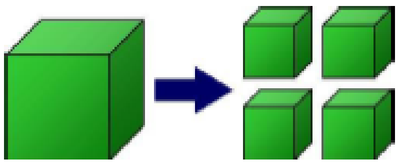
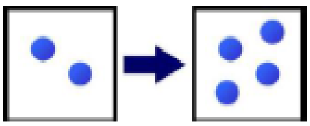
Learning Statement	Red	Amber	Green
A chemical reaction can be recognised by one of the following: ▪ A **colour change** *e.g. blue → red (always give start and end colour)* ▪ An **energy change** *e.g. a rise or fall in temperature* ▪ A **gas being given off** ▪ A **solid being formed**.			
In all chemical reactions a new substance is formed. This is called the **product**. The substances you started with are called the **reactants**. **Reactants → Products** **e.g. magnesium + oxygen → magnesium oxide**			
An EXOTHERMIC reaction is one in which energy EXits, which means the temperature of the surroundings increases.			
An ENDOTHERMIC reaction is one in which energy ENters, which means the temperature of the surroundings decreases.			
The rate of a reaction is a measure of the speed of the reaction.			
There are 4 ways we can change the rate of a chemical reaction: ▪ Changing Particle Size ▪ Changing Concentration ▪ Changing the Temperature ▪ Using a Catalyst			
Changing Particle Size If you decrease particle size, the rate of reaction will increase. This is because more surfaces are available for reactions to take place on.  If you increase particle size, the rate of reaction will decrease.			
Changing Concentration If you increase concentration, the rate of reaction will increase. This is because there are more reactant particles present to react.  If you decrease concentration, the rate of reaction will decrease. This is because there are fewer reactant particles present to react.			

Changing Temperature

If you increase temperature, the rate of reaction will increase. This is because particles have gained more energy and are more likely to collide and react.

If you decrease temperature, the rate of reaction will decrease. This is because particles have less energy and as a result are less likely to collide and react.

Using a Catalyst 

In some reactions another chemical can be added to the reaction which will speed up the reaction. This chemical is called a **catalyst**.

Catalysts do not get used up in the reaction so can be used again. For example if you put 2.73 g of catalyst into the reaction, you would be able to get 2.73 g of the catalyst back at the end of the reaction.

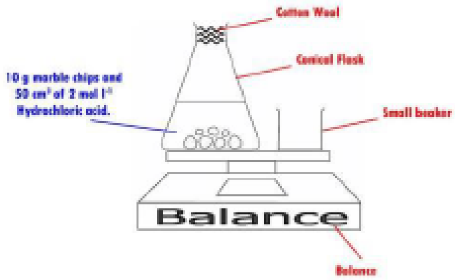
e.g. Platinum is used as a catalyst in catalytic converters in car exhausts.

Enzymes are biological catalysts. They catalyse reactions in living things.

The rate of a reaction can be measured by monitoring the mass loss of reactants or the volume of gas produced in a reaction.

Measuring Mass Loss of Reactants

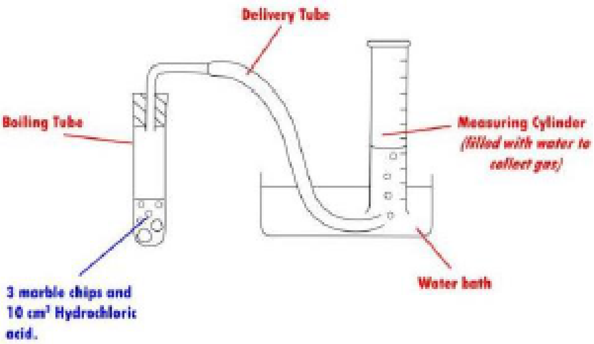
You could use the following apparatus to carry out this experiment.

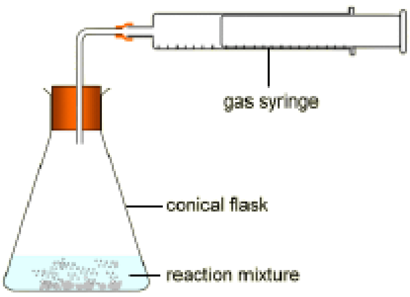


The reading on the balance decreases with time as a gas is being produced.

Measuring Volume of Gas Produced

You could use the following apparatus to carry out this experiment.

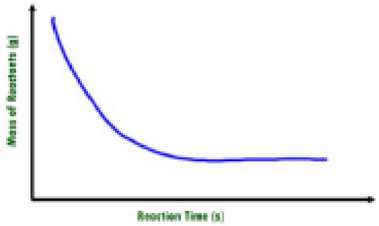




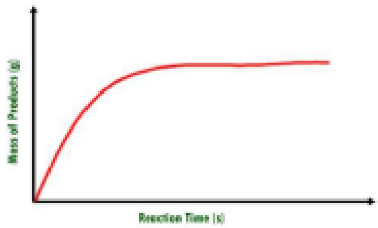
Graphs can be plotted showing the change in mass or volume against time. This gives you curves with the following shapes. The steeper the line is, the faster the reaction is. The graphs level off eventually as the reactants are used up.

1. **Mass of reactants or products against time**

Graph 1: Mass of Reactants with Time

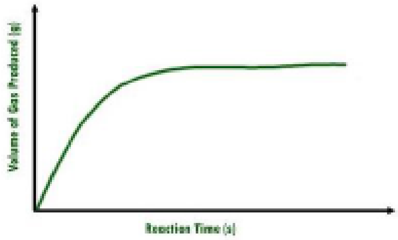


Graph 2: Mass of Products with Time



2. **Volume of gas produced against time**

Volume of Gas Produced with Time



The results from a *mass loss* or *volume of gas produced* experiment can be used to calculate the **average rate** of reaction.

Calculating Average Rate

$$Average\ Rate = \frac{Change\ in\ Mass\ or\ Volume}{Change\ in\ Time}$$

Units of Average Rate

1. Mass and Time

In this case rate is a measure of the mass loss over time therefore the unit of rate is **$g\ s^{-1}$** (grams per second).

$$\frac{g}{s} \rightarrow g\ s^{-1}$$

2. Volume and Time

In this case rate is a measure of the change in volume of gas over time therefore the unit of rate is **$cm^3\ s^{-1}$** (cubic centimetres per second).

$$\frac{cm^3}{s} \rightarrow cm^3\ s^{-1}$$

<table>
<tr><th>Learning Statement</th><th>Red</th><th>Amber</th><th>Green</th></tr>
<tr><td>The Periodic Table was devised by the Russian chemist Dmitri Mendeleev in 1869. It lists all the known elements. As of 2013, there are 118 known elements.</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>The Periodic Table lists these elements in order of increasing atomic number.</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>The Periodic Table is arranged into vertical columns called groups and horizontal rows called periods.</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Elements can be classified in different ways:
<table>
<tr><th>Classification Type</th><th>Examples</th></tr>
<tr><td>naturally occurring/made by scientists</td><td>All elements with atomic number above 92 are made by scientists and are not found naturally on Earth.</td></tr>
<tr><td>solid/liquid/gas</td><td>The majority of elements are solid. There are 11 gases (mainly found on the right of the Periodic Table). Two elements exist as liquids, they are bromine and mercury.</td></tr>
<tr><td>metal/non-metal</td><td>Metals are found on the left of the zig-zig line.
Non-metals are found on the right of the zig-zag line.</td></tr>
</table></td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Some groups have specific names:
<table>
<tr><th>Group</th><th>Name</th><th colspan="3">Examples</th></tr>
<tr><td>1</td><td>Alkali Metals</td><td>lithium</td><td>sodium</td><td>potassium</td></tr>
<tr><td>2</td><td>Alkaline Earth Metals</td><td>calcium</td><td>magnesium</td><td>strontium</td></tr>
<tr><td>Between 2 and 3</td><td>Transition Metals</td><td>gold</td><td>iron</td><td>copper</td></tr>
<tr><td>7</td><td>Halogens</td><td>fluorine</td><td>chlorine</td><td>bromine</td></tr>
<tr><td>0 or 8</td><td>Noble or Monatomic Gases</td><td>helium</td><td>neon</td><td>argon</td></tr>
</table></td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Elements in the same group in the Periodic Table have similar chemical properties.</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Group 1 - The Alkali Metals

The alkali metals are soft metals which are shiny when freshly cut but lose their shininess when exposed to air as a layer of metal oxide forms.
metal + oxygen → metal oxide</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>For this reason alkali metals are stored under oil to prevent contact with the air or water.</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Alkali metals react violently with water. When they react with water they form alkaline solutions, this is why they are called the Alkali Metals.
e.g. sodium + water → sodium hydroxide + hydrogen
(an alkali)</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Group 7 - The Halogens

The halogens and their compounds have many uses. Fluorine compounds are used in toothpaste to help avoid tooth decay. Chlorine is used in drinking water and swimming pools as it can kill harmful bacteria. Bromine is used in dyes and medicines. Iodine can be used as an antiseptic.</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Astatine is a radioactive element that does not occur naturally on Earth.</td><td>○</td><td>○</td><td>○</td></tr>
<tr><td>Like hydrogen, nitrogen and oxygen, the halogens are diatomic molecules. This means they exist in a molecule of two atoms. This means that we can write a chemical formula for them: F_2, Cl_2, Br_2, I_2, At_2.</td><td>○</td><td>○</td><td>○</td></tr>
</table>

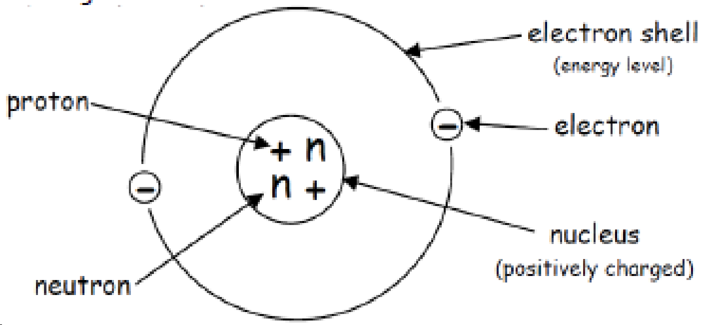
Group 0 (or 8) - The Noble (or Monatomic) Gases

The gases in this group are all colourless. They are unreactive and form almost no known chemical compounds. These gases can be used in lasers and lighting.

Every element is made up of small particles called **atoms**.

Each element contains one type of atom only, e.g. sulfur contains only sulfur atoms and copper contains only copper atoms.

Atoms have the following structure:



Atoms are made up of 3 different particles.

Particle	Location	Charge	Mass
proton	nucleus	+1	1 amu
neutron	nucleus	0	1 amu
electron	outside nucleus	-1	approximately zero

Atoms are neutral because they have the same number of protons and electrons. The positive charge of the nucleus is equal to the sum of the negative charges of the electrons.

Electrons are held in electron shells which are sometimes called energy levels.

1st energy level holds	2nd energy level holds	3rd energy level holds	4th energy level holds
2 electrons	8 electrons	8 electrons	8 electrons

** Electron arrangements for the first 20 elements can be found on page 6 of the data booklet**

Elements in the same group have the same number of outer electrons e.g. Group 1 – The Alkali Metals.

Lithium	Sodium	Potassium
2,1	2,8,1	2,8,8,1

Elements with the same number of electrons in their outer shell have similar chemical properties.

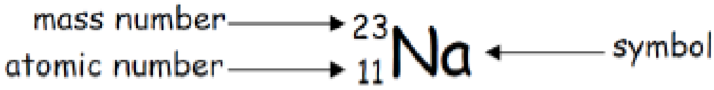
Each element in the Periodic Table has its own atomic number.

atomic number = number of protons

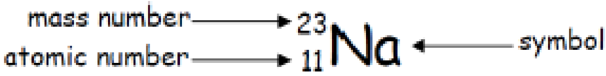
The mass of atom is due to the particles found in the nucleus, the protons and the neutrons.

mass number = no. of protons + no. of neutrons

The symbol of an element can be written showing the mass number and atomic number. This is called nuclide notation.



The number of protons, neutrons and electrons in **neutral atoms** can be calculated by doing the following:



Number of protons = atomic number = 11

Number of neutrons = mass number - atomic number = 23 - 11 = 12

Number of electrons = number of protons (for neutral atoms only) = 11

An **ion** is a charged particle. **For example, Ca^{2+}, Cl^{-}, Na^{+} or O^{2-}.**			
An **ion** is formed when a neutral atom loses or gains electrons. ○ Metals always lose electrons to form positive ions *e.g. Na^{+}, Mg^{2+}* ○ Non-metals gain electrons to form negative ions *e.g. F^{-}, O^{2-}*			
Nuclide notation can also be drawn for ions. The charge is written at the top right of the symbol. mass number → 23, atomic number → 11, $^{23}_{11}Na^{+}$, charge, symbol mass number → 18, atomic number → 8, $^{18}_{8}O^{2-}$, charge, symbol			
The number of protons, neutrons and electrons in **charged ions** can be calculated by doing the following: mass number → 23, atomic number → 11, $^{23}_{11}Na^{+}$, charge, symbol Number of protons = atomic number = 11 Number of neutrons = mass number - atomic number = 23 - 11 = 12 Number of electrons = number of protons - charge = 11 - (+1) = 10			
mass number → 18, atomic number → 8, $^{18}_{8}O^{2-}$, charge, symbol Number of protons = atomic number = 8 Number of neutrons = mass number - atomic number = 18 - 8 = 10 Number of electrons = number of protons - charge = 8 - (-2) = 10			
Isotopes are atoms of the same element which have: ○ the *same* atomic number but a *different* mass number ○ the *same* number of protons but a *different* number of neutrons. e.g. $^{35}_{17}Cl$ $^{37}_{17}Cl$			
The relative atomic mass (RAM) is the average mass of all the isotopes of an elements ○ RAM is rarely a whole number because it is an average.			
e.g. The RAM of chlorine is 35·5 ○ The two isotopes of chlorine are ^{35}Cl and ^{37}Cl ○ As the RAM is closer to 35 than 37, there must be more ^{35}Cl atoms in the sample than ^{37}Cl.			

Learning Statement

The chemical formula of a substance tells us which elements are present and how many of each element we have, e.g. CH_4, HBr.

The **valency** method can be used to work out a chemical formula. The valency of an element is how many bonds it can form. Valency is the number of unpaired electrons in the outermost shell.

Group	1	2	3	4	5	6	7	8 (or 0)
Valency	1	2	3	4	3	2	1	0

The **valency method** involves doing the following:

Write down element symbols	Write down Valency below each element's symbol	Put in *Cross-over Arrows*	Follow arrows and cancel down if necessary to get formula
Si O	Si O 4 2	Si O 4 2	Si_2O_4 ↓ SiO_2

Some chemical names contain a prefix in them, e.g. mono, di, tri, tetra, which tells us how many of the each element we have. This means we can write the formula for these without having to use the valency method.

Compound	carbon **mono**xide	carbon **di**oxide	sulfur **tri**oxide	carbon **tetra**chloride
Formula	CO	CO_2	SO_3	CCl_4
Meaning	mono = 1	di = 2	tri = 3	tetra = 4

Some formulae involve **group ions**. Group ions are ions that contain more than one element. A list of group ions can be found on page 8 of the data booklet.

Ion	Formula	Ion	Formula	Ion	Formula	Ion	Formula
ammonium	NH_4^+	ethanoate	CH_3COO^-	carbonate	CO_3^{2-}	phosphate	PO_4^{3-}
		hydrogencarbonate	HCO_3^-	chromate	CrO_4^{2-}		
		hydrogensulfate	HSO_4^-	dichromate	$Cr_2O_7^{2-}$		
		hydrogensulfite	HSO_3^-	sulfate	SO_4^{2-}		
		hydroxide	OH^-	sulfite	SO_3^{2-}		
		nitrate	NO_3^-	thiosulfate	$S_2O_3^{2-}$		
		permanganate	MnO_4^-				

The valency of a group ion is the number value of its charge, e.g. sulfate SO_4^{2-} has a valency of 2 as the charge is 2-.

The **ionic formula** of an ionic compound is a formula that contains ions, therefore has charges in it, e.g. Na^+Cl^- is the ionic formula for sodium chloride.

Reactions can be described using **word** and **formula** (or **chemical**) equations.

- **Word Equations**
  Describe chemical reactions using words. For example:

  calcium + nitric acid → calcium nitrate + hydrogen

- **Formula (or chemical) Equations**
  Describe chemical reactions using the chemical formulae for the substances involved. For example:

  $$Ca + HNO_3 \rightarrow Ca(NO_3)_2 + H_2$$

Formula (or chemical) equations can be balanced using the following method.

Write down correct chemical formula of all reactants before the arrow and all products after the arrow.

$$Na \quad + \quad O_2 \longrightarrow Na_2O$$

There are 2 oxygen atoms on left hand side but only 1 oxygen atom on right hand side. As the formula of Na_2O cannot be changed, double the number of Na_2O molecules by adding the number 2 *in front* of the formula

$$Na \quad + \quad O_2 \longrightarrow 2Na_2O$$

There is 1 sodium atom on the LHS but 4 sodium atoms on the RHS. As the formulae of Na and Na_2O are set and cannot be changed, we must add the number 4 in front of the Na on the LHS to balance the number of Na atoms

$$4Na \quad + \quad O_2 \longrightarrow 2Na_2O$$

Formula Mass

The formula mass of a substance is the relative atomic masses of all the elements present added together. A list of relative atomic masses can be found on page 7 of the data booklet.

C_3H_8

3 x C = 3 x 12 = 36
8 x H = 8 x 1 = 8
44

Formula mass has no units.

Gram Formula Mass (GFM)

The gram formula mass of a substance is the relative atomic masses of all the elements present added together. A list of relative atomic masses can be found on page 7 of the data booklet.

$Ca(OH)_2$

1 x Ca = 1 x 40 = 40
2 x O = 2 x 16 = 32
2 x H = 2 x 1 = 2
74 g

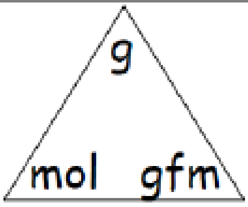
The unit of gram formula mass is **grams, g**.

The gram formula mass (GFM) of a substance is also known as 1 mole of a substance.

1 GFM = 1 mole

Calculations Involving No. of Moles, Mass and GFM

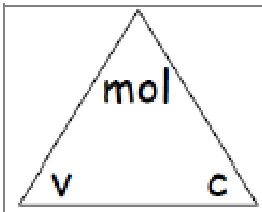
The number of moles, mass and GFM have the following relationship.



g = no. of grams	mol = no. of moles	GFM = gram formula mass
$g = mol \times gfm$	$mol = \frac{g}{gfm}$	$gfm = \frac{g}{mol}$

Calculations Involving No. of Moles, Volume and Concentration

The number of moles, volume and concentration have the following relationship.



mol = no. of moles	c = concentration (mol/l)	v = volume (litres)
$mol = v \times c$	$c = \frac{mol}{v}$	$v = \frac{mol}{c}$

N.B. Concentration has the unit mol l^{-1} (moles per litre) this means the volume in this equation must be in litres as well.

To convert from cm^3 to litres, divide by 1000.

e.g. 45 cm^3 = 45/1000 = 0.045 litres

Worked Example

Calculations involving concentration and number of grams of solid:

e.g. Calculate the concentration of a solution when 5.85g of NaCl is dissolved in 50cm^3 water.

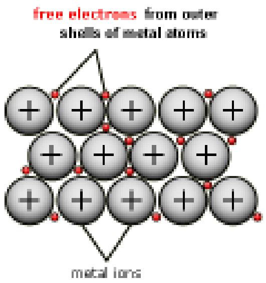
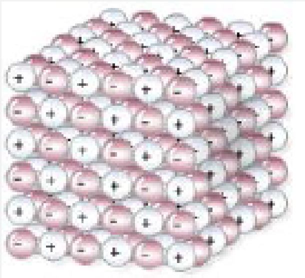
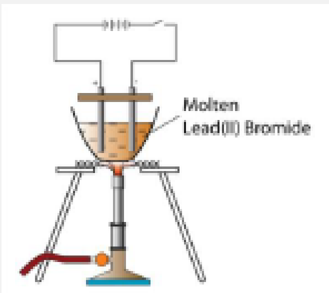
Calculate the gfm of NaCl

Na	1	×	23	=	23
Cl	1	×	35.5	=	35.5
			gfm	=	**58.5g**

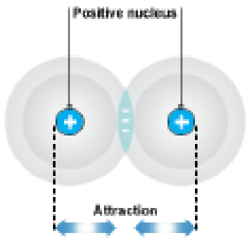
$$\text{no. of mol} = \frac{\text{no. of grams}}{\text{gfm}} = \frac{5.85}{58.5} = 0.1\text{mol}$$

$$\text{concentration} = \frac{\text{no. of moles}}{\text{volume}} = \frac{0.1\text{mol}}{0.05 \text{ litres}} = 2 \text{ mol/l}$$

NB Volume must be in litres!

Learning Statement	Red	Amber	Green
There are 3 types of bonding: o Metallic o Ionic o Covalent.	◯	◯	◯
Metallic Bonding o Occurs in metals. o Results from an electrostatic attraction between positively charged metal ions and a sea of delocalised (free) electrons. o Metallic bonds are strong. o Most metals are solids. Mercury is the only liquid metal. o As electrons can move from metal ion to metal ion, metals conduct electricity. 	◯	◯	◯
Ionic Bonding o Ionic bonds are the electrostatic forces of attraction between positive ions and negative ions. o Ionic bonds are strong. o Ionic compounds have a **lattice structure**. o Ionic compounds dissolve in water. When they dissolve in water the lattice breaks up. o Ionic compounds conduct electricity as a melt or a solution as the ions are free to move. As solids they do not conduct as the ions are not free to move. o Ionic compounds have high melting and boiling points. This means they are solids are room temperature. o The colour of an ionic compound comes from the ions present. *A list of ions colours can be found on page 6 of the data booklet.* 	◯	◯	◯
When an ionic compound is dissolved in water a solution called an **electrolyte** is formed. Electrolytes conduct electricity as the ions are free to move.	◯	◯	◯
Solutions of ionic compounds can be broken down using a process called **electrolysis**. **Electrolysis** is the breaking down of compound using electricity. 	◯	◯	◯

Covalent Bonding



- A covalent bond is a shared pair of electrons between atoms.
- Atoms share electrons to gain a full, stable outer shell of electrons.
- The atoms are held together in a covalent bond by the electrostatic attraction between the positively charged nuclei of each atom and the negatively charged electrons.
- Covalent substances do not conduct electricity. *The exception to this rule is carbon in form of graphite.*
- Most covalent substances do not dissolve in water. However, there are substances they do dissolve in e.g. acetone (nail varnish remover).
- Covalent molecules tend to be liquids or gases at room temperature as they have low melting and boiling points.

A **molecule** is a group of atoms held together by covalent bonds.

A **diatomic molecule** is one which is made up of two atoms.

di atomic molecule

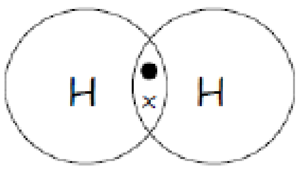
2 atoms atoms joined by covalent bond

Several elements exist as diatomic molecules.

Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine	Astatine
H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2	At_2

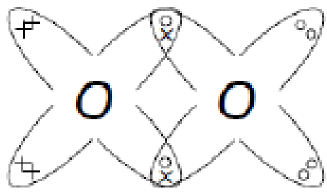
Diagrams can be drawn to show how the outer electrons in atoms are shared to form a covalent bond.

Hydrogen H_2 molecule



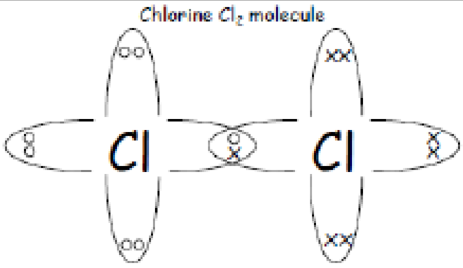
H—H

Oxygen O_2 molecule



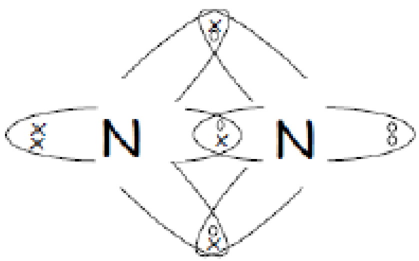
O=O

Chlorine Cl_2 molecule



Cl—Cl

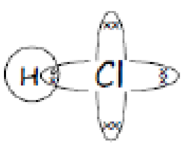
Nitrogen N_2 molecule



N≡N

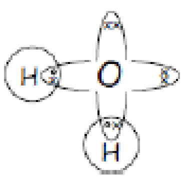
Hydrogen chloride

HCl



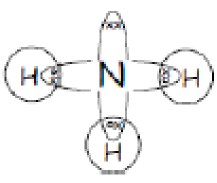
Water

H_2O



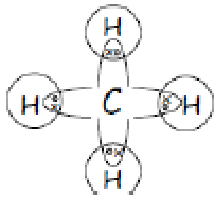
Ammonia

NH_3

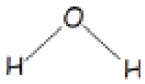
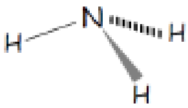
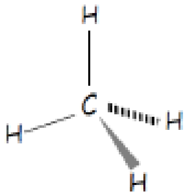


Methane

CH_4



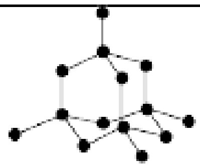
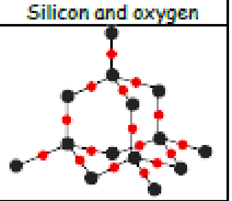
Shapes of Molecules

Linear	Bent	Pyramidal	Tetrahedral
			
HCl	H_2O	NH_3	CH_4
Also with same shape: HF, HBr, HI	Also with same shape: H_2S, H_2Se	Also with same shape: PH_3, NCl_3, PCl_3	Also with same shape: CCl_4, CF_4, SiH_4

Covalent substances can also exist as giant networks. We call these **covalent networks**.

Examples of covalent networks are: **diamond (carbon)** and **sand (silicon dioxide)**.

They only contain strong covalent bonds and therefore are solids and have extremely high melting and boiling points.

Name	Diamond	Sand
Elements present	Carbon	Silicon and oxygen
Structure		
Melting point (°C)	3550	1610
Boiling point (°C)	4827	2230

Bonding Summary

State at Room Temp	Solid	Liquid	Gas
Type of Bonding	Ionic or Covalent	Covalent	Covalent

Type of Bonding	Conduction as a Solid	Conduction as a Liquid	Conduction as a Solution
Metallic (Metals only)	✓	✓	metals do <u>not</u> ______ dissolve in water
Covalent (Non-metals only)	✗	✗	✗
Ionic (Metals + Non-metals)	✗	✓	✓

Learning Statement	Red	Amber	Green
The pH scale is a continuous range of numbers. 0 1 2 3 4 5 6 (Acids) – 7 (Neutral) – 8 9 10 11 12 13 14 (Bases) ○ it is possible to get values below 0 and above 14.	○	○	○
pH **below** 7: Acid; pH **equal to** 7: Neutral (including pure water); pH **above** 7: Base **An alkali is a soluble base.**	○	○	○
Examples of common household and laboratory acids and alkalis. Acids – Household: Vinegar, Lemon juice, Fizzy drinks; Laboratory: Sulfuric acid, Hydrochloric acid, Nitric acid Alkalis – Household: Baking soda, Caustic soda, Oven cleaner; Laboratory: Sodium hydroxide, Ammonia solution, Potassium hydroxide Examples of bases include metal oxides, metal carbonates or metal hydroxides.	○	○	○
Making Acids ○ Non-metal oxides dissolve in water to produce acidic solutions. ○ Carbon dioxide + water → carbonic acid ○ Sulfur dioxide + water → sulfurous acid ○ Nitrogen dioxide + water → nitrous acid	○	○	○
Problems with Acids Sulfur dioxide reacts with water in the atmosphere to produce **acid rain**.	○	○	○
The damaging effects of acid rain include: ○ damage to building rocks ○ damage to structures like metal bridges ○ acidifying soil which reduces crop growth ○ damage to the habitat of plant and animal life.	○	○	○
Making Alkalis Alkalis are soluble bases that are made by dissolving metal oxides or metal hydroxides in water. e.g. lithium oxide, sodium oxide, potassium oxide or magnesium oxide.	○	○	○
e.g. potassium hydroxide, sodium hydroxide or calcium hydroxide.	○	○	○
Acids and alkalis contain ions. ○ Acids contain the hydrogen ion, $\mathbf{H^+(aq)}$ ○ Alkalis contain the hydroxide ion, $\mathbf{OH^-(aq)}$ This means that solutions of acids and alkalis can conduct electricity.	○	○	○

The pH scale:

0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
← Acids →							Neutral	← Bases →						

pH **below** 7	pH **equal to** 7	pH **above** 7
Acid	Neutral (including pure water)	Base

Acids		Alkalis	
Household	**Laboratory**	**Household**	**Laboratory**
Vinegar	Sulfuric acid	Baking soda	Sodium hydroxide
Lemon juice	Hydrochloric acid	Caustic soda	Ammonia solution
Fizzy drinks	Nitric acid	Oven cleaner	Potassium hydroxide

Acidic and alkaline solutions contain the following ions:

Type	Ion(s) Present	Numbers of Ions
Acid	$H^+(aq)$ and $OH^-(aq)$	$H^+(aq) > OH^-(aq)$
Neutral	$H^+(aq)$ and $OH^-(aq)$	$H^+(aq) = OH^-(aq)$
Alkali	$H^+(aq)$ and $OH^-(aq)$	$H^+(aq) < OH^-(aq)$

Diluting solutions of acids or alkalis has the following effects.

Type of Solution	Effect of Dilution on pH	Effect of Dilution on Solution	Effect of Dilution on Ions
Acid	0 → 7	Acidity decreases	Decrease in the concentration of $H^+(aq)$ ions
Neutral	7 → 7	No change	No change in the concentration of $H^+(aq)$ or $OH^-(aq)$ ions. **$H^+(aq) = OH^-(aq)$**
Alkali	14 → 7	Alkalinity decreases.	Decrease in the concentration of $OH^-(aq)$ ions

When an acid reacts with a base a reaction called **neutralisation** occurs.

Neutralisation changes the pH of acids and bases.

Type of Substance	Effect on pH	Example of pH Change
Acid	Increases to 7	pH = 0 → pH = 7
Base	Decreases to 7	pH = 14 → pH =7

There are many everyday examples of neutralisation reactions.

- Reducing soil acidity by adding lime.
- The use of lime to reduce acidity in lakes caused by acid rain.
- Treatment of indigestion.
- Treating wasp or bee stings.

In neutralisation the hydrogen ions in acids react with the hydroxide ions found in alkalis to form water.

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

hydrogen ion + hydroxide ion → water

In reactions involving metal carbonates, carbon dioxide gas is also formed.

$$2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$$

hydrogen ions + carbonate ion → carbon dioxide + water

Acids react with bases and metals to form salts.

acid	+	alkali (metal hydroxide)	→	salt	+	water		
acid	+	metal oxide	→	salt	+	water		
acid	+	metal carbonate	→	salt	+	water	+	carbon dioxide
acid	+	metal	→	salt	+	hydrogen		

The chemical test for hydrogen gas is that it ignites with a squeaky pop.

The chemical test for carbon dioxide is that it turns lime water chalky.

A salt is a substance in which the hydrogen ion of an acid has been replaced by a metal ion.

- Ammonium ions (NH_4^+) can also replace hydrogen ions (H^+) to make salts.
- Most ionic substances are salts (except oxides and hydroxides).

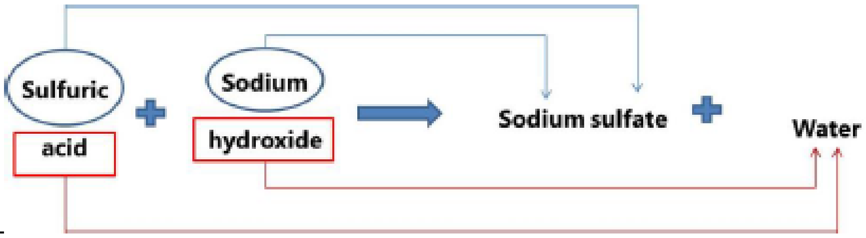
To name the salt formed in reactions, we have to use the name of the acid and base.

- The neutraliser provides the first name of the salt formed.

Neutraliser Name	Sodium hydroxide	Potassium oxide	Calcium Carbonate
First Name of Salt	Sodium	Potassium	Calcium

- The acid provides the second name of the salt formed.

Acid Name	Hydrochloric acid	Sulfuric acid	Nitric acid
Second Name of Salt	...chloride	...sulfate	...nitrate

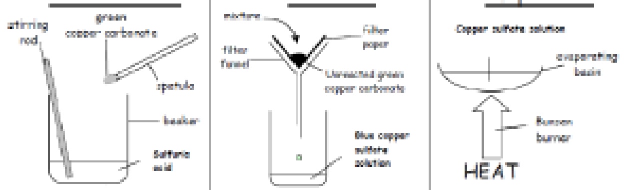


There are 2 types of salt: **soluble** and **insoluble**.

Making Soluble Salts

Soluble salts are made by **(1) Neutralisation, (2) Filtration and (3) Evaporation**.

1. Neutralisation | 2. Filtration | 3. Evaporation



1. Neutralisation
- Insoluble metal carbonate (or metal oxide) is used to neutralise the acid.
- When all acid has been neutralised, some excess carbonate or oxide will lie on the bottom of the beaker

2. Filtration
- Excess metal carbonate (or metal oxide) is removed from the solution by filtration
- The *residue* in the filter paper is unreacted metal carbonate
- The filtrate in beaker is the solution of salt you are making

3. Evaporation
- The salt solution can be returned to the solid salt by evaporating the water

NB: If the metal carbonate or metal oxide used is soluble
- the excess metal carbonate/metal oxide would dissolve in the water
- filtration would not remove the excess metal carbonate/metal oxide
- salt you are making would be contaminated by the reactants

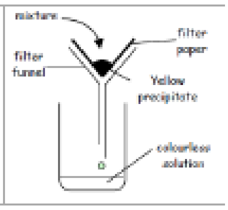
Making Insoluble Salts

Insoluble salts are made by a **precipitation** reaction. This involves mixing two solutions and forming a powdery solid called a precipitate.

****The precipitate is the insoluble salt****

When 2 solutions are mixed, there can be a chemical reaction where one of the products is insoluble in water.
- Insoluble solid product of chemical reaction is called a *precipitate*
- Insoluble salts can be *formed* by **precipitation** and *collected* by **filtration**

potassium iodide (colourless solution) + lead nitrate (colourless solution) ⟶ lead iodide (yellow insoluble solid) + potassium nitrate (colourless solution)



The insoluble solid formed in a precipitation reaction can be identified by:

Writing down the names of the reactants		Swap the names over		Check p5 of data book for solubility of products	
Potassium Iodide	Lead Nitrate	Potassium Nitrate	Lead Iodide	Potassium Nitrate is soluble ↓ Dissolved in solution	Lead Iodide is insoluble ↓ Precipitate on bottom

A special technique can be carried out to accurately work out how much base is needed to neutralise an acid. This technique is called a **titration**.

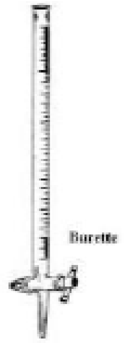
Titration experiments involve using the following apparatus:

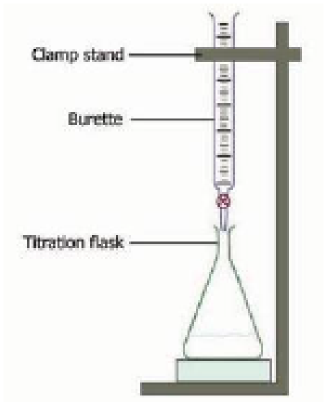
1. **A Pipette**

   A pipette is used to accurately measure out a volume of solution into a conical flask.

2. **A Burette**

   A burette is a graduated piece of glassware with a tap at the bottom of it. It can be used to release small volumes of a solution into a conical flask. Using small volumes of solution, sometimes drop by drop, allows a high degree of precision in this technique.





An **indicator** is also added to the titration flask, which will change colour when the neutralisation has taken place. In this technique you should always swirl the titration flask as you are running solution from the burette into it, this ensures thorough mixing of the chemicals. A white tile should also be placed under the titration flask to allow the colour change to be clearly seen.

In a titration experiment results must be **concordant**. This means that volume readings from the burette should be within 0.2 cm3 of each other.

Results should be recorded in a table like the following.

	Rough	Run 1	Run 2
Start Volume (cm^3)	0	16.2	32.3
End Volume (cm^3)	16.2	32.3	48.5
Titre (cm^3)	16.2	16.1	16.2

$$Average\ Titre = \frac{Run\ 1 + Run\ 2}{2}$$

Never use the rough titre value in any calculation