St. Andrew’s Academy



Chemistry

Chemical Changes and Structures

**CHEMICAL REACTIONS**

**By the end of this section I will be able to**

* Identify a chemical reaction by a change in appearance, an energy change or a release of gas
* State that a chemical reaction involves a new substance being formed.
* Give examples of everyday chemical reactions.
* State what is meant by an endothermic and an exothermicchemical reaction
* Identify a chemical change as one in which a new substance is formed
* State that a physical change does not result in a new substance being formed.
* Give examples of physical changes including changes of state, dissolving, distilling and crystallising.
* State the tests for oxygen, carbon dioxide and hydrogen gas.

**EVIDENCE OF CHEMICAL REACTION**

Chemical reactions are occurring around us all the time!

We are able to identify a chemical reaction by the following:

1. A change in appearance
2. An energy change
3. A precipitate forming
4. Production of a gas

Complete the table showing the **evidence** of reaction from the list.

(there may be more than one!)

|  |  |
| --- | --- |
| Experiment | Evidence |
| Burning magnesium |  |
| Adding chalk to hydrochloric acid |  |
| Mixing cobalt chloride solution and  potassium carbonate solution |  |
| Mixing hydrochloric acid and sodium  hydroxide solution |  |
| Mixing copper sulphate solution and ammonia solution |  |
| Heating copper carbonate |  |
| Mixing lead nitrate solution and  potassium iodide solution |  |

Conclusion

In every chemical reaction a new substance is formed.

**Everyday chemical reactions**

Chemical reactions occur around us all the time

Examples of everyday chemical reactions are

* Lighting a match
* A candle burning
* Rusting
* Making a cake

**Exothermic reactions**

* A Firework exploding is an example of a chemical reaction
* In this case the evidence of a chemical reaction is that energy is released in the forms of heat, light and sound. Gases are also produced.

What is meant by an **exothermic reaction?**

Further examples of exothermic reactions are:



Some examples of **endothermic** reactions include

* Photosynthesis (light energy is absorbed by chlorophyll)
* Eating sherbet (the chemicals in sherbet dissolve in saliva then react with each other causing a small drop in temperature
* Cool packs to reduce fever

What is meant by an **endothermic reaction**?

Watch the following video on endothermic and exothermic reactions:

<http://www.twigonglow.com/films/energy-change-of-reactions-1450/>

**Chemical and Physical Changes**

Physical changes do not involve the formation of a new substance. The substance may change state, but it is not a new chemical!

Chemical changes occur during chemical reactions and **ALWAYS** involve the formation of a new substance.

Decide if each of these represents a chemical or a physical change.

1. A puddle evaporating \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
2. Making toast \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
3. Bridge rusting \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
4. Breaking a window \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
5. Melting chocolate \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
6. Burning wood \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Now answer the following questions

1. Name 4 ways in which a chemical reaction can be identified.
2. Explain the difference between a chemical and a physical change.
3. Complete the following table to say whether the following changes are chemical, physical or you are unable to say.

A chemical burning Mixing two powders together

Grinding a lump of chemical to a powder An ice cube melting

A gas forming A solid forming when 2 liquids are mixed

A rise in temperature when liquids mix Evaporating salty water to leave salt

A solid forming when a liquid cools A chemical changing colour when heated

|  |  |  |
| --- | --- | --- |
| Chemical | Physical | Can’t say |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

1. What is meant by an exothermic reaction?
2. Give 3 examples of exothermic reactions
3. What is meant by an endothermic reaction?
4. Give 3 examples of endothermic reactions.

**Separating Mixtures**

Mixtures come in many forms and phases. Most of them can be separated, and the kind of

separation method depends on the kind of mixture it is. Below are some common separation

methods:

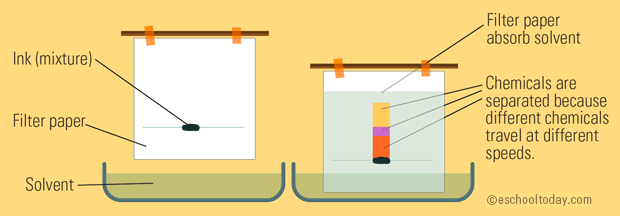
Separation of mixtures by paper chromatography

This method is often used in the food industry. It is used to identify chemicals (colouring

agents) in foods or inks. For example, if a scientist wants to know how many substances are

in a particular blob of ink, paper chromatography can be used.

Below is an illustration of using chromatography to separate and identify the different

chemicals in an ink (mixture)   
  


A blob of ink is smeared on a special paper called **filter paper**. The paper is placed in a

trough of solvent. The solvent used depends on the chemicals in the ink blob. As the paper

gets soaked upwards, it attracts the various chemicals in the ink blob. Because different

chemicals have different rates of attractions to the solvent, the chemicals will travel

upwards in different amounts. Sometimes, a chemical may not react at all. If a different

solvent is used, all the various chemicals may travel at a completely different rate



Once you have carried out your experiment, stick your chromatogram here:

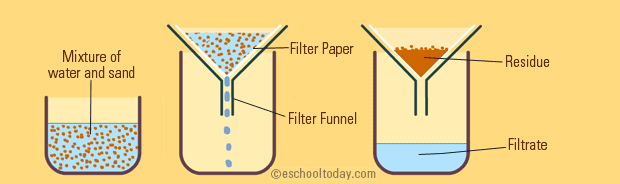
Filtration

This is a more common method of separating an insoluble solid from a liquid. An example of

such a mixture is sand and water. Filtration is used in water treatment plants, where water

from the river is filtered to remove solid particles.

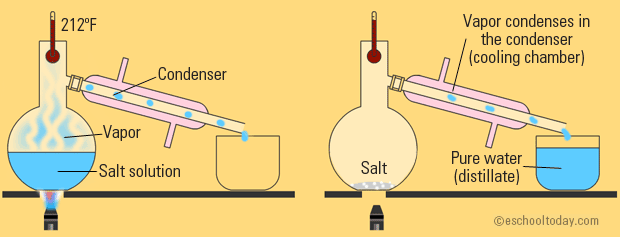
Here is a basic lab setup for filtration:

  
  
This process involves the use of filter **paper** placed in a filter funnel. The funnel is placed in a beaker and the mixture of water and sand is poured into the funnel. The liquid part drains through the filter paper into the beaker, leaving the solid sand particles trapped on the filter. In filtration, the liquid part collected is called the **filtrate** and the solid bit that   
remained on the filter paper is called the **residue**.

Evaporation   
Evaporation is the process by which water (and other liquids) changes from a liquid state to a vapour or gas state.   
  
Evaporation is great for separating a mixture (solution) of a soluble solid and a solvent. The process involves heating the solution until the solvent evaporates (turns into gas) leaving behind the solid residue.

Here is a simple example involving a mixture of salt and water.   
  
  
  
To get the salt back from the salt water, the solution is heated to boiling point. As it boils, the water escapes as vapour (gas). After some time, all the water evaporates, leaving a layer of salt at the bottom of the beaker.

**Distillation**  
  
This method is best for separating a liquid from a solution. In a way, the concept is similar to evaporation, but in this case, the vapour is collected by **condensation**.

For example: to separate pure water from a salt solution, it works like this:  
  
  
  
A beaker of the salt solution is heated to the boiling point of the liquid. As it boils, the liquid turns into vapour (gas).

The vapour is directed through tubes (condenser) connected to another beaker. As the vapour goes through the tube, it is cooled down by running cold water around   
the tubes. This forces the temperature of the vapour to fall, causing the gas to turn into liquid again (condensation). The liquid is pure at this point, as it is free from salt.

The process continues until all the liquid in the solution turns into vapour, leaving the salt residue. The distilled liquid is called a ‘**Distillate’**

**Separating mixtures - Questions**

**1** Which is the best way to get salt from salty water?

1. Evaporation
2. Filtration
3. Distillation

**2** Pure water can be separated from inky water by simple distillation because:

1. Water and ink have different boiling points
2. Water evaporates leaving the ink particles behind
3. Ink evaporates leaving the water behind

**3** What is the correct order for obtaining salt from a mixture of sand and salt?

1. Dissolving in water - filtration - evaporation
2. Evaporation - filtration - dissolving in water
3. Filtration - dissolving in water - evaporation

**4** Which method is usually used to separate coloured substances from each other?

1. Simple distillation
2. Evaporation
3. Chromatography

**5** How could you separate iron filings from a mixture of iron and sulfur?

1. Using a magnet
2. By adding water and filtering
3. By distillation

**6** In filtration, what name is used to describe the solid left in the filter paper?

1. Filtrate
2. Residue
3. Distillate

**7** If you wanted to make pure drinking water from sea water, what process would you use?

1. Filtration
2. Distillation
3. Evaporation

**8** Crude oil can be separated into several liquids that have different boiling points. What is the name of this process?

1. Simple distillation
2. Chromatography
3. Fractional distillation

**9** In chromatography, where are the spots of coloured substances placed?

1. Randomly on the piece of paper
2. In a vertical line on the paper
3. On a horizontal line on the paper

**10** What is the name of the piece of paper at the end of a chromatography experiment?

1. Chromatogram
2. Filtrate
3. Residue

**Preparation of Copper sulfate crystals**

**Aim** \_ To prepare copper sulfate crystals

**Method**

1. Collect the following:

**Eye Protection** Copper carbonateSulphuric AcidGlass Beaker

Conical Flask Measuring cylinder Spatula Glass Stirring Rod

Evaporating Dish Bunsen Burner Tripod Filter Paper Filter Funnel

1. Measure 20 ml of sulfuric acid into a beaker, using a measuring cylinder.
2. Add a spatula of green copper carbonate powder.
3. Stir with a glass rod.
4. Repeat steps 3 and 4 until no more gas is produced.
5. Filter the solution.
6. Add the filtrate to an evaporating basin.
7. Heat over a gentle blue flame until half of the original volume remains
8. Leave the solution to cool.

**Result**

**Draw a picture of your crystals here:**

**Conclusion:**

**Testing for gases**

**Oxygen**

1. Collect a test tube labelled “oxygen”
2. Light a splint using the Bunsen flame
3. Blow it out, leaving the splint glowing.
4. Insert the splint into the test tube.
5. What happens?

**Carbon Dioxide**

1. Collect a test tube labelled “carbon dioxide”.
2. Add limewater to the test tube.
3. What happens?

**Hydrogen**

1. Collect a test tube labelled “hydrogen”.
2. Light a splint from the Bunsen.
3. Insert the lit splint into the test tube.
4. What happens?

Match the gas to its chemical test

|  |  |
| --- | --- |
| Oxygen | Turns limewater cloudy |
| hydrogen | Relights a glowing splint |
| Carbon dioxide | Burns with a pop |

**Rates of Reaction**

By the end of this section I will be able to:

* Collect data from experiments and construct graphs showing the change of mass or volume against time.
* Calculate the average rate of a reaction using information from
* a graph.
* Describe how the following affect the rate of a chemical
* reaction:
* Particle size
* Concentration
* Temperature
* State that catalysts are substances which:
* Speed up some reactions
* Are not used up during the reaction
* Can be recovered chemically unchanged
* Give everyday examples of uses of catalysts, e.g. transition

Metals in car exhaust systems.

**Speed of Reactions**

As you came to school this morning you were probably too sleepy to notice that ALL AROUND YOU CHEMICAL REACTIONS WERE GOING ON! For example, metal fences and cars were rusting, stone buildings were weathering, wooden doors were rotting and so on - in fact the piece of paper this is written on is slowly changing in front of your eyes - in 100 years it will be yellow and brittle.

ALL of these are examples of **slow** reactions. However not all reactions are slow.

Write down any fast reactions you can think of.

Now use the words below to complete the table of reaction times.

**seconds weeks years hours minutes days**

|  |  |
| --- | --- |
| **Reaction Time** | **Example** |
|  | Instant glue setting |
|  | Making toast |
|  | Burning coal in a fire |
|  | Concrete setting |
|  | Steel rusting |
|  | Paper yellowing |

Chemists are interested in the **speed** at which chemical reactions take place and how they can make them go **faster** or **slower**. This is very important in the chemical industry since the amount of product a chemical manufacturer can produce will determine the amount of profit they make. Obviously, the greater the amount they can produce in a given time will result in a larger profit margin!

**How to measure the speed of a reaction**

The speed of a chemical reaction can be measured by how fast a new substance is produced. When marble chips are added to hydrochloric acid in a conical flask, carbon dioxide is produced. Since carbon dioxide is a gas it will escape from the mouth of the conical flask and as a result the flask will get lighter as the reaction proceeds. If the reaction is set up on a balance, we can determine the mass of carbon dioxide gas lost at regular intervals. The table on the next page shows some typical results for this reaction.

|  |  |
| --- | --- |
| Time / seconds | Mass of carbon dioxide produced / grams |
| 0  10  20  30  40  50  60  70  80  90  100 | 0.0  1.2  2.0  2.5  3.0  3.3  3.6  3.8  3.9  4.0  4.0 |



marble chips

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

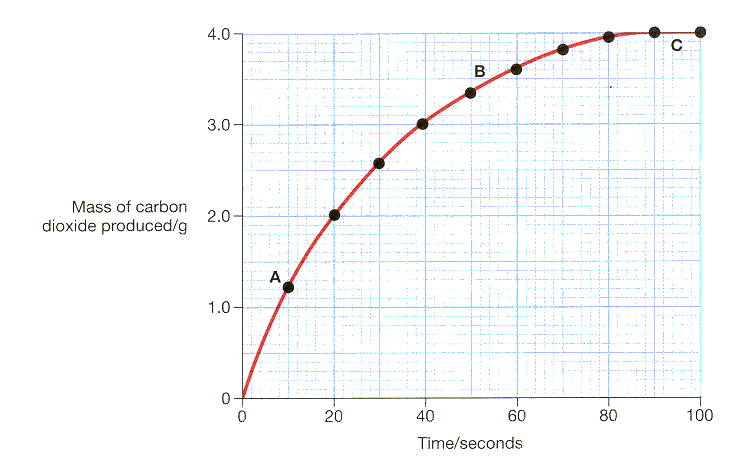
balance

loose plug of

cotton wool

## **These results can then be plotted on a graph as shown on the next page.**

The steepness of the slope of the graph gives us an indication of how fast the reaction is at various times.



## **A** - the slope is at its steepest here indicating that a lot of gas is being released every second. The reaction is at its fastest at this point.

**B** - the slope is less steep here indicating that the reaction is slowing down

**C** - the graph has levelled off. No more gas is being produced hence the reaction must have stopped

**Use the graph to answer these questions:**

1. What was the mass of gas produced between 0 and 15 seconds?

2. What was the mass of gas produced between 15 and 30 seconds?

3. At which of the two time intervals (0 - 15 or 15 - 30) is the reaction fastest?

4. Why do you think this is the case?

5. How long did it take for the reaction to stop?

**Average Rate of Reaction**

The average rate of a reaction can be calculated by:

**Average Rate = Change in measured quantity**

**Change in time**

The measured quantity could be *volume, mass or concentration*.

***Example:***

*A reaction produced 60cm3 of gas in the first 30 seconds.*

*Calculate the average reaction rate, in cm3s1.*

Average Rate = Change in volume

Change in time

= 60

30

= 2 cm3s-1.

(You will always be given the units to use at National 5)

Using the graph on the previous page, answer these questions:

1. Calculate the average rate for the first 20 seconds of the reaction.
2. Calculate the average rate of the reaction between 20 and 40 seconds.
3. A second reaction had an average rate of 0.05 cm3s-1 for the first 100 seconds.

Is the second reaction faster or slower than the reaction in the graph?

## **Changing the Speed of a Reaction**

Imagine you are an athlete and you want to improve your speed. You decide to buy new running shoes, change your diet and embark on a weight training programme. After a few weeks you find that you are indeed running faster! However, since you changed three things at the same time in order to try and improve your speed you don't actually know which one helped you to improve. The only way to find out is to change one thing at a time and then measure your speed.

The same applies when you try to find out what might speed up the rate of a chemical reaction. The things that can be changed in a chemical reaction are called **variables**. It is very important that you only change one variable at a time, keeping all the others the same in order to make the test fair.

**Particle Size**

The following experiment should be carried out to see what affect particle size has on the speed of a reaction. The reaction to be investigated is that between calcium carbonate and hydrochloric acid. In this reaction carbon dioxide gas is produced and so by comparing the rate at which the bubbles of carbon dioxide are released we can determine which reaction is faster.

To make a fair comparison between the experiments only the particle size must be changed. Calcium carbonate chips and calcium carbonate powder are used. Other variables such as concentration of acid and the temperature of the reactants must be kept the same in both experiments.

j0186164**Activity 1**

1. Measure 25cm3 of hydrochloric acid into each of two small beakers.

2. Weigh out 2g of powdered calcium carbonate and 2g of calcium carbonate chips into two weighing boats.

3. Add the powdered calcium carbonate to one beaker and the lumps to the other.

4. Observe the rate at which the bubbles of carbon dioxide are released.

**Questions**

1. What was the difference between the two experiments?

2. To keep the experiment fair what was kept the same in each experiment?

3. What reacted faster - lumps or powder? Explain your answer.

**WHY DOES A POWDER REACT FASTER?**

Before a chemical reaction can take place, the particles of the reacting chemicals must come together or COLLIDE. With a solid, only the particles on the surface can collide, so the larger the surface area (i.e. the smaller the particle size) the faster the reaction.

**Chopping this marble chip into two pieces gives EXTRA surface area at the cut part.**

**If you were to grind this chip into a hundred tiny pieces, the EXTRA SURFACE AREA would be huge.**

Although it was easy to determine which of the two experiments reacted faster just by observation it would have been more accurate to carry out the experiment on a balance in order to determine the mass loss due to carbon dioxide at regular time intervals.

balance

balance

25cm3 dilute hydrochloric acid

2g lumps of calcium

carbonate

2g powdered calcium

carbonate

Using the set of results in the following table plot line graphs of 'mass of CO2(g) produced' versus 'time' for both the powdered calcium carbonate and the calcium carbonate lumps, using the graph paper.

|  |  |  |
| --- | --- | --- |
| 0  1  2  3  4  5  6  7  8  9  10 | 0∙0  1∙5  2∙7  3∙5  4∙1  4∙6  4∙8  5∙0  5∙0  5∙0  5∙0 | 0∙0  0∙9  1∙8  2∙5  3∙2  3∙8  4∙3  4∙6  4∙9  5∙0  5∙0 |

**Time Mass of CO2 /g (powder) Mass of CO2 /g (lumps)**

**Questions**

From the graph determine the mass of carbon dioxide lost in each reaction after a period of 2∙5 minutes.

From the graph determine how long it would take for 4∙5grams of carbon dioxide to be produced.

Which reaction was fastest?

The steepness of the slope indicates the speed of the reaction. Why do you think the slope of any rate curve is always most steep at the beginning?

Explain why both curves level off at the same point.

On your graph draw a dotted line to represent the reaction between 25cm3 dilute hydrochloric acid and 1 gram of calcium carbonate.

**Concentration of Solutions**

To investigate the effect of concentration on reaction speed all other variables must be kept the same in order to make a fair comparison. Only the concentration can be changed. The experiment you will carry out is the reaction between magnesium and hydrochloric acid. When magnesium reacts with an acid hydrogen gas is produced.

**Activity 2**

1. Add 2cm depth of the dilute solution of hydrochloric acid to one test tube and 2cm depth of the concentrated solution to the other.

2. Add 2cm of magnesium ribbon to each of two test tubes.

3.Observe the rate at which the bubbles of hydrogen gas are released.

**Now complete the diagram and answer the questions which follow.**

magnesium + dilute

hydrochloric acid

lots of bubbles

**Questions**

1. What variables were kept the same in order to ensure the experiment was fair?

( hint: there are four variables )

2. What reacted faster - concentrated or dilute acid?

3. The curve on the graph on the below represents the reaction between the magnesium and the dilute acid. Add a dotted line to represent the reaction between the magnesium and the concentrated acid.

Volume of hydrogen/ cm3

Time / s

dilute acid + magnesium

**WHY DOES A CONCENTRATED SOLUTION REACT FASTER?**

## Remember that a solution is formed when a solute is dissolved in a solvent. The hydrochloric acid used in the last experiment is actually a solution. The concentration of a solution depends on the amount of solute dissolved in a given volume of solvent. Obviously, the more solute that is dissolved the more concentrated the solution is.

**Dilute Solution**

**(Only a few solute**

**particles dissolved)**

**Concentrated Solution**

**(many solute particles dissolved )**

## 

## Remember the reactant particles must collide for a chemical reaction to take place. Since the concentrated solution has many more reactant particles present than the dilute solution there will be a far greater chance of collisions occurring hence the rate of the reaction will be faster.

**Temperature**

In order to study the affect of temperature on reaction speed, we will look at the reaction between sodium thiosulphate and hydrochloric acid. In this reaction sulphur is formed and the reaction mixture, which is at first clear, turns cloudy.

As with the previous two experiments all other variables must be kept the same to keep the experiment fair.

**Activity 3**

This experiment uses 2M acid. Wash off any splashes on your skin immediately. Wear goggles. Do not sniff reaction mixture as small

amounts of sulphur dioxide are produced.

1. Add about 80cm3 of sodium thiosulphate solution to a large glass beaker.
2. From the large beaker, measure out 20 cm3 of sodium thiosulphate solution into a small glass beaker.
3. Place the small beaker containing the sodium thiosulphate solution over a piece of paper with a cross-marked on it.
4. From a small syringe, add 1 cm3 of hydrochloric acid and at the same time start the stop clock.
5. Measure the time it takes for the cross to disappear as you look down into the solution. Take a note of the temperature of the reaction mixture.

6. Heat the remaining sodium thiosulphate solution in the large beaker until its temperature is about 40oC. Remove the bunsen.

7. Measure out 20cm2 of the warm sodium thiosulphate solution into a small glass beaker and place it over the cross. Add 1cm2 of hydrochloric acid and measure the time it takes for the cross to disappear. Take a note of the temperature of the reaction mixture.

8. Repeat the experiment once more after heating the sodium thiosulphate solution to about 60oC

**Complete the table below:**

|  |  |  |
| --- | --- | --- |
| Experiment | Temperature / **oC** | Time for X to disappear |
| 1 |  |  |
| 2 |  |  |
| 3 |  |  |

**Questions**

1. What was the only difference between the experiments?

2. What variables were kept the same in order to ensure the experiment was fair?

3. As the temperature increased, what happened to the speed of the reaction?

**WHY DOES A HIGHER TEMPERATURE LEAD TO A FASTER REACTION?**

## All substances have a certain amount of kinetic energy associated with them and as a result the particles move around. The higher the temperature of a substance the more kinetic energy the particles will have hence the faster they will be able to move. If the particles are moving faster this will increase the chance of collisions hence result in a faster reaction speed**.**

**OUTSIDE THE CLASSROOM**

As was mentioned at the start of this section, reactions are going on all around you. Sometimes you do things to make reactions go faster or slower.

**Questions**

1. Explain why cutting a potato into small pieces allows it to cook faster.

2. Why does milk go sour in the sunshine quicker than it does in the fridge?

3. Why does a blacksmith blow air onto a fire?

**CATALYSTS**

Sometimes even when particle size is decreased and concentration and temperature are increased, many reactions are still too slow. In these circumstances chemists add a substance called a **catalyst** to help speed up the reaction.

Hydrogen peroxide is a solution which decomposes (breaks down) very slowly to produce water and oxygen:

hydrogen peroxide water + oxygen

The reaction is so slow that there is almost no sign of anything happening at all!

**Activity 4**

Your teacher will show you what happens to the reaction when a piece of potato is added to the hydrogen peroxide.

Describe below what happened.

**Question**

From the result of the above experiment what **two** things can be said about a catalyst?

1.

2.

**Activity 5**

1. Add hydrogen peroxide solution to a test tube to a depth of 2cm.

2. Hold the glowing splint at the mouth of the test-tube.

3. Using a spatula add some manganese (IV) oxide to the hydrogen peroxide.

4. Retest using a glowing splint.

**Questions**

1. What is meant by decomposition?
2. What gas is formed when hydrogen peroxide decomposes?
3. What is the test for this gas?
4. What happens to be speed of the reaction when the catalyst is added?
5. Is the catalyst used up in the reaction?
6. How could you recover the catalyst at the end of the reaction?

## **CATALYSTS AND THE CHEMICAL INDUSTRY**

Catalysts are very important in the chemical industry where they are used to speed up a lot of reactions. If the reaction takes place more quickly then more products can be made hence the production is more profitable to the manufacturer.

**Ask your teacher to help you complete the table below:**

|  |  |  |
| --- | --- | --- |
| **Process** | **Used to make** | **Catalyst used** |
| Contact | Sulphuric acid |  |
| Haber | Ammonia |  |
| Otswald | Nitric acid |  |

## **CATALYSTS AND THE ENVIRONMENT**

One of the most common everyday chemical reactions is the combustion of petrol in a car engine. Unfortunately, the oxides of nitrogen produced are harmful. However, these days, all car exhausts are fitted with a catalytic converter. The converter consists of a ceramic honeycomb structure coated in platinum. The harmful gases are passed through the converter. The platinum acts as a catalyst to speed up the conversion of the harmful gases into less harmful gases before they are emitted from the exhaust.

**Questions**

1. Why are cars fitted with a catalytic converter?

2. What substance acts as the catalyst?

3. Why do you think the catalyst is coated over a ceramic honeycomb structure?

**Elements**

**Elements** are the building blocks of all the substances in the world just as letters are the building blocks of all the words in our language.

Some letters, like ‘a’, ‘e’, ‘s’ and ‘t’, are found in many words while other letters, like ‘x’ and ‘z’, are found in few words. In the same way, elements like carbon and hydrogen are found in many substances while others like argon and krypton are found in few substances.

You will find all the elements in a chart called the **Periodic Table**. About 94 elements are found in the ground or in the atmosphere. Chemists have been able to make new elements that otherwise would not exist and there are now over 100 known elements.

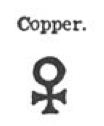
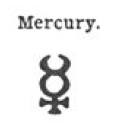
A property of a substance is something about the substance that is worth knowing … so that the substance can be used in particular ways. Although the properties of an element largely determine its use, other things must also be considered, e.g. gold does not corrode but no manufacturer is going to make cars from gold.

 ***Complete the following table for at least FIVE elements.***

|  |  |  |
| --- | --- | --- |
| **Element** | **Use** | **Reason for use** |
|  |  |  |

**Chemical symbols**

A chemical symbol is a shorthand way of representing an element. Each element has its own chemical symbol.

The ancient chemists (or alchemists) were the first to use symbols for elements in place of their names.



Modern symbols for elements consist of one or two letters. The first letter is always a capital letter; if there is a second letter, it is always a small letter.

A few elements have symbols that come from the Latin name for the element.

***REFER TO A PERIODIC TABLE.***

***Complete the following tables.***

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Element** | **Symbol** |  | **Element** | **Symbol** |
| hydrogen |  |  | chlorine |  |
|  | C |  |  | Mg |
| iodine |  |  | calcium |  |
|  | N |  |  | Sc |
| sulphur |  |  | argon |  |
|  | V |  |  | Si |

**Chemical symbols**

***What is the symbol for helium? Why is it not just H?***

***What is the symbol for cadmium? Why is it not just C or Ca?***

***Complete the following table.***

|  |  |  |
| --- | --- | --- |
| **Element** | **Symbol** | **Latin name** |
| silver |  |  |
| gold |  |  |
| iron |  |  |
| sodium |  |  |
| potassium |  |  |

***What is the origin of the word “plumber”?***

**Compounds**

A **compound** is a substance that is made up of two or more elements that are chemically joined together. Since they are joined together, it is difficult to separate out the elements that make up the compound … energy must be supplied to do this, e.g. silver oxide can be broken up into silver and oxygen by heat energy, electrical energy can be used to break up copper chloride solution.

Most of the substances that exist are not elements … they are compounds. A compound is very different from the elements that make it up, e.g. sugar (a white solid) is made from carbon ( a black solid) and hydrogen and oxygen (both colourless gases), salt (another white solid) is made up of sodium (a very reactive metal) and chlorine (a poisonous green gas).

***What is meant by a compound?***

**Co*mplete the following two tables.***

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Substance** | **Element or compound** |  | **Substance** | **Element or compound** |
| copper |  |  | sodium oxide |  |
| salt |  |  | aluminium |  |
| water |  |  | carbon chloride |  |
| sugar |  |  | hydrogen |  |
| iron |  |  | argon |  |
| oxygen |  |  | potassium nitrate |  |
| alcohol |  |  | chlorine |  |
| sulphur |  |  | calcium carbonate |  |

**Naming of compounds**

Compounds containing only two elements have names ending in **–ide**, e.g. hydrogen oxide is made up only of hydrogen and oxygen.

Compounds containing more than two elements, one of which is oxygen, have names ending in **–ate** or **–ite**, e.g. calcium nitrate is made up of calcium, nitrogen and oxygen, sodium sulphite is made up of sodium, sulphur and oxygen.

**Co*mplete the following table.***

|  |  |
| --- | --- |
| **Compound** | **Elements present** |
| iron sulphide |  |
| sodium chloride |  |
| magnesium nitride |  |
|  | hydrogen, fluorine |
|  | lithium, oxygen |
|  | calcium, iodine |
| copper sulphate |  |
| lead phosphate |  |
| aluminium sulphite |  |
| zinc carbonate |  |

***Name TWO possible compounds containing potassium, nitrogen and oxygen.***

Note that ‘hydroxide’, the second part of the name of some compounds, means that two elements, hydrogen and oxygen, are present … so compounds ending in hydroxide contain three elements, e.g. sodium hydroxide contains sodium, hydrogen and oxygen.

**Gathering Elements Together**

Everything in the world is made up of **elements**. There are now over 100 known elements. These can be gathered together in different ways. All the elements in the one ‘class’ have something in common. ***Complete each of the three following tables with at least SIX elements for each Column.***

|  |  |
| --- | --- |
| **Metal** | **Non - Metals** |
|  |  |

|  |  |  |
| --- | --- | --- |
| **Solid** | **Liquid** | **Gas** |
|  |  |  |

|  |  |
| --- | --- |
| **Natural** | **Made by Scientist** |
|  |  |

**There are only two elements which are liquid at 20 oC**

***1. To which side of the Periodic Table are the metals to be found? … the non-metals to be found?***

***2. Where in the Periodic Table are the gases to be found?***

***3. What is meant by the transition metals? Name SIX transition metals.***

***4. Name THREE elements that have been made by scientists. Where in the Periodic Table are these elements to be found.***

|  |  |  |
| --- | --- | --- |
| Atomic Structure | | |
| 1 | Every element is made up of very small particles called atoms |  |
| 2 | The atom has a nucleus, which contains protons and neutrons, with electrons moving around outside the nucleus |  |
| 3 | Protons have a charge of one-positive, neutrons are neutral and electrons have a charge of one-negative |  |
| 4 | An atom is neutral because the number of protons and electrons are equal |  |
| 5 | Protons and neutrons have an approximate mass of one atomic mass unit and electrons, in comparison, have virtually no mass |  |
| Important numbers | | |
| 6 | Atoms of different elements have different numbers of protons, called the atomic number |  |
| 7 | The electrons in an atom are arranged in energy levels |  |
| 8 | The elements of the Periodic Table are arranged in terms of their atomic number and chemical properties |  |
| 9 | Elements with the same number of outer electrons have similar chemical properties |  |
| 10 | An atom has a mass number which equals the number of protons plus neutrons |  |
| 11 | The number of protons, neutrons and electrons can be found from the atomic number and mass number, and vice versa |  |
| 12 | Atoms can be represented by nuclide notation |  |
| Isotopes | | |
| 13 | Isotopes are atoms with the same atomic number but different mass numbers |  |
| 14 | Most elements exist as a mixture of isotopes |  |
| 15 | The relative atomic mass of an element is rarely a whole number. |  |

# **Structure of the atom**

Elements are all made up of very small particles called atoms. Each element is made up of its own kind of atom. For example, copper is made up of copper atoms and helium is made up of helium atoms. Although there are many different types of atom they all have the same basic structure:

All atoms have a **nucleus** at its centre. The nucleus of an atom is **positively charged** because it contains **protons, (p).** Each proton has a **mass** of **1amu** and a **1+** charge. **Neutral** particles called **neutrons, (n)** are also found in the nucleus. Neutrons also have a **mass** of **1 amu**.

**Electrons, (e-)** are **negatively charged** particles which move around the nucleus. Electrons are so small that they have **practically no mass** and a charge of **1-**.

**Atoms** are electrically **neutral** since the **positive** charge on the **nucleus** is **cancelled** by the **negative** charge of the **electrons**. Since each proton has a 1+ charge and each electron has a 1- charge and the atom is electrically neutral it can be concluded that the number of **electrons** in an atom is **equal** to the number of **protons**.

Simple diagram of an atom:

nucleus   
(contains protons & neutrons)

electrons   
orbit the nucleus

Now complete the following table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Particle** | **Symbol** | **Charge** | **Mass** | **Location** |
| Proton  electron  neutron |  |  |  |  |

# **Atomic Number**

Earlier we said that all elements are made up of their own type of atom. So what makes atoms of one element different from atoms of the next? The answer is that the number of protons in the nucleus of each type of atom is different. In fact, the elements are arranged in the Periodic Table according to their **number of protons**. The **number of protons** in the element is **defined** as the **atomic number.**

The atomic number of the elements increases by one as you move from one element to the next across the table.

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| 1 |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  | 2 |
| 3 | 4 |  |  |  |  |  |  |  |  |  |  | 5 | 6 | 7 | 8 | 9 | 10 |
| 11 | 12 |  |  |  |  |  |  |  |  |  |  | 13 | 14 | 15 | 16 | 17 | 18 |
| 19 | 20 | 21 | 22 | 23 | 24 | 25 | 26 | 27 | 28 | 29 | 30 | 31 | 32 | 33 | 34 | 35 | 36 |
| 37 | 38 | 39 | 40 | 41 | 42 | 43 | 44 | 45 | 46 | 47 | 48 | 49 | 50 | 51 | 52 | 53 | 54 |
| 55 | 56 | 57-71 | 72 | 73 | 74 | 75 | 76 | 77 | 78 | 79 | 80 | 81 | 83 | 83 | 84 | 85 | 86 |
| 87 | 88 | 89-103 | 104 | 105 |  |  |  |  |  |  |  |  |  |  |  |  |  |

**Questions**

Use the information above and your data booklet to answer the following:

1. What is meant by the atomic number?

**MCj01345490000[1]**

2. What is the atomic number of hydrogen?

3. How many protons are there in an atom of carbon?

Complete the following table:

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Element** | Na | S |  |  |  |  |
| **Atomic Number** |  |  | 6 | 18 |  |  |
| **Number of Protons** |  |  |  |  | 1 | 8 |

Atoms are **electrically neutral** since the number of **negative electrons** is **equal** to the number of **positive protons**. This means that, as well as giving the number of protons, the atomic number also tells us the number of electrons present in the atom.

Complete the following table:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Atomic Number** | **Number of Protons** | **Number of Electrons** |
| carbon |  |  |  |
| hydrogen |  |  |  |
| magnesium |  |  |  |
|  | 5 |  |  |
|  | 7 |  |  |
|  |  |  | 20 |
|  |  | 11 |  |

# **Electron arrangement**

The electrons in an atom don't just move around the nucleus in a random fashion. They are arranged in layers called **electron shells**. The electron shell **closest to the nucleus is lowest in energy and can only hold a maximum of two electrons whereas subsequent shells can hold up to eight electrons**. The further a shell is from the nucleus the higher in energy it becomes.

Electrons

Nucleus   
(containing protons and neutrons)

Complete the table:

|  |  |
| --- | --- |
| **Energy Level** | **Maximum number of electrons** |
| 1 |  |
| 2 |  |
| 3 |  |

The electrons always fill up the electron shells in order from the lowest energy level to the highest energy level. This means that they fill the shell closest to the nucleus, working outwards away from the nucleus.

e.g.

an atom with **9** electrons would have the electron arrangement **2, 7**

an atom with **12** electrons would have the electron arrangement **2, 8, 2**

an atom with **19** electrons would have the electron arrangement **2, 8, 8, 1**

Complete the following table:

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Number of electrons | 4 | 11 | 13 | 15 | 2 | 6 | 17 |
| Electron arrangement |  |  |  |  |  |  |  |

Now complete the drawings of the atoms showing how the electrons are arranged e.g. lithium. Lithium has the electron arrangement **2, 1**

**Lithium  
2,1**

**Chlorine Phosphorus**

**Nitrogen Aluminium**

**Beryllium Argon**

# Properties of the Elements

The way an atom reacts depends on the number of electrons in the outside shell. **Since all the elements in group 1 have 1 electron in their outer shells they all react in a very similar fashion.** They all have similar chemical properties.

You will also notice that the number of electrons in the outer shell of an atom is the same as the group number that the element can be found in e.g. everything in **group 2** has **two electrons in their outer shells** and as a result of having the same number of electrons in their outer shell they all have **similar chemical properties**.

## Mass Number

We know that the mass of an atom is made up from the protons and neutrons it contains. Since each of these particles has a mass of 1amu, the **mass number** of an atom is the total **number of protons and neutrons in the atom**. For example, an atom of lithium contains 3 protons and 4 neutrons, therefore the mass number of lithium is 3 + 4 = **7.**

Now complete the following table:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Number of Protons** | **Number of neutrons** | **Mass Number** |
| carbon | 6 | 6 |  |
| calcium | 20 | 20 |  |
| sodium |  | 12 |  |
| fluorine |  | 10 |  |

## Using Mass Numbers and Atomic Numbers

Provided we know the mass number and the atomic number of an atom, we can work out the number of protons, neutrons and electrons present in the atom.

Remember:

**Atomic number = Number of protons**

Since atoms are neutral, the number of electrons is the same as the number of protons, hence the atomic number also tells us the number of electrons present in the atom.

**Mass Number = Number of protons + neutrons**

So, the atomic number tells us the number of protons and also the number of electrons in an atom. The number of neutrons can be calculated using the mass number.

We know that:

**mass number = protons + neutrons**

From this we can calculate the number of neutrons:

**neutrons = mass number - protons**

And since the number of protons is the same as the atomic number:

**neutrons = mass number - atomic number**

For example, fluorine has a mass number of 19 and an atomic number of 9. Therefore:

neutrons = 19 - 9

= 10

Chemists use a special system for writing the atomic number, mass number and symbol for an element. This is called **nuclide notation**.

**mass number**

where X is the symbol

**X**

**atomic number**

To find out if you really understand this, calculate the number of protons, neutrons and electrons in each of the following atoms of elements. The first one has been done for you.

7 23 39

Li Na K

3 11 19

protons : **3** protons : protons :

electrons : **3** electrons : electrons :

neutrons : **4** neutrons : neutrons :

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Questions**

Using nuclide notation, show:

(a) A helium atom which contains two protons and two neutrons.

(b) A sulphur atom which contains 16 neutrons and 16 electrons.

Now complete the following table:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Element** | **Atomic number** | **Mass number** | **Number of protons** | **Number of neutrons** | **Number of Electrons** |
| 12  C  6 |  |  |  |  |  |
| 22  Ne  10 |  |  |  |  |  |
| 18  O  8 |  |  |  |  |  |
|  | 20 | 40 |  |  |  |
|  |  | 7 |  | 4 |  |
|  |  | 9 |  |  | 4 |

# Ions

During a chemical reaction atoms can often ***gain*** or ***lose*** electrons. When this happens the atoms are no longer electrically neutral since the number of electrons is no longer equal to the number of protons. The atoms have now become **electrically charged** particles called ***ions.***

If an atom ***gains*** an electron during the reaction then it will become a ***negatively charged ion***. This can be shown in nuclide notation by adding a negative sign to the symbol. For example, if fluorine gains an electron and becomes negatively charged it can be written as:

**19**

**F**

**9**

If an atom ***loses*** an electron during the reaction then it will become a ***positively charged ion***. This can be shown in nuclide notation by adding a positive sign to the symbol. For example, if sodium gains an electron and becomes positively charged it can be written as:

**23**

**Na+**

**11**

Now calculate the number of protons, neutrons and electrons in each of the following ions:

Li

7

3

Cl

35

17

(a) (b) (c)

K

19

39

protons : protons : protons :  
neutrons: neutrons: neutrons:  
electrons: electrons: electrons:

15

7

N3-

40

20

Ca2+

16

32

S 2-

3

(d) (e) (f)

protons : protons : protons :  
neutrons: neutrons: neutrons:  
electrons: electrons: electrons:

# Isotopes (5)

Chemists soon discovered that elements had more than one type of atom.

How can two atoms of the same element be different? Consider the element lithium:

6 7

Li Li

3 3

Nucleus containing

3 protons

**3+**

**3+**

**MCj01345490000[1]Questions**

1. What is the atomic number of both atoms?

2. How many protons does each atom have?

3. Why can both atoms be considered to be atoms of lithium?

4. How many electrons does each atom have?

5. Explain why both atoms will react in exactly the same way.

6. What is the only thing that is different between the two atoms?

Since both atoms have a **different mass number** they must have a **different number of neutrons** in their nuclei.

**6 7**

Li Li

**3** **3**

mass number: mass number:

no. of neutrons: no. of neutrons:

Atoms with the **same atomic number** but **different mass numbers** are called **ISOTOPES**. They are atoms of the same element but have different numbers of neutrons.

Complete the table below:

|  |  |  |
| --- | --- | --- |
|  | 35  **Cl**  17 | 37  **Cl**  17 |
| Symbol |  |  |
| Atomic Number |  |  |
| Number of Protons |  |  |
| Number of electrons |  |  |
| Mass Number |  |  |
| Number of Neutrons |  |  |

|  |  |
| --- | --- |
| **Isotopes have the same** | **Isotopes have different** |
|  |  |
|  |
|  |  |
|  |

# Relative Atomic Mass (5)

Most elements are made up of isotopes, for example hydrogen has three isotopes:

1 2 3

H H H

1 1 1

Using a machine called a **Mass Spectrometer**, scientists were able to find out:

a) how many isotopes of an element there was and

b) how much of each isotope was present.

Information from a mass spectrometer for the element chlorine is shown in the graph:

Percentage of each isotope

Mass Number

35

37

75

50

25

It can be seen from the graph that chlorine has **two** isotopes. This means that there are two different masses for chlorine atoms. As a result an average is taken of both the isotopes. This average is called the **relative atomic mass**. The relative atomic mass of chlorine is, in fact, **35.5** and not 36 as you might expect. This is because there are **more 35Cl atoms present** than 37Cl so the average is nearer **35** than 37.

By definition, the ***relative atomic mass*** is:

***The average mass of all the isotopes of an element taking their*** abundance ***into account.***

**MCj01345490000[1]Questions**

Copper has a relative atomic mass of 63.5. It contains two isotopes:

63Cu and 65Cu

(a) How many neutrons are present in each kind of atom?

(b) Which isotope is most common in copper? Explain your answer.

# 

|  |  |  |
| --- | --- | --- |
| **Covalent Bonding** | | |
| **1** | Atoms can be held together by bonds |  |
| **2** | In forming bonds, atoms can achieve a stable electron arrangement |  |
| **3** | In a covalent bond atoms share pairs of electrons |  |
| **4** | The covalent bond is a result of two positive nuclei being held together by their common attraction for the shared pair of electrons |  |
| **5** | Covalent bonds are strong forces of attraction |  |
| **6** | Chemical formula gives the number of atoms of each element in a molecule of covalent substance |  |
| **7** | Usually only atoms of non-metal elements bond to form molecules |  |
| **8** | A diatomic molecule is made up of two atoms |  |
| **9** | Give examples of elements which exist as diatomic molecules |  |
| **10** | Identify shapes of molecules as linear, bent, pyramidal and tetrahedral |  |
| **11** | Draw a diagram to show how the outer electrons form a covalent bond |  |
| **12** | A covalent network structure consists of a giant lattice of covalently bonded atoms. |  |
| **13** | The formula for a covalent network substance gives the simplest ratio of atoms of each element. |  |

# **Bonding, Structure and Properties**

The Noble Gases can be found in Group 0 of the Periodic Table. They are the only elements that exist as individual atoms. Scientists believe that this is because the electron arrangements of the Noble gases is very stable; each with a full outer energy level. In order for any other element to achieve this stable electron arrangement they react with other elements in a chemical reaction.

We already know that when a chemical reaction occurs chemical bonds are formed between the atoms of the reactants and one or more new substances are always formed.

Scientists believe that the reason a chemical reaction occurs is for atoms to become more stable; to achieve a Noble Gas electron arrangement. It makes sense then, that when chemical bonds form between reactants the electrons in the outer energy level are involved. There are two ways in which the electrons can rearrange themselves to achieve a Noble gas arrangement. They can share their electrons with other atoms forming Covalent bonds or there can be a transfer of electrons between atoms. If this happens Ionic bonds are formed.

## **Covalent Bonds**

**Non-metal** elements form covalent bonds by sharing electrons.

Consider hydrogen; it only has one electron which is found in the orbital in the first electron shell. In order to achieve stability similar to that of helium, the first Noble gas, two atoms of hydrogen can overlap their half-filled orbital and share their electrons:

H

H

H

H

electron

atom atom molecule

Each hydrogen atom now has a share in an extra electron. This gives both atoms the **same electron arrangement as helium** and thus added stability.

Once the two orbitals overlap a force of attraction occurs between the **positive nuclei** of the two atoms and the negative **electrons** in the overlap region. This **force of attraction** holds the atoms together and is called a **Covalent Bond**.

**+**

**e**

**e**

Positive nucleus

Negative electrons

Force of attraction

**+**

**Positive**

**nucleus**

By definition, ***a molecule is a group of atoms held together by covalent bonds***. The hydrogen is therefore now a **molecule** containing **two atoms** bonded together. A covalent bond can be represented by a line so the molecule of hydrogen can be shown as **H-H,** which is its **full structural formula**. It is more common to use the **molecular formula** which is written as **H2**. Since the molecule only contains **two atoms** it is called a **diatomic** molecule.

### Consider Chlorine

**Chlorine** has the electron arrangement **2, 8, 7**. This means that the outer electrons are found in the **four orbitals** within the **third** energy level. Three of the four orbitals are filled however, the **fourth orbital is half-filled**. In order to become more stable the **half-filled** **orbital** overlaps and **shares** electrons.

**Cl**

**Cl**

**Cl**

**Cl**

**Cl**

**Cl**

Each chlorine atom now has a share in an extra electron. This gives both atoms the **same electron arrangement as Argon** and thus added stability. The formula for chlorine is therefore **Cl2**. Chlorine is also a diatomic molecule.

### Consider Oxygen

**Oxygen** has the electron arrangement **2, 6**. This means that the outer electrons are found in the four orbitals within the second energy level. Two of the four orbitals are filled however, **two orbitals are half-filled**. In order to become more stable the half-filled orbitals overlap and **share** electrons.

##### **O**

**O**

**O**

**O**

Each oxygen atom now has a **share** in an extra **two electrons**. This gives both atoms the **same electron arrangement as Neon** and thus added stability. To achieve the same arrangement as neon, oxygen forms a **double covalent bond**. The formula for oxygen is therefore **O2**. Oxygen is also a diatomic molecule.

There are seven elements in total which exist as diatomic; hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine.

Atoms of different non-metal elements can bond in exactly the same way.

### Consider hydrogen chloride

**Cl**

**H**

**Cl**

**H**

The **half-filled orbital** in the **hydrogen** atom **overlaps** with the **half-filled orbital** in the **chlorine** atom. This means that they now have a **share of two electrons** giving each atom a full outer shell of electrons. **Hydrogen** now has the **same electron arrangement as helium** and **chlorine** now has the **same electron arrangement as argon** thus both atoms are now more stable. The formula for hydrogen chloride is therefore **HCl**. Hydrogen chloride is also a diatomic molecule since it only contains two atoms however, since the atoms are different hydrogen chloride is a **compound**.

### Consider carbon hydride

Carbon hydride has the chemical formula **CH4**. If we consider the electron arrangements of both hydrogen and carbon and then draw a bonding diagram we can understand why this is the case.

###### H

###### H

###### H

###### H

##### C

###### H

##### C

###### H

###### H

###### H

A molecule of carbon hydride therefore consists of one carbon atom and four hydrogen atoms held together by four covalent bonds. **Carbon hydride** is also a **compound** since the **atoms** held together by the covalent bonds are **different** however; it is **not diatomic** since there are **5 atoms** in the molecule.

Draw bonding diagrams for water (hydrogen oxide) and ammonia (nitrogen hydride)

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## **Simpler Electron Diagrams**

By leaving out the boundaries of the orbitals bonding diagrams can be more simply drawn as electron diagrams:

e.g

H

H

H

H

**Cl**

**H**

**Cl**

**H**

In the space below draw electron diagrams for chlorine, oxygen, nitrogen, water (hydrogen oxide), ammonia (nitrogen hydride) and methane (carbon hydride), phosphorus hydride and tetrachloromethane (CCl4)

**Activity 1**

Collect a set of models. Using the key provided look very closely at an atom of carbon, hydrogen, oxygen and chlorine and complete the following table.

|  |  |  |
| --- | --- | --- |
| **Type of Atom** | **Colour of Model Atom** | **Number of Possible Bonds (valency)** |
| carbon |  |  |
| hydrogen |  |  |
| oxygen |  |  |
| chlorine |  |  |

Now using the models make molecules of hydrogen chloride, carbon hydride and carbon chloride and draw them in the boxes below. (use coloured pencils )

Hydrogen chloride Carbon hydride Carbon chloride

Chemists, being smart people, often use a shorthand way of drawing molecules.

Consider the water molecule:

where the lines between the atoms represents a covalent bond

**O**

**H**

**H**

**O**

**H**

**H**

becomes

Now draw molecules of hydrogen chloride, carbon hydride and carbon chloride using shorthand in the boxes below.

Hydrogen chloride Carbon hydride Carbon chloride

When a molecule is represented in this way it is called the FULL STRUCTURAL FORMULA.

Another, even shorter way, of representing the molecules is to write them without showing the covalent bonds between the atoms.

ie

**O**

**H**

**H**

becomes **H2O**

**H**

**H**

**H**

**H**

**C**

becomes **CH4**

Representations like H2O and CH4 are called CHEMICAL FORMULA.

**The chemical formula tells you the number of atoms of each element present in the covalent molecule.**

**Activity 2**

Collect a box of molecules. Using the key provided in the box, draw the full structural formula and chemical formula for the six molecules in the table below.

|  |  |  |
| --- | --- | --- |
| **Name of Compound** | **Full Structural Formula** | **Chemical Formula** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

**Shapes of molecules**

The chemical formula for a covalent substance indicates the number of atoms of each element in the molecule.



It does not give any information about the shape of the molecules.

The shapes of molecules are based on the tetrahedron arrangement of electrons.

In some molecules the atoms are arranged in a straight line (linear),



e.g. hydrogen chloride, HCl.



Some molecules are two-dimensional (flat), e.g. water, H2O.

Others are three-dimensional, e.g.





carbon tetrachloride, CCl4

nitrogen hydride, NH3

The **full structural formula** shows the shape of a molecule.

## **Diatomic Molecules**

The **noble gases** are the only elements which exist as **individual atoms**. These elements are described as being **'monatomic'** which literally means **'one atom'**.

atoms of neon

All other elements exist in structures in which the atoms are joined together in some way.

The **non-metal** elements (except the noble gases) exist as different sized groups of atoms held together by **covalent bonds**.

What do we call a group of atoms held together by covalent bonds?

The smallest molecules have just **two atoms** joined together. These molecules are described as being **diatomic**.

What does 'diatomic' mean?

There are **seven** non-metal elements which exist as **diatomic molecules**. This means that they exist as two atoms joined together by covalent bonds.

**Activity 3**

Using models complete the table on the next page for the seven diatomic elements.

|  |  |  |
| --- | --- | --- |
| Element | Chemical Formula | Structural Formula |
| hydrogen |  |  |
| nitrogen |  |  |
| oxygen |  |  |
| fluorine |  |  |
| chlorine |  |  |
| bromine |  |  |
| iodine |  |  |

You should have noticed that the atoms of all the elements except **oxygen** and **nitrogen** were held together by a **single** covalent bond. If you built the molecules correctly you would also have noticed that **oxygen** was held together by **two** covalent bonds and **nitrogen** was held together by **three**.

**O**

**O**

**N**

**N**

**double covalent bond**

**triple covalent bond**

It is not only elements that exist as diatomic molecules. Compounds can be diatomic too.

In the space below give examples of diatomic compounds.

**A diatomic molecule is made up of two atoms held**

**together by covalent bonds.**

**Covalent Network**

All of the covalent molecules we have considered so far are very **small molecules** and are described as being **covalent molecular**. Another type of covalent substance exists. These are giant structures containing many thousands of strong covalent bonds and are described as being covalent networks. Silicon dioxide (silica) is a covalent network structure. The formula is SiO2 however, this does not mean that it only has one silicon atom bonded to two oxygen atoms; it means that in the network structure the ratio of atoms is 1:2.

**Si**

**O**

**O**

**O**

**O**

**Si**

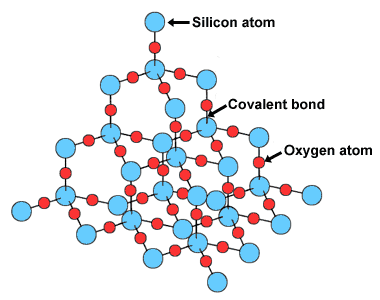
**Si**

The diagram only shows a tiny part of the network structure. In reality there are thousands of silicon atoms bonded to thousands of oxygen atoms in a ratio of 1:2

**O**

**O**

The actual network structure of silicon dioxide is shown in the diagram below. Again, the diagram represents only a small part of the vast network structure!



|  |  |  |
| --- | --- | --- |
| Ionic Bonding | | |
| **1** | Describe the formation of ions in terms of atoms losing and gaining electrons |  |
| **2** | Ionic compounds are usually formed when metals combine with non-metals |  |
| **3** | Positive ions are formed by metal atoms losing electrons and negative ions are formed by non-metal atoms gaining electrons |  |
| **4** | Ionic bonding is the electrostatic force of attraction between oppositely charged ions |  |
| **5** | An ionic structure consists of a giant lattice of oppositely charged ions |  |
| **6** | The formula for an ionic compound gives the simplest ratio of positive to negative ions |  |

# **Ionic Bonds**

# 

Atoms of elements can also achieve the stable electron arrangement of the noble gases by losing or gaining electrons. Metal atoms *lose* their outermost electrons to form *positively* charged particles and non-metal atoms *gain* electrons to form *negatively* charged particles. These charged particles are called *ions*. The *electrostatic force of attraction* between the oppositely charged ions is called an *ionic bond*.

Sodium chloride, which is common salt, contains ionic bonds. If we consider the electron arrangements of both sodium and chlorine we can see why.

Sodium

What is the electron arrangement of sodium?

If sodium were to lose one electron what would the new electron arrangement be?

Why is it desirable for sodium to lose this electron?

Which noble gas does the sodium ion have the same electron arrangement as?

Atoms are electrically neutral since the number of negative electrons is exactly the same as the number of positive protons. Sodium has 11 protons and 11 electrons. However, if the atom loses an electron it is no longer electrically neutral since it now has 11 protons and 10 electrons. The charged particle formed is called a positive ION and is written Na+.

Sodium atom Sodium ion

8e

2e

11+

**2e**

1e

8e

2e

11+

**2e**

11+

11 protons 11 protons  
11 electrons 10 electrons

Na 2, 8, 1 Na+  2, 8

Chlorine

What is the electron arrangement of chlorine?

If chlorine were to gain one electron what would the new electron arrangement be?

Why is it desirable for chlorine to gain this electron?

Which noble gas does the chloride ion have the same electron arrangement as?

Chlorine has 17 protons and 17 electrons however, if the atom gains an electron it is no longer electrically neutral since it now has 17 protons and 18 electrons. The charged particle formed is called a negative ION and is written Cl-.

Chlorine atom Chlorine ion

7e

8e

2e

17+

**2e**

8e

8e

2e

17+

**2e**

11+

17 protons 17 protons  
17 electrons 18 electrons

Cl 2, 8, 7 Cl -  2, 8, 8

The sodium atom achieves a stable electron arrangement by losing its outer electron and the chlorine atom achieves a stable electron arrangement by gaining the electron lost by the sodium. The formation of sodium chloride can be summed up as follows:

sodium atom + chlorine atom sodium ion + chloride ion

Na + Cl Na+ + Cl -

2, 8, 1 2, 8, 7 2, 8 2, 8, 8

The electrostatic force of attraction between the positive ion and the negative ion is called an **IONIC bond**.

Let’s consider the formation of **magnesium oxide**:

What is the electron arrangement of magnesium?

How many electrons would magnesium need to lose to achieve a stable arrangement?

Which noble gas does the magnesium ion have the same electron arrangement as?

What would the charge on the magnesium ion be?

What is the electron arrangement of oxygen?

How many electrons would oxygen need to gain to achieve a stable arrangement?

Which noble gas does the oxide ion have the same electron arrangement as?

What would the charge on the oxygen ion be?

The formation of magnesium oxide can be summed up as follows:

magnesium atom + oxygen atom magnesium ion + oxide ion

Mg + O Mg2+ + O2-

2, 8, 2 2, 6 2, 8 2, 8

The electrostatic force of attraction between the positive magnesium ion and the negative oxide ion is called an IONIC bond.

The charge which an ion will have depends on which group it belongs to in the Periodic Table. All atoms in group 1 can achieve a stable arrangement by *losing* one outer electron. This means that they all form singly charged positive ions such as Li+ and K+. The following table shows the charges on the ions for Groups 1 to 7.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Group No | 1 | 2 | 3 | 4 | 5 | 6 | 7 |
| Charge on ion | 1+ | 2+ | 3+ | No Charge | 3- | 2- | 1- |

**Note** that the metal elements always form positive ions and the non-metal elements always form negative ions. Elements in group 4 do not usually form ions.

## **Ion - Electron Equations**

The formation of ions can be shown by equations, where e is the symbol for an electron. The equation for the formation of the sodium ion is:

Na Na+ + e -

And for a magnesium ion:

Mg Mg2+ + 2e -

Careful consideration must be given when writing the ion-electron equation of chlorine. Chlorine is a diatomic element; its formula is Cl2. The equation for the formation of chloride ions shows that two Cl- ions are produced from one diatomic molecule:

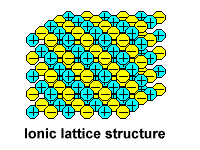
Cl2 + 2e - 2Cl-

Ion-electron equations are given on page 7 of the data booklet.

## **Ionic Lattices**

So far we have looked at the formation of individual ions and the ionic bond that is formed as a result of the electrostatic force of attraction between them. However, these forces of attraction are so strong that millions of ions cluster together. The positive ions surround the negative ions and the negative ions surround the positive ions. This produces a regular, ordered structure called an ionic lattice.

In the sodium chloride lattice, each chloride ion is surrounded by six sodium ions and each sodium ion is surrounded by six chloride ions.



**Structure**

Huge lattice of positively and negatively charged ions held together by forces of attraction called ionic bonds.

**Properties**

Due to the close packing of the ions in the structure, all ionic compounds are SOLIDS at room temperature.

Since thousands of strong ionic bonds must be overcome to change the physical state melting points and boiling points are very high.

Ionic substances conduct electricity when molten or in solution since the ions are no longer fixed in the lattice and are free to move.

**Colours of compounds**

Many ionic compounds are coloured. These compounds dissolve in water to give coloured solutions.

Some ions are colourless, e.g. Na+ and Cl-. In the solid form they appear white due to reflection of light from the crystals; in solution they are colourless.

Some ions are coloured, e.g. Cu2+. For an ionic compound **X**+ve **Y**-ve, the colour of the compound is determined by the colour of **X** and **Y**.

***Complete the following table.***

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Compound** | | **Positive ion** | | **Negative ion** | |
| **Name** | **Colour** | **Name** | **Colour** | **Name** | **Colour** |
| sodium chloride | none | sodium | none | chloride | none |
| potassium sulphate | none |  |  |  |  |
| sodium nitrate | none |  |  |  |  |
| copper chloride | blue |  |  |  |  |
| copper sulphate |  |  |  |  |  |
| copper nitrate |  |  |  |  |  |
| nickel nitrate | green |  |  |  |  |
| nickel sulphate |  |  |  |  |  |
| sodium dichromate | orange |  |  |  |  |
| potassium dichromate |  |  |  |  |  |
| potassium permanganate | purple |  |  |  |  |

Ions are not usually the same colour as the atoms of the element, e.g. copper ions are blue but copper atoms are brown, bromide ions are colourless but bromine molecules are brown.

# Chemical Formulae

When a molecule of hydrogen chloride is formed there is always one hydrogen atom and one chlorine atom in the molecule. In a molecule of water there is always one oxygen atom and two hydrogen atoms. This is true for all molecules; **the number of atoms making up the molecule is always the same.**

The chemical formula of a molecule shows the number of atoms of each element in the molecule. For example, the formula for carbon hydride is **CH4**. This tells us that there are **four hydrogen atoms** bonded to **one carbon atom**. If a molecule contains only one of a certain atom there is no need to put a '1' in the formula. For example, the formula of hydrogen chloride is HCl and not H1Cl1.

## Valency

Fortunately, we don't have to learn the formulae of the many millions of compounds in existence! We only need to learn a simple rule which will allow us to determine the correct formula for any compound. The rule is called the **VALENCY RULE**. The valency tells us how many bonds the atom will form.

The valency of many elements can be determined from their position in the Periodic Table.

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **Group Number** | **1** | **2** | **3** | **4** | **5** | **6** | **7** | **0** |
| **Valency** | **1** | **2** | **3** | **4** | **3** | **2** | **1** | **0** |

This means that all the elements in **group 1** have a **valency of 1** and so on.

Use your data booklet to complete the table below:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Symbol** | **Group** | **Valency** |
| Magnesium |  |  |  |
| Chlorine |  |  |  |
| Phosphorus |  |  |  |
| Sodium |  |  |  |
| Aluminium |  |  |  |
| Oxygen |  |  |  |
| Carbon |  |  |  |

### Valency Rule

Valencies can be used to work out the formulae for compounds using a 5 step method:

**Example 1: Write the formula for magnesium oxide.**

***Step 1*** write the symbols of the elements present. Mg O

***Step 2***  identify their valencies from their group number 2 2

***Step 3*** cross over the valencies Mg2 O2

***Step 4*** cancel any common factor Mg1 O1

***Step 5*** omit ‘1’ if present Mg O

**Example 2: Write the formula for phosphorus fluoride.**

***Step 1*** write the symbols of the elements present. P F

***Step 2***  identify their valencies from their group number 3 1

***Step 3*** cross over the valencies P1 F3

***Step 4*** cancel any common factor (not required in this case)

***Step 5*** omit ‘1’ if present P F3

Using the valency rule write the chemical formulae for the following:

(a) sodium chloride (b) lithium oxide (c) calcium sulphide

(d) hydrogen oxide (e) boron fluoride (f) potassium nitride

(g) aluminium oxide (h) sodium phosphide (i) calcium bromide

## Using Prefixes

If the name of a compound contains a prefix then we do not need to use the valency rule to determine the formula. The name tells us the formula. The following table tells us the meaning of each of the prefixes:

|  |  |
| --- | --- |
| **Prefix** | **Meaning** |
| Mono | one |
| Di | two |
| Tri | three |
| Tetra | four |
| Penta | five |
| Hexa | six |

Examples

Carbon **di**oxide: one carbon atom joined to **two** oxygen atoms.

**CO2**

Phosphorus **penta**chloride: one phosphorus atom joined to **five** chlorine atoms.

**PCl5**

Use the names of the following compounds to determine their formulae:

(a) silicon **di**oxide (b) carbon **tetra**chloride

(c) sulphur **tri**oxide (d) nitrogen **mon**oxide

(e) **di**fluorine **mon**oxide (f) aluminium **tri**chloride

(g) lead **di**oxide (h) **di**nitrogen **tetr**oxide

### Formulae for Ionic Compounds

The formula of an ionic compound shows the **ratio** of ions present and the charge on the ions. The formula of sodium chloride is Na+Cl-. This means that there is one Na+ ion for every Cl- ion.

**How to write ionic formula.**

Use the valency rule to write the simple chemical formula and then add the ionic charges.

|  |  |
| --- | --- |
| **Group number** | 1 2 3 4 5 6 7 |
| **Valency** | 1 2 3 4 3 2 1 |
| **Charge on ion** | 1+ 2+ 3+ 3- 2- 1- |

Here is how the formula for magnesium chloride is worked out:

**Step 1**: symbols Mg Cl

**Step 2**: valencies 2 1

**Step 3**: cross over valencies Mg1 Cl2

The simple chemical formula for magnesium chloride is MgCl2

Now to convert the chemical formula to an ionic formula:

1. Add the charges to the ions, for example magnesium chloride, MgCl2 becomes Mg2+Cl-2.
2. When there is more than one ion of a particular type, it must be put in brackets. For example, the chemical formula for magnesium chloride is MgCl2. The ionic formula is Mg2+(Cl-)2. Note that the 2 is placed outside the bracket.

Write ionic formula for the following compounds:

1. lithium bromide 2. potassium oxide

3. calcium fluoride 4. magnesium sulphide

5. aluminium chloride 6. sodium nitride

7. aluminium sulphide 8. lithium oxide

### Writing ionic formula for compounds containing Transition Metals

The transition metals are found in the Periodic Table between groups two and three. Transitions metals can have ions with more than one charge. The charge is always shown in Roman Numerals after the name.

eg In copper (I) oxide the charge on the copper ion is 1+.

In copper (II) oxide the charge on the copper ion is 2+

The ionic formula for compounds containing transition metals is worked out in exactly the same way.

eg copper (I) oxide copper (II) oxide

(Cu+)2O2- Cu2+O2-

Write ionic formula for each of the following compounds:

1. iron(III)chloride 2. copper(II)sulphide

3. copper(I)chloride 4. iron(II)bromide

5. lead(I)oxide 6. vanadium(V)oxide

## GROUP IONS – Ions containing more than one kind of atom.

So far we have looked at ions containing only one kind of atom, for example the chloride ion, Cl-. However, a number of ions consist of a group of atoms which tend to stay together during reactions. These are known as group ions. The sulphate ion, for example, contains sulphur and oxygen. The formula for the sulphate ion is SO42-. The charge is on the **whole group** and not on any particular ion.

SO42-

The formula of group ions can be found on page 5 of the data booklet.

Complete the following table by writing the formula and charge of each of the group ions.

|  |  |  |
| --- | --- | --- |
| **Group Ion** | **Formula** | **Charge** |
| carbonate | CO32- | 2- |
| nitrate |  |  |
| sulphate |  |  |
| phosphate |  |  |
| hydroxide |  |  |
| sulphite |  |  |
| ammonium |  |  |

The presence of a group ion can usually be recognised from the **–ate** or **–ite** name ending which indicates the presence of oxygen. The exceptions are the ammonium ion and the hydroxide ion.

*Apart from the ammonium ion all group ions have a negative charge and behave in the same way as non-metal ions. Ammonium is the only positively charged group ion and behaves in the same way as metal ions.*

### Writing ionic formula for compounds containing Group Ions

The rules for writing ionic formulae of compounds are exactly the same. Consider the compound potassium sulphate:

**Step 1** symbols K SO4

**Step 2** valency 1 2 note the *valency* is given

by the *charge* on the ion

**Step 3** cross over valency K2 (SO4)1

**note:** always put the formula of the group ion in brackets however, they can then be removed from the formula if there is only one present.

The simple chemical formula for potassium sulphate is K2SO4

Now to convert the chemical formula to an ionic formula:

1. Add the charges to the ions.

2. When there is more than one ion of a particular type, it must be put in brackets.

The ionic formula for potassium sulphate is (K+)2SO42-

Write ionic formula for each of the following compounds:

1. nickel(II) sulphate 2. copper(II) sulphite

3. magnesium hydroxide 4. ammonium chloride

5. sodium phosphate 6. ammonium carbonate

|  |  |  |
| --- | --- | --- |
| Properties of Substances | | |
| 1 | Metallic bonding is the electrostatic force of attraction between positively charged ions and delocalised electrons |  |
| 2 | Metal elements (solids & liquids) and carbon (graphite) are conductors of electricity because they contain free electrons |  |
| 3 | Covalent substances (solids, liquids & solutions) do not conduct electricity since they are made up of molecules which are uncharged |  |
| 4 | 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 |  |
| 5 | Discrete covalent substances have low melting and boiling points due to the weak forces of attraction that need to be overcome |  |
| 6 | Ionic compounds are usually soluble in water |  |
| 7 | Covalent substances which are insoluble in water may dissolve in other solvents |  |
| **Electrolysis** | | |
| 8 | An electric current is a flow of charged particles |  |
| 9 | Electrolysis is the flow of ions through solutions and molten compounds (electrolytes) |  |
| 10 | Electrolysis chemically changes the electrolyte and may lead to the break up of the compound |  |
| 11 | A DC supply must be used if the products of electrolysis are to be identified |  |
| 12 | Positive metal ions gain electrons at the negative electrode and negative non-metal ions lose electrons at the positive electrode |  |
| **Physical States** | |  |
| 13 | Metals, ionic compounds and covalent network substances have high melting and boiling points due to the strong forces of attraction which need to be overcome |  |

# 1. Properties of Substances – Conductivity

When an electric current flows through a substance, it is due to the movement of electrically charged particles.

***Activity 1* Conduction of electricity by Elements**

Set up the circuit shown in the diagram below and test the elements provided to see which ones conduct and put your results in the table below.

element to be bulb  
 tested

|  |  |
| --- | --- |
| Conductors | Insulators |
|  |  |

**Questions**

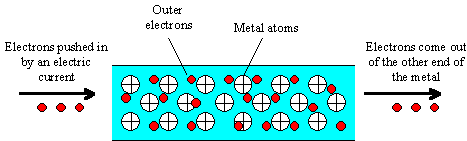
1. What type of elements conduct electricity?

2. What type of elements do not conduct electricity?

3. What is the exception to this rule?

## Why can metals conduct electricity?

In metals, the atoms are packed so close together that the outer electron levels of neighbouring atoms almost touch each other. As a result, the electrons in these outer shells can easily drift from one atom to another. These loose electrons are described as being delocalised which means they are free to move around within the structure. When more electrons are pushed into a piece of metal (by an electric current for example) the electrons just flow through the metal in between the metal atoms. It is this movement of electrons that is responsible for the conduction of electricity.

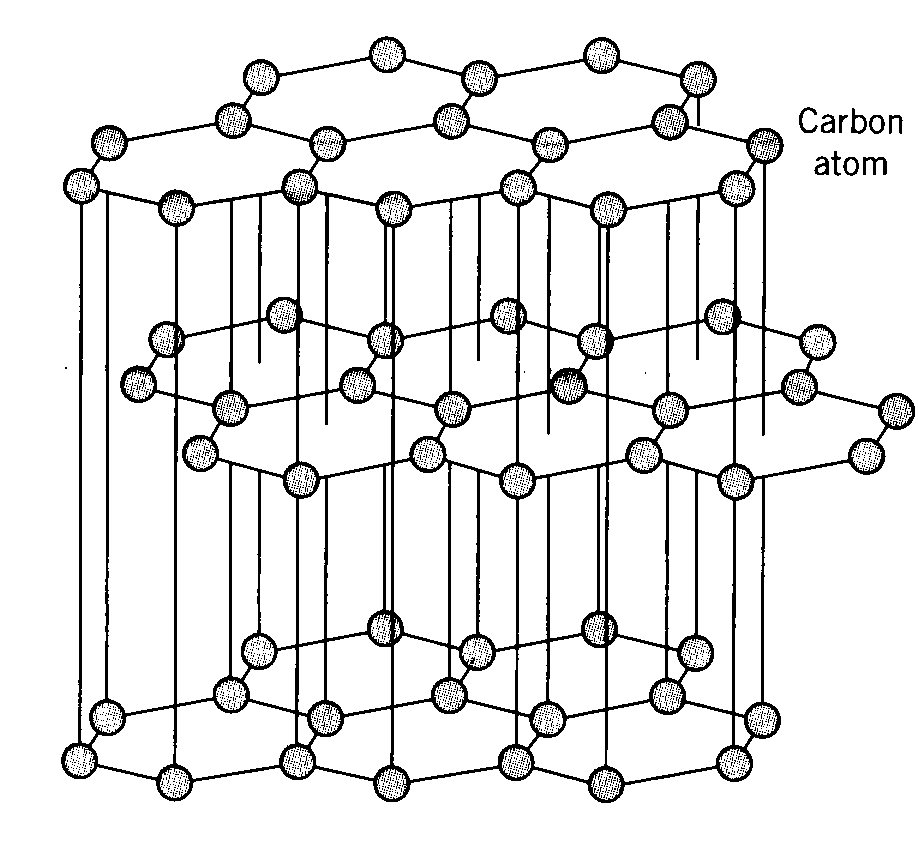


## Why can graphite conduct electricity?

Carbon is an unusual element in that it exists in three different forms; diamond, fullerene and graphite. Carbon has the electron arrangement 2, 4 which means it has four half-filled orbitals in its outer shell.

**C**

This allows carbon to form four covalent bonds since each of the half-filled orbitals can overlap and share electrons. In diamond and fullerene each carbon atom forms four covalent bonds and since all of the outer electrons are involved in bonding these substances do not conduct electricity. In graphite however, only three of the four valence (outer) electrons are involved in bonding leaving the fourth electron free to move throughout the structure. It is this movement of delocalised electrons that allows graphite to conduct electricity.



*When metals* ***and*** *graphite* ***conduct electricity it is*** *electrons* ***that are the charged particles that move through them!***

***Activity 2* Conduction of electricity by Covalent Compounds**

Test the covalent substances provided and put your results in the table below.

|  |  |  |
| --- | --- | --- |
| **Name of Substance** | **State** | **Did it Conduct?** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

**Questions**

1. Why are all the compounds you tested described as being covalent?

2. Did any of them conduct?

## Why do covalent substances not conduct electricity?

Remember, an **electric current** is a **flow of charged particles**. In metals and graphite the charged particles that flow are electrons. There are no such free electrons in covalent substances so they do not conduct electricity.

**Activity 3 Conduction of electricity by Ionic Compounds**

Set up the circuit as shown:

**A**

Ammeter

carbon electrodes

Test each of the solid ionic compounds provided to see if they conduct.

Now complete the table below:

|  |  |  |  |
| --- | --- | --- | --- |
| **Name** | **Formula** | **Does it Conduct?** | |
| **SOLID** | **SOLUTION** |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |

**Question**

What conclusion can you draw about the conductivity of ionic compounds?

Why can ionic substances conduct when in solution but not when solid?

Ionic substances are made of charged particles - positive and negative ions. In the solid state they are held very firmly in place in a **lattice** structure and are not free to move. However, when the ionic solid is dissolved in water, the bonds holding the ions in place in the lattice are broken. The ions can then move around freely. When an electric current is applied to an ionic solution the electricity is carried by the ions that are now able to move. In an ionic solution the electric current is a flow of ions.

# Electrolytes and electrolysis

An **electrolyte** is a liquid which can carry an electric current through it. Ionic solutions are electrolytes. **Electrolysis** describes the process which takes place when an ionic solution has electricity passed through it.

***Activity 4* Electrolysis of copper chloride solution**

Collect a power supply, 2 carbon rods, 2 leads, 2 crocodile clips, a small beaker and a bottle of copper chloride solution.

Set up the diagram as shown:

+

-

copper(II)chloride solution

Set your power supply to 6V D.C. Switch on the power supply for 2 minutes. Watch carefully to see what happens at each electrode.

Complete the table showing what happened at each electrode

|  |  |
| --- | --- |
| **Product at Positive Electrode** | **Product at Negative Electrode** |
|  |  |

**Questions**

Explain why copper was produced at the negative electrode.

Explain why chlorine was produced at the positive electrode.

Why was a D.C. current used?

***Activity 5* Conduction of electricity by Ionic Compounds**

Ionic Compounds also conduct electricity when molten. Draw on the diagram below what would be observed at each electrode if molten copper bromide was electrolysed.

+

-

## Why can ionic substances conduct when molten?

Heat

When the ionic solid is heated, the bonds holding the ions in place in the lattice are broken. The ions can then move around freely. When an electric current is applied to an ionic melt the electricity is carried by the ions that are now able to move. In an ionic melt the electric current is a flow of ions.

# Physical States

***Activity 6* Solids, Liquids or Gases**

The physical state of a substance can give us a clue about the type of bonding present in the substance.

The table noting the type of bonding present and whether the substance is a solid, liquid or a gas at room temperature.

|  |  |  |
| --- | --- | --- |
| ***Substance*** | ***Type of bonding*** | ***Solid/liquid/gas*** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

***Questions***

1. What did you notice about the physical state of ionic compounds at room temperature?

2. What did you notice about the physical state of covalent compounds at room temperature?

**Activity - Testing Melting Points of Ionic and Covalent Substances**

1. Put a spatula full of each solid into a test tube (if test tube has already been prepared– remove stopper).
2. Fill a beaker half way with water and place prepared test tubes of substances into water.
3. Heat beaker of water up on Bunsen burner and observe what substances happens. Record results in table below.

|  |  |  |
| --- | --- | --- |
| **Substance** | **Type of Bonding** | **Did it Melt?** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

***In summary, ionic compounds are always solid at room temperature whereas covalent compounds can be solids, liquids or gases.***

## Solubility

Water is the most common solvent. Some substances, however, do not dissolve in water. This does not necessarily mean that they are insoluble. They may be soluble in non-aqueous solvents.

***Activity 7* Solubility of Ionic and Covalent Compounds**

Collect four test-tubes, a test tube rack, the four substances provided and heptane.

1. Half fill a test tube with water, add a little of one of the substances provided, stopper and shake to see if the substance dissolves.

2. Repeat with the other 3 substances.

3. Now repeat the whole experiment but this time use heptane (non-aqueous) as your solvent instead of water. (ask your teacher how to dispose of the heptane – do not pour it down the sink)

Now complete the table below:

|  |  |  |
| --- | --- | --- |
| **Name of Substance** | **Solubility in Water?**  **(aqueous)** | **Solubility in**  **non-aqueous?** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

**Questions**

1. In which solvent did the ionic compounds dissolve?

2. In which solvent did the covalent compounds dissolve?

**Summary**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Type of Bonding** | **Solubility in Water** | **Melting Point** | **Boiling Point** | **Conductivity** | | |
| **Solid** | **Liquid** | **Solution** |
|  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |

Reaction Quantities and the Mole

REVISION

Earlier in this unit you learned how to write a chemical formula:

* For an element the formula is its symbol (eg Sodium Na), unless it is one of the seven diatomic elements (N2, O2, Cl2, Br2, F2, I2, H2)
* For a compound use valencies to work out the formula, unless the name of the compound tells you the formula (eg sulphur trioxide SO3)

**EQUATIONS**

A **formula** is a chemist’s shorthand way of writing down a chemical element or compound.

An **equation** is a chemist’s shorthand way of writing down what happens in a chemical reaction.

**WORD EQUATIONS**

A **word equation** is simply an equation made up of words.

For example, in a reaction in which a piece of **magnesium** is burned in **oxygen** a white powder called **magnesium oxide** is formed.

The names of the chemicals reacting (**the reactants)** are put on the left, and the chemicals produced **(the products)** are put on the right. An arrow is put in between to show that the reactants are changing into the products i.e.

**magnesium + oxygen magnesium oxide**

**(Reactants) (Product)**

**Exercise:**

Now try writing word equations for the following reactions

Hydrogen burns explosively in oxygen to produce water

............................... + ............................. ...........................

Carbon burns in a plentiful supply of oxygen to give a gas that turns lime water cloudy.

............................... + ............................. ...........................

Iron reacts with sulphur to give iron sulphide.

...................... + ............................. ........................................

**CHEMICAL (FORMULA) EQUATIONS**

A **Chemical (Formula) Equation** uses chemical formulae instead of names for the reactants and products.

So, far the example on the last page where magnesium burns in oxygen to form magnesium oxide you would write the formulae instead of the words i.e.

**magnesium + oxygen magnesium oxide**

*becomes*

**Mg + O2 MgO**

(You may have spotted something a bit odd! There are two atoms of oxygen on the left hand side of the equation and only one on the right – more about that later!!)

**Exercise:**

Now try writing chemical equations for the following reactions:

hydrogen + fluorine hydrogen fluoride

..................... + ................... ..................

sulphur dioxide + oxygen sulphur trioxide

..................... + ................... ..................

aluminium + oxygen aluminium oxide

..................... + ................... ..................

copper(II) oxide + carbon carbon dioxide + copper

..................... + ................... .................. + ................

**BALANCED CHEMICAL EQUATIONS**

A **Balanced Chemical Equation** is an equation which has the same number of atoms of each type on both sides of the arrow i.e.

**Mg + O2 MgO**

*becomes*

**2Mg + O2 2MgO**

**+**

**Exercise:**

Now try balancing the following equations:

**H2  + O2 H2O**

**CH4 + O2 CO2 + H2O**

**Mg + PbBr4 MgBr2 + Pb**

**H2 + N2 NH3**

**Mg + HCl MgCl2 + H2**

**P + O2 P2O5**

**C3H6O + O2 CO2 + H2O**

**FORMULA MASS**

The **Formula Mass** of a substance is the sum of the relative atomic masses of all the atoms shown in the formula.

e.g. calcium sulphate has the formula **CaSO4**

This represents one calcium atom, one sulphur atom and four oxygen atoms

**CaSO4**

Formula mass = 40 + 32 + (16 x4)

= 40 + 32 + 64

= 136

Remember, if there are brackets in the formula, everything inside the brackets is multiplied by the number after the bracket.

e.g. **Mg(OH)2**

Formula mass = 24.5 + (16 + 1)x2

= 24.5 + 17x2

= 24.5 + 34

= 58.5

**Exercise:**

Now calculate the formula mass of each of the following. Set out your working in the same way as in the examples above – do your working in your jotter.

1. CaCO3 f) Mg3P2
2. Na2CO3 g) Ca3(PO4)2
3. Al2S3 h) CaSO4
4. KOH i) (NH4)2CO3
5. C3H8 j) C3H6O

**THE MOLE**

One **Mole** of a substance is its formula mass expressed in grams.

So, to find the mass of **one mole** ofa **substance** work out the formula mass then put the symbol **‘g’** for grams after it i.e.

**CaSO4**

1 mole = 40 + 32 + (16 x4)

= 40 + 32 + 64

= 136g

**Exercise:**

Now calculate the mass of one mole of each of the following. Set out your working in the same way as in the example above – do your working in your jotter.

1. CH4 f) Li3P
2. K2Og) Mg3(PO4)2
3. PbCl4 h) Na2SO4
4. NH3 i) (NH4)2SO3
5. CuSO4 j) Al(NO3)3

**To find the mass of any number of moles of a substance multiply the mass of 1 mole by the required number of moles**

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

So, 2 moles of calcium sulphate (CaSO4) = 2 x 136g = 272g

3.7 moles of calcium sulphate = 3.7 x 136g = 503.2g

0.3 moles of calcium sulphate = 0.3 x 136g = 40.8g

**Exercise:**

Now calculate the mass of each of the following. Set out your working as before in your jotter.

1. 5 moles of ammonia (NH3) d) 0.4 moles of propane (C3H8)
2. 4 moles of carbon dioxide e) 0.3 moles of sodium hydroxide (NaOH)
3. 0.25 moles of water f) 15 moles of ethanol (C2H6O)

**To find how many moles are in a stated mass of a substance, divide the stated mass by the mass of one mole**

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

e.g. How many moles are there in 1292g of calcium sulphate?

**CaSO4**

1 mole = 40 + 32 + (16 x4)

= 40 + 32 + 64

= 136g

Number of moles = mass ÷ mass of one mole

= 1292 ÷ 136

= 9.5 moles

**Exercise:**

Now calculate the number of moles of the given substance in each of the following. Set out your working, as above, in your jotter.

1. How many moles are in 220g of carbon dioxide?
2. How many moles are in 24g of sulphuric acid (H2SO4)?
3. How many moles are in 1.62g of hydrogen bromide?
4. How many moles are in 2g of methane (CH4)?
5. How many moles are in 934.5g of aluminium bromide?

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**Solutions and Concentration**

A solution is made when a solute is dissolved in a solvent:

Solute + Solvent Solution

For the majority of solutions we deal with the **solvent is water**. We call these solutions **aqueous solutions** and show this by using the state symbol **(aq)**

e.g. sodium chloride solution can be written as **NaCl (aq)**

Let’s consider how the above solution is made:

Solute + Solvent Solution

NaCl(s) + H2O(l) NaCl(aq)

**It is useful when working with acid and alkali solutions to know what quantity of the acid or alkali is in solution i.e.**

how much solid is dissolved in what volume of liquid

*measured in* ***moles*** *measured in* ***litres***

The **quantity of a solute in a certain volume of solution** is called the **concentration** of the solution. We can therefore see that the concentration of a solution must be measured in **moles per litre**. The symbol for this is **mol l-1**

**Standard Solutions**

Your teacher will show you how to make aStandard Solution**.**

A Standard Solution is one which is made up accurately by dissolving a known mass of solute in a known volume of solvent (water).

Now complete the diagram below showing how to make a STANDARD SOLUTION.

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

A = weigh out 1 mole of the compound

B =

C = transfer the solution to the 1 litre standard flask

D =

E =

**More Calculations**

So far we have looked at calculations involving only solids. What about solutions? We now know how to make up a solution accurately and understand why the units of concentration are moles per litre (moll-1). If we apply this mathematically we find:

**Concentration = number of moles of solid**

**Volume of solvent (water) in litres**

Or, using symbols **C** = concentration

**V** = volume of solvent in litres

**n** = number of moles of solid

we find that:

|  |  |  |
| --- | --- | --- |
| **C =** | **n** |  |
| **V** |  |

We can use this equation in calculations to find the concentration of solutions.

**Example:**

Calculate the concentration of a potassium hydroxide solution containing 0.25moles of potassium hydroxide in 500ml of solution.

*First we need to note what information the question is giving us.*

*0.25 moles = n*

*500ml = 0.5 litres = V*

*Since we know two parts of the above equation we can now use it to find the answer to our problem.*

*C = n = 0.25 = 0.5 moll-1*

*V 0.5*

**Exercise:**

Calculate the concentration of each of the following solutions:

* 1. 1 litre of solution containing 2 moles of solute
  2. 500ml of solution containing 1 mole of solute
  3. 100ml of solution containing 0.25moles of solute

**We are now familiar with the “solutions” equation:** C = n

V

We can rearrange this equation to give two other equations which will help us in other calculations.

We can calculate the number of moles of solute in a solution using:

n = C x V

ORcalculate the volume of solvent we require to make a solution of a specific concentration using:

V = n

C

NOTE: The volume must always be in litres in these equations.

If it is not given in litres then it must be changed to litres

To change *ml* to *litres* we divide by 1000.

It is also important to realise that *1ml* is exactly the same as *1 cm3* .

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**Let’s look at a few examples of how to use these equations:**

**Example1**

Calculate the number of moles of sodium chloride required to make 250ml of a solution with a concentration of 2moll-1

*We have been given the following information in the question:*

*V = 250 ml = 0.25 litres*

*C = 2moll-1*

*We want to find the number of moles of sodium chloride so we must use the following equation:*

*n = C x V*

*n = 2 X 0.25 = 0.5 moles*

**Example 2**

Calculate the concentration of a solution of sodium chloride which contains 3 moles of solute dissolved in 250 ml of solution.

*We have been given the following information in the question:*

*The no. of moles, n = 3*

*The volume = 250ml BUT we must change this into LITRES*

*Therefore V= 250/1000 = 0.25 litres*

*We want to find the concentration so we must use the equation:*

*C = n*

*V*

*C = 3 / 0.25 = 12 mol l-1*

**Example 3**

A 0.1mol l-1 solution of hydrochloric acid contains 0.5 moles of solute.

What is the volume of the solution?

*We have been given the following information in the question:*

*C= 0.1 mol l-1*

*n = 0.5 moles*

*We have to calculate the volume so we must use the following equation:*

*V = n*

*C*

*V = 0.5 / 0.1*

*= 5 litres*

**Exercise:**

1. Calculate the concentration of each of the following in moll-1

1. 2 litres of solution containing 3 moles of solute
2. 200 ml of solution containing 0.5 moles of solute
3. 500 ml of solution containing 3 moles of solute

2. Calculate the number of moles of solute in each of the following:

1. 5 litres of a solution with a concentration of 0.25 moll-1
2. 250 ml of a solution with a concentration of 0.125 moll-1
3. 15 ml of a solution with a concentration of 5 moll-1

3. Calculate the volume of solution which could be made using:

1. 2 moles of solute in a solution of 0.8 moll-1 concentration
2. 4 moles of solute in a solution of 0.25 moll-1 concentration
3. 0.2 moles of solute in a solution of 1.5 moll-1 concentration

**Learning Outcomes Acids & Bases**

The pH scale

* State that acids have a greater number of hydrogen ions (H+(aq)) than hydroxide ions (OH-(aq)).
* State that alkalis have a greater number of hydroxide ions (OH-(aq)) than hydrogen ions (H+(aq)).
* State that in neutral solutions the number of hydrogen ions is equal to the number of hydroxide ions. H+(aq) = OH-(aq)
* Explain that when an acid/alkali is diluted, the concentration of H+(aq)/OH-(aq) decreases and the pH moves towards 7.

Reactions of Acids

* State the products of the following neutralisations;

acid and alkali(metal hydroxide), acid and metal oxide, acid and metal carbonate, acid and metal, and write balanced chemical equations for them, including state symbols and identifying spectator ions.

METAL OXIDE + ACID → SALT + WATER

METAL HYDROXIDE + ACID → SALT + WATER

METAL CARBONATE + ACID → SALT + WATER + CARBON DIOXIDE GAS

METAL + ACID → SALT + HYDROGEN

Volumetric Titrations

* Calculate unknown volumes or concentrations by titrations with acids or alkalis of known concentration.

Ionic Equations

* Rewrite balanced ionic equations for salt formation omitting spectator ions.

**Acids and alkalis**

What are acids and alkalis? How do we tell them apart? What are they used for?

Acids and alkalis are very common substances which are found in the home, in industry and in our bodies. Chemists need to be able to recognise the differences between acids and alkalis and be aware of their typical chemical reactions.

#### **Types of Solution**

All solutions in water can be **acidic**, **alkaline** or **neutral**.

We can tell them apart by using special **indicators**, which change colour according to the type of solution it has been added to. The indicator which we have used most often in science is called **universal indicator**. This indicator has a range of colours for both acids and alkalis. It is universal indicator which is soaked into paper to make **pH paper**.

#### **The pH Scale**

**All acids and alkalis are solutions**.

The pH scale measures how **acidic** **or alkaline** a solution is.

The scale runs from **below 0** to **14 and above**.

To find out which group a chemical belongs to **pH** **paper** or **universal indicator solution** is used.

##### **Activity 1: The pH Scale**

##### 

1. Put 1 cm depth of one of the solutions in a test tube.
2. Add 3 drops of universal indicator solution.
3. Note what colour the indicator turns and check the pH number against the colour chart.
4. Record your results in the table below.
5. Repeat steps 1 – 4 for all the other solutions in the kit.

|  |  |  |  |
| --- | --- | --- | --- |
| **Solution** | **Colour with Indicator** | **pH Number** | **Acid/Alkali/Neutral** |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |

**Summary**

All \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ have a **pH >7** and turn indicator solution \_\_\_\_\_\_\_\_\_\_\_\_

All \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ have a **pH < 7** and turn indicator solution \_\_\_\_\_\_\_\_\_\_\_\_

If a solution has a **pH =7** and turns indicator solution **green** it is a \_\_\_\_\_\_\_\_\_\_solution.

## Acids in Everyday life

Every day you will come across, and even drink, several different acids. Fizzy drinks contain carbonic acid. Vinegar is an acid. Even inside your mouth, bacteria change sugar into acids which can cause tooth decay. The diagram below shows a number of different substances which you encounter daily. Colour the diagram to show the correct range of colours in the pH scale.

**8**

n

C

V

**E**

**D**

**C**

**B**

**A**

Sample

mass

No of moles

Mass of 1 mole

**Number of moles = mass ÷ mass of one mole**

**Mass = number of moles x mass of one mole**

**Mg**

**Mg**

**O**

**O**

**Mg**

**Mg**

dilute hydrochloric acid

**9**

**10**

**11**

**12**

**13**

**14**

**1**

**6**

**5**

**4**

**2**

****

**7**

lemon juice

vinegar

water

salt solution

ammonia solution

milk of magnesia

14

13

12

11

10

9

8

7

6

5

4

3

2

1

Now complete the table below to show how these substances can be classified.

|  |  |  |
| --- | --- | --- |
| **Acid** | **Neutral** | **Alkali** |
|  |  |  |
|  |  |  |

**Making Acids**

When an **element** reacts with **OXYGEN** it forms an **OXIDE.** Many non-metals burn in oxygen to form **non-metal** oxides

e.g. Carbon burns to give carbon dioxide:

**C + O2 CO2**

If a **non-metal oxide** dissolves in water it will form a solution**.**

Watch while your teacher shows you how to make a solution from a non metal.

Write a **brief description** of how each experiment was carried out.

**Activity 2** **Burning carbon**.

**Activity 3** **Burning sulphur**

**Activity 4 Dissolving nitrogen dioxide in water.**

**Questions**

1. How did you know that the carbon and the sulphur were reacting with the oxygen in each gas jar?

2. What gas is formed when carbon burns in oxygen? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

## **3. What gas is formed when sulphur burns in oxygen? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

## **4. Are the gases formed in each experiment soluble or insoluble in water?**

5.When sulphur dioxide dissolves in water it forms sulphurous acid (H2SO3).

Write a word equation and a formula equation for this reaction.

**Complete the table below**

|  |  |  |  |
| --- | --- | --- | --- |
| **NON-METAL OXIDE** | **INDICATOR COLOUR** | **pH NUMBER** | **TYPE OF SOLUTION** |
| **Carbon dioxide** |  |  |  |
| **Sulphur dioxide** |  |  |  |
| **Nitrogen dioxide** |  |  |  |

**Acids are solutions which are formed when**

**\_\_\_\_\_ -\_\_\_\_\_\_\_\_\_ oxides dissolve in water.**

**Complete the following table and LEARN the formulae for the acids**

|  |  |
| --- | --- |
| Acid | Formula |
| Hydrochloric acid |  |
|  | HNO3 |
| Sulphuric acid |  |



**Question**

## **What element is common to** ALL **acids?**

## Acids in Everyday Life - Acid Rain

Soluble non-metal oxides are called **acidic oxides** as they form acids when they dissolve in water. As rainwater contains dissolved carbon dioxide from the atmosphere, this explains why it is naturally acidic. In some areas, however, other non-metal oxides like **sulphur dioxide** and **nitrogen dioxide** are present in the air. These oxides are mainly produced by the burning of fossil fuels and they dissolve in rainwater making it even more acidic. Some of the effects of acid rain are caused by the reactions of acids e.g.

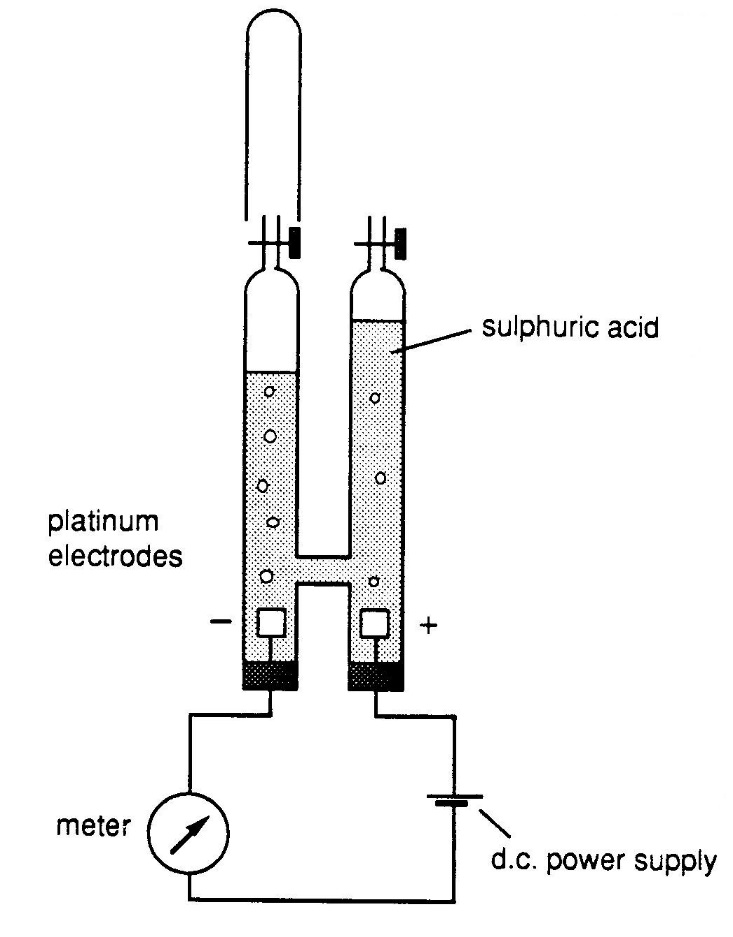
1. Acid rain reacts with metal oxides and so washes these essential minerals out of the soil. It also alters the pH of rivers and lakes. All of this affects plant growth and life in the rivers and lakes. Many lakes in Scandinavia, and some lochs in Scotland, now contain hardly any fish.
2. Acid rain damages buildings made of limestone and marble. These stones are mostly calcium carbonate and acids react with all metal carbonates.
3. Acid damages iron and steel objects like railings and car bodies, as acids react with most metals.

**Electrolysis of Acids**

In **Unit 1.6** you saw that ionic solutions can be **electrolysed** (broken down into elements using electricity). In this experiment we are going to pass electricity through an acid and collect any gas which may be given off at the **negative** electrode.

**Activity 5**

**** Your teacher will demonstrate the experiment below.

****

**Summary**

When electricity is passed through an ionic solution this is called \_\_\_\_\_\_\_\_\_\_\_\_\_.

When an acid is electrolysed \_\_\_\_\_\_\_\_\_\_\_\_\_\_ is produced at the negative electrode. This proves that acids conduct electricity and therefore must contain .

Since the hydrogen gas was produced at the negative electrode the hydrogen ions must have a \_\_\_\_\_\_\_\_\_\_\_\_ charge. Hydrogen gas burns with a \_\_\_\_\_!

**Making Alkalis**

Many **metals** burn in oxygen to form **metal oxides**.

e.g. Magnesium + oxygen magnesium oxide

Some **metal oxides** will react with **water** to produce substances called **metal hydroxides:**

sodium + water sodium hydroxide

When a metal oxide or hydroxide dissolves in water, an **alkaline solution** is produced.

At the start of this topic you found out that some solutions have a **pH greater than 7.** These solutions are called **ALKALIS.**

**Activity 6**

Watch while your teacher shows you how to make some alkalis.

Record your observations in the table below.

|  |  |  |  |
| --- | --- | --- | --- |
| **Name of metal oxide/hydroxide** | **Soluble/ insoluble** | **Appearance of Oxide** | **Effect of solution in indicator** |
|  |  |  |  |

**Activity 7** **– Making an Alkali**



magnesium

tongs

HEAT

burning magnesium

universal indicator

magnesium ash

1. ignite a little 2. When lit remove 3. Drop the white ash

magnesium ribbon from flame and into a dimple-tile.

in a bunsen flame. allow to burn. Check the pH of its

solution in water.

**Now complete the following**

**Metal + oxygen metal oxide**

**Mg** **+** \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_

**Metal oxide + water metal hydroxide**

**MgO** **+** ­­­\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_

**Questions**

What type of solutions are formed when metal oxides react with water?

What other kind of compound dissolves in water to form an alkali?

**Do All Metal Oxides Dissolve In Water To Form an Alkali Solution?**

**NO** **- not all metal oxides dissolve in water to form alkalis.** The table of solubilities on page 5 of your data book indicates which metal oxides dissolve. Use this table to place the metal oxides below in the correct column.

**(a)** sodium oxide **(b)** copper (II) oxide **(c)** iron (III) oxide **(d)** potassium oxide

|  |  |
| --- | --- |
| **Soluble Oxides forming Alkali Solutions** | **Insoluble Oxides** |
|  |  |

**Common Alkalis:**

| **Alkali** | **Formula** | **Ions Present** |
| --- | --- | --- |
| sodium hydroxide  calcium hydroxide  magnesium hydroxide | KOH | Na**+** and OH**-**  Ca**2+** and OH**-** |



**Question**

What ion do all alkalis have in common? **Activity 8 - The pH Scale**

Prepare the 8 solutions of acid (or alkali) as shown in the table below.

Add a few drops of Universal Indicator to each solution and note the colour and pH number in the table.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  |  | **ACID** | |  | **ALKALI** | |
| **TUBE** | **SOLUTION** | **COLOUR** | **pH** |  | **COLOUR** | **pH** |
| **1** | **10cm3 of the acid or alkali** |  |  |  |  |  |
| **2** | **1cm3 from tube 1 + 9cm3 water** |  |  |  |  |  |
| **3** | **1cm3 from tube 2 + 9cm3 water** |  |  |  |  |  |
| **4** | **1cm3 from tube 3 + 9cm3 water** |  |  |  |  |  |
| **5** | **1cm3 from tube 4 + 9cm3 water** |  |  |  |  |  |
| **6** | **1cm3 from tube 5 + 9cm3 water** |  |  |  |  |  |
| **7** | **1cm3 from tube 6 + 9cm3 water** |  |  |  |  |  |
| **8** | **1cm3 from tube 7 + 9cm3 water** |  |  |  |  |  |

As an acidic solution is diluted the pH of the solution\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

As an alkaline solution is diluted the pH of the solution\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

When 1cm3 of an acid is diluted to 100cm3, the \_\_\_\_\_\_ ions will be more spread out.

This means the concentration of **H+(aq)** ions is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ as we dilute an acid with water until we eventually reach pH7.

Similarly, diluting an alkali reduces the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ of **OH** (aq) until the pH reaches 7.

Complete the chart at the top of the next page.

**8**

**9**

**10**

**11**

**12**

**13**

**14**

**1**

**6**

**5**

**4**

**2**

**3**

**7**

Increasing

Alkalinity

Increasing

Concentration

14

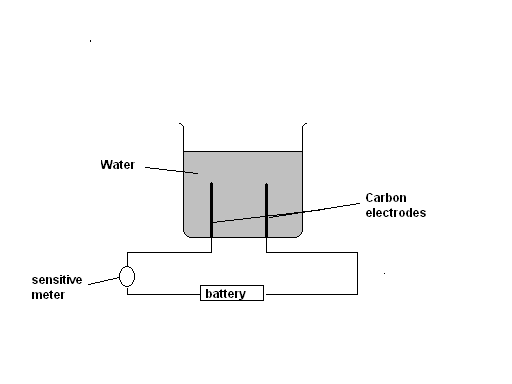
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7

**j0186164**

**Does water contain ions?**

**Activity 9** - Electrolysing Water

Your teacher will demonstrate the following experiment

**Questions**

1. What visible signs are there that a reaction is taking place?

2. Is there a reading on the meter? If so, note the reading here.

3. What type of bonds does water contain?

4. What is Electrolysis?

5. Would you expect water to be able to conduct electricity?

6. From the results of your experiment does water conduct electricity?

7. What kind of particles must be present in water to allow it to conduct electricity?

For water to conduct electricity it must contain **IONS**. From your work in Unit 1.6 you discovered that ionic compounds can be broken down by electricity . We called this **ELECTROLYSIS**. The ions present in ionic compounds allowed the substance to conduct electricity as a solution or a liquid. Our experiment has shown that water conducts electricity. This means that **water must contain some ions** despite the fact that it is a covalent compound. The small reading on the ammeter suggests that there are only a very small number of ions present.

**Where do the ions come from?**

Pure water is a neutral liquid. It is made up almost entirely from covalent molecules. The result of the above experiment suggests that ions must be present.

Water does contain a very small concentration of ions. There is about one ion for every 250 million water molecules. The ions come from the water molecules themselves.

Read on to find out how ….

A water molecule can split up to form a hydrogen ion and a hydroxide ion as shown in the equation:

**H2O H+(aq) + OH-(aq)**

**H+** and **OH-** ions are found in acid and alkali solutions. Water, however, is a neutral liquid with a pH of 7. This is because **water contains exactly the same number of H+ and OH- ions**.

Acid solutions have a pH< 7 and therefore contain more hydrogen ions than water. Technically speaking we should say that **acid solutions have a greater concentration of H+ ions than water.** We can also state that **alkali solutions have a pH > 7 and have a greater concentration of OH- ions than water.**

It is also very important to remember that both acids and alkalis contain H+ and OH- ions since they are solutions which have been made using water as the solvent. Their pH is different because they do not have equal concentrations of each ion. **In acids the concentration of H+ ion is greater than the concentration of OH-. In alkalis the concentration of OH- is greater than H+.**

**Summary**

|  |  |  |
| --- | --- | --- |
| **Type of solution** | **Comparison of H+ and OH- Concentration** | **Effect of Dilution on pH** |
| **Water and neutral solutions** | **H+ concentration is**  **\_\_\_\_\_\_\_\_\_\_ \_\_\_**  **OH- concentration** | **No effect.**  **pH remains at 7** |
| **Acid solutions** | **H+ concentration is**  **\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_**  **OH- concentration.** | **pH moves towards \_.**  **Concentration of H+**  **\_\_\_.** |
| **Alkali solutions** | **H+ concentration is**  **\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_**  **OH- concentration.** | **pH moves towards \_.**  **Concentration of OH-**  **\_\_\_\_\_\_\_\_\_\_\_\_\_\_.** |

**Neutralisation**

Acid rain forms when non-metal oxides such as carbon dioxide and nitrogen dioxide dissolve in the moisture in the atmosphere. Fish, plants and animals in lakes are dying because of the increased acidity in the water as a result of acid rain. Lime is added to the lakes to try to reduce the acidity of the water. Lime is a neutraliser.

Having too much acid in the stomach causes indigestion. Indigestion tablets, like Rennies, contain neutralisers such as calcium carbonate. These neutralise the stomach acid and so relieve the indigestion.

#### **Bases and Alkalis**

A **base** is a substance that **neutralises** an **acid** producing **water**.

Bases include **metal hydroxides**, **metal carbonates**, **metal oxides** and **ammonia**. Alkalis are a type of base. An **alkali** is a **soluble base**. The carbonates, oxides and hydroxides of all group 1 metals are dissolved easily in water forming alkaline solutions.

In this part of the unit we will look at the neutralisation of acids by three different bases and also by metals.

# Neutralising Acids with Alkalis

Earlier we learned that an acid could be neutralised by an alkali. During this neutralisation process the **OH-** ions from the alkali combine with the **H+** ions from the acid producing **water**. A **salt** is also produced in the reaction.

**Acid + Alkali Salt + Water**

## Naming the salt

The chemical name for a salt is derived from the **negative ion** from the **acid** and the **positive ion** from the **neutraliser**.

For example **hydrochloric acid** produces **chloride** salts and **potassium hydroxide** produces **potassium salts**. So, neutralising hydrochloric acid with potassium hydroxide produces **potassium chloride** salt.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Acid** | **Salt** |  | **Alkali** | **Salt** |
| Hydrochloric acid | … chloride |  | Sodium hydroxide | Sodium … |
| Nitric acid | …nitrate |  | Potassium hydroxide | Potassium … |
| Sulphuric acid | …sulphate |  | Calcium hydroxide | Calcium … |

****

**Questions**

1. Write a word equation and a chemical equation for the neutralisation of nitric acid with sodium hydroxide.

+ +

+ +

1. Write a word equation and a chemical equation for the neutralisation of sulphuric acid with calcium hydroxide.

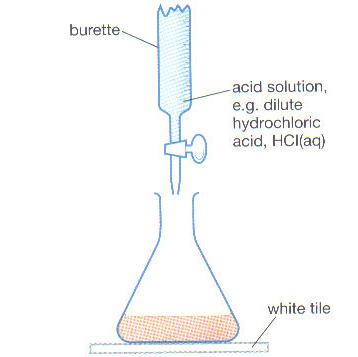
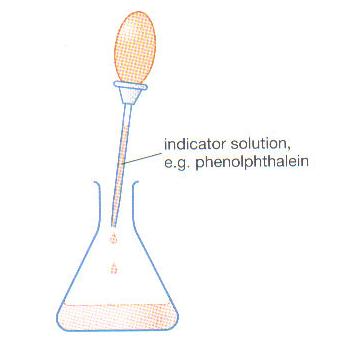
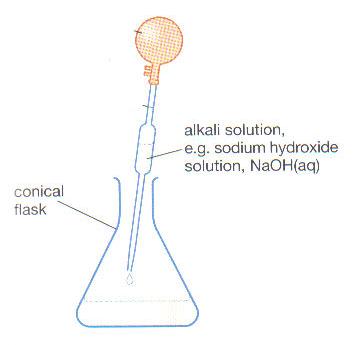
+ +

+ +



## Activity 10

Watch whilst your teacher carries out a titration to neutralise hydrochloric acid with sodium hydroxide.





## Questions

1. Name the salt produced in the neutralisation reaction.

2. Write a word equation for the reaction.

+ +

3. Write a chemical equation for the reaction.

+ +

4. Explain why phenolphthalein indicator was used during the titration.

5. The salt produced is a soluble salt. How can the salt be obtained from the solution?

### *Neutralising Acids with Metal Carbonates*

Many rocks used for building, like limestone, contain carbonates. When acid rain falls on this type of rock the carbonate reacts with the acid and the building gradually wears away.

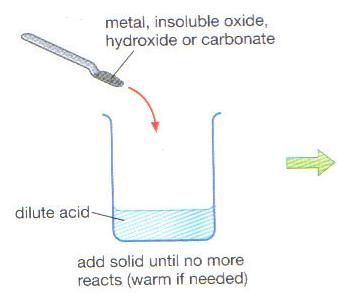
**Metal carbonates** neutralise **acids** to form a **salt**, **water** and **carbon dioxide gas**. The salt is named in exactly the same way as before ie, the **positive ion** from the **neutraliser** combines with the **negative ion** of the **acid.**

**Acid + metal carbonate salt + water + carbon dioxide**

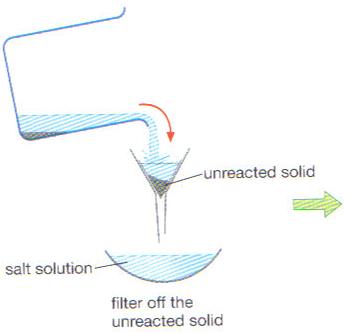
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## Activity 11

Carry out the following experiment neutralise an acid with a metal carbonate.



copper carbonate



On cooling, salt crystals form



## Questions

1. Name the salt produced in the neutralisation reaction.

2. Write a word equation for the reaction.

+ + +

3. Write a chemical equation for the reaction.

+ + +

4. How did you know that the reaction had stopped and that the acid had been neutralised?

5. How was the excess copper carbonate removed from the reaction mixture?

6. How was the soluble salt be obtained from the solution?

### *Neutralising Acids with Metal Oxides*

Gardeners and farmers often have soil, which is too acidic for their crops to grow properly. They use lime to neutralise the acidity in the soil. Lime is a metal oxide. Its chemical name is calcium oxide. Like all metal oxides it will react with acids to form salt and water.

The salt is named in exactly the same way as before ie, the **positive ion** from the **neutraliser** combines with the **negative ion** of the **acid.**

**Acid + metal oxide salt + water**

## Activity 12

**What to do:**

1. Add sulphuric acid to a boiling-tube until it is about one third full.

2. Add a spatula of copper oxide and gently shake the test-tube.

3. Add copper oxide until no more dissolves.

4. Pour the mixture through filter paper into an evaporating basin.

5. Leave the salt solution to evaporate overnight.

## Questions

1. Name the salt produced in the neutralisation reaction.

2. Write a word equation for the reaction.

+ +

3. Write a chemical equation for the reaction.

+ +

4. How did you know that the reaction had stopped and that the acid had been neutralised?

5. How was the excess copper oxide removed from the reaction mixture?

6. How can the soluble salt be obtained from the solution?

### *Neutralising Acids with Metals*

Acid rain attacks iron. If a bridge or a car or any structure made of iron is not protected by paint or plastic or any other coating then the acid rain falling on it will react with the iron. This is another example of a **neutralisation** reaction. In this reaction the **metal** combines with the **negative ion** from the acid and the **hydrogen ions** in the acid solution are converted into **hydrogen gas**.

***Acid + Metal Salt + Hydrogen***

**NOTE:** Not all metals react with acids however, it is easy to remember the ones that will. They are called the MAZIT metals ie, magnesium, aluminium, zinc, iron & tin.

***Activity 13***

1. Add hydrochloric acid to the test-tube until about a quarter full.

2. Add a piece of magnesium ribbon to the acid and collect the gas produced in an upside-down test tube.

3. Once the reaction is over, carry out the chemical test for hydrogen on the gas you have collected.

## Questions

1. Name the salt produced in the neutralisation reaction.

2. Write a word equation for the reaction.

+ +

3. Write a chemical equation for the reaction.

+ +

4. How did you know that the reaction had stopped?

5a. In your experiment all the magnesium was used up.

How could you change the experiment to ensure that all the ACID had been neutralised?

5b. How could the soluble salt then be obtained from the solution?

In each of the four examples above where an acid was neutralised the salt produced was always a soluble salt and could be obtained from the solution by evaporation.

Not all salts are soluble. **Insoluble salts** can be prepared by **Precipitation**.

### *Preparation of Insoluble Salts by Precipitation*

An insoluble salt is one where less than 1g of the salt can be dissolved in 100g of water at 20oC. We can make insoluble salts by precipitation.

A precipitation reaction is the reaction of two solutions to form an insoluble product called a precipitate. For example:

**Silver nitrate + sodium iodide silver iodide + sodium nitrate**

The **silver iodide** is the **insoluble salt** whereas the sodium nitrate is soluble and could be obtained by evaporation in the usual way.

The ‘solubility’ table on page 7 of the data booklet should be used to determine insoluble salts.

## Activity 14

1. Measure 10cm3 of copper nitrate solution into a small beaker.

2. Add 10cm3 of potassium carbonate and stir.

3. Pour the mixture through filter paper into an evaporating basin.

## Questions

1. Use the solubility table in the data booklet to determine the insoluble salt in the precipitation reaction.

2. Write a word equation for the reaction.

The method selected to make a salt depends on the salt itself. The following flow chart can be used to determine which method should be used to make salts.

# Summary on Making Salts

**Is the salt soluble or insoluble?**

**Soluble?**

**Insoluble?**

**Prepare by precipitation**

**Use solubility table on p7 of data book**

**Neutralise an acid**

**Which acid?**

**Which neutraliser?**

**Sulphate**

salt

**Nitrate**

salt

**Chloride**

salt

**Hydrochloric Acid**

acid

**Nitric Acid**

acid

**Sulphuric Acid**

acid

**Does metal have a soluble hydroxide?**

**Yes**

B

**Use metal hydroxide as neutraliser**

**Use metal oxide or metal carbonate.**

**No**

****

**Activity 15 – Preparing a salt**

**Your task is to prepare a dry sample of a salt and produce a written report (which should take the form of an A3 or A2 poster) with diagrams showing how you prepared your salt sample.**

**Your teacher will tell you which salt your group has to make.**

**Use the flow chart on the previous page, your data book, and one or more of the following references, to help you decide the correct method to use and the chemicals you will require.**

**Standard Grade Chemistry (Green Book) Pages 90-91**

**Chemistry Counts (Red Book) Pages 159-161**

**BEFORE STARTING ANY PRACTICAL WORK YOU MUST GET YOUR TEACHER’S APPROVAL**

#### **Spectator Ions**

#### **Acids + Alkalis**

#### We found out earlier that acidic solutions contain an excess of hydrogen ions, H+ and alkalis contain an excess of hydroxide ions, OH-. During neutralisation these two ions react to form water.

#### **H+ + OH- H2O**

#### The other ions in the reaction do not react. Using an ionic equation where all the ions are shown separately can show this.

#### Consider the neutralisation reaction between hydrochloric acid and sodium hydroxide:

#### **H+(aq) + Cl-(aq) + Na+(aq) + OH-(aq) Na+(aq) + Cl-(aq) + H2O**

#### You can see that the **Na+** and the **Cl-** ions are not changed. They appear in exactly the same form on both sides of the equation. We call ions such as these, which do not take part in the chemical reaction, **spectator ions**.

#### If we remove the spectator ions from the equation then the **ESSENTIAL** equation representing the changes that took place is:

#### **H+(aq) + OH-(aq) H2O(l)**

#### For the following examples

#### write a word equation

#### write an ionic equation with the ions separated out.

#### identify the spectator ions by circling them

#### 4. write the essential equation

#### (a) Nitric acid and Potassium hydroxide

#### (b) Sulphuric acid and calcium hydroxide

#### **Acids + Metal Carbonates**

#### We found out earlier that when an acid is neutralised by a metal carbonate salt, water and carbon dioxide are produced.

#### **2HCl(aq) + CaCO3(s) CaCl2(aq) + H2O(l) + CO2(g)**

#### If we write the ionic equation again separating out the ions we can see that the actual reaction involves only the **H+ ions** from the acid and the **CO32- ions** from the neutraliser. The other ions in the reaction do not react.

#### **2H+(aq) + 2Cl-(aq) + Ca2+CO32-(s) Ca2+(aq) + 2Cl-(aq) + H2O + CO2(g)**

#### You can see that the **Ca2+** and the **Cl-** ions are not changed. They appear in exactly the same form on both sides of the equation. We call ions such as these, which do not take part in the chemical reaction, **spectator ions.**

#### If we remove the spectator ions from the equation then the **ESSENTIAL** equation representing the changes that took place is:

#### **2H+(aq) + CO32-(s) H2O(l) + CO2(g)**

#### 

#### For the following examples

#### write a word equation

#### write an ionic equation with the ions separated out.

#### identify the spectator ions by circling them

#### 4. write the essential equation

#### (a) Nitric acid and magnesium carbonate

#### (b) Hydrochloric acid and lithium carbonate

#### **Acids + Metal**

#### We know from our general level knowledge that when an acid is neutralised by a MAZIT metal salt and hydrogen are produced.

#### **2HCl(aq) + Mg(s) MgCl2(aq) + H2(g)**

#### If we write the ionic equation again separating out the ions we can see that the actual reaction involves only the **H+ ions** from the acid and the metal. The other ion in the reaction does not react.

#### **2H+(aq) + 2Cl-(aq) + Mg(s) Mg2+(aq) + 2Cl-(aq) + H2(g)**

#### You can see that the **Cl-** ion is not changed. It appears in exactly the same form on both sides of the equation. We call ions such as these, which do not take part in the chemical reaction, **spectator ions**.

#### If we remove the spectator ions from the equation then the **ESSENTIAL** equation representing the changes that took place is:

#### **2H+(aq) + Mg(s) Mg2+(aq) + H2(g)**

#### For the following examples

#### write a word equation

#### write an ionic equation with the ions separated out.

#### identify the spectator ions by circling them

#### write the essential equation

#### (a) Nitric acid and Zinc

#### (b) Hydrochloric acid and Aluminium

#### **Top tip:** if you find it difficult to identify spectator ions just learn that the spectators are always the ions in the soluble salt (in the reaction of a metal it is just the negative ion that is the spectator)

#### **Titrations**

#### Earlier in this topic your teacher demonstrated how to neutralise an acid by an alkali by titration.

#### A **titration** is a very accurate technique which allows you to determine the concentration of a solution.

**Activity 16** Your teacher will let you carry out a titration.

#### Usually, but not always, a titration involves a neutralisation reaction between an acid and an alkali where the concentration of one solution is known, and the concentration of the other can be found using the results of the titration.

#### **Calculations Based on Titrations**

#### The following equation can be rearranged to calculate the unknown concentration of a reactant in a titration.

#### **Cac x Vac = Cal x Val**

**nac nal**

#### **Where:** **Cac** is the **concentration of acid**

#### **Vac** is the **volume of acid**

#### **nac** is the **number of moles of acid in the equation**

#### and similarly **Cal , Val , nal** refer to the concentration, volume and number of moles of alkali in the equation.

#### **Worked example**

#### **In a titration, 20cm3 of sodium hydroxide solution with a concentration of 0.2moll-1 was neutralised by 25cm3 of dilute sulphuric acid. Calculate the concentration of the acid solution in moll-1.**

#### **Step 1** Write the balanced equation for the reaction.

#### **H2SO4 + 2 NaOH Na2SO4 + 2 H2O**

#### **Step 2** Put values into the formula for everything you have been given in the question.

#### **Cac x Vac** **=** **Cal x Val**

**nac nal**

#### 

so, **Cac x 25 = 0.2 x 20**

**1 2**

so, **25 Cac = 2**

so, **Cac = 2 / 25**

**= 0.08 moll-1**

#### **calculatorQuestions**

#### 25cm3 of 0.2moll-1 sodium hydroxide solution was neutralised by 16cm3 of hydrochloric acid. Calculate the concentration of the acid.

#### 10cm3 of potassium hydroxide solution was neutralised by 12cm3 of 0.1moll-1 nitric acid. Calculate the concentration of the potassium hydroxide solution.

#### 20cm3 of 0.5moll-1 lithium hydroxide solution was neutralised by 30cm3 of sulphuric acid. Calculate the concentration of the sulphuric acid.

#### 20cm3 of sodium hydroxide was neutralised by 20cm3 of 0.2moll-1 sulphuric acid. Calculate the concentration of the sodium hydroxide solution.

1. What volume of hydrochloric acid (concentration 0.1 mol/l) is required to neutralise 100cm3 of sodium hydroxide solution (concentration 0.5 mol/l)?
2. What volume of potassium hydroxide solution (concentration 2 mol/l) is required to neutralise 50cm3 of sulphuric acid (concentration 1 mol/l)?
3. What volume of nitric acid (concentration 0.5 mol/l) is required to neutralise 25cm3 of potassium hydroxide solution (concentration 4 mol/l)?
4. What volume of sulphuric acid (concentration 2 mol/l) is required to neutralise 25cm3 of potassium hydroxide solution (concentration 4 mol/l)?
5. If 25cm3 of hydrochloric acid is neutralised by 20cm3 of sodium hydroxide solution (concentration 2 mol/l), what is the concentration of the acid?

10. If 50cm3 of potassium hydroxide solution is neutralised by 25cm3 of sulphuric acid (concentration 2 mol/l), what is the concentration of the alkali?

**Salt Production**

**Four types of neutralisation reaction in which a salt is produced are:**

**acid + alkali salt + water**

**acid + metal oxide salt + water**

**acid + reactive metal salt + hydrogen**

**acid + metal carbonate salt + water + carbon dioxide**

**Soluble salts**

1. For each of the following reactions, write a word equation and underline the soluble product.
   1. dilute hydrochloric acid + sodium hydroxide solution
   2. potassium hydroxide solution + dilute sulphuric acid
   3. dilute nitric acid + magnesium hydroxide
   4. calcium carbonate + dilute hydrochloric acid
   5. dilute nitric acid + sodium hydroxide solution
   6. dilute sulphuric acid + lead (ll) oxide
   7. dilute hydrochloric acid + zinc
   8. copper (ll) oxide + dilute hydrochloric acid.
2. Name the acid and the alkali which can be used to prepare the following soluble products.

(a) sodium sulphate (b) potassium nitrate (c) lithium chloride.

1. Name the acid and the metal oxide which can be used to prepare the following soluble products.

(a)magnesium chloride (b)iron(II)nitrate (c)copper(II) sulphate

4. Name the acid and the metal carbonate which can be used to prepare the following soluble products.

(a)zinc (II) sulphate (b)magnesium chloride (c)calcium nitrate.

1. Name the acid and metal used to prepare the following soluble products.
2. Zinc (II) chloride (b) aluminium nitrate (c) tin sulphate

**Spectator ions**

1. For each of the following equations, cancel out the spectator ions to show the ions which actually react.

(a) K+(aq) and OH-(aq) + H+(aq) and NO3- (aq) K+(aq) and NO3- (aq) + H2O

(b) Zn(s) + 2H+(aq) and SO42-(aq) Zn2+(aq) and SO42-(aq) + H2(g)

(c) Ba2+(aq) and 2Cl-(aq) + 2Na+(aq) and SO42-(aq) Ba2+SO42-(s) + 2Na+(aq) and 2Cl-(aq)

2. For each of the following equations, write the formula of each substance to show the ions which are present and cancel out the spectator ions to form the ion equation.

1. magnesium + hydrochloric magnesium + hydrogen(g) + carbon

carbonate(s) acid(aq)  chloride(aq) dioxide(g)

1. potassium hydroxide solution with dilute hydrochloric acid