4th Form Chemistry

2017/18

Copper sulfate crystals

Notes taken in class; typed up by Timothy Langer

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Kinetic Theory

Kinetic theory is the theory that say “everything is made up of particles, and, unless at absolute zero the particles are constantly moving.

Kinetic theory helps us to explain why diluting a coloured solution makes the solution lighter in colour, and also that gas or liquid particles move by diffusion.

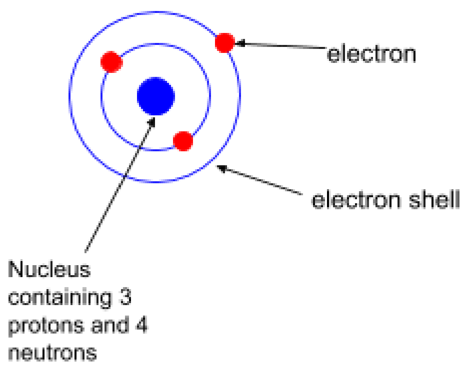
In Chemistry we are particle-arly concerned with two types of particle:

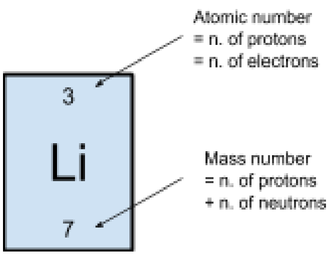
1. Atoms: the smallest particles that can be identified as a particular element.
2. Molecules: are two or more atoms chemically bonded together.

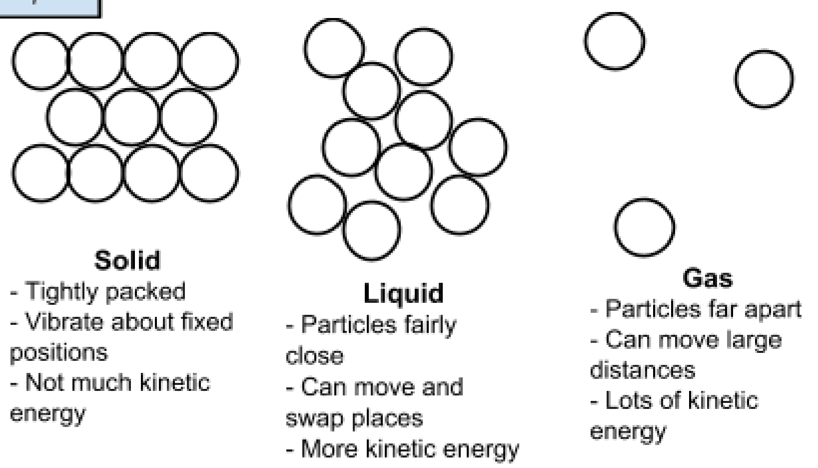
Atoms

Based on a model proposed by Rutherford. He fired positively charged particles at gold foil. Most of the particles passed straight through, showing that most of a gold atom is empty space. A few particles turned around and came back, showing that they had hit something small and positive.

We now have a well established model for atomic structure. This is a drawing of a lithium atom. It has three protons, four neutrons and three electrons. This information can be found on the periodic table.







Protons and neutrons are found inside the nucleus and make up most of the atom's mass. Protons and neutrons have approximately the same mass, while electrons are lightweight. Electrons are arranged in rings; they orbit the nucleus and are negatively charged. Protons have a positive charge. Neutrons, as per the name, are neutral and have no charge.

Molecules

A molecule is two or more non-metal atoms that are chemically bonded together, not necessarily the same elements. For example, liquid water from your tap is made of lots and lots of water molecules. Each water molecule is made of two hydrogen *atoms* and one oxygen *atom.* A compound is a type of molecule, but not all molecules are compounds.

Compounds

Is formed when atoms of different elements react together, forming a molecule. All chemical compounds can be represented by chemical formulae. A chemical formula tells us which atoms are present in a compound. For example:

CO_2 1 carbon & 2 oxygen atoms

H_2O 2 hydrogen & 1 oxygen atom

NaCl 1 atom of sodium & 1 atom of chlorine

How do we know what the formula is? The answer lies in something called valencies. Valencies are on the back of your periodic table. Valencies need to match in a formula.

Mixtures

A mixture is a material made up of two or more different substances which have been mixed and have not reacted. Air is a mixture of gases, as explained on page 14. When we mix two substances, there is no chemical change, so we can (in most cases) obtain the previous materials. We also need to know the following terms:

Solvent the liquid in which a solute is dissolved to form a solution. For example, some tea.
Solute the minor component in a solution, dissolved in the solvent. For example, the sugar which you stir into the tea.
Solution a liquid mixture in which the solute is uniformly distributed within the solvent. For example, the sweet tea mixture.
Saturated solution a saturated solution is a chemical solution containing the maximum concentration of a solute dissolved in the solvent. Additional solute will not dissolve in a saturated solution. For example, when you put a whole 500g pack of sugar in your cup of tea and it doesn't dissolve any more.

Simple distillation

Simple distillation is a method for separating the solvent from a solution. For example, water can be separated from salt solution by simple distillation. This method works because water has a much lower boiling point than salt. When the solution is heated, the water evaporates. It is then cooled and condensed into a separate container. The salt does not evaporate and so it stays behind. In simple distillation we often use a Liebig condenser. Every pure substance has its own particular melting point and boiling point. One way to check the purity of the separated liquid is to measure its boiling point. For example, pure water boils at 100°C. If it contains any dissolved solids, its boiling point will be higher than this, thus we can prove whether water is pure. We rely on a similar idea of dissolving solids to alter the freezing point when we throw salt on ice and watch the ice melt, since salty ice has a lower freezing point.

Fractional distillation

Fractional distillation is a method for separating a liquid from a mixture of two or more liquids by identifying their different boiling points. For example, we can separate a mixture of liquid water and ethanol by fractional distillation. This method works because ethanol has a boiling point of 78.37 °C, while water boils at 100 °C. When we heat the mixture, the ethanol boils and evaporates before the water.

Filtration

If a substance does not dissolve in a solvent, we say that it is insoluble. For example, sand does not dissolve in water – it is insoluble. Filtration is a method for separating an insoluble solid from a liquid. When a mixture of sand and water is filtered:

- the sand remains in the filter paper (it becomes the **residue**)
- the water passes through the filter paper (it becomes the **filtrate**)

This method works because the tiny particles of water can easily pass through the minute gaps in the filter paper, however the much larger particles of sand cannot pass.

Crystallisation through evaporation

Evaporation is used to separate a soluble solid from a liquid. Just like in simple distillation, when the water evaporates, it leaves the crystallised solute behind. If we don't heat the mixture to make it evaporate, but instead leave it to evaporate slowly at room temperature, then larger crystals will form.

For example, copper sulfate is soluble in water and its crystals dissolve in water to make the water have a blue colour. During evaporation, water leaves the mixture and leaves copper sulfate crystals behind. If we leave it to evaporate slowly, large blue crystals of *hydrous* copper sulfate will be formed. On the other hand, if we boil the mixture, small white crystals of *anhydrous* copper sulfate will be formed. We can use white anhydrous copper sulfate crystals as a test for water, since they will turn blue when water is present.

Paper Chromatography

Chromatography is one of the ways of separating **a mixture** of substances. A mixture is two or more substances mixed together but without being chemically bonded.

We separate a mixture by identifying a property of the substances which differs for each individual substance. Indicators differ by solubility level and by colour. In order to identify the specific substances we have separated using chromatography we can either compare them with known substances or we can use a numerical factor called the retention factor or R_f value to compare our substances with data from a database.

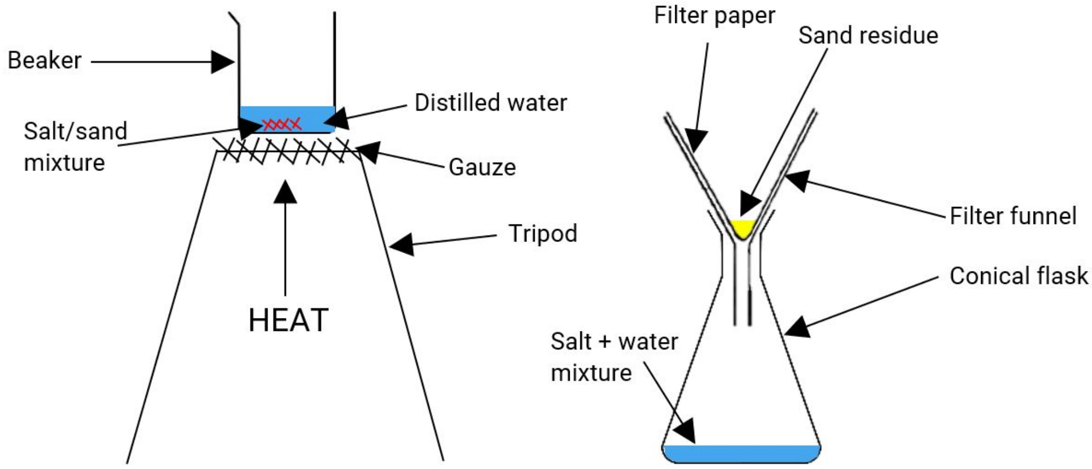
The R_f value is specific for a specific substance in a given solvent.

$$Rf\ value = \frac{Distance\ the\ chemical\ travelled\ up\ the\ paper}{Distance\ the\ solvent\ travelled\ up\ the\ paper}$$

R_f values vary with temperature, because solubility varies with temperature.

Definition	**Word**
The part of the apparatus that the solvent and dyes move up	Stationary phase
A ratio that Compares the distance moved by a dye to the distance Moved by The solvent	R_f value
The solvent that moves up the paper in chromatography	Mobile phase
The pattern of spots produced when an ink separate into it's different dyes	Chromatogram

Separating Salt and Sand



We added the salt/sand mixture to water and stirred until all the salt had dissolved; we heated it to speed up the dissolving. We then poured the water/salt mixture and sand into a filter funnel with filter paper and waited until all the water had drained along with the salt. Only sand remained. Our teacher then tested the remaining sand in silver nitrate solution, if it went cloudy, salt still remained in the sand residue.

Results

When tested by the teacher, our silver nitrate solution appeared partially cloudy.

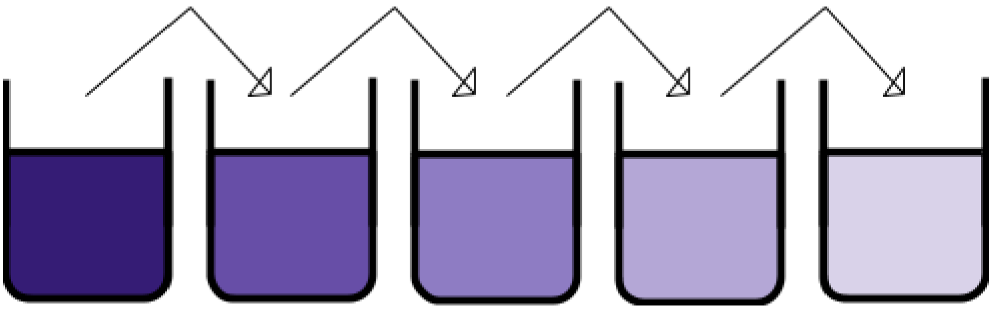
Evaluation

Our experiment was partly successful – the silver nitrate solution did not turn completely cloudy, but only partially in comparison to others'. When stirring to dissolve all salt, we should have stirred more thoroughly in order to completely do so.

Proving existence of atoms

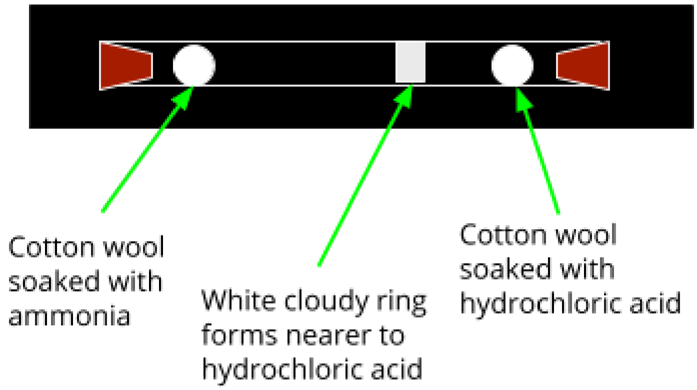
Atoms and molecules are generally too small to be seen. We can show that they are there and exist through a series of experiments.

Dilution of potassium permanganate



The potassium permanganate has invisible purple particles. We cannot see them but we can see their colour. A small amount of potassium permanganate crystals are able to give colour to a large amount of water, because the potassium permanganate breaks down to individual particles when dissolved in water. Since even the last beaker is somewhat colour, we can say that these individual particles must be very small.

Diffusion of gases



The cotton wool soaked in ammonia provides ammonia gas. The cotton wool soaked in hydrochloric acid provides hydrogen chloride gas. When we place the gases at opposite ends of the tube the gases spread out. A white ring, marking where the diffusing gases meet. The white solid (ammonium chloride) forms closer to the hydrochloric acid end, since ammonia gas particles are lighter and travel faster than the hydrogen chloride.

Isotopes

Isotopes are atoms of the same element with different numbers of neutrons. They have the same atomic number, however, they have different mass numbers. Most elements have isotopes, the most famous being chlorine, which has two isotopes.

	$^{35}_{17}Cl$	$^{37}_{17}Cl$
Protons	17	17
Neutrons	18	20
Electrons	17	17

Calculating relative atomic mass

Most periodic tables don't have a mass number; instead, they have the relative atomic mass, which is calculated from the masses of the different isotopes and their **abundances.**

- **75%** of Chlorine is $^{35}\mathrm{Cl}$
- **25%** of Chlorine is $^{37}\mathrm{Cl}$

RAM (Relative Atomic Mass) of Cl

$$= (\tfrac{75}{100} \times 35) + (\tfrac{25}{100} \times 37)$$
$$= 26.25 + 9.25$$
$$= 35.5$$

Another example: **RAM** of Sb

- **57%** of antimony is $^{121}\mathrm{Sb}$
- **43%** of antimony is $^{123}\mathrm{Sb}$

$$= (\tfrac{57}{100} \times 121) + (\tfrac{43}{100} \times 123)$$
$$= 121.86$$

Diatomic elements

The following elements are diatomic: [**I** **H**ave **N**o **Cl**ever **O**r **Br**ight **F**riends]

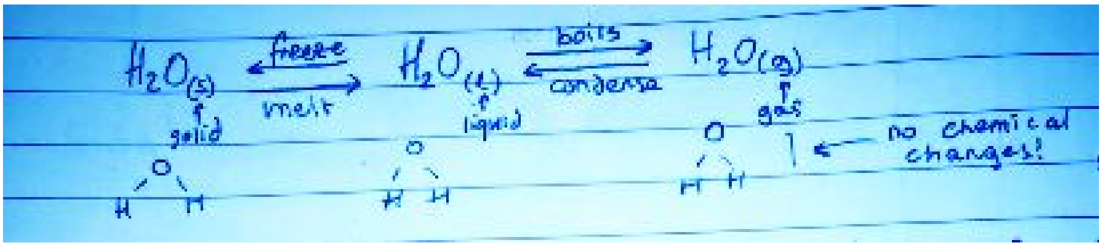
Bromine (Br_2)
Iodine (I_2)
Nitrogen (N_2)
Chlorine (Cl_2)
Hydrogen (H_2)
Oxygen (O_2)
Fluorine (F_2)

These elements being diatomic means that we **never** find just one of these atoms on its own, but instead these atoms join up to form a molecule of two atoms of that element.

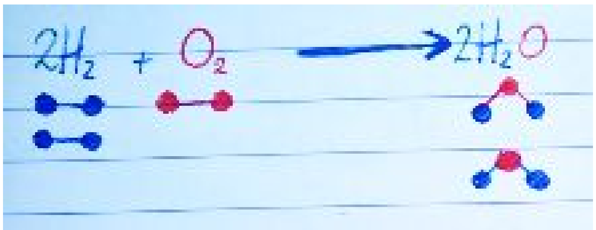
Chemical and Physical Change

Changes of state are **physical** changes. Particles gain or lose energy. They can get closer together or further apart. Bonds are formed or broken between the particles, but no **chemical bonds** are broken or made.

Consider water:



For a chemical change to happen, we need to break and make chemical bonds. Consider making water from hydrogen and oxygen.



When this reaction occurs, we need to break some bonds. Bond-breaking <u>requires</u> energy. It is an **endothermic** process.

The product (water) is formed when new chemical bonds are made. **Bond making** releases energy. It is **exothermic**!

To sum up:

- Breaking bonds *requires* energy
- Making bonds *releases* energy

Laboratory Preparation of Hydrogen Gas

Zinc + Hydrochloric acid → Zinc chloride + Hydrogen gas

$Zn + 2HCl \rightarrow H_2 + ZnCl_2$

Our observations

When performing this displacement reaction, the zinc fizzed and turned black. The bubbles were presumably of hydrogen gas. The solution turned cloudy

Method

Hydrochloric acid is added to granulated zinc. The zinc reacts and turns into zinc chloride, while the escaping hydrogen gas is collected over water into a test tube. The hydrogen gas is tested with a lit splint to produce a squeaky pop.

Electrons and their arrangement

Electrons are found in shells or orbitals which surround the nucleus. As you get further from the nucleus, the shells get bigger. Bigger shells can hold more electrons.

Shell number	**Maximum number of electrons**
1	2
2	8
3	18 (8 at GCSE)
4	32 (8 at GCSE)

Electron shells always fill from the first shell outwards. The arrangement of electrons in an atom has many different names, but they all mean the same:

- Electron arrangement
- Electron configuration
- Electronic arrangement
- Electronic configuration

It looks like this: ^{12}Mg has an electron arrangement of 2, 8, 2.
[12 is the atomic number, meaning it has 12 protons and 12 electrons]

We need to know the electron arrangement for the first 20 elements.

^{1}H	1	^{6}C	2, 4	^{11}Na	2, 8, 1	^{16}S	2, 8, 6
^{2}He	2	^{7}N	2, 5	^{12}Mg	2, 8, 2	^{17}Cl	2, 8, 7
^{3}Li	2, 1	^{8}O	2, 6	^{13}Al	2, 8, 3	^{18}Ar	2, 8, 8
^{4}Be	2, 2	^{9}F	2, 7	^{14}Si	2, 8, 4	^{19}K	2, 8, 8, 1
^{5}B	2, 3	^{10}Ne	2, 8	^{15}P	2, 8, 5	^{20}Ca	2, 8, 8, 2

All chemical reactions occur because electrons move from one atom to another. Sometimes they are lost or gained. Sometimes they are shared. If electrons are lost from an atom, a positive **ion** is formed.

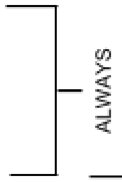
Solubility

Chemical compounds vary in their solubility. Some compounds dissolve in water and can form solutions: these compounds are soluble. Other compounds are insoluble, they do not form solutions and so can never have the (aq) sign.

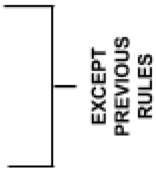
Luck does not work on solubilities! Solubilities need to be learnt.

There are a series of rules:

1. All nitrates (NO_3) are soluble
2. All ammonium (NH_4) compounds are soluble.
3. All group I compounds (Li, Na, K, Rb and Cs) are soluble.



4. Most carbonates (CO_3) are insoluble
5. Most oxides (O) are insoluble
6. Most hydroxides (OH) are insoluble *except Barium Hydroxide (BaOH)*



7. Most sulfates (SO_4) are soluble *except Ag_2SO_4 $BaSO_4$ $PbSO_4$*
8. Most Group VII compounds (F, Cl, Br, I, and At) are soluble *except those containing Ag or Pb*

Solubility Curves

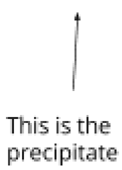
Soluble substances dissolve in water

A saturated solution is one which cannot dissolve any more substance. By looking at the mass of a solid that is dissolved in a saturated solution at different temperatures, we can draw a graph.

Precipitation Reactions

A precipitation reaction is one in which a solid product is formed from two soluble reactants. On example is to produce $PbI_{2\,(s)}$

$$Pb(NO_3)_{2\,(aq)} + 2KI_{(aq)} \rightarrow PbI_{2\,(s)} + 2KNO_{3\,(aq)}$$



Making Salts

A salt is formed when the hydrogen in an acid replaced by a metal or ammonium. Salts can be prepared in multiple different ways. The simplest method is to prepare a soluble salt from an insoluble base or carbonate.

Soluble salt from an insoluble base or metal and acid - $CuSO_4$

1. Warm 50cm^3 of H_2SO_4 - turn the Bunsen burner off as soon as the acid starts to boil
2. Add a spatula of CuO. Watch to see if it dissolves.
3. Continue to add the CuO a spatula at a time until it no longer dissolves.
4. Filter your mixture into an evaporating basin.
5. Heat the mixture until the volume has halved or until it starts to spit.
6. Leave it on the windowsill to crystallise.

Balanced symbol equation:

$H_2SO_4 + CuO \rightarrow CuSO_4 + H_2O$

The acid is heated is to speed up the reaction, and excess copper oxide is added to make sure that the reaction is complete and all of the acid has been reacted. Only half of the water is boiled away because crystal creation requires water. If more is boiled, anhydrous copper sulfate will be produced.

Soluble salt from acid and alkali - NaCl

In this method, we make a <u>soluble</u> salt from two soluble reactants.

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

1. Measure out 25cm^3 of hydrochloric acid into a measuring cylinder.
2. Pour the hydrochloric acid into a conical flask
3. Add 6 drops of phenolphthalein
4. Fill burette with NaOH
5. Record the initial burette reading
6. Add NaOH to the HCl until the indicator goes <u>**pale**</u> pink. (If dark pink, restart)
7. Record the new burette reading
8. Calculate the amount of NaOH used.
9. Repeat without indicator, but with the same amount of NaOH.
10. Pour the neutral solution into an evaporating basin, then heat it until ½ - ⅔ of the solvent has evaporated.
11. Leave on the windowsill to crystallise.

Insoluble salt from two soluble reactants - PbI_2

The balanced symbol equation below is a double displacement reaction, producing the insoluble salt as a precipitate. The equation assumes Lead **(II)** Nitrate and Lead **(II)** Iodide.

$Pb(NO_3)_{2\,(aq)} + 2KI_{(aq)} \rightarrow PbI_{2\,(s)} + 2KNO_{3\,(aq)}$

1. Using a measuring cylinder measure out $25cm^3$ of $Pb(NO_3)_{2(aq)}$
2. Using a <u>different</u> measuring cylinder, measure out $25cm^3$ of $KI_{(aq)}$
3. Pour both solutions into a 100 ml beaker
4. Put the beaker on the tripod and heat it until the mixture just **starts** to boil. Turn the Bunsen burner off.
5. Carefully (using blue paper towel like an oven glove) pour the mixture through some filter paper into a conical flask.
6. Wash the residue with distilled water. Carefully lift out the filter paper with residue and leave it to dry.

Combustion

Combustion means burning or exploding. A substance reacts with oxygen, releasing energy (light+heat). For example, magnesium, which burns very well in oxygen with a bright white flame.

$2Mg + O_2 \rightarrow 2MgO$

Non-metals can also undergo combustion. Sulfur burns well in oxygen with a bright blue flame.

$S + O_2 \rightarrow SO_2$

Melting and decomposition

When heated, thermoplastics and materials with crystalline structures such as metals will melt. On the other hand, some compounds will break down by **thermal decomposition** when heated, forming two or more products from one reactant.

Metal carbonates can often decompose, for example green copper carbonate can thermally decompose to form black copper oxide and carbon dioxide gas.

$CuCO_3 \rightarrow CuO + CO_2$

Thermal decomposition is an endothermic reaction, since breaking bonds requires energy, as explained on page 7.

Testing for gases

Hydrogen

A lit wooden splint will produce a squeaky ‘pop’ in a test tube of hydrogen. The squeaky ‘pop’ is actually a small explosion of hydrogen and oxygen, creating water.

Oxygen

A glowing wooden splint relights in a test tube of oxygen, because the concentration of oxygen is higher in the test tube than in the air.

Carbon dioxide

Bubbling carbon dioxide through limewater (calcium hydroxide / slaked lime) forms calcium carbonate and turns the limewater milky white.

Ammonia

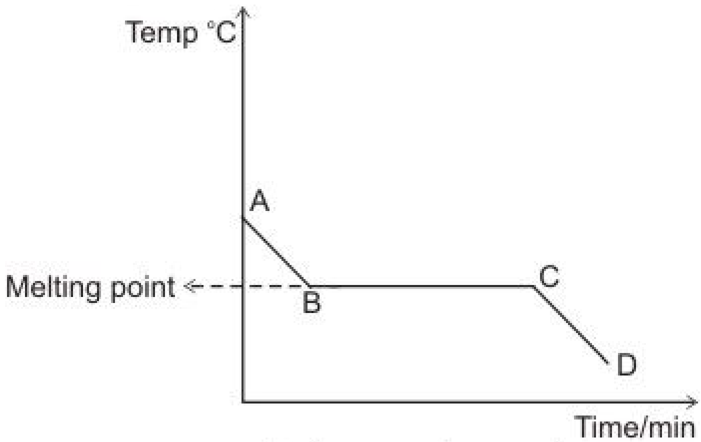
Ammonia extinguishes a lit splint and turns damp red litmus paper blue. A drop of hydrochloric acid into ammonia produces a white smoke, ammonium chloride.

Chlorine

Chlorine has a characteristic sharp, choking smell. It also makes damp litmus paper turn red, eventually bleaching it white. Also, chlorine makes damp starch-iodide paper turn blue-black.

Cooling Curves

A cooling curve shows the temperature of a pure substance as it cools. These curves have a special shape.



Cooling curve of a pure substance

The flat bit happens at the melting or freezing point. As the liquid cools, the liquid particles lose energy more and more slowly. The means the temperature falls.

When the liquid freezes, new bonds are formed between the particles and the particles stick together in a solid.

When these bonds are formed, energy is released. Because this energy is released, it stops the substance cooling.

Once the solid is formed, no more bonds are made and the cooling process continues.

Acids and pH

A substance with a pH lower than 7 can be considered acidic. It will turn universal indicator paper red/orange/yellow and turns litmus paper red.

Standard GCSE definition for an acid: *An acid is defined as a substance that releases H^+ in water.*

Acids are important laboratory chemicals and we need to know five of them.

Hydrochloric acid	HCl
Ethanoic acid	CH_3COOH
Sulfuric acid	H_2SO_4
Nitric acid	HNO_3
Phosphoric acid	H_3PO_4

How do we know if a solution is acidic?

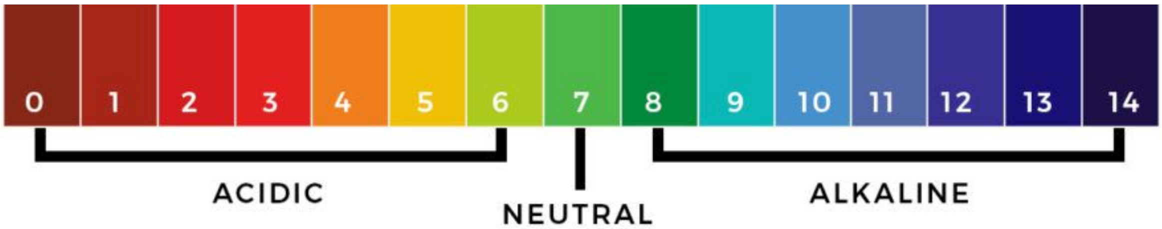
We can use a substance called an indicator, which changes colour in the presence of an acid. There are three common indicators.

Name	Color in Acid	Color when Neutral	Color in Alkali
Litmus	Red	Purple	Blue
Methyl Orange	Pink	Orange	Yellow
Phenolphthalein	Colourless	Pale Pink	Bright Pink

Universal indicator is a mixture of many different indicators. This means that it can have different colours and indicate different strengths of acid and alkali.

The different colours correspond to different pH values and at GCSE, the pH scale goes from 1 through to 14.

pH 1 very strong acid
pH 7 neutral
pH 14 very strong alkali



Whilst universal indicator is useful in telling us whether something is a strong or weak acid or alkali, the actual distinction between different pH values is often difficult to see.

Acids react with bases and alkalis. What is the difference between a base and an alkali?

ALL bases will react with an acid to neutralise it.
ALL metal oxides and metal hydroxides are bases.

Reactions of Acids

When acids react, they form compounds called salts.

Different acids form different types of salt:

hydrochloric acid	chlorides
sulfuric acid	sulfates
nitric acid	nitrates

There are some reactions that all acids have in common.

1. acid + metal → salt + hydrogen gas
2. acid + base → salt + water
3. acid + carbonate → salt + carbon dioxide + water

Composition of the atmosphere

Atmosphere has evolved over millions of years. The first atmosphere contained ammonia carbon dioxide and water vapour. Over millions of years chemical reactions occurs and our current atmosphere came into being:

Nitrogen	78%
Oxygen	21%
Argon	0.9%
Carbon dioxide	0.04%

The other gases make up only a small percentage and vary by location.

An experiment to show the percentage of oxygen in the air

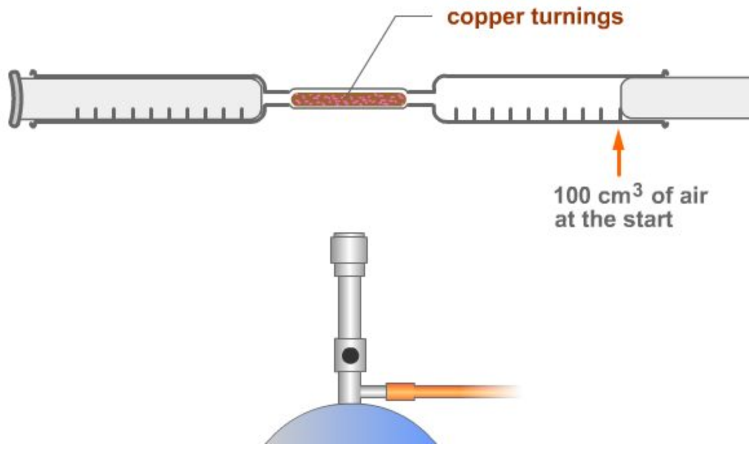
The percentage of oxygen in the air can be measured by passing a known volume of air over hot copper, and measuring the decrease in volume as the oxygen reacts with it. There is an excess of copper turnings so that all the possible oxygen can react. The volume of air will stop decreasing once all the oxygen has reacted.

Note that there is some air in the tube with the copper turnings. The oxygen in this air will also react with the hot copper, causing a small error in the final volume recorded. It is also important to let the apparatus cool down at the end of the experiment, otherwise the final reading will be too high. There could also be a leakage and some air has escaped from the apparatus. Here are the word and symbol equations for this experiment:

copper + oxygen → copper oxide

$2Cu + O_2 \rightarrow 2CuO$

Gas syringes are used to measure the volume of gas in the experiment.



Results

Volume of air at start 100cm^3

Volume of air at end 75cm^3

$\frac{100-75}{100} \times 100 = 25\%\ oxygen$

Some air escaped from the apparatus in our experiment. This means the final volume ended up being *less* than it should be.

Writing Chemical Equations

All chemical reactions can be represented by chemical equations. Word equations are common in prep school. For example,

Aluminium + oxygen → aluminium oxide

At GCSE we need to write this as a balanced symbol equation. This happens in two steps, as follows.

1. Replace words with symbols using valencies on the back of your periodic table
   a. $Al + O_2 \rightarrow Al_2O_3$
2. Put BIG numbers to balance all the atoms
   a. $4Al + 3O_2 \rightarrow 2Al_2O_3$

Any chemical that ends in -ide USUALLY has only two elements.

e.g. sodium iodide
NaI
e.g. magnesium oxide
MgO

e.g. aluminium sulfide
Al_2S_3

The only exception to this rule is hydroxide.

Formula	OH
Valency	1

e.g. sodium hydroxide
$Na^{+1}OH^{-1}$ = NaOH

BUT magnesium hydroxide

$MgOH_2$ is **incorrect**

We need brackets, like so

$Mg(OH)_2$

Many other compound names end in "-ate". This ending means oxygen is present as well as another element. For example

Sodium sulfate = Na_2SO_4
Potassium nitrate = KNO_3
Magnesium carbonate = $MgCO_3$
Aluminium phosphate = $AlPO_4$

Highlighted in green are the "-ate"s. We need to learn the following ones:

NO_3	valency 1
CO_3	valency 2
SO_4	valency 2
PO_4	valency 3