less reactive metals. The metals used are **MAZIT** metals: **M**agnesium, Aluminium,

Zinc,

Iron,

Tin

Example: Add excess magnesium ribbon to dilute HCI. This will produce Hydrogen gas. In this reaction, hydrogen ions from HCI gain electrons from the metal atoms.

$$\mathrm{Mg} + 2\,\mathrm{HCl} \longrightarrow \mathrm{MgCl}_2\,\mathrm{(aq)} + \mathrm{H}_2$$

Acid + base

Acids and bases can neutralize each other in a reaction, forming water and a salt.

Example:

$$\rm HCl + NaOH \longrightarrow NaCl + H_2O$$

In this reaction, the hydrogen ion $(\mathrm{H^+})$ and the hydroxide ion $(\mathrm{OH^-})$ react to form water.

$$\mathrm{H^{+}}\left(\mathrm{aq}
ight) + \mathrm{OH^{-}}\left(\mathrm{aq}
ight) \longrightarrow \mathrm{H_{2}O}\left(\mathrm{l}
ight)$$

Acid + carbonate

Acids react with carbonate compounds to produce carbon dioxide gas, water, and a salt.

Carbon dioxide gas can be tested by passing it through limewater, if it turns milky, then it is indeed carbon dioxide.

Example:

$$2 \operatorname{HCl} + \operatorname{Na}_2 \operatorname{CO}_3 \longrightarrow 2 \operatorname{NaCl} + \operatorname{H}_2 \operatorname{O} + \operatorname{CO}_2$$

Base reactions

Bases usually react with:

- acids
- ammonium salts

Base + Acid

This is the same neutralization react as the one described above (acid + base).

Base + ammonium salt

Bases can react with ammonium salts to produce ammonia gas (NH3), water, and a salt.

The Haber process is preferred over this process in industry as the haber process has a higher yield.

$$Ca(OH)_2(s) + 2 NH_4Cl \longrightarrow CaCl_2 + H_2O + 2 NH_3$$

Ammonia can be detected by its pungent odour and by turning red litmus blue.

7.2 Oxides

Oxides can be acidic, basic or amphoteric.

Acidic oxides are usually the non-metal oxides, for example sulfur dioxide (SO_2) and carbon dioxide (CO_2)

Basic oxides are usually the metal oxides, for example copper(II) oxide (CuO) and calcium oxide (CaO).

Amphoteric oxides are oxides which can behave as both acids and bases and can react with both to produce a salt and water. Examples include aluminium oxide (Al_2O_3) and zinc oxide (ZnO).

7.3 Preparation of salts

Before preparing salts, we must know which salts are soluble and insoluble.

Soluble Salts
All Sodium, Potassium and Ammonium salts are soluble.
All nitrates are soluble.
All chlorides are soluble except lead and silver.
All sulfates are soluble except barium, calcium and lead
All carbonates and hydroxides are insoluble except sodium, potassium and ammonium
Calcium hydroxide is partially soluble

Preparing soluble salts

There are many ways to prepare soluble salts.

Acid + alkali (soluble base; titration)

Choose the acid and the alkali.

For example, to form sodium chloride (NaCl), use hydrochloric acid (HCl) and sodium hydroxide (NaOH).

Add a known volume of the acid to a flask and add a few drops of an indicator like phenolphthalein or universal indicator.

Then, titrate by slowly adding the alkali from a burette until the indicator changes color, signaling that the acid has been neutralized.

The resulting solution contains the soluble salt along with water, now you can just seperate the salt by evaporating the water using an evaporating basin.

Acid + excess metal

This is the same as the Acid + metal reaction described in 7.1, with one extra step which is to filter out the salt from the unreacted reactants.

Acid + excess insoluble base

Combine the acid (e.g., hydrochloric acid, HCI) with an excess of an insoluble base (e.g., copper oxide, CuO).

The base reacts with the acid to form the soluble salt and water.

Filter the solution to filter out the leftover insoluble base.

Now evaporate the water to get salt crystals.

Acid + excess insoluble carbonate

This is the same as Acid + carbonate in 7.1, except there are 2 more steps: Filter the solution to filter out the leftover insoluble carbonate. Now evaporate the water to get salt crystals.

Preparing insoluble salts

The method of preparing an insoluble salt is called precipitation.

In this method, 2 solubles salt (aqueous) react to make 1 insouble salt (solid) and 1 soluble salt (aqueous).

soluble salt
$$\rightarrow$$
 insoluble salt \rightarrow insoluble salt $+$ soluble salt

for example,

$$BaCl_{2}(aq) + Na_{2}SO_{4}(aq) \longrightarrow BaSO_{4}(s) + 2NaCl(aq)$$

Then the insouble salt is filtered out, washed with water and then left to dry.

Water of crystallisation

A hydrated substance is a material that is chemically bonded to water molecules, while an anhydrous substance contains no water.

Water of crystallization is the water trapped in crystals. For example

$CuSO_4\cdot 5\,H_2O$

This shows that copper(II) sulfate has five water molecules as water of crystallisation. This water is trapped within the crystals.

To remove water of crystallization from a compound, you can heat it gently, causing the water molecules to evaporate and leaving the anhydrous form of the compound.

	The Periodic Table of Elements																
								Gr	oup								
1	11											Ш	IV	V	VI	VII	VIII
				Key			1 H hydrogen 1						97				2 He helium 4
3 Li lithium 7	4 Be beryllium 9		ato	atomic numbe omic sym name ative atomic m	r bol ass			-				5 B boron 11	6 C carbon 12	7 N nitrogen 14	8 O oxygen 16	9 F fluorine 19	10 Ne neon 20
11 Na sodium 23	12 Mg magnesium 24					1						13 Al aluminium 27	14 Si silicon 28	15 P phosphorus 31	16 S sulfur 32	17 Cl chlorine 35.5	18 Ar argon 40
19 K potassium 39	20 Ca calcium 40	21 Sc scandium 45	22 Ti titanium 48	23 V vanadium 51	24 Cr chromium 52	25 Mn manganese 55	26 Fe iron 56	27 CO cobalt 59	28 Ni nickel 59	29 Cu copper 64	30 Zn zinc 65	31 Ga gallium 70	32 Ge germanium 73	33 As arsenic 75	34 Se selenium 79	35 Br bromine 80	36 Kr krypton 84
37 Rb rubidium 85	38 Sr strontium 88	39 Y yttrium 89	40 Zr zirconium 91	41 Nb niobium 93	42 Mo molybdenum 96	43 Tc technetium	44 Ru ruthenium 101	45 Rh rhodium 103	46 Pd palladium 106	47 Ag silver 108	48 Cd cadmium 112	49 In Indium 115	50 Sn tin 119	51 Sb antimony 122	52 Te tellurium 128	53 I iodine 127	54 Xe xenon 131
55 Cs caesium 133	56 Ba barium 137	57–71 Ianthanoids	72 Hf hafnium 178	73 Ta tantalum 181	74 W tungsten 184	75 Re rhenium 186	76 Os osmium 190	77 Ir iridium 192	78 Pt platinum 195	79 Au ^{gold} 197	80 Hg mercury 201	81 T <i>I</i> thallium 204	82 Pb lead 207	83 Bi bismuth 209	84 Po polonium	85 At astatine	86 Rn radon
87 Fr francium	88 Ra radium	89–103 actinoids	104 Rf rutherfordium	105 Db dubnium	106 Sg seaborgium	107 Bh bohrium	108 Hs hassium	109 Mt meitnerium	110 Ds darmstadtium	111 Rg roentgenium	112 Cn copernicium		114 F <i>l</i> flerovium		116 Lv livermorium		
lanthand	bids	57 La Ianthanum 139	58 Ce cerium 140	59 Pr praseodymium 141	60 Nd neodymium 144	61 Pm promethium	62 Sm samarium 150	63 Eu europium 152	64 Gd gadolinium 157	65 Tb terbium 159	66 Dy dysprosium 163	67 Ho holmium 165	68 Er erbium 167	69 Tm thulium 169	70 Yb ytterbium 173	71 Lu Iutetium 175	
actinoid	S	89 Ac actinium	90 Th thorium 232	91 Pa protactinium 231	92 U uranium 238	93 Np neptunium	94 Pu plutonium	95 Am americium	96 Cm curium	97 Bk berkelium	98 Cf californium	99 Es einsteinium	100 Fm fermium	101 Md mendelevium	102 No nobelium	103 Lr Iawrencium	

8.1 Arrangement of elements

The volume of one mole of any gas is 24 dm³ at room temperature and pressure (r.t.p.).

Fig. 1: The perioidic table

The Periodic Table is an arrangement of elements in **periods** and **groups** and in order of increasing proton number / atomic number.

Periods in the periodic table

Properties of elements moving across a period (horizontal) in the perioidic table:

- Gradual change from metal to non-metal
- Increase in number of electron
- Change in structure (Giant ionic to giant covalent to simple molecular)

Groups in the periodic table

Groups are vertical while periods are horizontal.

There are many groups in the periodic table, such as:

Group 1 - alkali metals

- Group 2 alkaline earth metals
- Group 6 chalcogens

Group 7 - halogens Group 8 - noble gases

The group numbers tell us about the oxidation number of an ion of that element and helps us figure out the charge on the ion of the element.

For example, the oxidation number of Lithium (Li) is 1, and therefore the charge on its ion would be +1. This is only applicable for the oxidation numbers: 1, 2, 3 and 4.

For oxidation numbers above 4 like 5, 6, 7, we simply subtract 8 from the oxidation number to calculate its charge. For example, oxygen (O) has an oxidation number of 6 which is obviously above 4. 6-8 = -2 which means that the charge on an ion of oxygen will be -2.

Elements with the oxidation number 8 are unreactive and do not form ions.

Elements of a certain group have similar **chemical properties** because they all have a similar electronic configuration. The electrons in the last shell of the atom is equal to the group number.

Predicting the properties of elements using the position in the periodic table

It is possible to predict the properties of an element using its position in the perioidic table, for example:

- The mass number tells us about how heavy the element is.
- The period number tells us about the number of shells an atom of that element has.
- The group number tells us about the element's chemical properties and if it is a metal or not.

8.2 Group I properties

They are good conductors of heat and electricity, are soft and have a low density.

When they are freshly cut, they're very shiny

They have a low melting and boiling point.

They burn in oxygen or air with specific flame colors. They react vigorously with water and with halogens to form metal halides.

As you go down the group:

- the melting point decreases
- the density increases
- the reactivity increases

8.3 Group VII properties

- These elements are colored and become darker down the group
- They exist as diatomic molecules
- At r.t.p, they show a gradual change from gas (Cl_2) , to liquid (Br_2) to solid (I_2)
- They form covalent compounds with non-metallic elements (such as (HCl))
- They react with hydrogen to form hydrogen halides, which dissolve in water to form acidic solutions.
- They react with metals to produce ionic metal halides such as $2 \operatorname{Fe}(s) + 3 \operatorname{Cl}_2(g) \longrightarrow 2 \operatorname{FeCl}_3(s)$

Chlorine gas can bleach moist indicator paper.

Colors of some halogens:

Halogen	Color
Chlorine	Yellow-green gas
Bromine	Red-Brown liquid
lodine	Gray-black solid

As you go down the group, the density increases and the reactivity decreases.

Displacement reactions

More reactive elements displace (replace) less reactive elements when added to a substance containing the less reactive element.

Example:

 $Potassium iodide + Chlorine \longrightarrow Potassium chloride + Iodine$

 $2 \operatorname{KI}(\operatorname{aq}) + \operatorname{Cl}_2(\operatorname{g}) \longrightarrow 2 \operatorname{KCl}(\operatorname{aq}) + \operatorname{I}_2(\operatorname{aq})$

8.4 Transition elements

The transition elements are metals that:

- have high densities
- have high melting points (except Mercury which is liquid at r.t.p)
- · form coloured compounds
- often act as catalysts as elements and in compounds

They are also:

harder, stronger and have a higher density than the metals in group I and II

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As discussed in earlier chapters (2.4), transition metals can have ions with multiple different oxidation numbers.

8.5 Noble gases

This is the most unusual group of elements, called the noble gases. It is made up of Helium, Neon, Krypton, Xenon, and Radon (radioactive).

- They are colorless gases
- They are monatomic gases
- They are very unreactive

Use of Argon: the gas used to fill light bulbs to prevent the tungsten filament from reacting with air.

Use of neon: Advertising signs and lasers

These gases are unreactive as their last shell is complete.

9.1 Properties of metals

Property	Metal	Non-metal
Conduct Heat and Electricity		×
Malleable		ig ackslash - usually soft or brittle
Ductile		ightarrow - usually soft or brittle
Brittle		\checkmark
Melting and boiling point	Usually high	Usually low

Metal reactions

Metals + Dilute Acids

Most metals react with dilute acids (e.g., hydrochloric acid, sulfuric acid) to produce salt and hydrogen gas.

During such a reaction, you will see **effervescence** which is caused by bubbles of hydrogen gas being formed as the reaction proceeds.

A simple reaction between magnesium and hydrochloric acid can be taken as an example:

$$Mg\left(s\right)+2\operatorname{HCl}\left(aq\right)\longrightarrow MgCl_{2}\left(aq\right)+H_{2}\left(g\right)$$

We can figure out which metals are the most reactive and least reactive depending upon their reaction with acids.

Metals + Oxygen

Many metals react with oxygen to form oxides.

For example, reacting magnesium with oxygen by burning it in excess oxygen to form magnesium oxide:

$$2 \operatorname{Mg}(s) + \operatorname{O}_{2}(g) \longrightarrow 2 \operatorname{MgO}(s)$$

Metals + Cold Water/Steam

Very reactive metals

Very reactive metals such as potassium, sodium and calcium react with **cold water** to make metal hydroxides and hydrogen gas. Example:

$$2 \operatorname{Na}(s) + 2 \operatorname{H}_2 O(l) \longrightarrow 2 \operatorname{NaOH}(aq) + \operatorname{H}_2(g)$$

Less reactive metals

Less reactive metals such as magnesium, zinc and iron react slowly with water, however they react faster with steam.

When metals react with steam, they produce metal oxide and hydrogen.

$$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$$

These experiments can be dangerous. Some precautions include:

• Wear eye protection

Using these experiments, we can figure out an order of reactivity showing us which metals are more reactive. A figure of the reactivity series can be seen in 9.4.

9.2 Uses of metals

Less reactve metals usually have more uses. Some uses of metals include:

- aluminium in the manufacture of aircraft because of its low density
- aluminium in the manufacture of overhead electrical cables because of its low density and good electrical conductivity
- aluminium in food containers because of its resistance to corrosion
- · copper in electrical wiring because of its good electrical conductivity and ductility

9.3 Alloys and their properties

An alloy is a mixture of a metal with other elements, designed to enhance or modify its properties for specific applications.

Examples of alloys include:

- brass, a mixture of copper and zinc
- stainless steel, a mixture of iron along with chromium, nickel and carbon

Alloys can be harder and stronger than the pure metals and are more useful.

These properties can be explained simply by the fact that alloys are made up of metals and different elements. All these different elements have atoms of different sizes, and these atoms being different in sizes means that it is harder for the layers in alloy to slide over each other.

Alloys have many uses, one of the major uses of stainless steel is in cutlery because of its hardness, and resistance to rusting.



Fig. 1: The structure of an alloy

In the structure above, we can see that there are two different sized atoms (blue ones and purple ones) which means that the structure is of an alloy.

9.4 Reactivity series

The reactivity series is a ranking of metals in order of their reactivity with water and acids, from most reactive (e.g., potassium) to least reactive (e.g., gold).

Reactivity series	Reaction with dilute acid	Reaction with air/oxygen	Reaction with water	Ease of extraction	
Potassium (K)		Burn very brightly	Produce hydrogen gas when reacted with cold water	Difficult to extract	
Sodium (Na)			with decreasing strength		
Calcium (Ca)					
Magnesium (Mg)	Produce hydrogen gas with				
Aluminium (Al)	decreasing strength as you go down	Burn to form an oxide with decreasing strength as you go down	React with steam with decreasing strength	Easier to extract	
Carbon (C)					
Zinc (Zn)					
Iron (Fe)					
Hydrogen (H)		React slowly to form			
Copper (Cu)		an oxide	Do not react with either		
Silver (Ag)	Do not react with dilute acids	Do not react	cold water or steam	Found as the element	
Gold (Au)		Do not readt			

However, aluminium is special. It appears to be unreactive but actually isn't. When exposed to air/oxygen, it reacts very quickly with the oxygen and forms a protective oxide layer on it's surface.

Displacement reactions

Displacement reactions in metals involve a more reactive metal displacing a less reactive metal from a compound, leading to the formation of a new compound.

For example, a reaction between zinc and copper(ii) nitrate. Zinc is more reactive than copper which means that it will displace copper and form zinc nitrate.

 $ext{zinc} + ext{copper}(ext{II}) ext{ nitrate} \longrightarrow ext{zinc nitrate} + ext{copper}$

 $\operatorname{Zn}(s) + \operatorname{Cu}(\operatorname{NO}_3)_2(\operatorname{aq}) \longrightarrow \operatorname{Zn}(\operatorname{NO}_3)_2(\operatorname{aq}) + \operatorname{Cu}(s)$

This can be seen as there is a color change from blue to colorless.

Experiment To find out the reactivity series from a set of metals

To do this, we can perform displacement reactions.

We can take a metal solution, for example zinc nitrate.

Now we can add various metals to this to see if the metal we added is more reactive than zinc or not, we can repeat this with solutions of other metals as well to make a reactivity series.

9.5 Corrosion of metals

Corrosion is when metals slowly rust or wear away due to chemical reactions with their surroundings. Basically it is when metals/alloys are chemically *attacked* by oxygen and water in the environment.

The more reactive a metal is, the faster it will corrode. This is why highly reactive metals like sodium and potassium are stored in oil as they will instantly corrode when exposed to air and moisture.

Rust is a form of corrosion that specifically refers to the reddish-brown, flaky coating that develops on **iron or steel surfaces** when they react with oxygen and moisture in the presence of air. Rust usually consists of hydrated iron(III) oxide (Fe₂O₃ $\cdot x$ H₂O) which is reddish-brown in color.

The requirements for rusting are water (in the form of moisture or water itself) and oxygen (from the air).

Both of the above are required for rusting. If either of them is not present, the metal (iron/steel) will not rust.

Note that rusting only occurs on iron or steel surfaces.

Preventing rust

Rusting is a problem because it **weakens the structural integrity** of metal objects and structures. To prevent rusting, it is important to stop either oxygen or moisture (or both) from coming into contact with the metal.

All of the methods involve isolating the metal from the surroundings using a protective shield (paint, oil, plastic etc.).

Painting

Applying a layer of paint to the metal surface forms a protective barrier, preventing direct contact with oxygen and moisture. A positive to this is that the object will look better however, if the paint is scratched, the metal beneath it will start to rust which will eventually spread under the paint.

Greasing

Coating the metal with grease or oil creates a **protective layer** that repels water and air, reducing the chances of rust formation.

Greasing is often used for moving parts, machinery, and equipment.

Coating with plastic

Plastics can be used as a coating to isolate the metal from the corrosive elements, preventing rust. This is usually done in pipes.

Galvanising

Zinc is applied as a protective coating over the iron or steel surface, forming a physical barrier. To do this, the object is dipped into molten zinc and a thin layer of zinc is formed. Zinc is more reactive than iron/steel and this layer also corrodes but when this zinc layer corrodes, it transfers/loses electrons to the iron, protecting it. This also protects the iron even if the zinc layer has been mostly scratched away.

Sacrificial protection

Bars of zinc are usually attached to ships to prevent iron from rusting. This is done because zinc is more reactive than iron and is corroded instead of iron.

So as long as some of the zinc stays in contact with the iron, the iron will be protected. When the zinc runs out however, it must be replaced.

This is the same concept as galvanisation, where when the zinc corrodes, it transfers/loses electrons to the iron, protecting it as shown by the equation below:

$${\rm Zn}\,(s)+{\rm Fe}^{2+}\,({\rm aq}) \longrightarrow {\rm Zn}^{2+}\,({\rm aq})+{\rm Fe}\,(s)$$

9.6 Extraction of metals

Metals are found in the Earth's crust primarily as ores because of how reactive they are. Some unreactive metals such as gold and silver can be found as pure metals. Some common ores include:

- Bauxite (aluminium ore)
- Copper pyrites (copper ore)
- Hematite (iron ore)
- Rock salt (sodium ore)
- Zinc blende (zinc ore)

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(and yes, ores have names fellow minecraft players, better learn them)

To figure out how to extract the metal from these ores, we can refer to the metal's reactivity along with the table below.

Position in the reactivity series	Method
Under hydrogen	Exist as pure metals
Above hydrogen but under carbon	Metal oxide reacted with carbon
Above carbon	Electrolysis

Extraction of iron from hematite

Iron is extracted from its ore, hematite, in a blast furnace.

This process uses the following:

- carbon in the form of coke (made by heating coal)
- limestone
- hematite (iron ore)



Fig. 2: A blast furnace being used to extract iron from hematiteJ Navas, CC BY-SA 3.0, via Wikimedia Commons

• Hot air is sent in blasts through the bottom of the furnace (see diagram above) which heats up the iron ore as the coke burns in the hot air, forming carbon dioxide.

 $\circ \ C\left(s\right) + O_{2}\left(g\right) \longrightarrow CO_{2}\left(g\right)$

• The limestone then begins to decompose due to the temperature

$$\circ \ CaCO_{3}\left(s\right) \longrightarrow CaO\left(s\right) + CO_{2}\left(g\right)$$

- The carbon dioxide produced earlier reacts with more hot coke/carbon in the furnace, producing carbon monoxide in an endothermic reaction.
 - $\circ \ \mathrm{CO}_{2}\left(\mathrm{g}\right) + \mathrm{C}\left(\mathrm{s}\right) \longrightarrow 2 \, \mathrm{CO}\left(\mathrm{g}\right)$
- Carbon monoxide is a reducing agent, which means that it will reduce the iron(III) oxide in the ore and oxidise itself to form carbon dioxide and also forms iron.

 \circ iron(III) oxide + carbon monoxide \longrightarrow iron + carbon dioxide

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- $\circ \ Fe_{2}O_{3}\left(s\right)+3\,CO\left(g\right) {\longrightarrow}\,2\,Fe\left(s\right)+3\,CO_{2}\left(g\right)$
- This molten iron goes to the bottom of the furnace due to its high density.
- The calcium oxide produced earlier by the decomposition of limestone is a base and reacts with acidic impurities such as silicon(IV) oxide in the iron to form something called slag. Slag is basically calcium silicate.
 - $\circ \operatorname{CaO}(s) + \operatorname{SiO}_2(s) \longrightarrow \operatorname{CaSiO}_3(s)$
 - This slag also goes to the bottom of the furnace, but it is above iron since it is less dense than iron.

Molten iron and slag is then piped out separately at regular intervals.

The waste gases produced such as nitrogen and oxides of carbon escape from the top of furnace. These waste gases also heat the air and reduce energy costs.

Slag produced by this reaction can be used for foundations of buildings and roads.

Extraction of aluminium

The main ore of aluminium is bauxite. Aluminium is extracted by the process of electrolysis. This has already been covered in chapter 4 <u>here</u>.

Chapter 10

10.1 Water

Water has many uses apart from sustaining life. Such as:

- Home
 - Cooking
 - Cleaning
 - Drinking
- Industry
 - As a solvent
 - As a coolant
 - For cleaning
 - As a chemical reactant

Pure Water is a colorless liquid which (at r.t.p) boils at 100°C and freezes at 0°C.

Note that tap water is NOT pure water. It contains many impurities, therefore does not boil at 100°C.

Distilled water is extremely pure and is therefore used in experiments over tap water.

We can also use the information above to test for the purity of water.

Test for the presence of water

There are two methods to test for the presence of water.

- Anhydrous cobalt(II) chloride
- Anhydrous copper(II) sulfate

Anhydrous cobalt(II) chloride

Anhydrous cobalt(II) chloride can be used to test for the presence of water. When anhydrous cobalt(II) chloride (blue in color) comes into contact with water, it undergoes a chemical reaction (hydration) and forms a pink-colored hydrated cobalt(II) chloride. This change in color indicates the presence of water.

Anhydrous copper(II) sulfate

Anhydrous copper(II) sulfate can also be utilized to test for the presence of water. When anhydrous copper(II) sulfate (white in color) is exposed to water, it readily absorbs the water molecules and forms a blue-colored hydrated copper(II) sulfate. The change in color from white to blue indicates the presence of water.

Unique Properties of water

Water is a pretty unique substance. It is an excellent solvent for many ionic substances and has many unusual properties such as

- It has a fairly high boiling point which is unusual for it's relatively low molecular mass.
- It has greater specific heat capacity (energy required to increase temperature basically) than almost any other liquid.
- Its density **decreases** when it freezes.

Water pollution and treatment

Water from natural sources/tap water can have things like oxygen, metals, plastics, sewage, harmful germs, nitrates, and phosphates. We need to make sure water is clean and safe.

Some of these substances have benefits:

- a. Dissolved oxygen is good for aquatic life to survive.
- b. Some metal compounds provide essential minerals for life.

However, some substances can be harmful:

- a. Certain metal compounds can be toxic.
- b. Plastics can harm aquatic life.
- c. Sewage contains harmful microbes that cause disease.
- d. Nitrates and phosphates can reduce oxygen in water and harm aquatic life.

Treating water

We need to treat water so that we can remove the harmful substances and make it drinkable.

- 1. Water is filtered to remove debris and solid particles.
- 2. Aluminum sulfate is added to clump together small particles.
- 3. Coarse sand filtration further removes particles.
- 4. Water enters a sedimentation tank for settling.
- 5. Fine sand filtration removes remaining impurities.
- 6. Chlorine is added to kill microbes and ensure safety.

10.2 Fertilisers

Before even starting this topic, you need to know what fertilisers are.

In simple words, fertilizers are substances that help plants grow better by providing them with the nutrients they need.

They basically enrich the soil with essential nutrients.

NPK elements

These essential nutrients are the NPK elements (Nitrogen, Phosphorus and Potassium).

Each NPK element is responsibile for the healthy growth of the plant but in different ways.

Nitrogen (really important for exams)

As you may know that the process of photosynthesis requires chlorophyll, a green pigment which absorbs sunlight. Nitrogen is a major constituent of chlorophyll. Therefore, without nitrogen, the plant would lose it's green color (chlorophyll) and will not able to produce food, and therefore will die.

Phosphorus

Plants have roots which help secure the plant into the ground. Phosphorus promotes the growth of roots, therefore making stronger roots.

Potassium

Potassium basically contributes to the overall health and growth of the plant.

Artificial fertilisers

Even though the soil naturally contains some nutrients, it may not always have enough for healthy growth of the plants.

Therefore, we might need to add artificial fertilisers.

We use ammonium salts and nitrates as fertilisers.

The most commonly known artificial fertiliser is ammounium nitrate (NH_4NO_3) which is formed by the following reaction:

 $ammonia + \ nitric \ acid \longrightarrow \ ammonium \ nitrate$

 $NH_{3}(g) + HNO_{3}(aq) \longrightarrow NH_{4}NO_{3}(aq)$

Ammounium nitrate is a nitrogenous fertilizers.

As you may have guessed, nitrogenous fertilizers are fertilizers that contain nitrogen 💀.

Some nitrogenous fertilizers include:

- Ammonium Nitrate - NH_4NO_3

- Ammonium Phosphate $(NH_4)_3PO_4$
- Ammonium Sulfate $(NH_4)_2SO_4$
- Urea $CO(NH_2)_2$

10.3 The air

Composition of the air

Clean, dry air is a mixture of several compounds **and** elements. With 78% of the air being Nitrogen, 21% being Oxygen, 0.9% Argon, 0.04% Carbon dioxide and the rest is a mixture of noble gases.

Air pollution

The air can become polluted in multiple ways but the syllabus states 5 air pollutants, therefore we'll cover only those five.

Carbon Dioxide

The primary source of carbon dioxide in the air is from the **complete** combustion of fossil fuels such as coal, oil, natural gas, petroleum etc.

Higher levels of carbon dioxide leads to a rise in global warming, therefore causing climate change.

Carbon Monoxide and particulates

Carbon monoxide and particulates are formed due to the **incomplete** combustion of a carboncontaining fuel (fossil fuel). The combustion is incomplete due to the lack of oxygen.

Carbon monoxide is a toxic gas.

Breathing in carbon monoxide is dangerous because it sticks to our red blood cells and stops them from carrying oxygen to our body parts, which can harm important organs like the brain and heart.

While particulates increase the risk of respiratory problems and cancer

Methane

Methane is produced through the decomposition of vegetation and the digestion process in animals.

Similarly to carbon dioxide, higher levels of methane leads to a rise in global warming, therefore causing climate change.

Carbon dioxide and methane, as greenhouse gases, contribute to global warming by trapping heat in the atmosphere. They act like a blanket, absorbing and re-emitting thermal energy, which prevents it from escaping into space. This leads to an increase in global temperatures as more heat is retained rather than being released back into space.

Nitrogen oxides

Car engines produce nitrogen oxides (NOx), and catalytic converters help reduce these emissions.

Nitrogen oxides can still contribute to environmental problems like the formation of acid rain and photochemical smog.

These pollutants can also cause respiratory problems for people who are exposed to high levels of NOx.

Sulfur Dioxide

When fossil fuels containing sulfur compounds are burned, sulfur dioxide (SO_2) is released.

Sulfur dioxide contributes to the formation of acid rain when it reacts with moisture in the atmosphere.

Reducing the effect of air pollution

Climate Change

- Planting trees: trees carry out photosynthesis, which uses carbon dioxide and produces oxygen, you will read more about this below.
- Decreasing use of fossil fuels: Use cleaner energy sources like wind, solar.

Acid Rain

- Catalytic converters: Installing catalytic converters in all cars, you'll read more about this below.
- Reducing emissions of sulfur dioxide: Using low sulfur fuels and flue gas desulfurization units (more about them below).

Flue Gas Desulfurization Units

Flue gas desulfurization (FGD) is a method that removes sulfur dioxide (SO2) from industrial emissions by using substances like calcium oxide or calcium carbonate to capture and remove the sulfur dioxide before it is released into the atmosphere. This helps reduce the formation of acid rain.

Catalytic converters

Catalytic converters are devices installed in car exhaust systems that contain catalysts, such as platinum, to facilitate chemical reactions that convert harmful pollutants, like nitrogen oxides and carbon monoxide, into less harmful substances.

Example equation:

 $Carbon monoxide + Nitrogen(II) oxide \longrightarrow Carbon dioxide + Nitrogen$

 $2 \operatorname{CO} + 2 \operatorname{NO} \longrightarrow 2 \operatorname{CO}_2 + \operatorname{N}_2$

Photosynthesis

Photosynthesis is the reaction between carbon dioxide and water to produce glucose and oxygen in the presence of chlorophyll and using energy from light.

 $Carbon\; dioxide + Water \longrightarrow Glucose + Oxygen$

 $6\,\mathrm{CO}_2 + 6\,\mathrm{H}_2\mathrm{O} \longrightarrow \mathrm{C}_6\mathrm{H}_{12}\mathrm{O}_6 + 6\,\mathrm{O}_2$

Chapter 11

11.1 Formulae, functional groups and terminology

Displayed Formulas

Displayed formulas show the arrangement of molecules and bonds in a compound. An example is shown below.



Fig. 1: Displayed formula of ethane ()NEUROtiker, Public domain, via Wikimedia Commons

From this displayed formula, we can figure out that ethane has 2 Carbon atoms and 6 Hydrogen atoms and can also deduce it's formula (C_2H_6).

We can also see that only single bonds are being used. Double bonds are denoted by \equiv while triple bonds are denoted by \equiv .

Homologous series

A homologous series is a family of compounds which have a similar structure and similar name endings, they also behave chemically in a similar way due to the presence of the same functional group.

In simpler words, a homologous series is a **family of similar compounds** with **similar chemical properties** due to the presence of the **same functional group**.

All the members of such a series can be represented by a formula.

The members of homologous series have similar characteristics, such as:

• having the same functional group

- having the same general formula
- differing from one member to the next by a -CH2- unit
- displaying a trend in physical properties
- sharing similar chemical properties

There are 4 such series:

- Alkanes (e.g. methane)
- Alkenes (e.g. ethene)
- Alcohols (e.g. ethanol)
- Carboxylic acids (e.g. ethanoic acid)

Each have a different general formula. For each of the formulas below, n is the number of carbon atoms present.

Alkanes

Alkenes

 $C_n H_{2n}$

 C_nH_{2n+2}

Alcohols

 $C_nH_{2n+1}OH$

Carboxylic acids

 $\mathbf{C}_{n}\mathbf{H}_{2n+1}\mathbf{COOH}$

Functional groups

A functional group is an atom or group of atoms that determines the chemical properties of a homologous series.

Some functional groups include: -C00H (carboxylic acids), -OH (alcohols).

Structural formulas

A structural formula provides a clear and precise representation of the arrangement of atoms in a molecule.

For example, the structural formula of C_2H_6 (ethane) will be CH_3CH_3

Other important examples include $CH_2 = CH_2$ (ethylene), CH_3CH_2OH (ethanol), and CH_3COOCH_3 (methyl acetate).

Structural isomers

Structural isomers are compounds with the **same molecular formula** (i.e., the same number and types of atoms) but **different structural formulae** (i.e., different arrangements of atoms within the molecules).

- 1. C_4H_{10} has two structural isomers:
 - $\circ CH_3CH_2CH_2CH_3$ (butane)
 - $\circ \ CH_3CH(CH_3)CH_3$ (isobutane)





Fig. 3: Isobutance

- 1. C_4H_8 also has two structural isomers:
 - $CH_3CH_2CH=CH_2$ (1-butene)
 - CH₃CH=CHCH₃ (2-butene)



Fig. 4: Structure of 1-butene



Fig. 5: 2-butene

The 1- and 2- show the position of the double bond. Now you may ask from which side, left or right? The answer is whichever one is the smallest. So you can start counting from left if the bond is near to the left side else you can start counting from right if the bond is near right.

Saturated vs unsaturated compounds

A saturated compound is a compound in which all carbon–carbon (C-C) bonds are single bonds.
On the other hand, an unsaturated compound is a compound in which one or more carbon–carbon bonds are not single bonds which means that there is at least one double (or a triple) bond between the carbon atoms.



Fig. 6: A saturated hydrocarbon, ethane



11.2 Naming organic compounds

The prefix of a compound refers to the number of carbon atoms in that compound.

- meth- means 1 carbon atom
- eth- means 2 carbon atoms
- prop- means 3 carbon atoms
- but means 4 carbon atoms and so on...

for example, simply by the name of a compound known as butane, we can now tell that there are 4 carbon atoms by the prefix. We can also figure out the general formula of it as it ends in -ane indicating that it is an alkane, which means that its general formula would be:

 $\mathrm{C}_n\mathrm{H}_{2n+2}$

and since n = 4, the formula of the compound can be calculated which would be:

C_4H_{10}

We can also do this with other compounds which do not end in -ane, such as:

- -ene, alkene
- -o1, alcohol
- -oic, carboxyllic acid

Here is a list of some displayed formulas:



Fig. 8: Ethane, C₂H₆



Fig. 10: Ethene, C₂H₄



Fig. 9: Methane, CH₄



Fig. 11: Ethanol, C₂H₅OH



Fig. 12: Ethanoic acid, CH₃COOH

Alcohols





Fig. 14: Propan-2-ol

Notice the position of the -OH functional group in the alcohols above? The position of this functional group determines the name of the compound.

11.4 Alkanes

Bonding in alkanes consists only of single, covalent bonds which is why all alkanes are saturated hydrocarbons.

Alkanes are usually unreactive, except in reactions like combustion and substitution by chlorine.

Combustion

Gasesous alkanes, such s methane will easily burn in excess air to produce carbon dioxide, water and heat energy.

```
methane + oxygen \longrightarrow carbon dioxide + water + \DeltaH
```

$$\mathrm{CH}_{4}\left(\mathrm{g}
ight)+2\,\mathrm{O}_{2}\left(\mathrm{g}
ight)\longrightarrow\mathrm{CO}_{2}\left(\mathrm{g}
ight)+2\,\mathrm{H}_{2}\,\mathrm{O}\left(\mathrm{g}
ight)$$

You'll learn more about this later in this chapter.

Substitution

In a substitution reaction, one atom or group of atoms is **replaced** by another atom or group of atoms.

For example, **chlroine** will react with **methane** in the **presence of UV light/sunlight**. This is a photochemical reaction which means it requires sunlight/UV light. This light is also the source of activiation energy for the reaction.

This reaction would form chloromethane and hydrogen chloride (hydrochloric acid, but gaseous), the equation would be:

methane + chlorine \longrightarrow chloromethane + hydrogen chloride

 $CH_{4}(g) + Cl_{2}(g) \longrightarrow CH_{3}Cl(g) + HCl(g)$



Fig. 15: Chloromethane

Bonding in alkenes includes a double carbon–carbon covalent bond which means that alkenes are unsaturated hydrocarbons.

Examples of alkenes include: ethene, propene, butene.

Important

Note that there's nothing known as methene, this is because a minimum of 2 carbon atoms are required to make a alkene, while meth- means only 1 carbon atom is present.

Catalytic cracking

Very few alkenes are found in nature. To obtain alkenes, we break down larger, less useful, alkane molecules obtained from the **fractional distallation of petroleum** (later in this chapter).

This breaking down of larger alkanes is known as catalytic cracking.

In this process, the large alkane molecules are passed over a **mixture of aluminium and chromium oxides** heated to 550°C.

An example of this is:

 $\mathrm{dodecane} \longrightarrow \mathrm{decane} + \mathrm{ethene}$

 $C_{12}H_{26}\left(g\right) \longrightarrow C_{10}H_{22}\left(g\right) + C_{2}H_{4}\left(g\right)$

As the name suggests, a catalyst is also required for this reaction.

Now why do we do this? because larger alkanes are not very useful while smaller hydrocarbons are more useful and therefore in higher demand.



Fig. 16: Cracking of paraffin in a lab

Addition reactions

Addition reactions are reactions in which only one product is formed.

Hydrogenation

To make an alkane from an alkene, we can react the alkene with hydrogen in an addition reaction.

The conditions necessary for this reaction include:

- A nickel catalyst
- A temperature of 200°C

Example:

 $ethene + hydrogen \longrightarrow ethane$

$$\mathrm{C_{2}H_{4}\left(g
ight)+H_{2}\left(g
ight)}\longrightarrow\mathrm{C_{2}H_{6}\left(g
ight)}$$





Hydration

This reaction involves reacting an alkene with steam to form an alcohol.

The conditions necessary for this reaction are:

- A catalyst of phosphoric (V) acid
- A temperature of 300°C
- A pressure of 6000 kPa or 60 atmospheres

These conditions are chosen so that the highest yield possible is produced (favour the forward reaction).

This process is used in the manufacture of ethanol.

 $ethene + steam \underbrace{\xrightarrow{300^\circ \text{C}, \, 60 \text{ atm}}}_{\text{phosphoric(V) acid}} ethanol$

 $C_{2}H_{4}\left(g\right)+H_{2}O\left(g\right)\Longrightarrow C_{2}H_{5}OH\left(g\right)$

This is a reversible reaction as shown by the symbol \rightleftharpoons .

The percentage yield of this reaction is 96%.

This is also a continuous process which means reactants are continuously fed.

Halogenation

Halogenation is basically a test we use to check if a compound is saturated or unsaturated.

In this reaction, we add a few drops of aqueous bromine to a hydrocarbon and stir it. If the compound is a saturated compound, nothing happens and there is no color change.

However, if it is an unsaturated compound, then the color of the bromine water (aqueous bromine) which is red-brown disappears which means a reaction has occured.

 $ethene+bromine \longrightarrow dibromoethane$

 $C_{2}H_{4}\left(g\right) + Br_{2}\left(aq\right) \longrightarrow C_{2}H_{4}Br_{2}\left(aq\right)$



Fig. 18: Halogenation of ethene

11.6 Alcohols

Alcohols is a homologous series with the functional group -OH. Examples include:

- Methanol
- Ethanol
- Propanol
- Butanol

Manufacture of ethanol

There are two methods to manufacture ethanol:

- Fermentation of glucose
- Hydration of ethene (catalytic addition of steam)

Fermentation of glucose

Ethanol can be formed by the fermentation of aqueous glucose. The conditions for this reaction are:

- A temperature of 25–35°C
- A catalyst of yeast
- Absence of oxygen

This reaction will not occur at temperatures above 35°C because the yeast (catalyst) will get denatured at a higher temperature.

 ${\rm glucose} \xrightarrow{{\rm yeast}} {\rm ethanol} + {\rm carbon} \ {\rm dioxide}$

 $C_6H_{12}O_6(aq) \longrightarrow 2 C_2H_5OH(l) + 2 CO_2(g)$

This process is not continous and therefore has to be done in batches. Once the reaction is completed, ethanol needs to be separated from the mixture by fractional distillation.

The percentage yield of this reaction is approximately 15%.

Hydration of ethene

This reaction has already been covered above in 11.5, click here.

Advantages and disadvantages of hydration and fermentation

Method	Advantages	Disadvantages
Fermentation	 Low energy consumption Uses easily available materials Uses renewable materials 	 Slow process Not continuous, needs to be done in batches Less yield
Hydration of ethene	 Fast process Continuous process High yield 	 High energy usage Needs ethene which is obtained through catalytic cracking Ethene is non-renewable

Uses of ethanol

There are many uses of ethanol, some notable ones are:

- as a fuel
- as a solvent (it can dissolve a wide range of substances)

As a fuel (combustion of ethanol)

Ethanol combusts in excess air to produce heat energy along with carbon dioxide and water.

 $ethanol + oxygen \longrightarrow carbon \; dioxide + water + \Delta H$

 $\mathrm{CH}_{3}\mathrm{CH}_{2}\mathrm{OH}\left(l\right)+3\,\mathrm{O}_{2}\left(g\right)\longrightarrow2\,\mathrm{CO}_{2}\left(g\right)+3\,\mathrm{H}_{2}\mathrm{O}\left(g\right)+\Delta\mathrm{H}$

Remember CH_3CH_2OH is the <u>structural formula</u> for ethanol.

Some countries mix ethanol with petrol as a fuel for cars. It is also used for heating and cooking using domestic burners.

11.7 Carboxylic acids

Carboxylic acids is a homologous series with the functional group (-COOH). Examples include:

- Methanoic acid
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- Ethanoic acid
- Propanoic acid
- Butanoic acid



methanoic acid

ethanoic acid

cid propanoic acid

butanoic acid

Fig. 19: Examples of carboxylic acids

Carboxyllic acid reactions

Ethanoic acid + base

Even though ethanoic acid is a weak acid, it will still react with a base to make a salt.

 $ethanoic \ acid + sodium \ hydroxide \longrightarrow sodium \ ethanoate + water$

$$CH_3COOH(aq) + NaOH(aq) \longrightarrow CH_3COONa(aq) + H_2O(l)$$

Ethanoic acid + metal

 $ethanoic acid + magnesium \longrightarrow magnesium ethanoate + hydrogen$

$$2 \operatorname{CH}_{3} \operatorname{COOH}\left(\operatorname{aq}\right) + \operatorname{Mg}\left(\operatorname{s}\right) \longrightarrow (\operatorname{CH}_{3} \operatorname{COO})_{2} \operatorname{Mg}\left(\operatorname{aq}\right) + \operatorname{H}_{2}\left(\operatorname{g}\right)$$

Ethanoic acid + carbonate

 $ethanoic\ acid + sodium\ carbonate \longrightarrow sodium\ ethanoate + carbon\ dioxide + water$

$$2 \operatorname{CH}_{3} \operatorname{COOH}\left(\operatorname{aq}\right) + \operatorname{Na}_{2} \operatorname{CO}_{3}\left(\operatorname{s}\right) \longrightarrow 2 \operatorname{CH}_{3} \operatorname{COONa}\left(\operatorname{aq}\right) + + \operatorname{CO}_{2}\left(\operatorname{g}\right) + \operatorname{H}_{2} \operatorname{O}\left(\operatorname{l}\right)$$

Formation of ethanoic acid by the oxidation ethanol

Bacterial oxidation to form vinegar

Vinegar is the common name for ethanoic acid, which is also known as acetic acid.

Vinegar is manufactured by the oxidation of alcohol.

In this process, the bacteria in the solution oxidize ethanol using oxygen from the air, producing ethanoic acid (CH_3COOH) and water (H_2O) as the end products.

The equation for this reaction is:

 $ethanol + oxygen \longrightarrow ethanoic \ acid + water$

$CH_3CH_2OH + O_2 \longrightarrow CH_3COOH + H_2O$

Using acidified aqueous potassium manganate(VII)

Ethanol can also be oxidised to ethanoic acid by using an oxidising agent such as acidified aqueous potassium manganate(VII).

The acidified aqueous potassium manganate(VII) purple in color, but it loses its color when it reacts.

acidified aqueous potassium manganate(VII) is also known as potassium permanganate, $\rm KMnO_4$

 $ethanol + oxygen \xrightarrow{heat} ethanoic acid + water$

 $CH_{3}CH_{2}OH(l) + 2[O] \longrightarrow CH_{3}COOH(aq) + H_{2}O(l)$

Note that the oxygen is in square brackets ([O])? This is because the oxygen is from $KMnO_4$ and is acting as an oxidizing agent. The square brackets denote that oxygen is acting as an oxidizing agent.

Esters

Esters are organic compounds made when a carboxyllic acid reacts with an alcohol. In this reaction, a catalyst of concentrated sulfuric acid is used (a few drops.)

For example, a reaction between ethanoic acid and ethanol:

 $ethanoic acid + ethanol \xleftarrow{H_2SO_4} ethyl ethanoate + water$

$$CH_3COOH(l) + C_2H_5OH(l) \Longrightarrow CH_3COOC_2H_5 + water$$

This reaction is known as **esterification**.

Esters usually have strong and sweet/pleasant smells and are used to make perfumes along with food flavourings. Fats and oils are naturally occuring esters.

Naming esters

The name of an ester contains 2 parts.

The first part of the name is the alcohol and the second part of the name is the acid.

For example, propyl ethanoate means that the alcohol used was propanol and acid was ethanoic acid.

However, when writing the formula for esters, we reverse this. which means that in the formula, we write the acid part first and then the alcohol part.



Fig. 20: Structure of propyl ethanoate



Fig. 21: Structure of propyl methanoate

11.8 Polymers

Polymers are large molecules built up from many smaller molecules called **monomers**.

When multiple monomers join together to form polymers, it is known as polymerisation.

Poly(ethene), also known as polythene, is a plastic. It is made by heating ethene to a relatively high temperature under a high pressure in the presence of a catalyst.



Fig. 22: Polymerisation of ethene

n respresents the number of ethene monomers. A polymer chain consists of many thousands of monomer units and a single piece of plastic would contain millions of polymer chains.

Since in this example, only one product is being formed, we can call this **addition polymerisation**.

Poly(ethene) is useful as it is easy to mould and it is an excellent insulator. However there are many disadvantages of using plastics, such as:

- It is not affected by weather and does not corrode making it hard to dispose
- They are non-biodegradable, which means they don't decompose naturally.
- It is also hard to dispose plastic in landfills as they don't decompose and end up taking a lot of space.
- Plastic can gather in water bodies causing harm to marine life
- Burning plastics releases toxic gases

Addition and condesation polymerisation

There are two types of polymers:

- Addition polymers
- Condensation polymers

Addition polymers

These polymers are made up of only one type of monomer combined in a chain. The structure of an addition polymer is like:

where A = monomer.

Condensation polymers

Condensation polymers are made up of two or more types of monomers.

An example of a condensation polymer is **nylon**.

The structure of nylon is like

where A and B are different types of monomers.

The monomers for nylon are a lot more complex than that of poly(ethene) and are 1,6diaminohexane and hexanedioic acid.

There will be more than one product of a reaction between these monomers, and that other product in this case would be water, which would be lost.

Hexanedioic acid + 1,6-diaminohexane

$$\mathrm{HOOC}(\mathrm{CH}_2)_4\mathrm{COOH} + \mathrm{H}_2\mathrm{N}(\mathrm{CH}_2)_6\mathrm{NH}_2$$

 $H_2N(CH_2)_6$ **CONH** $(CH_2)_4COOH + H_2O$

Notice the highlighted **CONH** in the product? That is called an **amide link** (also known as a peptide link). Because an amide link is being formed, we can call nylon a **polyamide**.



Fig. 23: The structure of nylon

Using the structure above, we can figure out the repeat unit for polyamides which is:



Fig. 24: The repeat unit for a polyamide

As the name suggests, the repeat unit is repeated to form a polymer.

Proteins are also an example of polyamides.

Polyesters

An example of a popular polyester is PET.

PET stands for polyethylene terephthalate (you're not required to learn the full name of PET).

To produce PET, we react ethane-1,2-diol with benzene-1,4-dicarboxylic acid (terephthalic acid).

ethane-1,2-diol + benzene-1,4-dicarboxylic acid

$$HO(CH_2)_2OH + HOOC(C_6H_4)COOH$$

 \downarrow

$$HO(CH_2)_2 \text{ OCO } (C_6H_4)COOH + H_2O$$

Notice the 0C0 in the product? that is an ester link which means that PET is a polyester.



Fig. 25: Structure of a polyester



Fig. 26: The repeat unit of a polyester

PET can be converted back into monomers and can also be re-polymerised which makes it recyclable.

Fats are also an example of polyesters.

Natural polyamides

As mentioned earlier, proteins are **natural polyamides**. They are formed from amino acid monomers which have a **general** structure of:



Fig. 27: The general structure of an amino acid

R represents different types of side chain.



Fig. 28: General structure of a protein

11.3 Fuels

Fuels which were formed millions of years ago from dead plants and animals are known are fossil fuels. They are hydrocarbon-based energy sources which can be combusted to release energy. Examples include:

- Coal
- Natural gas
- Petroleum

Hydrocarbons are compounds that contain hydrogen and carbon only.

The main constituent of natural gas is methane (CH_4) .

Petroleum is a mixture of hydrocarbons.

Fractional distillation of petroleum

Fractional distillation is a process used to separate petroleum into its useful fractions based on differences in boiling points

In this process, petroleum is heated to 400°C to vaporise all parts of the mixture.

Fractional Distillation Table

Fraction Name	Temperature
Refinery gas	30°C
Gasoline/petrol	110°C
Naphtha (Chem feedstock)	180°C
Kerosene/paraffin	260°C
Diesel	300°C
Fuel oil (ship and home heating)	300°C
Lubricating	340°C
Bitumen (making roads)	350°C

Viscosity increases as you move down the table making **Bitumen** the most viscous fraction.

Boiling point also increases as you move down the table.

The length of the hydrocarbon chain (and the number of carbon atoms) increase as you go down the table as well.

However, the volatility decreases as you go down the table

Uses of these fractions

Each fraction can be used for a different purpose, such as:

• refinery gas fraction for gas used in heating and cooking

- gasoline / petrol fraction for fuel used in cars
- naphtha fraction as a chemical feedstock
- kerosene / paraffin fraction for jet fuel
- diesel oil / gas oil fraction for fuel used in diesel engines
- fuel oil fraction for fuel used in ships and home heating systems
- lubricating oil fraction for lubricants, waxes and polishes
- bitumen fraction for making roads

Chapter 12

12.1 Apparatus used for measurement in chemistry

It is important to use the correct apparatus for measuring something.

Measurement of time

Experiments involving rates of reaction will require a stopwatch - one that is accurate and measures time to a 100th of a second.

Measurement of temperature

The most commonly used thermometers in a lab are alcohol-in-glass thermometers. Mercury in-glass thermometers can also be used, but should be handled with care as the mercury inside them is poisonous.

It is also important that the eye level is equal to the liquid's meniscus to make sure readings are accurate.

Measurement of a mass

A weighing scale/balance is used to measure mass.

The precision of an electronic balance is the size of the smallest mass that can be measured on the scale settings you're using.

Measurement of volume of liquids

1 litre = $1 \text{ dm}^3 = 1000 \text{ cm}^3$

burettes volumetric pipettes measuring cylinders

are the 3 instruments used to measure volume that are in the syllabus.

Measurement of volume of gas

The volume of a gas can be measured using a gas syringe They have a maximum volume of 100 $\rm cm^3$

Definitions

- solvent is a substance that dissolves a solute
- solute is a substance that is dissolved in a solvent
- solution is a mixture of one or more solutes dissolved in a solvent
- saturated solution is a solution containing the maximum concentration of a solute dissolved in the solvent at a specified temperature
- residue is a substance that remains after evaporation, distillation, filtration or any similar process
- filtrate is a liquid or solution that has passed through a filter

12.2 Acid–base titrations

An acid-base titration is a method to measure the amount of acid or base in a solution by reacting it with a solution of known concentration until the reaction is complete.

In an acid-base titration, we use a **burette** to slowly add a solution of known concentration into a flask kept below it. The flask contains the other solution we want to measure, which is called the **analyte**. To pour this analyte into the flask, we use a **volumetric pipette** because it is extremely accurate. We also add a pH indicator such as universal indicator, litmus or phenolphthalein to the flask so that we can know when the reaction has ended. This is because the pH should neutral when the reaction ends, universal indicator for example, turns green when the pH is neutral (7).



Fig. 1: Acid-Base titration

12.3 Chromatography

A method used for the seperation of dissolved substances is known as chromatography. There are multiple types of chromatography but they all follow the same principles.

The simplest type is paper chromatography. To seperate the different coloured dyes in a sample of black ink, a pencil line is drawn about 1 cm from the bottom of a piece of chromatography paper. This is called the baseline. **DO NOT USE A PEN TO DRAW THE BASELINE**. Then a spot of ink will be put on the baseline on the chromatography paper.

This paper is then put in a suitable container, such as a beaker containing a suitable solvent, ensuring that the level of solvent does not reach the pencil line.



Fig. 2: Paper chromatography

As the solvent moves up the paper, the dyes are carried with it and begin to seperate. They seperate because the substances have different solubilities in the solvent and are absorbed to different degrees by the chromatography paper.

An R_f value is defined as the ratio of distance travelled by the solute to the distance travelled by the solvent from the pencil line.

$$R_f = \frac{\text{distance travelled by solute}}{\text{distance travelled by solvent}}$$

If the substance is colorless, a locating agent is sprayed at the chromatogram. The locating agent will react with the colorless substances to form a coloured product. However, in some situations, ultraviolet light may be used to locate the position of the substances.

12.4 Separation and purification

Filtration

Filtration is the proess of seperating a solid from a liquid using a fine filter paper which does not allow the solid to pass through.

For example, when seperating a mixture of sand and water using a filter paper. The sand is called the **residue** while the water is called the **filterate**.

Evaportation

If a solid has been dissolved in a liquid, it can be seperated using the process of evaportation. The solution is heated so that the liquid evaporates completely and leaves the solid behind.

Crystallistion

A saturated solution of salt and water is known as a brine.

A saturated solution is a solution that contains the maximum concentration of the solute dissolved at a given temperature. (or simply, it is a solution in which no more solute can be dissolved)

The process of forming crystals from a liquid is known as crystallistion.

Simple distillation

If we want to obtain the solvent (liquid) from a solution, then we can use this process. The solution is put inside a beaker which has a condensor attached to it (diagram below).



The solution is then heated till all the water has evaporated and then condensed into the flask. The condensor is known as the Liebig condensor.

Cold water is passed through the "water in" and "water out" pipes of the condensor.

Seperating liquid/liquid mixtures

The liquids which do not mix easily are known as immiscible.

If two liquids are immiscible, they can be seperated using a seperating funnel. In a seperating funnel, the liquids are arranged by their density

To seperate miscible liquids, we use the process of fracational distillation. It relies on the different boiling points of liquids.

An example of this is the fracational distillation of crude oil in Chapter 11

Criteria for purity

A pure substance is one which contains a single element or compound not mixed with any other substance. Pure substances have a very specific melting and boiling point and will not have a range of melting/boiling points.

Another way to check for purity is to do a chromatography of the substance. Pure substances will produce **only one spot** on the chromatogram.

12.5 Identification of ions and gases

2023 onwards, a sheet of qualitative analysis will be given in the A.T.P. exam, you might still want to learn these as they might come in the theory paper as well.

Notes for use in qualitative analysis

Tests for anions

anion	test	test result
carbonate, CO32-	add dilute acid, then test for carbon dioxide gas	effervescence, carbon dioxide produced
chloride, C <i>l</i> - [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
bromide, Br⁻ [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	cream ppt.
iodide, I⁻ [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	yellow ppt.
nitrate, NO ₃ ⁻ [in solution]	add aqueous sodium hydroxide, then aluminium foil; warm carefully	ammonia produced
sulfate, SO₄²- [in solution]	acidify with dilute nitric acid, then add aqueous barium nitrate	white ppt.
sulfite, SO ₃ ²⁻	add a small volume of acidified aqueous potassium manganate(VII)	the acidified aqueous potassium manganate(VII) changes colour from purple to colourless

Tests for aqueous cations

cation	effect of aqueous sodium hydroxide	effect of aqueous ammonia
aluminium, Al ³⁺	white ppt., soluble in excess, giving a colourless solution	white ppt., insoluble in excess
ammonium, NH ₄ +	ammonia produced on warming	-
calcium, Ca ²⁺	white ppt., insoluble in excess	no ppt. or very slight white ppt.
chromium(III), Cr ³⁺	green ppt., soluble in excess	green ppt., insoluble in excess
copper(II), Cu ²⁺	light blue ppt., insoluble in excess	light blue ppt., soluble in excess, giving a dark blue solution
iron(II), Fe ²⁺	green ppt., insoluble in excess, ppt. turns brown near surface on standing	green ppt., insoluble in excess, ppt. turns brown near surface on standing
iron(III), Fe ³⁺	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
zinc, Zn ²⁺	white ppt., soluble in excess, giving a colourless solution	white ppt., soluble in excess, giving a colourless solution

Tests for gases

gas	test and test result	
ammonia, NH ₃	turns damp red litmus paper blue	
carbon dioxide, CO ₂	turns limewater milky	
chlorine, Cl ₂	bleaches damp litmus paper	
hydrogen, H ₂	'pops' with a lighted splint	
oxygen, O ₂	relights a glowing splint	
sulfur dioxide, SO ₂	turns acidified aqueous potassium manganate(VII) from purple to colourless	

Flame tests for metal ions

metal ion	flame colour
lithium, Li+	red
sodium, Na⁺	yellow
potassium, K*	lilac
calcium, Ca2+	orange-red
barium, Ba ²⁺	light green
copper(II), Cu2+	blue-green



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