

Cambridge GCSE Notes  
5070 Chemistry

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# 1 Stoichiometry

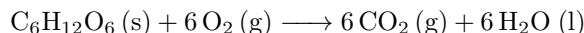
## 1.1 Formulae

All elements have symbols. These elements bond to form molecules, which can also be denoted by symbols.

The molecular formula of a compound shows the numbers of and the types of different elements in a molecule.  $C_6H_{12}O_6$  is an example of such a formula, that of glucose. The empirical formula of a compound is the simplest whole number ratio of the different atoms or ions in a compound.  $CH_2O$  is the empirical formula of glucose.

The overall charge in a molecule tends to be zero, unless stated otherwise. The numbers of atoms of each element and their corresponding charge can hence be deduced. Consider the case of a molecule of aluminium and chlorine,  $Al^{3+}$  and  $Cl^-$  respectively. Each atom of aluminium has a charge of +3, to cancel out which three chlorine atoms, each of charge  $-1$  is required. See that,  $3 + 3(-1) = 0$ .

Chemical reactions can be denoted by means of symbol equations. An example follows:



The above shows two reactants,  $C_6H_{12}O_6$  and  $O_2$  and two products,  $CO_2$  and  $H_2O$ .

Note that, the number of atoms of each element on either side of the “ $\longrightarrow$ ” symbol is equal. This is said to be a balanced symbol equation. The letters inside of the parentheses represent the state of matter in which each reactant and product is in, specifically: (s), solid; (l) liquid; (g) gas.

## 1.2 Relative masses of atoms and molecules

A carbon-12 atom has 6 protons and 6 neutrons. It is this isotope of carbon that is held as the standard for the atomic masses. That is, each proton and neutron, since both have masses with negligible difference, is said to have a mass of 1/12th of a carbon-12 atom. The relative atomic mass of an element is the average atomic mass of all the isotopes of the element, with respect to the amount of the isotope present. The relative atomic mass of an element is denoted  $A_r$ .

Consider the case of chlorine, 25% of all the world’s chlorine has an atomic mass of 37, and the remaining 75% has an atomic mass of 35. Thus:

$$A_r = (75\%)(35) + (25\%)(37) = 35.5$$

The relative molecular mass,  $M_r$  of a substance is the sum of all the relative atomic masses of its component elements.

## 1.3 The mole and the Avogadro constant

The mole (mol, in short), is the unit of amount of substance. Each mole contains  $N_A = 6.02 \times 10^{23}$  particles which are atoms, ions or molecules.  $N_A$  is the Avogadro constant or Avogadro’s number.

Understand that, the  $A_r$  of a substance is its molar mass, which is the grams of mass per mole of that substance. Knowing this,

$$\text{amount of substance (mol)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

Notice that the units cancel out to give (mol).

For gases, each mole of a gas has a volume of  $24 \text{ dm}^3$  at room temperature and pressure.

The concentration of a substance is the amount of it per unit volume. It can be measured as grams of substance per unit volume or moles of substance per unit volume. To convert between the two units, divide by molar mass and multiply molar mass to go from  $\text{g/dm}^3$  and  $\text{mol/dm}^3$  and back, respectively:

$$c (\text{mol dm}^{-3}) = (c (\text{g dm}^{-3}))(M_r)$$

hence,

$$c (\text{g dm}^{-3}) = \frac{c (\text{mol dm}^{-3})}{M_r}$$

The litre is the unit for measuring liquids. It is a measure of volume and is equivalent to the cubic decimetre ( $\text{dm}^3$ ). Small amounts of liquids are measured in  $\text{cm}^3$ , equivalent to the millilitre (ml). To convert from cubic decimetre and cubic centimetre, multiply by  $1000 = 10^3$  and vice versa.

Given the masses of element in a compound, or the percentage of each element per molecule of compound, we can derive the empirical formula of the compound. If we know the  $M_r$  of the compound, the molecular formula of the compound can be found.

Consider the case of a molecule where 33% of it is carbon and the rest is oxygen.

## 2 Electrochemistry

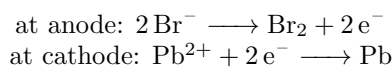
### 2.1 Electrolysis

Electrolysis is the decomposition of an ionic compound, in molten or aqueous form, by passage of electric current.

A simple electrolytic cell consists of two electrodes, the cathode and anode, connected to either side of a battery and an electrolyte. The anode is positively charged whereas the cathode is negatively charged. The electrolyte is the aqueous or molten substance that undergoes decomposition through electrolysis.

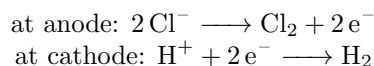
In the external circuit consisting of wires connected to the batteries, electrons flow from the battery to the cathode, making the cathode negatively charged. Here electrons are lost to substances in the electrolyte which are oxidised. The anode, being positively charged, attracts the negative ions from the ionic compound, oxidising them, by taking their excess electrons. Electrons are gained at the anode from negative ions, and are lost at the cathode to positive ions. The electrodes can be made of any conductive material, such as metals or graphite. Graphite electrodes are said to be inert as they do not take part in the reaction. Note that, oxidation occurs at anodes and reduction occurs at cathodes.

In case of the substance, lead (II) bromide,  $\text{PbBr}_2$ , the following happens during electrolysis with inert electrodes at either electrode:



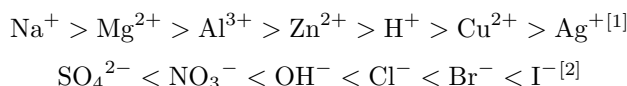
It is observed that a silvery solid accumulates at the cathode, i.e. lead (II) and a red-brown gas is given off at anode. Note that, from here we can judge that for molten ionic compounds, the metal part of the compound will always form at the cathode, and the non metal will form at the anode.

During electrolysis of concentrated aqueous sodium chloride, using inert electrodes, the following reactions occur at the electrodes:

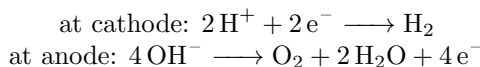


It is observed that a colourless gas is given off at cathode and a green gas is given off at the anode.

The reason why the metal, i.e., sodium is not discharged at the cathode is because it is not the least reactive amongst the cations present in the electrolyte. The ions present in the electrolyte are:  $\text{H}^+$ ,  $\text{Na}^+$ ,  $\text{Cl}^-$  and  $\text{OH}^-$ . Of these, at the cathode the ion discharged is the least reactive of the cations, and at the anode the ion discharged is the ion that loses electrons more readily, and is hence easier to discharge. The reactivity series of cations and the preferential discharge rule of anions are given below:

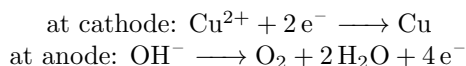


During the electrolysis of dilute sulfuric acid, the ions in the electrolyte are:  $\text{H}^+$ ,  $\text{OH}^-$ ,  $\text{SO}_4^{2-}$ . The only cation,  $\text{H}^+$  is released at the cathode and, by preferential rule,  $\text{OH}^-$  is the one which is oxidised at anode.



Colourless gases are given off at both electrodes, but the volume of oxygen produced is half that of hydrogen. This is a result of the fact that, the mole ratio of oxygen to hydrogen in the electrolyte is 1:2.

Consider the case of the electrolysis of aqueous copper (II) sulfate using inert electrodes. Uncharged copper (Cu) has a pinkish-orange colour, whereas copper ions,  $\text{Cu}^{2+}$  have a blue colour. The following happens at the electrodes in this electrolysis:

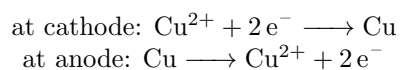


Notice that, the copper ions are reduced to copper, reducing the number of copper ions in electrolyte. It is observed that the blue colour of solution fades and a pinkish orange solid accumulates on the cathode.

An active electrode is that which takes part in the electrolytic process. Consider the case of the electrolysis of aqueous copper (II) sulfate with copper electrodes. It is seen that copper deposits on the cathode and copper is dissolved at the anode.

[1] where the ion to the left of  $>$  is more reactive than the ion to the right.

[2] where the ion to the left of  $<$  is less likely to be discharged than that to the right.

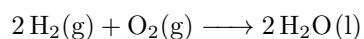


This property of active electrodes and metals comes in use during electroplating, which is a process of electrolysis in which a metal object is coated/plated with a layer of another metal.

To electroplate a metal object, we must place it as the cathode, and have the anode be the metal with which we want to plate the metal object. The electrolyte must be the solution of a salt with the metal to plate with.

## 2.2 Hydrogen-oxygen fuel cells

It is possible to use chemical reactions to produce electrical energy. Everyday batteries work in this way. A more efficient method would be using a fuel cell. A hydrogen oxygen fuel cell uses the following chemical reaction



Here, the hydrogen is considered to be a non-polluting fuel.

As a fuel, hydrogen is more efficient than any other, and the only waste product formed is water. However, due to it being difficult to produce and unsafe to store, it has failed to gain precedence.

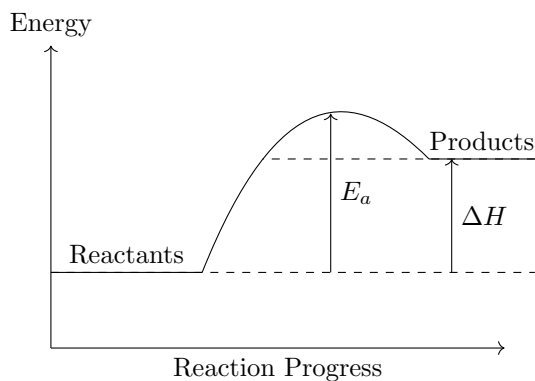
### 3 Chemical energetics

#### 3.1 Exothermic and endothermic reactions

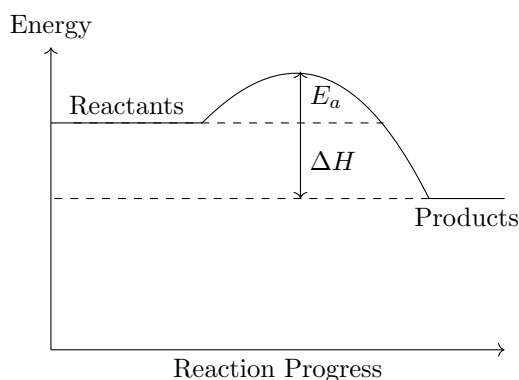
An exothermic reaction transfers thermal energy to the surroundings leading to an increase in the temperature of the surroundings. An endothermic reaction takes in thermal energy from the surroundings leading to a decrease in the temperature of the surroundings.

The transfer of energy during a reaction is called the enthalpy change,  $\Delta H$  of the reaction. The value is negative for exothermic reactions and positive for endothermic reactions.

Activation energy,  $E_a$  is the minimum energy that colliding particles must have to react.



Endothermic Reaction



Exothermic Reaction

Above are reaction pathway diagrams, showing the energy of a reaction as it progresses. The energy of products is greater than that of the reactants in endothermic reactions and the opposite is true for exothermic reactions. The differences in energy and their correspondences are labelled.

Remember that M-EXO (making is exo) and B-ENDO (breaking is endo).

Bond breaking is an endothermic process and bond making is an exothermic process. When a reaction, overall, has bonds broken with more energy than bonds made, it is endothermic. The opposite is also applies.

Enthalpy change can be calculated as:

$$\text{enthalpy} = (\text{energy needed to break bonds}) - (\text{energy needed to make bonds})$$

$$\Delta H = E_b - E_f$$

where  $E_b$  is the energy needed to break bonds and  $E_f$  is that needed to form bonds.

## 4 Chemical reactions

### 4.1 Physical and chemical changes

In a physical change, the substances present remain chemically the same and no new substances are formed. Such changes are often easy to reverse.

Changes such as melting and boiling are endothermic as heat is taken in, changes such as condensing and freezing are exothermic as heat is given out.

Chemical changes are those where a new substance is formed. Most chemical changes happen through chemical reactions, almost all of which are exothermic, very few are endothermic.

### 4.2 Rate of reaction

A reaction progresses when effective collisions occur between reactants to form a molecule of a product. The rate of effective collisions is hence the rate of the reaction. This rate of effective collisions can be influenced. These effective collisions can only occur if the colliding particles have a minimum energy ( $E_a$ ).

Increasing the number of particles per unit volume means there are more particles per unit volume, when that is the case the particles are more likely to react.

Increasing the kinetic energy (KE) of particles by applying heat makes them move faster, meaning more particles are likely to collide more frequently and with energy greater than  $E_a$ .

A catalyst is that which increases the rate of a reaction, decreases the  $E_a$  and is unchanged at the end of the reaction.

Changing solution concentration changes particles per unit volume, refer above.

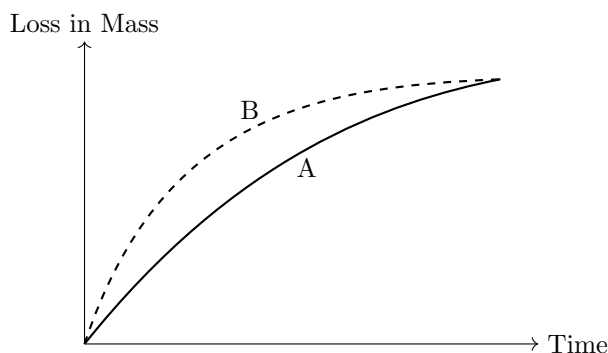
Changing gas pressure does the same, refer above.

Increasing surface area of solids means there are more exposed surfaces with which reactants can effectively collide, meaning effective collisions are more likely. The opposite is also true.

Increasing temperature increases KE of particles, refer above.

Addition of a catalyst, including enzymes, increases reaction rate and decreases  $E_a$ .

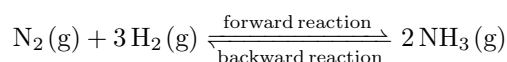
Investigating the rate of a reaction comes down to measuring the rate at which reactants are used up or the rate at which products are formed. The information can be represented in a graph:



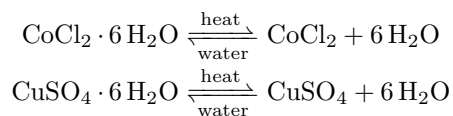
In reaction B, the reactants were powdered, with increased surface area and hence the reaction had a faster rate than A with just solid reactants. Notice that the loss in mass is equal but the rate of loss of mass is the difference.

### 4.3 Reversible reactions and equilibrium

Most reactions are one-way, i.e., reactants react to form products and that's it. Some reactions are reversible, in such reactions, reactants react to form products and those products also react to form reactants. Such reactions are shown with a " $\rightleftharpoons$ ":



Changing the direction of a reversible reaction can depend on the conditions applied to it. Such is the case for hydrating and making anhydrous salts.



Note that,  $\text{CuSO}_4$  is white while the hydrated form  $\text{CuSO}_4 \cdot \text{H}_2\text{O}$  is blue. Hydrated cobalt (II) chloride is pink whereas the anhydrous form is blue.

A closed system is that where no reactants or products can escape from the reacting system. In such a system, a reversible reaction is in equilibrium when the rate of forward reaction equals the rate of backward reaction and the concentrations of reactants and products are not changing. The reagents are reacting but the concentrations do not change.

The position of equilibrium tells us the rate of forward reaction compared to backward reactions. That means, if the position of equilibrium is to the right, the rate of forward reaction is greater than that of the backward reaction and vice versa.

Note that, when a change is made to the conditions of a system in dynamic equilibrium, the system moves so as to oppose that change.

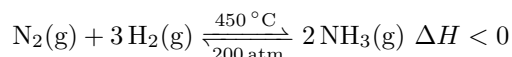
Changing temperature affects the reaction depending on the enthalpy of the reaction. Changing the temperature will cause the equilibrium to shift in the direction that will reverse that change. Increasing the temperature in a reaction where the forward reaction is exothermic will shift equilibrium to left, as the system will oppose the change by increasing rate of endothermic reaction to lower temperature. The opposite is also true.

Changing pressure only affects reactions whose reagents are all gaseous. Increasing pressure moves equilibrium toward the side which has less gaseous moles. The opposite is also true.

Increasing concentration of reactants moves equilibrium to right and more products are formed and vice versa.

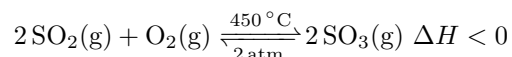
Using a catalyst does not affect equilibrium position, but the speed at which the system reaches equilibrium is affected by catalyst.

Below is the symbol equation for the Haber process.

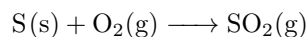


The nitrogen is gotten from fractional distillation of liquid air and hydrogen is gotten from cracking of crude oil. The above reaction requires an iron catalyst.

In the Contact process, sulfur dioxide is converted into sulfur trioxide.



The Contact process requires a vanadium (V) oxide,  $\text{V}_2\text{O}_5$  catalyst. Sulfur dioxide is gotten from burning or roasting sulfide ores:



and oxygen is gotten from air.

The conditions in the above industrial processes are such that to optimise economical costs and safety.

## 4.4 Redox

Roman numbers are used to indicate the oxidation number of an element in a compound (iron (II), vanadium (V), etc.).

A redox reaction is where reduction and oxidation occurs simultaneously.

The oxidation of a compound is its gain of oxygen, loss of electrons or increase in oxidation number.

The reduction of a compound is its loss of oxygen, gain of electrons or decrease in oxidation number.

Rules for identifying oxidation number:

- It is zero in uncombined states (B, Mg).
- It is the same as the charge on an ion ( $\text{B}^{+2}$ ).
- Sum of oxidation numbers in a compound is zero.
- Sum of oxidation numbers in an ion is equal to the charge in the ion.

Acidified potassium manganate(VII),  $\text{KMnO}_4$ , is an oxidising agent with a purple colour, which it loses when it is added to a reducing agent. Aqueous potassium iodide, KI, is a reducing agent which turns brown in presence of an oxidising agent.

An oxidising agent is that which oxidises another substance and is itself reduced. A reducing agent as a substance that reduces another substance and is itself oxidised.

Using changes in oxidation numbers, oxidising and reducing agents can be identified.



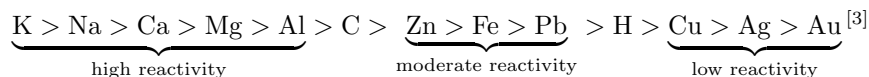
## 5 Metals

### 5.1 Properties of metals

A metal is an element whose ion forms as a result of loss of electrons, i.e., elements whose ions are positively charged are called metals. They exist between groups two to three in the Periodic Table.

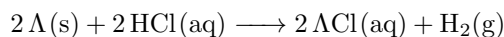
In comparison to non-metals, metals are more thermally and electrically conductive. They are more malleable, which means they can be beaten into many shapes. They are ductile as they can be wrought out into wires. They tend to have high melting and boiling points.

The reactions of metals depend entirely on the metals themselves, or rather, the reactivity of the metals, which follow:



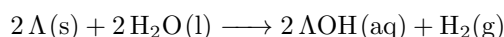
Note the presence of the two non-metals: carbon and hydrogen. Though these are out of place, it is these non-metals whose displacement or otherwise determined the reactivity of the metals. That is, the metals more reactive than carbon are show high reactivity, those less reactive are moderate and low reactive. Those that are more reactive than hydrogen show moderate and high reactivity, while those are less reactive show low reactivity.

Metals with high and moderate reactivity react with dilute hydrochloric acid, displacing the hydrogen in HCl to give the corresponding metal chloride and hydrogen gas. So, lets consider the reaction of a metal  $\Lambda$  which is more reactive than hydrogen, with dilute hydrochloric acid:



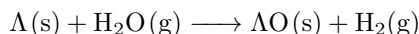
Note that, if  $\Lambda$  is highly reactive, the above reaction is very strong and often infeasible.

Highly reactive metals, react with cold water directly, giving respective hydroxides and hydrogen. An example with  $\Lambda$  as the metal follows:

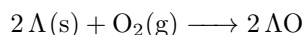


Note that the ratio of OH to  $\Lambda$  will depend on the electronic charges on the metal itself.

Moderately reactive metals react with water in the gaseous form, i.e., steam. The metal itself must also be heated. Such a reaction always gives the metal oxide and hydrogen gas. In this case, magnesium and aluminium act as moderately reactive metals. Once again, with  $\Lambda$  representing such a metal:



Metals with high and moderate reactivity both react with oxygen<sup>[4]</sup>. The product is that metal's oxide:



Note that, iron only reacts in the above reaction as powder or wool. Iron also reacts with air and moisture forming a layer of rust, which is hydrated iron(III) oxide. Low reactive metals, such as copper and lead react very slowly to give an oxide layer, under heat.

### 5.2 Uses of metals

Aluminium is a metal that shows low reactivity as a result of its oxide layer, which forms due to its reaction with the air. It is a metal with very low density and hence is used in manufacturing aircraft. It is also used in overhead electrical cables as it is a good electrical conductor and low density. It is also used in food containers as it does not corrode easily as a result of the oxide layer.

Copper is also a good electrical conductor, it is also ductile and is hence used in smaller scale wiring.

### 5.3 Alloys and their properties

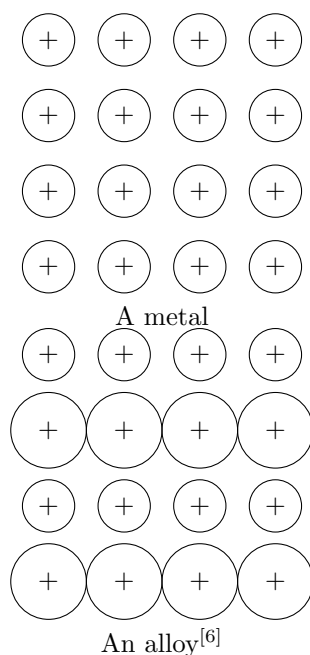
An alloy is a mixture of a metal with other elements. Brass is such a mixture consisting of copper and zinc. Stainless steel is a mixture of iron and other elements such as chromium, nickel and carbon.

The structure of a metal consists of layers of metal cations with a sea of delocalised electrons<sup>[5]</sup>. These layers are arranged such that they can cascade and slide over each other, which is why metals are malleable and ductile. However, in alloys, the cations are differently sized due to the different atoms present. This stops the layers from sliding easily, making alloys stronger and harder than pure metals. A structure of a metal and an alloy follows:

<sup>[3]</sup>where the metal to the left of > is more reactive than the ion to the right.

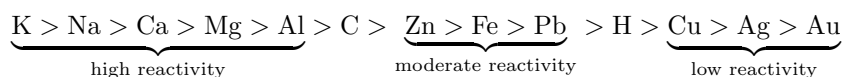
<sup>[4]</sup>Any reaction with oxygen is called burning.

<sup>[5]</sup>A delocalised electron is that which is not bound to any atom and can move freely

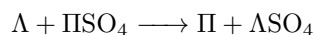


## 5.4 Reactivity series

As shown at the beginning of the chapter, the order of reactivity of the metals is:



The reactivity of a metal is said to be the tendency of the metal to form a positive ion. The relative reactivity of each metal is seen in displacement reactions involving the metals, where the more reactive metal displaces the less reactive metal from its compound. An example follows, where  $\Lambda$  is a more reactive metal than  $\Pi$ .



Note that, the less reactive metal is reduced and the more reactive metal is oxidised. The reactions of metals with water, hydrochloric acid and steam have all been discussed, some further observations are given below.

Potassium, sodium and calcium have very strong reactions with cold water, where the potassium and calcium are seen to skip over the water's surface while melting. Calcium reacts strongly. Magnesium, only reacts with cold water very slowly, but burns with a brilliant flame when steam is passed over heated magnesium.

Metals which are highly reactive have very violent reactions with dilute acid, so it is unsafe to add them directly.

Magnesium reacts strongly, disappearing into the acid while forming bubbles of gas. The result is a colourless solution.

Aluminium is slow to react in cold acid, but bubbles form on heating, causing the aluminium to disappear into the acid resulting in a colourless solution.

Zinc, disappears in cold acid, producing bubbles of gas and a colourless solution forms. Same is the case with iron, only a pale green solution is formed.

Copper silver and gold, i.e., metals with low reactivity have no reaction with acids at all.

Aluminium, when kept exposed to air, reacts with the surrounding oxygen in the air resulting in a layer of aluminium oxide around the sample of aluminium. This layer is unreactive, so aluminium is *apparently* unreactive as the layer is non porous and no oxygen can reach the underlying aluminium.

Given a set of experimental observations, using the strength of reactions described and the knowledge from the reactivity series of metals, we can deduce the reactivity of given metals.

## 5.5 Corrosion of metals

Metals corrode under certain conditions as results of certain reactions.

Iron and steel corrode in presence of oxygen or water producing a layer of iron (III) oxide, called rust. Prevention of this comes down to many strategies, including barrier protection, where the metal is either coated

<sup>[6]</sup>These diagrams are far too bad to be acceptable answers to O-Level questions, please refer to prescribed book.

by another or painted so as to prevent contact of water and oxygen with the metal. Barriers used to protect metals from corrosion can be painting, greasing or coating the metal with plastic.

Sacrificial protection consists of another more reactive metal being placed adjacent the metal to be protected. The moisture and air will then react with the more reactive metal in precedence to the relatively less reactive metal. In other words, the higher reactive metal loses electrons more readily, as a result, that metal is oxidised and the lower reactive metal is not corroded.

The two strategies can be combined in that zinc can be used to electroplate any metal, where since zinc is more reactive, it will be reduced in precedence to the metal being coated. The disadvantage to barrier protection is that when the coating comes off, that area will immediately be corroded. This method prevents that as even if the layer is scratched, the moisture will react with the zinc instead. This method is called galvanising.

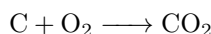
## 5.6 Extraction of metals

Metals can be extracted easily depending on their reactivity.

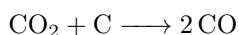
Iron is found naturally in an ore called haematite, which is iron (III) oxide. Iron is extracted from iron (III) oxide in a blast furnace, by means of reduction by carbon monoxide.

Raw materials, iron ore, coke (carbon made from coal) and the mineral<sup>[7]</sup> limestone. The furnace has blasts of hot air sent near the bottom, allowing a series of chemical reactions to occur resulting in the production of pure iron.

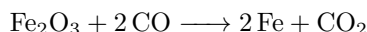
Firstly, the coke burns in the air blast and the furnace gets very hot from this exothermic reaction:



As carbon dioxide rises through the furnace it reacts with more carbon and is reduced to carbon monoxide:

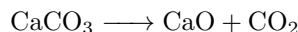


The most important reaction is the reduction of haematite by this carbon monoxide

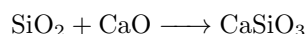


This produced iron flows to the bottom of the furnace where it can be tapped off.

Iron ore contains the impurity sand, which is silicon (IV) oxide, also called silica. Limestone, calcium carbonate is added to the blast furnace to rid the iron of this impurity. Firstly, the calcium carbonate decomposes in the heat of the furnace:



Subsequently, this calcium oxide reacts with silica to form calcium silicate, also called slag.

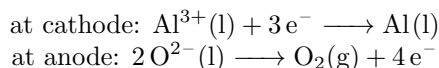


Aluminium is a metal that shows high reactivity, as a result reduction cannot be used for its extraction, rather, electrolysis is employed.

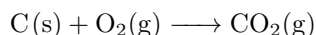
Cryolite is used as a solvent for aluminium as it reduces the melting point of the substance allowing the mixture to be liquid at lower temperatures which is economical. It also conducts electricity and as a result can be used in electrolysis.

The ore for aluminium is bauxite, which is simply aluminium oxide.

Bauxite, dissolved in cryolite is electrolysed using carbon electrodes. The cathode reactions are as follows:



At the very high temperature of the electrolysis conditions, the oxygen being formed at the anode burns the carbon in the anode,



forming carbon dioxide and causing the anodes to corrode and be regularly replaced.

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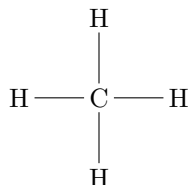
<sup>[7]</sup>A naturally occurring rock containing a particular compound

## 6 Organic chemistry

### 6.1 Formulae, functional groups and terminology

Structural formulae are unambiguous descriptions of the way atoms in a molecule are arranged, showing double bonds between carbon atoms and each molecule bonded to each carbon, eg.,  $\text{CH}_2=\text{CH}_2$ ,  $\text{CH}_3\text{CH}_2\text{OH}$ .

Displayed formulae show all the atoms and bonds between them in a molecule. An example, in the form of methane follows:



The general formulae of homologous series follow:

- Alkanes:  $\text{C}_n\text{H}_{2n+2}$ .
- Alkenes:  $\text{C}_n\text{H}_{2n}$ .
- Alcohols:  $\text{C}_n\text{H}_{2n+1}\text{OH}$ .
- Carboxylic acids:  $\text{C}_n\text{H}_{2n+1}\text{COOH}$ .

Structural isomers are compounds with same molecular formulae but different structural formulae. For example,  $\text{C}_4\text{H}_8$  has two structural isomers:  $\text{CH}_3\text{CH}=\text{CHCH}_3$  and  $\text{CH}_2=\text{CHCH}_2\text{CH}_3$ .

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

All members of a homologous series has the following properties:

- Have the same functional group.
- Have the same general formula.
- Differ by the next member by one  $-\text{CH}_2-$  unit.
- Display a trend in physical properties.
- Sharing similar chemical properties.

A saturated organic compound is that in which all carbons are singly bonded, in other words, all carbon-carbon bonds in a saturated compound are single bonds. An unsaturated organic compound is that which has atleast one carbon-carbon double bond.

### 6.2 Naming organic compounds

The number of carbon atoms in a molecule of an organic compound determines its name, so does the homologous series it is a part of. The pattern is:

number of carbons	name
1	meth-
2	eth-
3	prop-
4	but-

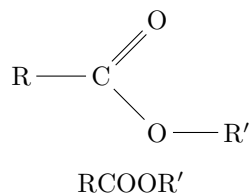
The next in the series follow the geometric pattern of nomenclature (five-penta; six-hexa; ...).

The end of the name of an organic compound follows, which depends on the homologous series it is a part of:

homologous series	name
alkane	-ane
alkene	-ene
alcohol	-ol
carboxylic acid	-oic acid

So a carboxylic acid with four carbons is called *butanoic acid*.

Alcohols and carboxylic acids react to form compounds called esters (see Section 11.7). They are composed of one alkyl part and another acidic part, and their nomenclature comes down to being *alkyl acidoate*. An ester formed from ethanol and butanoic acid will be called ethyl butanoate. The general structural and displayed formula for esters follow:



where R is carbon side of the acid part and R' is the carbon side of the alcohol part.

## 6.3 Fuels

There are three fossil fuels: coal, natural gas and petroleum. Methane is the main constituent of natural gas.

Hydrocarbons are organic compounds containing carbon and hydrogen only. Petroleum is a liquid mixture of such hydrocarbons. All of these hydrocarbons all have different uses and they must be separated. Fractional distillation is used to do so, using a fractionating column that separates the constituent hydrocarbons by rising them up depending on their properties.

So, as we observe the products from the bottom to the top of the fractionating column:

- Compound chain length decreases.
- Compound volatility increases.
- Boiling points lower.
- Viscosity<sup>[8]</sup> lowers.

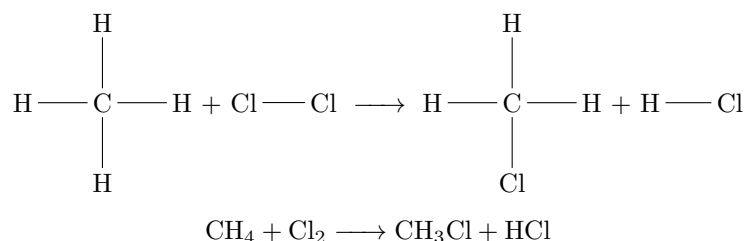
The topmost fraction that is separated from petroleum is refinery gas, which is used for heating and cooking. Petrol, also called gasoline, is the next, used as motor fuel. Subsequently, naphtha is collected which is used as a chemical feedstock. Paraffin is used as a fuel in jet engines and heating oil. Diesel oil is used as fuel in diesel engines which trucks and large vehicles use. Fuel oil is used in ships and home heating. Lubricating oil is used to make waxes and polishes. Lastly, the bottommost fraction is called bitumen, which is used to surface roads.

## 6.4 Alkanes

The bonding between all atoms in alkanes is single covalent, and hence they are saturated compounds <sup>[9]</sup>. Alkanes are generally unreactive compounds, except when combusted and substituted by chlorine.

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

The substitution of alkanes by chlorine is a photochemical reaction, where the activation energy is provided by ultraviolet light. An example of the reaction of chlorine and methane follows:



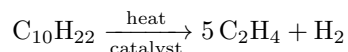
where the products are chloromethane, and hydrogen chloride.

<sup>[8]</sup>Viscosity is the resistance of flow of a fluid.

<sup>[9]</sup>Note that they are saturated due to the carbons being singly bonded, the bonds between other atoms are not of importance in this case.

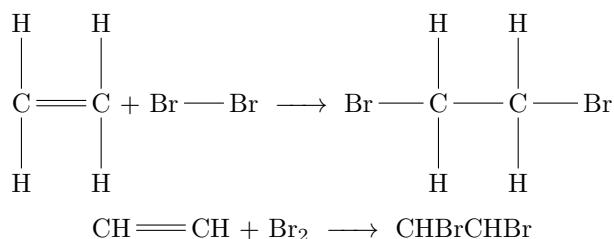
## 6.5 Alkenes

The bonding in alkenes includes a carbon-carbon double bond, and hence they are unsaturated hydrocarbons. Alkenes are produced by the *cracking* of longer chained alkanes, with hydrogen as a byproduct, in presence of a catalyst at high a high temperature of around 500 °C. For example:



The fractional distillation of petroleum produces alkanes, but not enough to meet consumer demands. This is why catalytic cracking is performed on large chain alkanes, to meet the demand of the shorter chained alkenes.

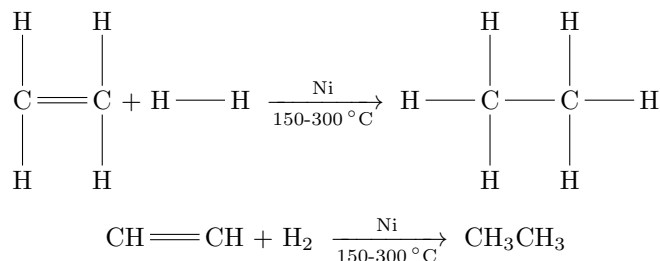
Aqueous bromine reacts with unsaturated compounds, losing its brown colour, this reaction does not happen with saturated compounds, however. The reaction of bromine with an unsaturated compound, ethene, follows:



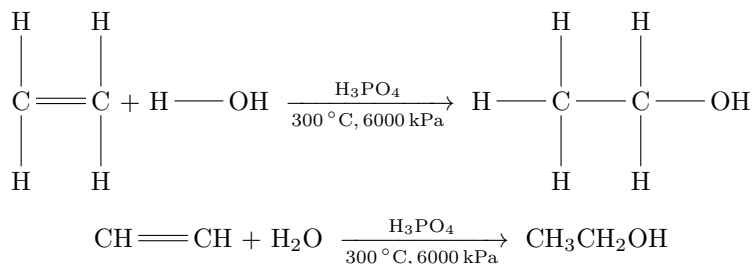
where the product is called dibromoethane.

The above is an example of an addition reaction, that which is where two or more reactants form one product. Alkenes undergo two more addition reactions.

**Hydrogen in presence of nickel.** An alkene will react with hydrogen in heated conditions, (150-300°C), in presence of a nickel catalyst, to form a corresponding alkane.



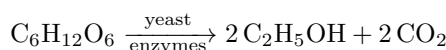
**Steam in presence of phosphoric acid.** An alkene reacts with steam at a temperature of 300 °C, at a pressure of 6000 kPa in presence of phosphoric acid as a catalyst to form the corresponding alcohol.



## 6.6 Alcohols

Ethanol is an alcohol that can be produced by two methods.

**Fermentation of glucose.** Aqueous glucose can be fermented at 25-30 °C in the presence of yeast and in the absence of oxygen. The reaction is:

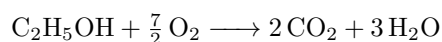


**Catalytic addition of steam.** see the end of Section 11.5.

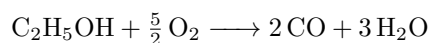
The ethanol produced by fermentation comes from a renewable source, in this way, when the ethanol is used as a fuel it can be said to be “carbon neutral”. It uses relatively simple large apparatus. The ethanol in this method is made by a batch process, where the process needs to be started anew each time. It is a relatively slow process, and the ethanol produced must be subsequently purified through distillation.

The ethanol by cat-hydration comes from a non-renewable source in the form of petroleum. It requires small scale apparatus, which can withstand pressure. The ethanol is produces in a continuous reaction with a high reaction rate, yielding very pure ethanol. However, it is a sophisticated and complex method.

Alcohols completely combust, i.e., react with oxygen<sup>[10]</sup>, to form water and carbon dioxide. When they combust incompletely, they form carbon monoxide and water.



is an example of complete combustion.

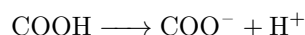


is an example of incomplete combustion.

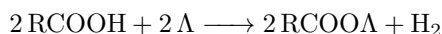
Ethanol is used as a solvent for paints, glues, perfumes, aftershaves and printing inks. The combustion of ethanol produces a lot of heat, which can be used in engines as fuel when ethanol is mixed with petrol.

## 6.7 Carboxylic acids

All carboxylic acids have a  $\text{—COOH}$  group, which is what gives each acid its acidic property as it is this group that disassociates to give the hydrogen ion:

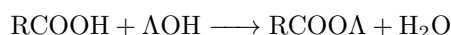


**Reactions with metals.** As with acids, the reaction of carboxylic acids with metals forms a metal salt and hydrogen gas.

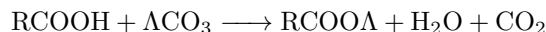


where  $\Lambda$  is a metal, R is a placeholder for the remaining part of the carboxylic acid.

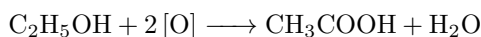
**Reactions with bases.** As with acids, the reactions of carboxylic acids with bases forms a metal salt and water.



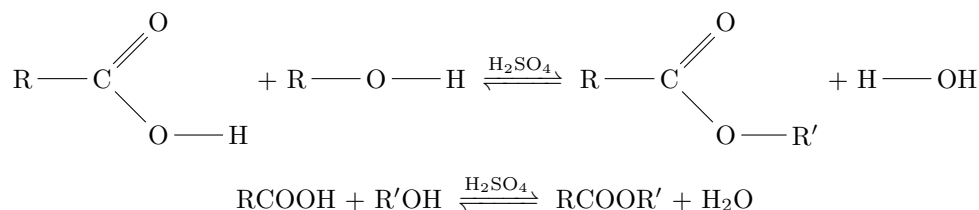
**Reactions with carbonates.** As with acids, the reactions of carboxylic acids with bases forms a metal salt, water and carbon dioxide.



Ethanoic acid can be formed from ethanol by two methods. When ethanol reacts with an oxidising agent, it forms ethanoic acid. Bacteria may also oxidise vinegar, which is ethanol to form ethanoic acid:



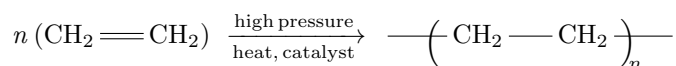
When carboxylic acids and alcohols react in presence of sulfuric acid as a catalyst, an ester and water is formed.



## 6.8 Polymers

A polymer is a large molecule built up from smaller constituent molecules, which are called monomers. The process by which monomers form polymers is called polymerisation, which is of two types.

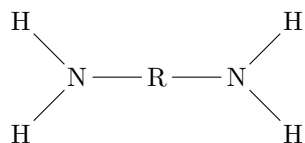
**Addition polymerisation.** In addition polymerisation, the monomers join up to form the polymer and nothing else, which is a form of addition reaction and hence the name. Alkenes, under high pressure, heat and a catalyst will react with each other to form a polymer. An example with ethene is given, where  $n$ , a very large number of ethene molecules react to give an  $n$ -long chain of ethene called poly(ethene).



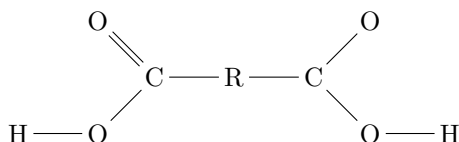
<sup>[10]</sup>Any reaction with oxygen is said called burning in oxygen.

The  $\text{CH}_2=\text{CH}_2$  is called the repeat unit of poly(ethene), because its repetition is what forms poly(ethene). Note the loss of  $\text{C}=\text{C}$  after polymerisation, which applies to any and all alkenes undergoing the process.

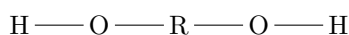
**Condensation polymerisation.** In condensation polymerisation, water is formed usually, as the byproduct. There are two condensation polymers to study: polyamides and polyesters. The amine functional group is  $\text{NH}_2$  —; that for carboxylic acids is —  $\text{COOH}$  and alcohols have —  $\text{OH}$  as their functional group. For polymers formed through condensation, there are three different monomers:



is a diamine (two amine groups).



is a dicarboxylic acid (two carboxyl groups).



is a diol (two alcohol groups).

In the above, the R is a variable carbon containing group. Note that, when reacting, a carboxylic acid group loses —  $\text{OH}$ , and the alcohol and amine groups lose —  $\text{H}$ .

**Polyamides.** The polymerisation of dicarboxylic acids and diamines form polyamides, with amide linkages.



## 7 Experimental techniques and chemical analysis

### 7.1 Experimental design

TODO.

### 7.2 Acid-base titrations

An acid base titration is an experimental method to find the concentration of an acid or alkali. It involves the use of a pH indicator, a burette and a volumetric pipette. The procedure follows:

1. Wash burette with distilled water and the chemical to be used.
2. Wash pipette with distilled water and the chemical to be used.
3. Wash conical flask with distilled water.
4. Fill burette with chemical of known concentration, note reading on burette.
5. Take known volume chemical of known concentration using volumetric pipette into a conical flask. Put two drops of indicator into the conical flask.
6. Titrate (open tap of burette) burette's liquid into conical flask until indicator changes colour.
7. Note the reading on the burette, and subtract initial reading from this. The result is the titre volume.
8. Repeat until concordant values<sup>[11]</sup>
9. Average volumes gotten, which gives titre volume.

The titre volume is that needed to just neutralise the used volume of the substance with unknown concentration. Using molar ratio, we can find the moles needed for one substance to react with the other. Since we know volume of both and concentration of one, we can derive concentration of other by mole calculation.

### 7.3 Chromatography

Paper chromatography is used to separate mixtures of soluble substances using a suitable solvent. These substances must be coloured, and they travel down a filter paper at varying distances depending on how soluble they are in a solvent or how attracted they are to the filter paper. The ratio of the distance travelled on the chromatogram by the solvent to that travelled by a certain soluble substance is called its  $R_f$  value. As such:

$$R_f = \frac{\text{distance moved by the substance}}{\text{distance moved by the solvent front}}$$

For a given solvent at a given temperature, the  $R_f$  value is the same for a particular substance, which can be used to identify the substance.

Colourless substances can be separated by use of *locating agents*. These are substances that react with colourless substances to give a coloured spot or spots that glow under ultraviolet light.

### 7.4 Separation and purification

For a mixture of substances, one of which is soluble in a particular solvent and the other isn't, we can separate the substances by dissolving it into this solvent.

We may then filter out the insoluble substance by making use of filter paper, where the insoluble substance will collect as residue, and the soluble part will pass through to form the filtrate.

For a given salt solution, we can form crystals of the salt by the following procedure:

1. Heat solution in evaporating dish until it becomes saturated, which can be seen using a glass rod dipped into the solution. If crystals form on the rod, the solution is saturated.
2. Cooling this saturated solution will form hydrated crystals of the salt.
3. These crystals can be dried using paper towels or tissues.

---

<sup>[11]</sup> Readings with a difference of  $\pm 0.1 \text{ cm}^3$ .

For a substance with a dissolved solid, the liquid into which the solid was dissolved usually has a much lower boiling point. To separate the two, we take the solution in a round bottom flask, with a thermometer attached at one side, and a condenser at the other. The thermometer measures the temperature of the flask, which, when equal to the boiling point of the substance, will cause it to boil into the condenser. The condenser has cold water running around a tube, which causes the substance to condense and this condensed liquid is collected in a vessel at the end of the condenser.

A solution of different liquids which all have different boiling points can be separated using fractional distillation. The substance is taken in a round bottom flask which heats it. The substance with the lowest boiling point first boils into the fractionating column which is a glass column filled with glass beads which maximise surface area to condense the liquid with the higher boiling point which may have evaporated. The liquid whose boiling point the temperature is at the passes through the condenser and collects in a vessel.

A pure substance has a fixed boiling point. The further the boiling point of a substance from its fixed one, the less pure it is.

## 7.5 Identification of ions and gases

The anion in a substance can be tested with the following procedures:

- Carbonate,  $\text{CO}_3^{2-}$ : Addition of hydrochloric acid effervesces colourless gas, which is carbon dioxide<sup>[12]</sup>.
- Chloride,  $\text{Cl}^-$ ; Bromide  $\text{Br}^-$ ; Iodide  $\text{I}^-$  (in solution): Acidification with dilute nitric acid and subsequent addition of silver nitrate gives white, cream and yellow precipitates respectively.
- Nitrate,  $\text{NO}_3^-$ : Addition of aqueous sodium hydroxide and aluminium foil, with subsequent warming will produce ammonia gas<sup>[13]</sup>.
- Sulfate,  $\text{SO}_4^{2-}$ : Adding barium nitrate to acidified sulfate solution gives white precipitate of barium sulfate.
- Sulfite,  $\text{SO}_3^{2-}$ : Adding aqueous potassium manganate (VII) solution to acidified sulfite solution de-colourises the purple potassium permanganate solution.

Reactions with sodium hydroxide give the following observations on the following cations:

- Aluminium,  $\text{Al}^{3+}$ : Gives white precipitate, which dissolves in excess to give a colourless solution.
- Ammonium,  $\text{NH}_4^+$ : Gives off ammonia gas on warming, produces no precipitate.
- Calcium,  $\text{Ca}^{2+}$ : White precipitate which is insoluble in excess reagent.
- Chromium (III),  $\text{Cr}^{3+}$ : Green precipitate which dissolves in excess sodium hydroxide.
- Copper (II),  $\text{Cu}^{2+}$ : Light blue precipitate forms, insoluble in excess.
- Iron (II),  $\text{Fe}^{2+}$ : Green precipitate, insoluble in excess. Turns brown near the surface (rusts) when left to stand.
- Iron (III),  $\text{Fe}^{3+}$ : Red brown precipitate forms, insoluble in excess.
- Zinc,  $\text{Zn}^{2+}$ : White precipitate that dissolves in excess NaOH to give a colourless solution.

Reactions with aqueous ammonia give the following observations on the following cations:

- Aluminium,  $\text{Al}^{3+}$ : Gives white precipitate, insoluble in excess.
- Ammonium,  $\text{NH}_4^+$ : No reaction.
- Calcium,  $\text{Ca}^{2+}$ : No reaction (very slight white precipitate).
- Chromium (III),  $\text{Cr}^{3+}$ : Green precipitate, insoluble in excess.
- Copper (II),  $\text{Cu}^{2+}$ : Light blue precipitate forms, soluble in excess to give a dark blue solution.
- Iron (II),  $\text{Fe}^{2+}$ : Green precipitate, insoluble in excess. Turns brown near the surface (rusts) when left to stand.

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<sup>[12]</sup>Which can be tested with limewater.

<sup>[13]</sup>To be tested with moist red litmus.

- Iron (III),  $\text{Fe}^{3+}$ : Red brown precipitate forms, insoluble in excess.
- Zinc,  $\text{Zn}^{2+}$ : White precipitate that dissolves in excess  $\text{NH}_3(\text{aq})$  to give a colourless solution.

The following gases can be tested as follows:

- Ammonia,  $\text{NH}_3$ : Turns damp red litmus paper blue.
- Carbon dioxide,  $\text{CO}_2$ : Makes limewater ( $\text{CaCO}_3(\text{aq})$ ) cloudy.
- Chlorine,  $\text{Cl}_2$ : Bleaches damp litmus paper.
- Hydrogen,  $\text{H}_2$ : Makes a pop sound with a lighted splint.
- Oxygen,  $\text{O}_2$ : Relights a glowing splint.
- Sulfur dioxide,  $\text{SO}_2$ : Decolorises acidified aqueous potassium permanganate.

Non transition metals do not form coloured compounds. Their identities can be analysed using a flame test since they all burn to give differently coloured flames.

A nichrome wire is heated up in a Bunsen burner before being dipped into concentrated acid. This wire is then dipped into the salt to be tested and the colour of the flame is observed. Salts with their respective flame colours follow:

Lithium, $\text{Li}^+$	Red
Sodium, $\text{Na}^+$	Yellow
Potassium, $\text{K}^+$	Lilac
Calcium, $\text{Ca}^{2+}$	Orange red
Barium, $\text{Ba}^{2+}$	Light green
Copper (II), $\text{Cu}^{2+}$	Blue green