Name________________________ **Class**____________

✓ I am confident that I understand this and I can apply this to problems ? I have some understanding but I need to revise this some more ✗ I don't know this or I need help because I don't understand it

National 4 outcomes are in lighter font. **National 5 outcomes are in bold**

Rates of Reaction	Covered (✓)	How well can you do this?
1. Identify a chemical reaction by a change in colour		✗ ? ✓
2. Identify a chemical reaction by a change in energy		✗ ? ✓
3. When energy is released to the surroundings, it is called an exothermic reaction.		✗ ? ✓
4. When energy is taken in from the surroundings, it is called an endothermic reaction.		✗ ? ✓
5. Identify a chemical reaction by a gas given off		✗ ? ✓
6. Identify a chemical reaction by a precipitate being formed		✗ ? ✓

	Covered (✓)	How well can you do this?
7. Describe how changing the particle size affects the speed of a reaction		✗ ? ✓
8. The collision theory can be used to explain the effects of particle size and surface area on reaction rates		✗ ? ✓
9. Describe how changing the concentration affects the speed of a reaction		
10. The collision theory can be used to explain the effects of concentration on reaction rates		✗ ? ✓
11. Describe how changing the temperature affects the speed of a reaction		✗ ? ✓
12. The collision theory can be used to explain the effects of temperature on reaction rates		✗ ? ✓
13. Describe how reactions can be followed by measuring changes in mass		✗ ? ✓

	Covered (✓)	How well can you do this?
14. Describe how reactions can be followed by measuring changes in volume		✗ ? ✓
15. Average rate of reaction, or stage in a reaction, can be calculated from initial and final quantities and the time interval		✗ ? ✓
16. Catalysts are substances which speed up a chemical reaction but are not used up during the reaction		✗ ? ✓
17. Catalysts can be classified as either heterogeneous or homogeneous		✗ ? ✓
18. There are many everyday examples of uses of catalysts.		✗ ? ✓

Atomic structure	Covered (✓)	How well can you do this?
1.An element is a simple substance containing only one type of atom.		✗ ? ✓
2. State that everything in the world is made from about 100 elements		✗ ? ✓
3. State that each element has a name and a symbol		✗ ? ✓
4. Elements are classified in different ways, including naturally-occurring/made by scientists, solid/liquid/gas, metal/non-metal.		✗ ? ✓
5. State that chemists have classified elements by arranging them in the Periodic Table		✗ ? ✓
6. A group is a column of elements in the Periodic Table		✗ ? ✓
7. A period is a row of elements in the Periodic Table.		✗ ? ✓
8. Identify the following families of elements: • the halogens • the alkali metals • the noble gases • the transition metals		✗ ? ✓

	Covered (✓)	How well can you do this?
9. State that elements in one group of the Periodic Table show similar chemical properties		✗ ? ✓
10. State that every element is made up of very small particles called atoms		✗ ? ✓
11. Describe the atom as having a vary small positively charged nucleus with negatively charged electrons moving around outside the nucleus describe the location and charge of the proton, neutron and electron		✗ ? ✓
12. State the relative masses of the proton, neutron and electron		✗ ? ✓
13. State that an atom is neutral because • the positive charge of the nucleus is equal to the sum of the negative charges of the electrons • the numbers of protons and electrons are equal		✗ ? ✓

Nuclide Notation, Ions and Isotopes	Covered (✓)	How well can you do this?
14. State that atoms of different elements are different and have a different number on the Periodic Table called the atomic number state that atoms of different elements have a different number of protons, called the atomic number		× ? ✓
15. State that an atom has a mass number which equals the number of protons plus neutrons		× ? ✓
16. Calculate the number of n, p and e from the mass number and atomic number, and vice versa		× ? ✓
17.State that an atom which has lost or gained electrons is known as an ion.		× ? ✓
18. Calculate the number of n, p and e from nuclide notation, including ions, eg **37** **C1-** **17**		× ? ✓
19. State what is meant by isotopes		× ? ✓
20. State that most elements exist as a mixture of isotopes		× ? ✓

	Covered (✓)	How well can you do this?
21. State that the relative atomic mass of an element is the average mass taking into account all the isotopes present.		× ? ✓
22. The formula mass of a substance can be calculated from the relative atomic mass of the elements.		× ? ✓

Chemical Formula and Equations	Covered (✓)	How well can you do this?
23. State that the chemical formula of a compound tells us what elements are present and how many atoms of each.		✗ ? ✓
24. Formulae can be written for 2 element compounds		✗ ? ✓
25. Formulae can be written for names using prefixes, including mono-, di-, tri-, tetra-.		✗ ? ✓
26. Formulae can be written for compounds which include Roman numerals in their names, eg iron (III) chloride.		✗ ? ✓
27. Formulae can be written for compounds involving group ions but not requiring brackets, eg Na_2SO_4.		✗ ? ✓
28. Formulae requiring brackets can be written for compounds, eg $Mg(OH)_2$.		✗ ? ✓
29. Chemical reactions can be described using word equations		✗ ? ✓
30. Chemical reactions can be described using chemical symbol equations.		✗ ? ✓
31. State that there are two types of compounds		✗ ? ✓

Covalent Molecular,Covalent network and Ionic Lattices	Covered (✓)	How well can you do this?
32. State that atoms can be held together by bonds		× ? ✓
33. State that atoms can achieve a stable electron arrangement		× ? ✓
34. Describe the covalent bond in terms of atoms sharing pairs of electrons		× ? ✓
35. State that a molecule is a group of atoms held together by covalent bonds		× ? ✓
36. State that (usually) only atoms of non-metal elements bond to form molecules		× ? ✓
37.A diatomic molecule is made up of two atoms There are 7 diatomic elements in the periodic table.		× ? ✓
38. Explain the covalent bond as a situation in which two positive nuclei are held together by their common attraction for the shared pair of electrons		× ? ✓

	Covered (✓)	How well can you do this?
39. Atoms can share more than one pair of electrons leading to the presence of double and triple bonds.		× ? ✓
40. Draw a diagram to show how the outer electrons form a covalent bond		× ? ✓
41. Draw diagrams to show the shape of simple two element molecules		× ? ✓
42. Covalent substances can exist as small molecules known as covalent discrete molecules		× ? ✓
43. Certain covalent substances can exist as a giant network structure		× ? ✓
44. A covalent network structure consists of a giant lattice of covalently bonded atoms		× ? ✓
45. Discrete covalent substances have low melting and boiling points due to the weak forces of attraction that need to be overcome		× ? ✓

	Covered (✓)	How well can you do this?
46. Covalent network substances have high melting and boiling points due to the strong covalent bonds which need to be broken.		× ? ✓
47. Ionic bonding is the electrostatic force of attraction between oppositely charged ions		× ? ✓
48. Ionic compounds are usually formed when metals combine with non-metals		× ? ✓
49. An ionic structure consists of a giant lattice of oppositely charged ions		× ? ✓
50.Ionic compounds have high melting and boiling points as strong ionic bonds need to be broken to break down the lattice		× ? ✓
51.Describe experiments to test the electrical conductivity of various compounds as solids, liquids, gases and solutions		× ? ✓

	Covered (✓)	How well can you do this?
52. Metal elements (solids, liquids) and carbon (graphite) are conductors of electricity		✗ ? ✓
53. Covalent substances (solids, liquids, solutions) do not conduct electricity since they are made up of molecules which are uncharged		✗ ? ✓
54. Ionic compounds do not conduct electricity in the solid state since the ions are not free to move, but these compounds do conduct electricity when dissolved in water or when molten as the ions are now free to move		✗ ? ✓
55. Different ionic compounds form different shaped lattice structures.		✗ ? ✓

Calculations involving Gram Formula mass, balanced equations and concentration	Covered (✓)	How well can you do this?
Balanced equations		
56. Formulae equations can be balanced to show the relative number of moles of reactant(s) and product(s)		
57. The gram formula mass of any substance is known as the Mole		× ? ✓
58. The number of moles can be calculated from the mass of a substance and vice versa		× ? ✓
59. The mass of a reactant or product can be calculated using a balanced equation		× ? ✓

Acids and Bases **i pH Scale**	Covered (✓)	How well can you do this?
1. The pH scale is a continuous range from below 0 to above 14.		× ? ✓
2. Acids have a pH of less than 7; alkalis have a pH of more than 7; pure water and neutral solutions have a pH equal to 7		× ? ✓
3. Non-metal oxides which dissolve in water produce acid solutions		× ? ✓
4. Metal oxides and hydroxides which dissolve in water produce alkaline solutions		× ? ✓
5. Insoluble metal oxides do not affect the pH of water		× ? ✓
6. In water and neutral solutions, the concentration of hydrogen ions (H^+) is equal to the concentration of hydroxide (OH^-) ions		× ? ✓
7. A very small proportion of water molecules will dissociate into an equal number of hydrogen ions (H^+) and hydroxide ions (OH^-)		
8. An acidic solution contains more hydrogen ions (H^+) than hydroxide ions (OH^-)		× ? ✓

	Covered (✓)	How well can you do this?
8. An alkaline solution contains more hydroxide ions (OH^-) than hydrogen ions (H^+)		× ? ✓
9. The effect of dilution on the pH of an acid or alkali is explained in terms of the decreasing concentration of hydrogen and hydroxide ions.		× ? ✓
10. A 10x dilution changes the pH number by 1		× ? ✓
11. pH is a measure of hydrogen ion (H^+) concentration		× ? ✓
12. Neutral solutions have an equal concentration of H^+ and OH^- ions		× ? ✓
13. CO_2, SO_2, NO_2 are produced when fossil fuels are burnt		× ? ✓
14. CO_2 is also produced during the manufacture of cement.		× ? ✓
15. Soluble metal oxides can have a major impact on our environment, such as acid rain, global warming and ocean acidification.		× ? ✓

	Covered (✓)	How well can you do this?
16. Lightning storms produce much needed nitrates in the soil of rain forests.		× ? ✓
17. Acids have an important role to play in the food and drink industry e.g. as preservatives or acidity regulators		× ? ✓
18. Acids in foodstuffs can have an impact on human health		× ? ✓
19. A neutralisation reaction involves an acid reacting with a base to form water.		× ? ✓
20. Neutralisation reactions also produce a salt		× ? ✓
21. Bases are metal hydroxides, metal oxides or metal carbonates		× ? ✓
22. Neutralisation of an acid with either a metal hydroxide(alkali) or a metal carbonate involves spectator ions		× ? ✓
23. A spectator ion is present during a chemical reaction but does not take part in the reaction		× ? ✓

	Covered (✓)	How well can you do this?
24. The course of a neutralisation reaction can be followed using a pH indicator		× ? ✓
25. An indicator is a chemical which changes colour as a reaction takes place.		× ? ✓
26. Titration is an analytical technique used to determine volumes involved in chemical reactions such as neutralisation.		× ? ✓
27. Indicators are often used to show the end-point of a titration.		× ? ✓
28. Be able to name the salt formed during a neutralisation reaction		× ? ✓
29. Word equations can be written for neutralisation reactions		× ? ✓

Chemical Reactions

Chemical reactions always involve at least **one new** substance being formed.
Chemical reactions can be detected by one or more of the following

- Formation of a **precipitate** (solid formed when 2 solutions are mixed)
- **Effervescence** (gas produced)
- **Colour** change
- **Energy** change (often heat)

A reaction which **releases energy** to the surrounding is **exothermic (temperature/thermometer reading goes up … feels warmer)** whereas a reaction which **takes in energy** is **endothermic (temperature/thermometer reading goes down …feels colder)**

Rate of Reactions

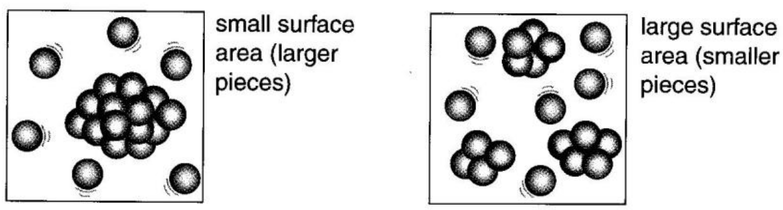
Before a reaction can occur the reactant molecules must collide and the collisions must have sufficient energy to produce a product.

There are four factors that affect the rate of reactions

- Particle size
- Concentration
- Temperature
- Catalyst

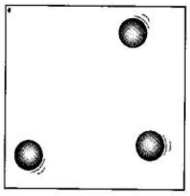
Particle size

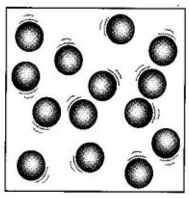
The **smaller** the particle size the **faster** the rate of reaction.
This is a result of smaller particles allowing a greater surface area of reacting molecules to be in contact and therefore a greater chance of a successful collision.



Concentration

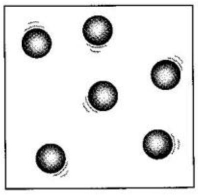
The **higher** the concentration, the **faster** the rate of reaction.
This is because higher concentrations have greater numbers of reacting molecules and therefore a greater chance of successful collision.

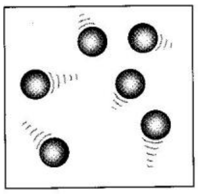

low concentration of particles


high concentration of particles

Temperature

The **higher** the temperature, the **faster** the rate of reaction.
When a substance is heated the molecules are all given more energy. As a result they move faster; increasing the chance of a successful collision.


A low temperature slows down reaction rate.


A high temperature speeds up reaction rate.

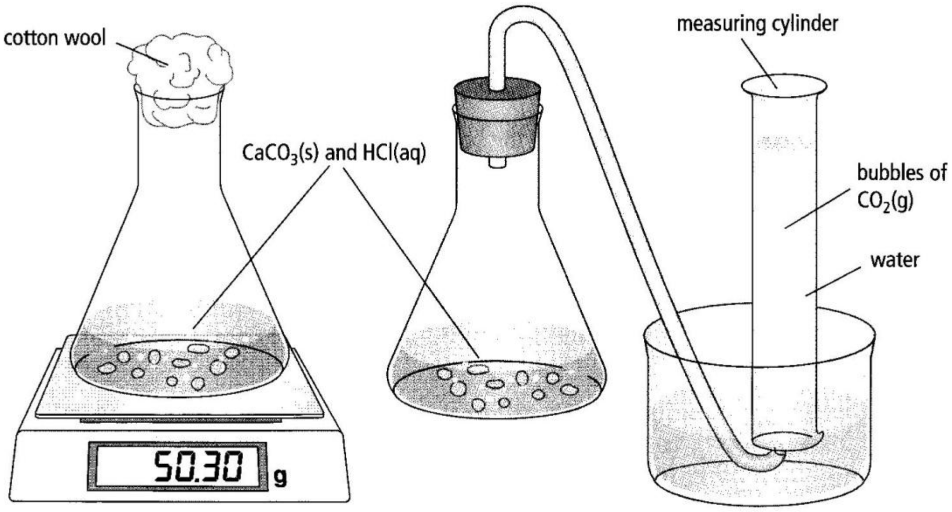
Catalyst

A catalyst **speeds up** the rate of reaction. A catalyst will remain unchanged during the reaction. The mass of catalyst remaining at the end of a reaction is the **SAME** as the mass used at the start.

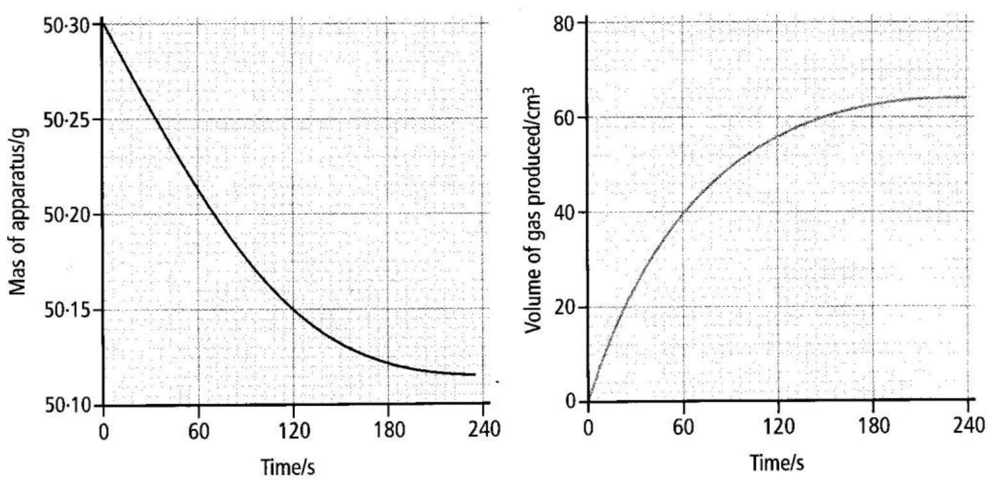
Following the Progress of a Reaction

Reactions can be followed by measuring changes in concentration, mass or volume of reactants or products.

In the reaction between calcium carbonate and hydrochloric acid the progress of the reaction can be followed by measuring the mass of carbon dioxide lost or the volume of carbon dioxide produced.

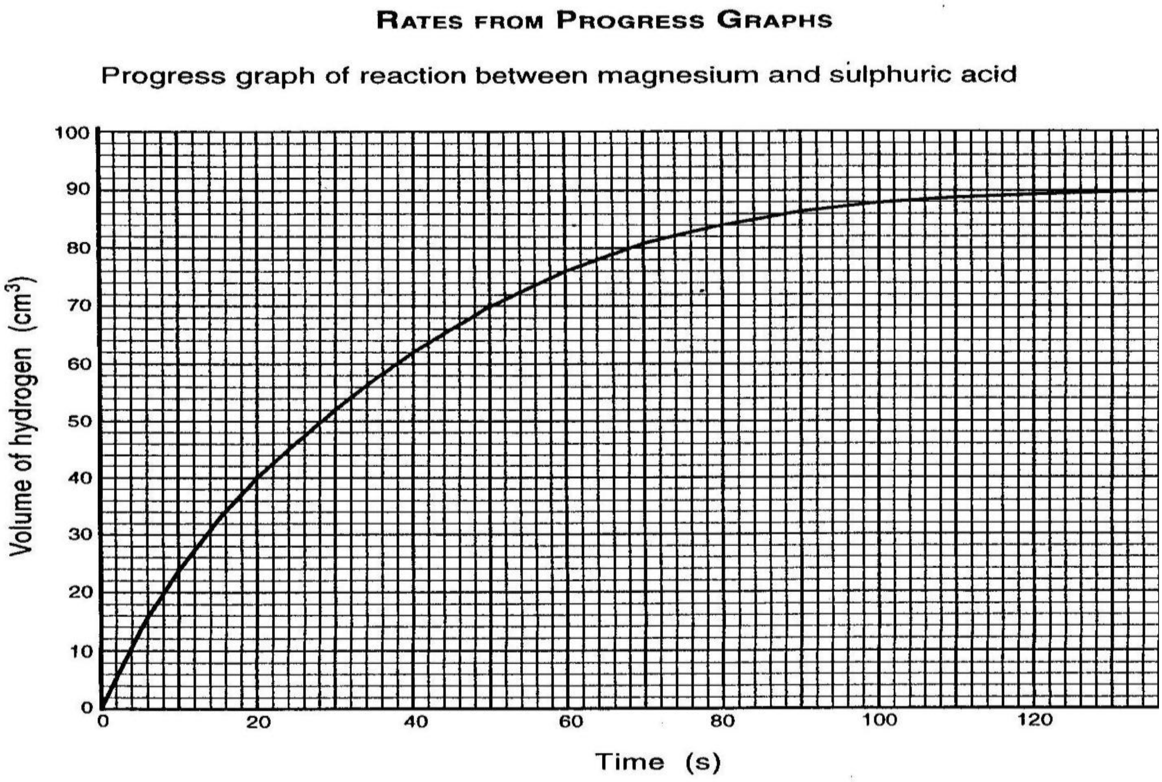


Examples of graphs are given below:



The slope of the graph gives an indication of the reaction rate. The steeper the slope, the faster the reaction. The slope is at its greatest at the start and then the reaction slows down. When the reaction is over when the line is horizontal.

Calculating the average rate of reaction



It is difficult to measure the rate at any one time as the rate is constantly changing. However, it is possible to work out the average rate over a period of time.

$$\text{average rate} = \frac{\text{change in quantity}}{\text{change in time}}$$

In the example above the change in quantity is the change in volume of hydrogen. To calculate the average rate for the first 20seconds

$$\text{average rate} = \frac{\text{change in volume}}{\text{change in time}}$$

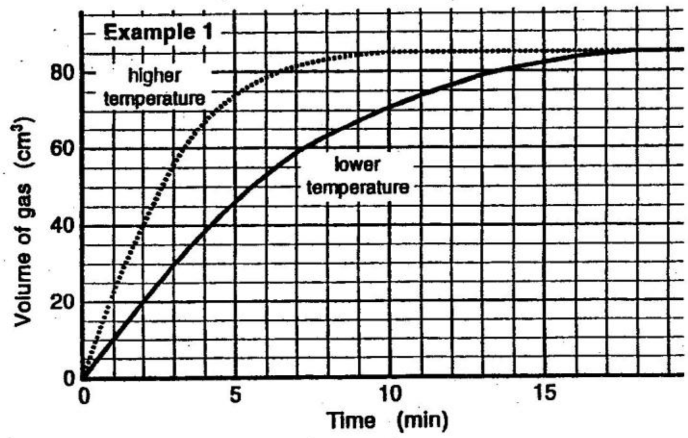
$$= \frac{40\text{cm}^3}{20\text{s}}$$

$$= \mathbf{2cm^3s^{-1}}$$

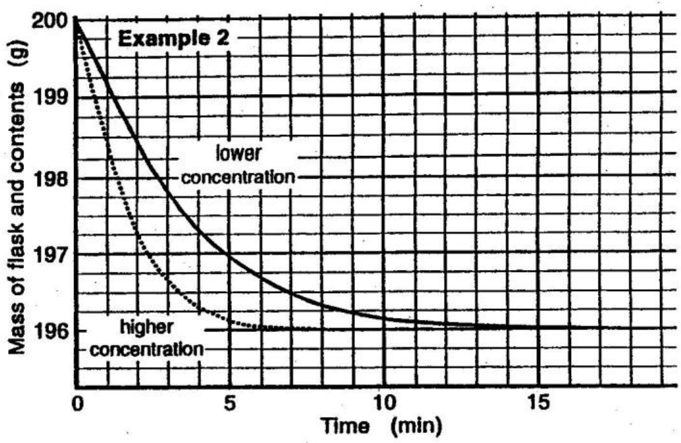
Comparing reaction progress

It is possible to compare the progress of the same reaction taking place under different conditions e.g. keeping all variables the same, the reaction repeated at a higher temperature.

COMPARING REACTIONS



Or with a greater concentration:



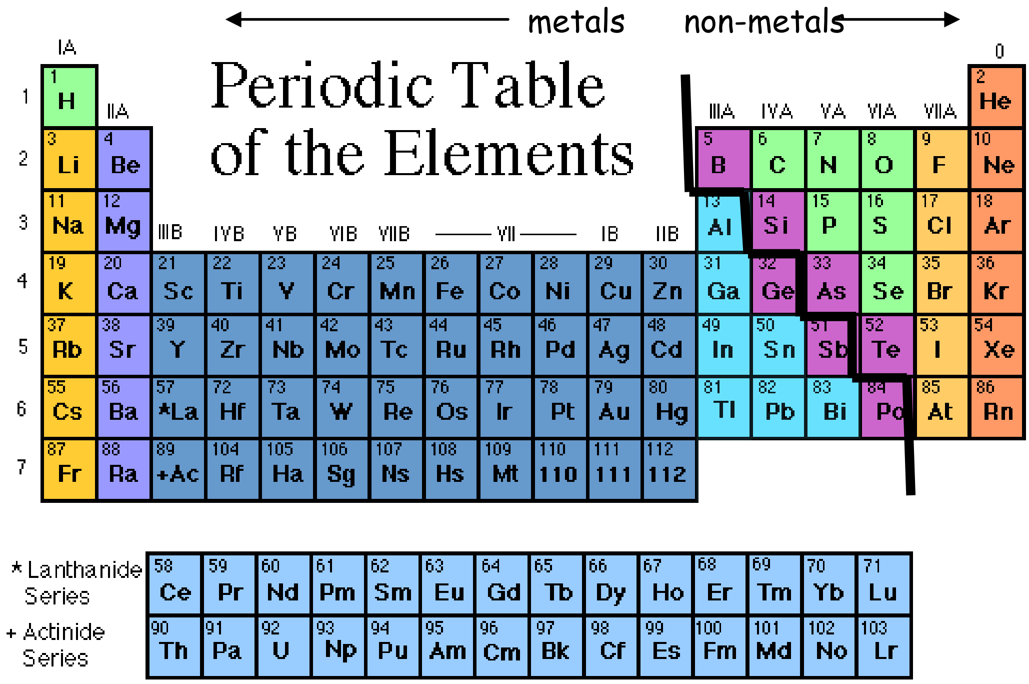
Atomic Structure

Atoms and elements

Everything in the world is made up from about 100 elements. Every element is made up of very small particles called atoms. An element is a substance in which **all the atoms are of one kind only**. An element is a substance which cannot be broken into any simpler substances. Elements can be classified in different ways

- naturally occurring or man made
- solid, liquid or gas
- metals or non-metals

The elements are arranged in the Periodic Table in increasing atomic number. Metals are all on the left side of the zigzag line whereas non-metals are on the right. The Periodic Table can be divided into **rows called periods** and **columns called groups**.



Elements with similar chemical reactions are arranged in the same column or group.

- Group 1 the Alkali Metals are very reactive and are stored under oil
- Group 7 the Halogens are reactive non-metals
- Group 0 the Noble Gases are very unreactive
- Transition Metals are the group in the middle of the Periodic Table between group 2 and 3

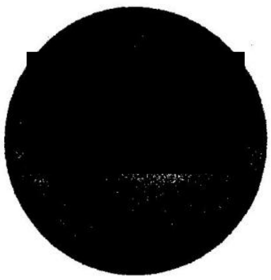
Compounds and mixtures

Compounds are formed when different elements react together. A compound contains two or more **different elements chemically joined together.**

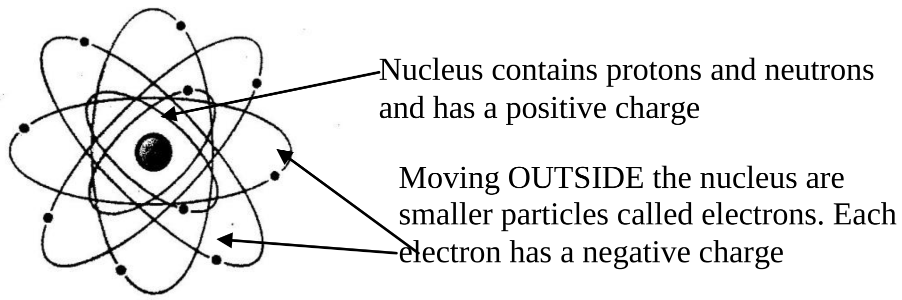
When two substances are mixed together but do not actually react (do not chemically join) a mixture is formed.

Structure of Atoms

Every element is made up of very small particles called atoms. Atoms have a very small positively charged nucleus with negatively charged electrons moving around outside the nucleus.



DALTON Model



RUTHERFORD Model

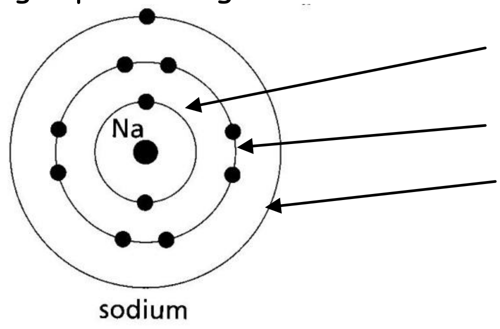
- protons have a positive charge and a mass of 1
- neutrons have no charge and also have a mass of 1
- electrons have a negative charge and a mass of 0

Particle	Location	Charge	Mass
Proton	In the nucleus	+1	1
Neutron	In the nucleus	0	1
Electron	Outside the nucleus	-1	0

An atom is neutral because the negative charge of the electrons is equal to the positive charge of the nucleus, i.e. the number of electrons is equal to the number of protons.

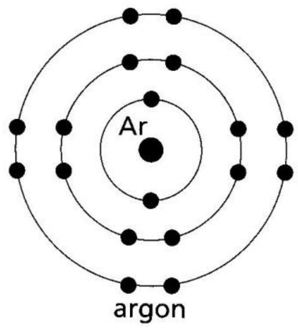
Electron arrangement

Electrons orbiting the nucleus, in the electron clouds, are arranged in energy levels (shells). The first energy level may contain up to 2 electrons, the second energy level up to 8 electrons and the third up to 18 electrons. The electron arrangement of an atom can be shown clearly using a target picture e.g.



1st energy level
2nd energy level
3rd energy level

Electron arrangement for Sodium is 2,8,1



Electron arrangement for Argon is 2,8,8

The chemical behaviour of elements is related to their electron arrangement. Elements in the **same group** of the Periodic Table have the **same number of outer shell electrons** and **similar chemical properties.**

Atomic and Mass Number

Atoms of different elements are different and have a different number on the Periodic Table called the **atomic number**.

The atomic number is the number of protons in an atom.

The sum of the protons and neutrons in an atom is known as the **mass number.**

Consider the information below about sodium in which the sodium atom is represented in nuclide notation.

$^{23}_{11}Na$

- Sodium has an atomic number of 11, so it has 11 protons (positive charges).
- The sodium atom has no overall charge so it must have 11 electrons (negative charges).
- The number of neutrons is given by the **mass number minus the atomic number** therefore sodium has **23-11=12 neutrons.**

Isotopes and Relative Atomic Mass

Isotopes of an element are atoms with the same number of protons but a different number of neutrons therefore **isotopes have the same atomic number but a different mass number.** Chlorine has two common isotopes

${}^{35}_{17}Cl$	${}^{37}_{17}Cl$
ATOMIC No = 17	ATOMIC No = 17
MASS No = 35	MASS No = 37
17 protons	17 protons
17 electrons	17 electrons
18 neutrons (35-17)	20 neutrons (37 – 17)

The isotopes will react chemically the same way because they have identical numbers of electrons. Most elements exist as a mixture of different isotopes.

The **relative atomic mass** of an element is the **average atomic mass** taking into account the proportion of each isotope. The relative atomic mass of chlorine is **35.5.**

Chorine has two isotopes ${}^{35}Cl$ and ${}^{37}Cl$ and the average mass is **35.5 is closer to 35 than 37**. **This tells us that ${}^{35}Cl$ is more abundant(large amount of) than ${}^{37}Cl$.**

Ions

Ions are formed when atoms **gain electrons to make negative ions** or when atoms **lose electrons to form positive ions**.

${}^{35}_{17}Cl^-$ (Gained e-)	${}^{23}_{11}Na^+$ (Lost e-)
17 protons	11 protons
18 electrons (17+1)	10 electrons (11-1)
18 neutrons (35-17)	12 neutrons (23-11)

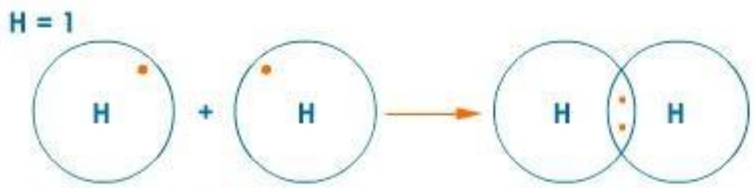
Bonding

National

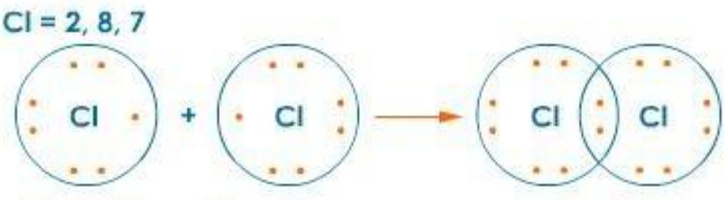
Compounds are formed when atoms of different elements join together. These atoms are held together by bonds. The atoms form bonds to achieve a full outer electron arrangement, this is also known as a stable electron arrangement. This stable arrangement can be achieved by two separate methods giving rise to two types of compounds, COVALENT and IONIC.

Covalent Molecules (Discrete Molecules)

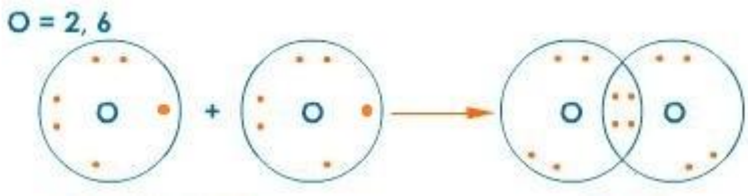
Most covalent compounds exist as molecules, which are a group of usually non metal atoms, held together by a covalent bond.
Covalent bonds are formed when electrons are shared between t National ve nuclei.
It is the attraction between the negative pair of electrons and the positive nuclei that hold the atoms together.
The seven diatomic elements exist as molecules with pairs of electrons shared between their atoms.



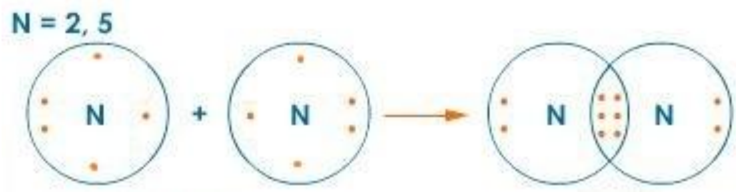
The diagram above shows how the unpaired outer electron in a hydrogen atom links up with another hydrogen atom to form a hydrogen molecule with a shared pair of electrons. Only unpaired electrons can form bonds as shown in the formation of chlorine below.



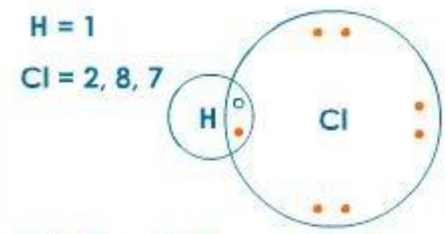
Oxygen has two unpaired electrons in its outer electron arrangement and therefore can form a double bond as shown.



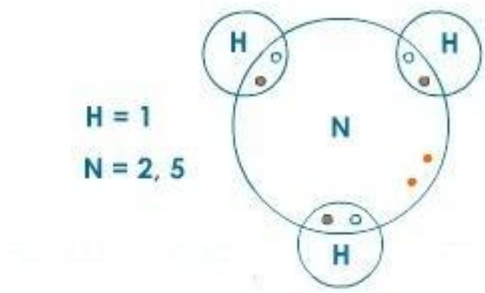
The nitrogen atom has three unpaired electrons and can therefore form a triple bond



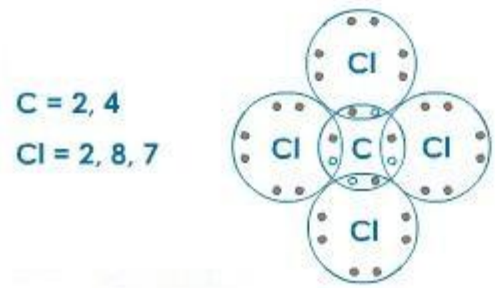
These diagrams can also be used to show how outer electrons form compounds. Again it is only unpaired outer electrons that can form bonds some examples are shown below.



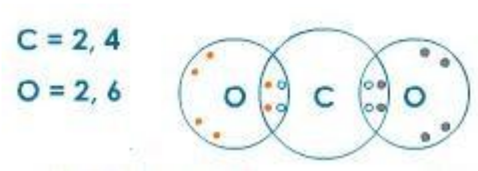
Hydrogen chloride



Nitrogen Hydride (Ammonia)



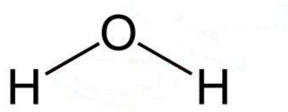
Carbon tetrachloride



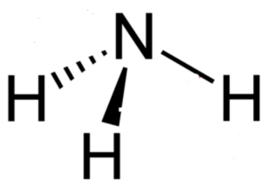
Carbon dioxide shows double bonds being formed in a compound.
Molecules also have distinctive shapes and fall into four main structures as shown below.

Linear structure include hydrogen chloride and carbon dioxide, H-Cl and O=C=O, both of these structures are described as 2 D and flat.

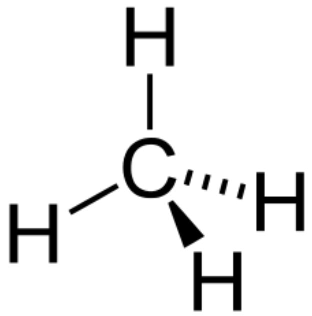
V shaped include water and hydrogen sulphide, this is again a 2 D shape and flat.



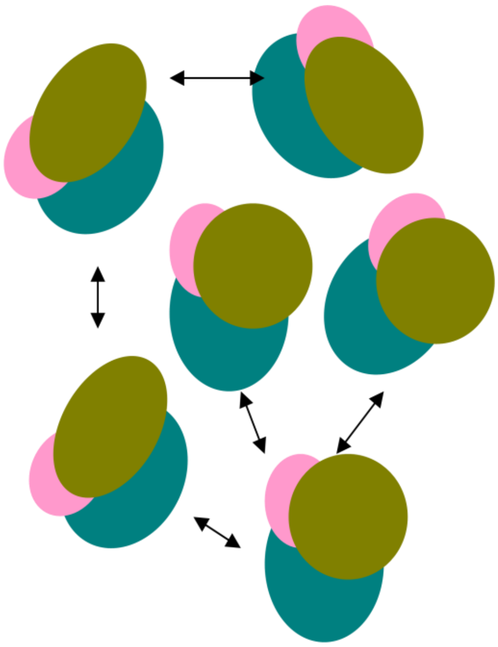
Pyramidal is the shape that nitrogen hydride has and is described as 3 D. The solid wedge bond is coming towards you and the dotted bond is going away from you.



The final shape is tetrahedral and again its described as 3D. Carbon hydride (methane) has this shape.



Most covalent molecular compounds have low melting and boiling points, and are usually liquid or gas at room temperature. When these compounds changes state it is not the strong covalent bonds inside the molecules that are broken but weak forces of attractions between the molecules. This requires a lot less energy and hence the melting and boiling points are low.

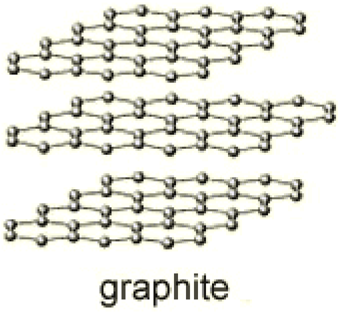


The bonds between these molecules are weak and easily broken.

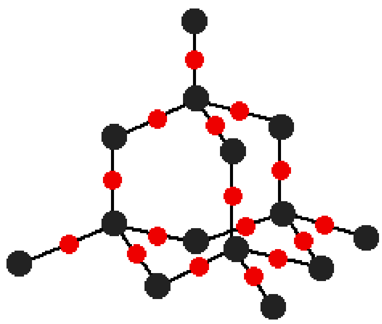
<u>Covalent Network</u>

The second type of structure covalent compounds can have is network. This is a giant lattice of millions of atoms joined together with strong covalent bonds. Unlike molecules these structures have very high melting and boiling points. When these strutures change state a lot of energy is required as its covalent bonds that are being broken and not weak forces of attraction. Substances that have this network structure include diamond, graphit and silicon dioxide. Some of the structures are shown below.

Tetrahedral structure of carbon atoms



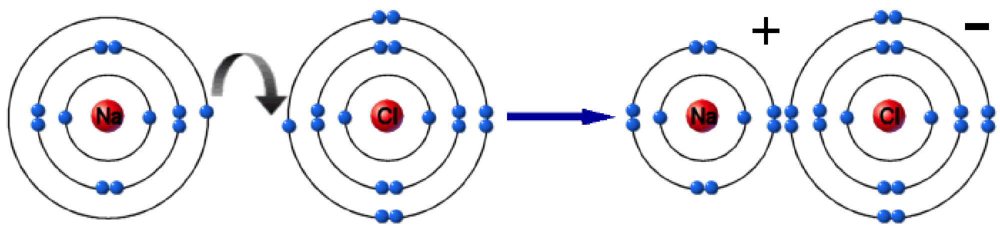
Layers of carbon with delocalised electrons between them.



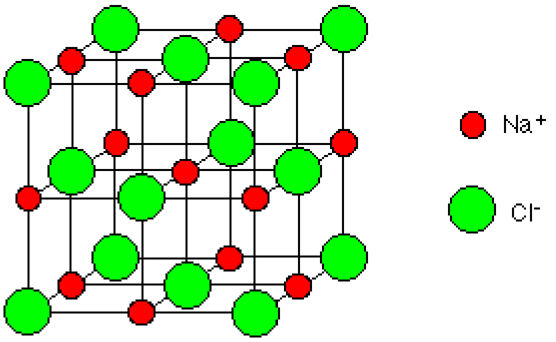
Silicon dioxide, very similar to diamonds structure.

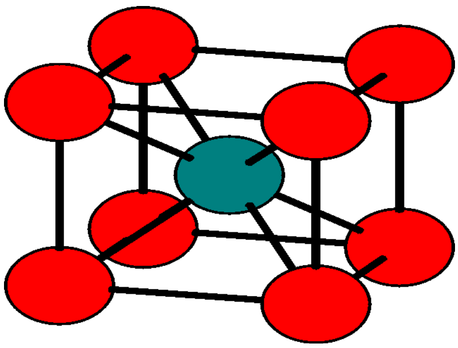
Ionic

Ionic compounds are usually made from a metal and a non metal. The bond is made when electrons are transferred from the metal to the non metal. This allows both to have a stable electron arranmgement. The metal atom becomes positivelky chared and the non metalk becomes negatively charrged. The electrostatic force of attraction betrween the metal ion amnd the non metal ion holds them together. The diagram below shows how sodium chloride is formed.



In the solid state the ions are arranged in a giant lattice of oppositely charged ions. These compounds have very high melting and boiling points as the strong ionic bonds within the lattice must be broken. All ionic compounds are solid at room temperature. There are two main shapes of an ionic(crystal) lattice and these are shown below.

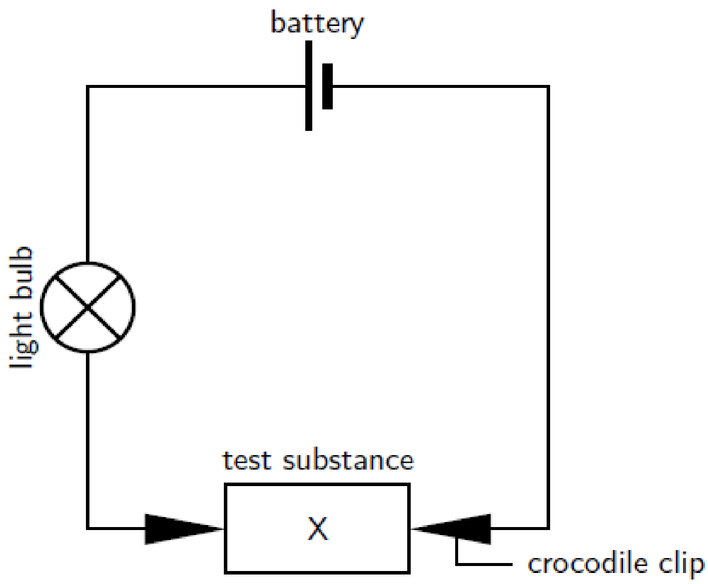




Conductivity

National 4

Testing the conductivity of a substance can indicate the type of bonding it conntains. We can use the experimental set up shown below to test the conductivity of simple substances.



If testing solutions or liquids graphite rods can be used as electrodes.

Metals are the only solid substances to conduct electricity as they have delocalised electrons that can move.
Graphite(carbon) is the only non metal to conduct as it too has free moving electrons. Molecules and other covalent compounds never conduct as they do not contain charged particles.

Ionic solids can not conduct electricity as their ions are held in the lattice and are not free to move. If we melt an ionic solid or dissolve it in water the lattice breaks allowing the ions to move and hence conduct electricity.

Compounds and Formula

Compounds are formed when two or more elements are joined together. When naming compound containing two elements, the element furthest to the left comes first in the name. The other element comes second and ends in "ide".

e.g. when copper joins with chlorine the compound formed is called copper chloride.

If compounds have more than two elements and one is oxygen they will usually end in "ate" or "ite".

e.g. Copper Carbonate contains copper, carbon and oxygen and sodium sulphite contains sodium, sulphur and oxygen.

Chemical Formula will allow us to find out the elements it contains, and the number of atoms of each element in a molecule or the ratio of the elements in larger structures such as ionic compounds.

CO_2 is the formula of carbon dioxide and tells us that it contains 1 carbon atom joined to two oxygen atoms. The formula for sodium chloride however is NaCl, this tells us for every one sodium ion there is 1 chloride ion.

Some elements have a chemical formula. There are seven diatomic element, these consist of molecules containing only two atoms. These are Hydrogen, H_2, Nitrogen, N_2, Oxygen O_2 and the halogens excluding astatine, F_2, Cl_2, I_2, and Br_2.

Chemical formulae of Two Element Compounds

Some two element compounds have meaningful names that allow us to simply write down the formula. If a compounds name contains a prefix like mono, di, tri or tetra the name tells us the formula. These prefixes equate to a number as shown in the table.

Prefix	Meaning
Mono	One
Di	Two
Tri	Three
Tetra	Four
Penta	Five
Hexa	Six

When writing the formula for these compounds if there is no prefix before the element it means there is only one of those elements in the formula.

Sulphur trioxides formula would contain 1 sulphur atom and three oxygen atoms, as we don't put the number 1 into formula we have SO_3.

Dinitrogen tetroxide would have 2 nitrogen atoms and 4 oxygen atoms giving a formula of N_2O_4.

Carbon Terachloride would contain I carbon atom and 4 chlorine atoms giving a formula of CCl_4.

Not all two element compounds will contain prefixes. This means we have to use a different method to work out the formulae of these compounds. This method requires the use of valency, this is the number of bonds an atom can make, sometimes called the combining power.

Writing Formulae Using Valency Rules.

The valency of an atom of an element, is equal to the number of bonds it can make with another atom. This can be worked out from its position in the periodic table. All elements in the same group have the same valency, e.g. all elements in group 1 have a valency of 1, and all elements in group 6 have a valency of 2.

The table below shows the valency of each group.

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

Group 8 elements, the noble gases have a zero valency as they have a full outer electron arrangement and don't form bonds.

Complete the table below writing in the valency of each of the elements.

Element	Valency
Calcium	
Phosphorus	
Caesium	
Gallium	
Arsenic	
Radon	
Iodine	

As transition metals are not in the groups we use a different rule for these elements. All transition metals have a valency of two unless the name contains a roman numeral. This numeral will indicate the valency of the metal e.g. Iron (III) chloride tells us the valency of iron is three whereas if the name was iron chloride the valency of the iron would be 2.

(I = 1, II = 2, III = 3, IV = 4, V = 5 and VI = 6.)

Using Valency Rules

When working out chemical formula we should follow these steps;

1. Write down the symbols for the elements, in the same order that they appear in the name.
2. Underneath each symbol write the element's valency
3. Cancel down (simplify) valencies if possible
4. Swap valencies.

Examples

Work out the formula of sodium oxide

Symbols Na O

Valency 1 2 Swap valencies

Formula Na_2O

Work out the formula of Iron sulphide

Symbols Fe S

Valency ~~2~~ ~~2~~ Cancel down
1 1 Swap Valencies

Formula FeS

Work out the formula for Iron (III) oxide

Symbols Fe O

Valency 3 2 can't cancel down swap valencies

Formula Fe_2O_3.

Work out the formula for Copper (I) Nitride

Work out the formula for Manganese (IV) oxide

Formulae with more than two elements.

Compounds ending in "ate", "ite", containing hydroxide or ammonium have more than two elements. All of these compounds contain group ions which can be found in the data book.

The formula given in the data booklet will have a charge, this charge indicates the valency of the group ion, and not the individual elements present. The table below shows some of the group ions and their valencies

Group ion	Formula	Valency
Hydroxide	OH^-	One
Carbonate	CO_3^{2-}	Two
Nitrate	NO_3^-	One
Sulphate	SO_4^{2-}	Two
Phosphate	PO_4^{3-}	Three
Hydrogensulphite	HSO_3^-	One
Ammonium	NH_4^+	One
Hydrogencarbonate	HCO_3^-	One

To work out the formula of a compound containing a group ion, you deal with the group formula in the same way as you would use the symbol of an element.

The same rules apply as before, swapping valencies. If the valency being swapped to the group ion is greater than one, the formula of the group ion must go inside a bracket. We do not include the charges on the group ion at this stage.

Write the formula for sodium carbonate

Symbols Na CO_3

Valency 1 2 cant cancel swap valencies

Formula Na_2CO_3

Write the formula for Calcium Hydroxide

Symbols Ca (OH)

Valency 2 1 Swap valencies bracket needed

Formula $Ca(OH)_2$ Number being swapped goes outside bracket

Write the formula for Iron (III) sulphate

Symbols Fe (SO_4)

Valency 3 2

Formula $Fe_2(SO_4)_3$

Work out the formula for Ammonium Phosphate below

Ionic Formula

Ionic formula shows the charges on the ions contained in an ionic compound. The formula is initially worked out as normal and then the charges added. The value of the charge is equal to the valency with metal ions being positive and non metal ions being negative. The charges on the group ions are shown on page 4 of the data book. Unlike valency we do not cancel charges and they should balance out to give an overall neutral charge. If we have more than one of an individual ion, the ion goes in a bracket with the charge.

Using the examples from the previous page we can write the ionic formulae.

Sodium Carbonate

Normal Formula Na_2CO_3

Ionic Formula $(Na^+)_2CO_3^{2-}$

The sodium in inside a bracket as there is two of them. You can also see that the overall charge is neutral.

Calcium Hydroxide

Normal Formula $Ca(OH)_2$

Ionic Formula $Ca^{2+}(OH^-)_2$

Chemical Equations

Chemists have devised a shorthand way of showing what happens in chemical reactions. They use word equations to show what happens and these can then be changed into formula equations.

In any equation there are two sides. On the left hand side of a word equation are the reactants, the substances that we start with. On the right hand side are the products, the substances we have made.

Reactants $\longrightarrow$ Products

When writing a word or formula equation we do not use "and", but use a "+" instead. An arrow is used to separate each side instead of an "=" sign as both sides of the equation are no longer equal.

Read the following worked examples.

In a reaction sodium has reacted with oxygen to form sodium oxide. Write the word and formula equation for this reaction.

Word Equation

Sodium + Chlorine $\longrightarrow$ Sodium
Chloride

When writing word equation the second part of the name should go on the line below the first part.

Formula Equation

$Na + Cl_2 \longrightarrow NaCl$

The formula for each reactant and product is worked out using the rules previously discussed.

Copper carbonate will break down into Copper Oxide and Carbon Dioxide when heated. Write a word and formula equation for this reaction.

Word Equation

Copper Carbonate → Copper oxide + Carbon Dioxide

Formula Equation

$Cu(CO_3) \longrightarrow CuO + CO_2$

Write word and formula equation for magnesium burning in oxygen to produce magnesium oxide.

Word Equation

Magnesium + Oxygen → Magnesium Oxide

Formula Equation

$Mg + O_2 \longrightarrow MgO$

Write the word and formula equations for the following examples.

Potassium chloride is formed when chlorine gas comes into contact with potassium metal.

Word Equation

Formula Equation

Nitrogen and hydrogen react to make nitrogen hydride

Word Equation

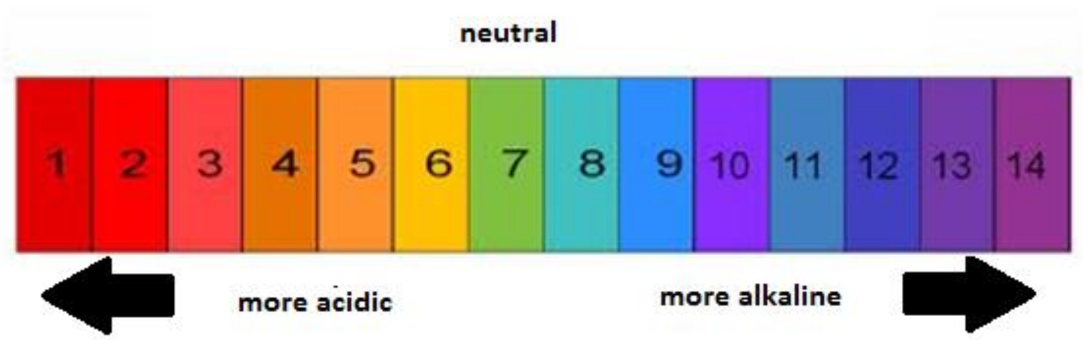
Formula Equation

The pH Scale

National 4

The pH scale allows us to determine whether a solution is acidic.
It is a continuous scale which ranges from below 0 to above 14.

Acidic solutions have pH number of less than 7.
Neutral solutions have a pH number equal to 7.
Alkaline solutions have pH number greater than 7.



There are also colours that correspond to the pH numbers of the pH scale.
Red, pink, orange and yellow indicates acidic solutions
Green indicates neutral solutions.
Blue, indigo, purple indicate alkaline solutions.

What are Acids and Alkalis?

National 4

Burning metals such as calcium and magnesium in air produces metal oxides.

calcium + oxygen $\rightarrow$ calcium oxide

If metal oxides are dissolved in water the solutions formed are alkaline, pH > 7.

Bases are compounds that react with acids to produce salts, we will look at these in detail later. However, metal oxides and metal hydroxides are bases and if soluble, will always produce alkaline solutions.

Soluble non-metal oxides can have a major impact on our environment, such as acid rain, global warming and ocean acidification. Insoluble oxides have no effect on the pH of water.

Burning non-metals such as carbon or sulfur in air produces non-metal oxides.

carbon + oxygen → carbon dioxide

If these non-metal oxides dissolve in water, acidic solutions are formed.
Solutions of non-metal oxides can have serious impacts on the environment.

Uses of Acids and Alkalis

National 4

Possibly without even noticing you will have come into contact with many acids and alkalis both at home and in the lab.

The common lab acids are hydrochloric acid, HCl, sulfuric acid, H_2SO_4 and nitric acid, HNO_3.

The most common lab alkalis are sodium hydroxide, NaOH, potassium hydroxide, KOH and calcium hydroxide, $Ca(OH)_2$.

At home you will have used vinegar, fruit juice and cola which are all acidic. Things like bleach, baking soda and washing up liquid are all alkaline.

Acids have an important role to play in the food and drink industry e.g. as preservatives or acidity regulators. Acids in foodstuffs can have an impact on human health e.g. decay of teeth.

Conductivity of Acids and Alkalis

National 4/5

Both acids and alkalis are good conductors of electricity. This means that both acids and alkalis are ionic compounds… they contain free moving ions

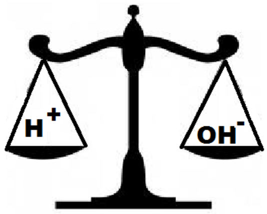
The ion common to all acidic solutions is the hydrogen ion, H^+.
The ion common to all alkaline solution is the hydroxide ion, OH^-.

Water has a very low conductivity. This means there are very few ions present. Water is made up of mainly covalently bonded H_2O molecules, a few of which break up or dissociate into H^+ and OH^- ions.

$$H_2O \rightleftharpoons H^+ + OH^-$$

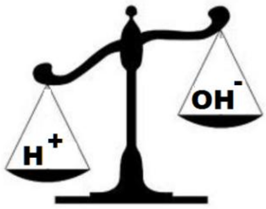
Concentration of Ions

National 5



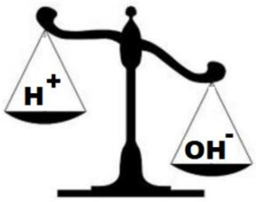
NEUTRAL

Water and neutral solutions contain equal concentrations of H^+ and OH^- ions. This means that the concentration of H^+ and OH^- ions is the same.



ACIDIC

Acidic solutions contain more H^+ ions than OH^- ions.



ALKALINE

Alkaline solutions contain more OH^- ions than H^+ ions.

As you dilute an acid, the pH <u>rises</u> towards neutral, 7, and the acidity decreases/ the concentration of H^+ ions decreases

As you dilute an alkali, the pH <u>falls</u> towards neutral, 7, the alkalinity decreases/ the concentration of hydroxide ions decreases.

Acids have more hydrogen ions than pure water
Alkalis have more hydroxide ions than pure water

Acid Rain National 4

All fossil fuels contain a little sulphur as an impurity and when the fuels are burnt, the sulphur also burns. This produces the gas sulphur dioxide which dissolves in rainwater to make acid rain.

Other gases such as nitrogen dioxide, which is produced by the sparking of air in car engines or during thunderstorms and carbon dioxide, which is produced during cement manufacture, can also produce acid rain.

Acid rain is harmful in several ways:

- It damages buildings made of carbonate rocks such as limestone and marble
- It increases the rate of rusting of iron and steel
- It increases the acidity of soil, preventing growth or harming plants
- It is harmful to plant and animal life, particularly in oceans, rivers and lochs.

Neutralisation National 4

Neutralisation is the cancelling out of an acid by another substance. The other substance is called a base

Neutralisation moves the pH of an acid up towards 7 or alkali down towards 7.

As discussed earlier, a base is a substance that neutralises an acid to form a salt. Metal oxides, metal hydroxides and metal carbonates are all bases.

Acid + Alkali National 4 & 5

Soluble metal hydroxides are alkalis. All of the reactions between acids and alkalis follow the general equation:

ACID + ALKALI → SALT + WATER

For example:

Hydrochloric acid + sodium hydroxide → sodium chloride + water

$HCl + NaOH \rightarrow NaCl + H_2O$

We can write the equation showing the ions involved:

$H^+ Cl^- \quad + \quad Na^+OH^- \quad \rightarrow \quad Na^+ Cl^- \quad + \quad H_2O$

Not all of these ions are involved in the reaction. Looking at the above equation, there are 2 ions that do not change during the reaction. These are known as spectator ions because they do not take an active part in the reaction. We can write the equation again with the spectator ions, Na^+ and Cl^- removed.

$H^+ \quad + \quad OH^- \quad \rightarrow \quad H_2O$

So in the reaction between an acid and an alkali, the hydrogen and hydroxide ions react to form water

Acid + Metal oxide **National 4 & 5**

Metal oxides which neutralise acids are often referred to as basic oxides. All reactions between acids and metal oxides follow the general equation:

ACID + METAL OXIDE → SALT + WATER

For example:

Hydrochloric acid + magnesium Oxide → magnesium chloride + Water

$2HCl \quad + \quad MgO \quad \rightarrow \quad MgCl_2 \quad + \quad H_2O$

The equation showing the ions involved is:

$2H^+ Cl^- \quad + \quad Mg^{2+} O^{2-} \quad \rightarrow \quad Mg^{2+} (Cl^-)_2 \quad + \quad H_2O$

The equation with spectator ions omitted is:

$2H^+ \quad + \quad O^{2-} \quad \rightarrow \quad H_2O$

So in the reaction between acids and metal oxides, the hydrogen and oxide ions react to form water.

Acid + Metal Carbonates **National 4 & 5**

Many metal carbonates are insoluble and as such are useful neutralisers. This means that when all of the acid has been neutralised, any unreacted metal carbonate settle to the bottom of the test tube or beaker and can be removed by filtration.

All reactions between acids and metal carbonates follow the general equation:

ACID + METAL CARBONATE → SALT + WATER + CARBON DIOXIDE

For example:

Hydrochloric acid + magnesium carbonate → magnesium chloride + water + carbon dioxide

$$2HCl + MgCO_3 \rightarrow MgCl_2 + H_2O + CO_2$$

The equation showing the ions involved is:

$$2H^+Cl^- + Mg^{2+}CO_3^{2-} \rightarrow Mg^{2+}(Cl^-)_2 + H_2O + CO_2$$

The equation with spectator ions removed is:

$$2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$$

So in the reaction between acids and metal carbonates, the hydrogen and carbonate ions react to from water and carbon dioxide.

Acid + Metals **National 4 & 5**

The reactivity of a metal is important in determining whether a metal will neutralise an acid. **Only metals which are above hydrogen in the electrochemical** series, found in the data book, will react with acids. Metals such as magnesium, aluminium and zinc react with dilute acids. Metals such as copper, silver and gold do **not** react with dilute acids

All reactions between acids and suitable metals follow the general equation:

ACID + METAL → SALT + HYDROGEN

hydrochloric Acid + magnesium → magnesium chloride + hydrogen

$$2HCl + Mg \rightarrow MgCl_2 + H_2$$

The equation showing the ions involved is:

$$2H^{+}Cl^{-} + Mg \rightarrow Mg^{2+}(Cl^{-})_2 + H_2$$

The equation with spectator ions removed is:

$$2H^{+} + Mg \rightarrow Mg^{2+} + H_2$$

So in the reaction between acids and metals, the hydrogen ions receive electrons from the metal to form hydrogen molecules.

What is a Salt? National 4/5

Salts are formed during neutralisation reactions. They are formed when the hydrogen ion of the acid is replaced by metal ions from the neutraliser/base

So in the reaction between hydrochloric acid and sodium chloride the H^{+} ion of the hydrochloric acid is replaced by the Na^{+} ion from the sodium hydroxide. This results in the salt, sodium chloride, NaCl, being formed.

Naming Salts National 4/5

The name of a salt is derived from the metal ion of the neutraliser and the negative ion of the acid.

Salts from Hydrochloric acid are called chlorides.
Salts from Nitric acid are called nitrates.
Salts from Sulfuric acid are called sulfates.

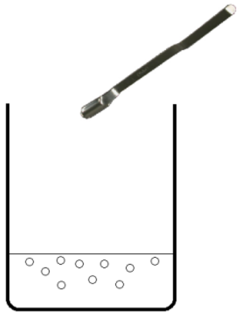
For example:

Name of the Acid	Neutraliser	Name of the Salt
Hydrochloric	Sodium hydroxide	Sodium chloride
Hydrochloric	Copper(II) oxide	Copper(II) chloride
Nitric	Zinc(II) carbonate	Zinc(II) nitrate
Nitric	Potassium hydroxide	Potassium nitrate
Sulfuric	Magnesium carbonate	Magnesium sulfate
Sulfuric	Aluminium	Aluminium sulfate

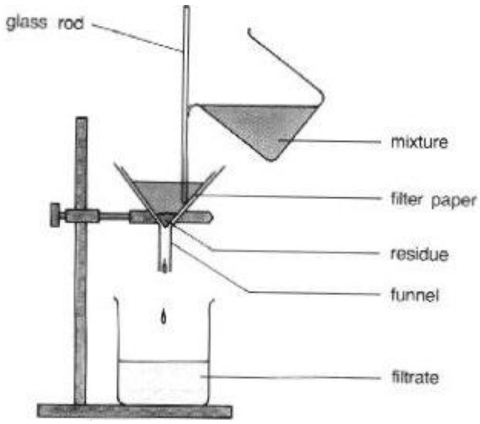
Preparing a Soluble Salt **National 4/5**

Soluble salts can be produced by carrying out a neutralisation reaction and then removing any water formed by evaporation.

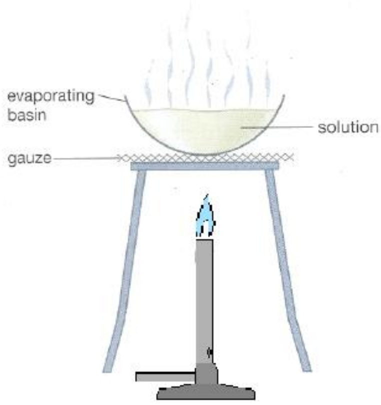
Add a spatula of calcium carbonate to 25cm^3 of hydrochloric acid and you will see effervescence as carbon dioxide is released. Continue to add small quantities of calcium carbonate until the effervescence stops and no more gas is released.



Any unreacted neutraliser must be removed by filtration.



The final stage is to remove the water from the salt solution by evaporation.



Everyday Neutralisations

National 4

Indigestion caused by excess acid in your stomach can be treated using over the counter remedies such as Milk of Magnesia, Gaviscon or Rennies. These products contain bases like magnesium hydroxide, aluminium oxide and calcium carbonate which will neutralise the excess acid.

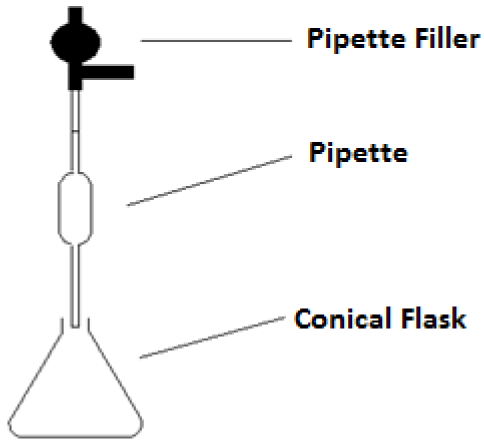
Soil which is too acidic to grow crops can be neutralised by adding lime which has the chemical name calcium hydroxide.

Lime is also used to reduce the acidity in rivers and lochs caused by acid rain.

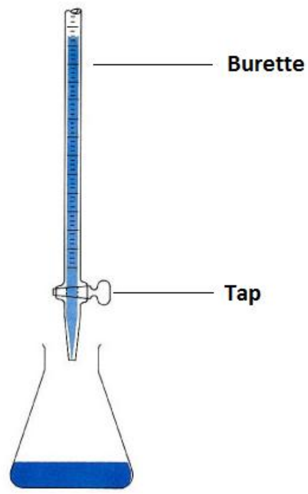
Acid-Base Titrations

National 5

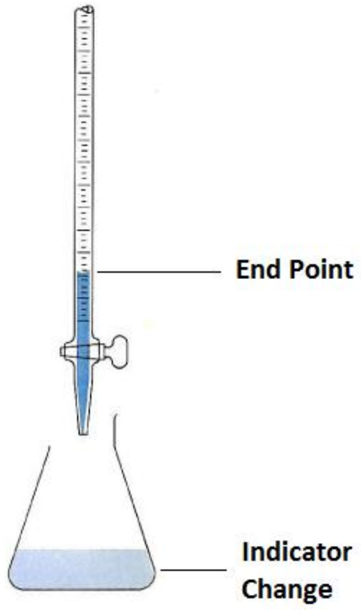
Titration is an analytical technique which allows a solution of an alkali to be neutralised accurately by the addition of an acid using an indicator to determine when the solution becomes neutral. The apparatus used is set up as follows:



A conical flask is filled with a known volume and concentration of alkali by using a pipette.



A burette is filled with a known concentration of acid. The tap is used to control the flow of acid into the conical flask containing the alkali.



When the indicator changes, signifying that the solution is now neutral, this is known as the end point or equivalence point. The volume of acid added is noted from the burette.

Titrations are carried out multiple times to ensure that the results are accurate. The first attempt is known as the Rough titration as this gives a rough estimate of the expected results. The titration is repeated until the results are concordant, meaning that results within 0.2ml of each other are obtained.

Volumetric Titration Calculations **National 5**

What volume of 0.1 mol l^{-1} hydrochloric acid is required to neutralise 25cm^3 of 0.2 mol l^{-1} sodium hydroxide?

To do this calculation we can use the formula:

$$P_{acid} \times V_{acid} \times C_{acid} = P_{alkali} \times V_{alkali} \times C_{alkali}$$

Where: P = Power (the number of H+ or OH- ions the acid/alkali has)
V = Volume, in cm^3
C = Concentration, in mol l^{-1}

STEP 1 Write down the formula of acid and the alkali and work out their power

Hydrochloric acid has the formula HCl and so has P = 1
Sodium hydroxide has the formula NaOH and so has P = 1

STEP 2 Write down the volume and concentration information given for both acid and alkali in the question

HCl has a concentration of 0.1 mol l^{-1}
NaOH has a volume of 25cm^3 and a concentration of 0.2 mol l^{-1}

STEP 3 Put all of the values into the equation

$$P_{acid} \times V_{acid} \times C_{acid} = P_{alkali} \times V_{alkali} \times C_{alkali}$$

$$1 \times V_{acid} \times 0.1 = 1 \times 25 \times 0.2$$

$$0.1\ V_{acid} = 5$$

$$V_{acid} = 5/0.1$$

$$V_{acid} = 50cm^3$$

So, $50cm^3$ of 0.1 mol l^{-1} hydrochloric acid is neutralised by $25cm^3$ of 0.2 mol l^{-1} sodium hydroxide.