GCSE

PRACTICAL BOOKLET

SINGLE AWARD SCIENCE: CHEMISTRY

Rewarding Learning

2.1 Acids, bases and salts

2.1.3 Investigate how indicators can be obtained from natural dyes that can be extracted from plants, such as red cabbage or beetroot.

Introduction

Humans have learned to use natural and artificial indicators to test acids and bases in solutions. Some common indicators seen in labs are litmus and universal indicator. They are used to test acids and bases by observing the change in colour.

Some fruits and vegetables contain natural indicators. Indicators are substances that can detect changes in properties such as pH. Examples of natural edible pH indicators include blueberries, beetroot and blackcurrant. These indicators change colour depending on the pH of the solution.

Apparatus

- Red cabbage
- Knife
- Bunsen burner
- Safety goggles
- Water
- Filter paper/funnel
- Glass beaker
- Boiling tubes (or spotting tiles)
- Hydrochloric acid
- Ethanoic acid
- Limewater
- Sodium hydroxide

Risk Assessment

The cabbage must be chopped with care using a knife. It may be appropriate to have this carried out beforehand by a technician. Care must be taken when boiling the water and filtering the solution which must be allowed to cool down sufficiently. Eye protection must also be worn when handling acids and alkalis (though relatively weak concentrations of acid and alkali such as 0.5M should be fine for this practical).

Method

1. Chop the cabbage into small pieces.
2. Place the cabbage in a large beaker and add boiling water to cover the cabbage.
3. Heat for at least ten minutes for the colour to leach out of the cabbage. This liquid is at about pH 7.
4. Allow to cool and filter the solution to remove cabbage pieces.
5. Place a few drops of the various solutions in boiling tubes (or spotting tiles).
6. Test the various solutions with the indicator and record the colour changes.

Results

Name	pH	Colour of indicator
Hydrochloric acid	1	Dark pink/red
Ethanoic acid	4	Light pink
Water	7	Blue
Limewater	9	Green
Sodium hydroxide	14	Yellow

What colour does cabbage indicator turn in a weak acid?

What colour does cabbage indicator turn in a strong alkali?

What colour does cabbage indicator turn in a neutral solution?

Fresh red cabbage is purple in colour. Why is pickled red cabbage, red in colour?

Additional information

Video of how to do the practical and expected results:
https://www.youtube.com/watch?v=T-OxyB9dsJ0
Video clips, a detailed lesson plan and student activities is available at:
http://www.bbc.co.uk/schools/teachers/bang/videos/lesson1_red_cabbage_indicator.shtml

2.1.11 Carry out practical work to follow a neutralisation reaction by monitoring pH (Prescribed Practical C1)

Introduction

Acids and bases are chemical opposites. When added together they react and cancel each other out. A neutralisation reaction occurs when an acid reacts with a base/alkali. The products are generally a salt and water. As water is a neutral liquid, this reaction is known as neutralisation.

Acid + base → salt + water

For example when hydrochloric acid reacts with sodium hydroxide (alkali), sodium chloride (salt) and water are formed:
Word equation: hydrochloric acid + sodium hydroxide → sodium hydroxide + water

Symbol equation: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

There are two ways of telling when just the right amount of acid has been added to an alkali to neutralise it

1. Using an indicator
2. Using a pH sensor

Apparatus

- Goggles
- 1M hydrochloric acid
- 1M sodium hydroxide
- Burette
- Pipette
- Pipette filler
- Magnetic stirrer
- Funnel
- pH meter
- Conical flask
- Universal indicator solution (optional)

Risk Assessment

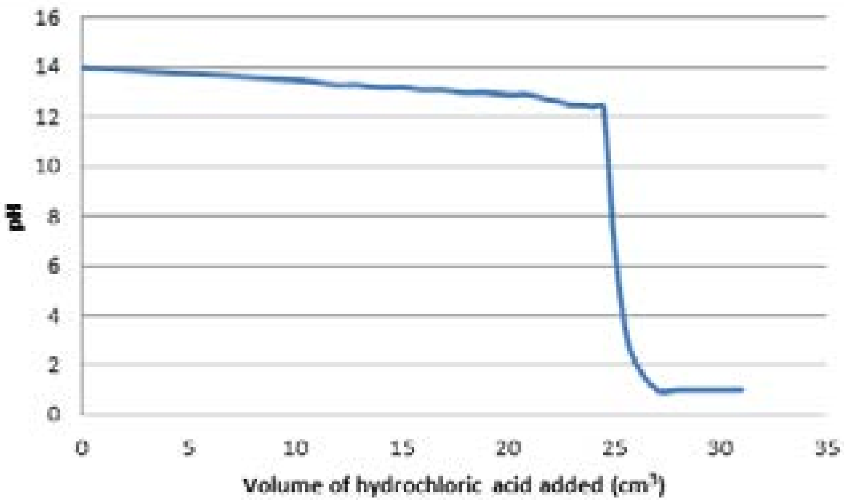
Care should be taken when transporting, using and filling burettes and pipette due to their delicate nature. Pipettes should always be filled using a pipette filler and never by mouth. Appropriate care should also be taken with using acids and alkalis. If any are spilt on the skin, the area should be rinsed with water.

Method

1. Rinse out the pipette with sodium hydroxide, and then carefully fill the pipette with $25cm^3$ of 1M sodium hydroxide solution. Add to the conical flask. Record the pH. (Optional: add a few drops of universal indicator solution)
2. Rinse out the burette with hydrochloric acid, and then carefully fill the burette with 1M hydrochloric acid.
3. Add $1cm^3$ of hydrochloric acid to the flask and record the pH.
4. Repeat step 3 until approximately $22cm^3$ of acid has been added. Then add in $0.5cm^3$ increments until after the end point is reached.
5. Continue adding acid until the pH of the solution is strongly acidic
6. Plot a titration curve of your results.

Results

The titration curve should look similar to the one below:



If desired, the experiment could be repeated using a weak acid such as ethanoic acid and the shapes of the curves compared.

What is an advantage of a pH meter over universal indicator?

How could we make our results more reliable?

Why do farmers sometimes do a similar reaction with their soil?

Additional information

Video of how to set up burette:

https://www.youtube.com/watch?v=ZHuuF8ADvYg

Video of how to use a pipette and how to do a titration:

https://www.youtube.com/watch?v=wd-tYaMGNxc

Alternative neutralisation reaction:

Apparatus

- Goggles
- 2M hydrochloric acid
- Sodium hydrogencarbonate
- Balance
- Spatula
- Weighing boat
- Conical flask
- pH meter

Risk Assessment

Appropriate care should also be taken with using acids and alkalis. If any are spilt on the skin, the area should be rinsed with water.

Method

1. Measure out $25cm^3$ of 2M acid, add to a conical flask.
2. Weigh out the given amount of sodium hydrogencarbonate.
3. Carefully add the sodium hydrogencarbonate to the acid.
4. Measure and record the pH using the pH meter.
5. Repeat the previous 3 steps, until 7g of sodium hydrogencarbonate has been added.

Results

The acid should be neutralised when approximately 4-4.5g of sodium hydrogen carbonate has been added

Amount of baking soda added (g)	pH of solution
0	
1	
2	
3	
3.5	
4.0	
4.5	
5	
6	
7	

Baking soda is an old fashioned remedy for indigestion. How did this work?

The word equation for this reaction is:

sodium hydrogencarbonate + hydrochloric acid → sodium chloride + water + carbon dioxide

Sometimes this was referred to as “getting your wind up.” Look at the equation – which product is “your wind”?

2.2 Elements, compounds and mixtures

2.2.7 investigate how mixtures can be separated using filtration, crystallisation, paper chromatography and simple distillation

Introduction

When elements react together, compounds are made. When elements or compounds are mixed together, a mixture is made. Sometimes, it is possible to separate the original substance out by using a variety of methods such as filtration, crystallisation, paper chromatography or simple distillation. These are collectively known as separation techniques.

Experiment 1: Filtration/Crystallisation

Apparatus

- Mortar and pestle
- Rock salt
- Filter funnel and paper
- Evaporating basin
- Bunsen burner
- Goggles
- Tripod and gauze
- Heatproof mat

Risk assessment

If using Bunsen burner, take care when evaporating the solution. When a little amount of water is left, the solution may "spit" some hot salt. Turn off the Bunsen to avoid this.

Method

1. Grind the rock salt with mortar and pestle until well mixed.
2. Add a small amount of warm water and stir.
3. Carefully filter this solution and collect the filtrate in an evaporating basin.
4. **Either** leave the filtrate to evaporate at room temperature in order to obtain larger sodium chloride crystals.
5. **Or** evaporate the water off using a Bunsen burner in order to obtain smaller salt crystals.

Results

When water is added, the soluble salt (solute) dissolves into the water (solute) to make a salt solution. The insoluble solids are left as a residue in the filter paper and the filtrate contains salt solution. The salt and water can then be separated by crystallisation in air or by evaporation of the water.

Why do we use warm water?

What is a residue?

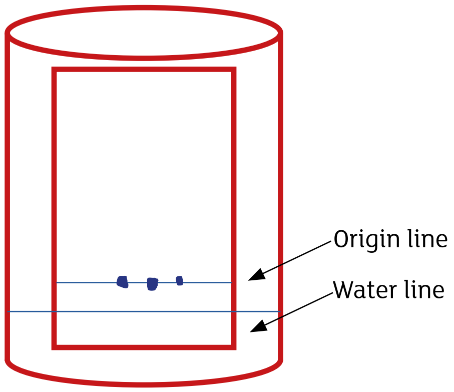
What is a filtrate?

Why do we use rock salt on the roads in winter?

Experiment 2 Paper chromatography

Apparatus

- Chromatography paper
- Felt tip pens
- Beaker
- Pencil



Method

1. Place a small amount of water in the bottom of the beaker.
2. On a piece of chromatography paper, draw a line in pencil approximately 1cm from the bottom of the paper.
3. Place a dot on different parts of the line with 3 or 4 different colour of pens.
4. Carefully place the chromatography paper in the beaker of water. The top may be attached to a pencil to provide support. Ensure that the water in the beaker does not cover the pencil line.
5. Leave for 15 minutes. Calculate how many dyes (and the R_f values) for each colour.

The distance travelled relative to the solvent is called the R_f value. For each dot (compound), it can be worked out using the formula:

$$R_f = \frac{\text{distance travelled by compound}}{\text{distance travelled by solvent}}$$

Results

Colour of dye	Number of dyes contained	Rf values of each dye

Which colour had most dyes?

Which pen is most soluble in water?

How can this experimental technique be used to fight crime?

Experiment 3 – simple distillation

Apparatus

- Liebig condenser
- Heating mantle/Bunsen burner
- nky water/salt solution
- Beaker

Risk assessment

The apparatus will get very hot when used, so do not dismantle until cooled down.

Method

Teacher will set up apparatus as shown in picture below. Suitable solutions include inky water or salt solution.

Results

The solution will boil and the solvent (water) will evaporate. This will be collected by the process of condensation in the Liebig condenser. The pure solvent is collected, while the solute is left behind in the reaction vessel. If this is heated to dryness, the solute may also be collected.

Additional information

Filtration: Worksheet for experiment may be found here:
http://www.fofweb.com/onfiles/seof/chemistry_experiments/3-07.pdf

Crystallisation: Cartoon explanation at https://www.youtube.com/watch?v=nztV4w0DtOo

Paper chromatography: Information on this may be found at:
http://www.bbc.co.uk/schools/gcsebitesize/science/add_edexcel/covalent_compounds/seperationrev2.shtml

Simple distillation: Cartoon explanation at https://www.youtube.com/watch?v=V5ep0-ojPGw

Simple distillation of salt water: https://www.youtube.com/watch?v=N0f73tbGCRE

2.7 Qualitative analysis

2.7.1 describe how to test for hydrogen gas: apply a lighted splint and a popping sound results (equation for reaction required)

2.7.2 describe how to test for carbon dioxide: limewater (calcium hydroxide solution) will change from colourless to milky if the test is positive

2.7.3 describe how to test for oxygen gas: apply a glowing splint and it relights in the presence of oxygen

Introduction

Earth's atmosphere is approximately 78% nitrogen, 21% oxygen, 0.9% argon, and 0.03% carbon dioxide with very small percentages of other elements. Our atmosphere also contains water vapour. Many of these gases have similar physical properties. They are colourless and odourless and so are difficult to identify. Fortunately, oxygen, carbon dioxide and hydrogen can be identified by a specific chemical test. Consequently, nitrogen could be identified by a process of elimination.

Apparatus

- Gas jars and cover slips
- Wooden splints
- Limewater
- Cylinders of hydrogen, oxygen and carbon dioxide (if available)

Risk assessment

Hydrogen is very flammable. Oxygen and hydrogen combined is explosive. Extreme care must be taken with cylinders of these gases. Pupil experiments: hydrogen – wear goggles when using acid and take care when lighting a splint. Hydrogen is flammable so only a small amount should be prepared. Carbon dioxide – pupils should blow gently into limewater to avoid displacing the limewater from the test tube and should take care not to ingest limewater. If they do they should rinse their mouth immediately with water.

Method

Hydrogen: Fill a gas jar full of hydrogen. Apply a lighted splint and a popping sound will occur.

Carbon dioxide: Bubble carbon dioxide gas through limewater (calcium hydroxide solution). If carbon dioxide is present, the limewater will turn from colourless to milky/cloudy.

Oxygen: Fill a gas jar full of oxygen. Light a splint and blow it out. Place the glowing splint in the gas jar of oxygen and it will relight.

If gas cylinders of gases are unavailable, oxygen may be prepared by adding a small amount of the catalyst manganese dioxide to hydrogen peroxide solution (teacher demo only – exothermic reaction). Hydrogen may be prepared by reacting a metal such as zinc or magnesium with hydrochloric acid and carbon dioxide may be prepared by the thermal decomposition of a metal carbonate.

Pupil experiments

Hydrogen

Method

1. Place a small piece of magnesium ribbon in a test tube.
2. Add a small amount of 1M hydrochloric acid (no more than half fill test tube with acid).
3. Invert a second dry test tube and hold tightly over the mouth of the first test tube.
4. After a short period of time, remove the inverted test tube, turn it around and apply a lighted splint. A squeaky pop will be heard.

Carbon dioxide

1. Half fill a test tube with limewater.
2. Insert a straw and blow gently into the test tube for a few minutes
3. The limewater will turn milky.

Results

Hydrogen: the following reaction occurs when a lighted splint is introduced

Hydrogen + oxygen → water

$2H_{2\ (g)} + O_{2\ (g)} \rightarrow 2H_2O_{\ (l)}$

Oxygen: oxygen supports combustion which is why it relights the glowing splint. To prove this, if a lighted splint is placed in a gas jar of oxygen, it will burn more brightly.

Carbon dioxide: The limewater turns milky as the carbon dioxide reacts with the calcium hydroxide solution to produce insoluble calcium carbonate. (NB if excess carbon dioxide is added, the cloudiness will disappear due to formation of soluble calcium hydrogencarbonate)

A student is given 4 gas jars containing carbon dioxide, oxygen, hydrogen and nitrogen. Describe how they could identify which gas was in each gas jar?

Additional information

Preparation of oxygen diagram and info: http://www.bbc.co.uk/education/guides/zjwnb9q/revision

Collecting gases over water:
http://www.bbc.co.uk/schools/gcsebitesize/science/edexcel_pre_2011/chemicalreactions/preparinggasesrev2.shtml

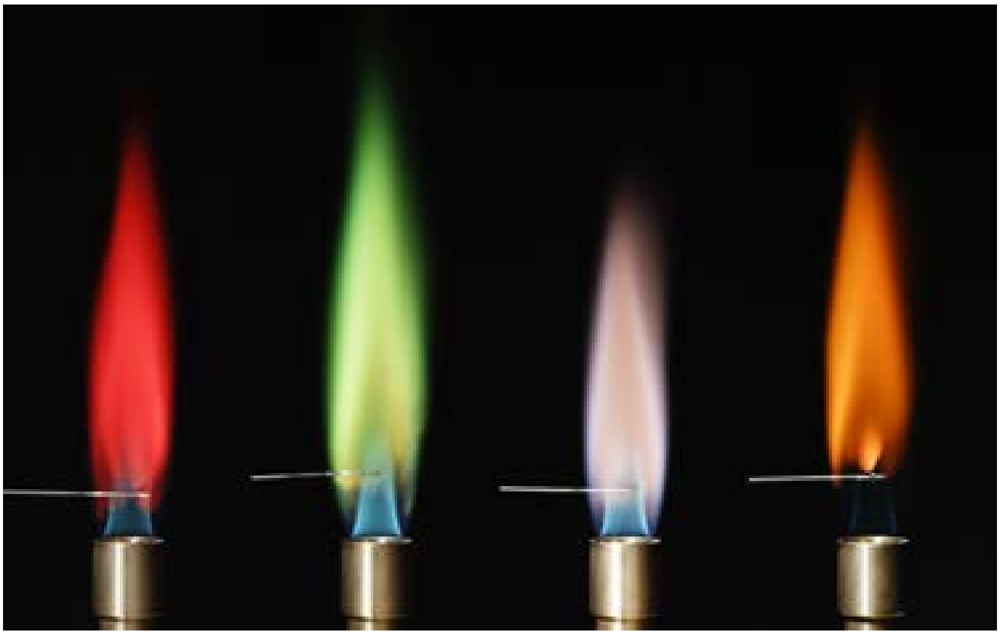
Thermal decomposition of copper carbonate: https://www.youtube.com/watch?v=D9amrlph-rA

Hydrogen and oxygen explosion: https://www.youtube.com/watch?v=DjcztiNGg_8

2.7.4 Investigate how a flame test can be carried out with a nichrome wire and concentrated acid using metal chlorides to identify metal ions

Introduction

Flame tests can be used to identify metals present at a crime scene, as different metals produce different colours when heated in a flame. Nichrome wire is used because it is an alloy with a high melting point that will not affect the colour of the Bunsen flame.



Apparatus

- Nichrome wire
- Metal chloride salts/solutions (lithium, sodium, potassium, calcium and copper chlorides)
- 2M hydrochloric acid
- Safety goggles

Risk assessment

Care should be taken when using strong mineral acids. If splashed on skin, wash off immediately. The wire will be very hot after being placed in a Bunsen flame and must never be touched. If a student is burned, place the affected area under cold water straight away.

Method

1. If the metal salt is solid, dissolve it in water to make a solution.
2. Clean the wire loop by dipping it in hydrochloric acid, and then hold it in the Bunsen flame to dry.
3. Dip the loop into one of the metal solutions.
4. Place the loop in the flame and record the colour.
5. Clean the loop between each metal by dipping it in hydrochloric acid.

Results

Metal ion	Colour
Lithium	Crimson
Sodium	Yellow/orange
Potassium	Lilac
Calcium	Red
Copper	Blue-green

Additional Information

Method, diagrams and pictures:

http://www.bbc.co.uk/schools/gcsebitesize/science/triple_aqa/further_analysis/analysing_substances/revision/1/

Flame test colours: https://www.youtube.com/watch?v=NEUbBAGw14k

2.8 Metals and the reactivity series

2.8.4 Carry out practical work to investigate the reactivity of metals (Prescribed Practical C2) (See also 2.3.14, 2.3.15, 2.8.1, 2.8.2)

Introduction

The reactivity series is an arrangement of metals in decreasing order of their reactivity. The most reactive metals are at top while the least reactive metals at the bottom. The order of reactivity can be determined by reacting different metals with the same substance and analysing the results of the reaction. In order to do this fairly, the only variable which should be changed each time is the metal. Other variable such as the mass of the metal or the amount and concentration of the other reactant should be kept the same.

Experiment 1: Reaction with water (Teacher demo only)

Apparatus

- Tongs
- Scalpel
- Water trough
- Safety screen
- Goggles
- Alkali metals
- Universal indicator paper

Risk assessment

The alkali metals are very reactive, especially sodium and potassium. Their behaviour sometimes can be unpredictable. Only a small piece of metal should be added to water and students should be a safe distance away, behind a safety screen. Note how these metals are kept. Metals should be stored in a safe, dry place when not in use. Under no circumstances should metals be touched as they are corrosive and will react with moisture on your skin.

Method

1. Check the pH of the water using universal indicator
2. With care, cut a **small** piece of lithium using a scalpel and add it to the water. Record observations.
3. Repeat using pieces of sodium and potassium.
4. Check the pH at end of the experiment to show that a strong alkaline solution has formed.

Results

Observations may be recorded as below. There are several similar observations which can be made for each metal. Below is only a selection of a few possible observations for each alkali metal.

Metal	Observations
Lithium	1. Floats on surface 2. Dissolves 3. Fizzing
Sodium	1. Floats on surface 2. Melts into ball 3. Fast reaction
Potassium	1. Very fast reaction 2. Lilac flame 3. Alkali solution formed

The general equation for the reaction is:

Metal + water → metal hydroxide + hydrogen

For example:

sodium + water → sodium hydroxide + hydrogen

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

How could you tell that this reaction is exothermic?

How did we keep this reaction as a fair test?

What is the order of reactivity for these metals?

How does the reactivity of these metals change as you go down group 1?

Additional information

Overview of group 1 and their reactivity: http://www.bbc.co.uk/education/guides/zvydmp3/revision/1

Reactions of Group 1: https://www.youtube.com/watch?v=ptvSrhwHp7s

Rubidium and caesium only: https://www.youtube.com/watch?v=t2uwzNZAp-s

Experiment 2: Reaction with acid

Apparatus

- Boiling tubes
- Boiling tube rack
- Measuring cylinder
- Balance
- Weighing boat
- 1M hydrochloric acid
- Magnesium, zinc, iron, copper powders
- Spatula

Risk assessment

A low concentration of acid and a small amount of powder should be used to avoid an overly vigorous reaction between magnesium and hydrochloric acid which possibly may fizz over the top of the boiling tube. Each protection must be worn when handling these chemicals.

Method

1. Measure out 0.5g of magnesium powder on the top pan balance and add to the boiling tube.
2. Measure out 20cm^3 of hydrochloric acid and add to the boiling tube. Record observations.
3. Repeat steps 1 and 2 using other metals, adding the same mass of metal and the same volume, concentration and type of acid each time.

Results

Metal	Observations
Magnesium	Fast reaction, lots of bubbles, heat evolved
Zinc	Steady reaction, many bubbles of gas
Iron	Slow reaction, bubbles of gas
Copper	No reaction

What is the order of reactivity for these metals?

How did we make this a fair test?

How could we make our results reliable?

Additional information

Overview of the reactivity series and reactions with water and acid:

http://www.s-cool.co.uk/gcse/chemistry/metals-the-reactivity-series/revise-it/reactions-of-metals

Test tube reactions of metals and acid: https://www.youtube.com/watch?v=l0U7VDSxGHk

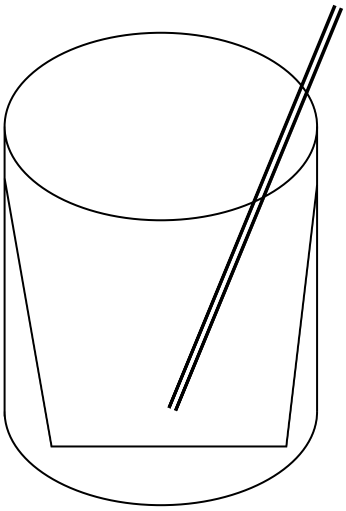
2.8 Energetics

2.8.7 Carry out practical work to investigate the temperature changes which occur during a reaction (Prescribed Practical C3)

Introduction

Reactions can either be endothermic or exothermic. Energy is required to break bonds in the reactants, but energy is given out when new bonds are made in the products of the reaction. Endothermic reactions must take in more energy to break bonds in the reactants than is released when new bonds are made. Exothermic reactions give out more energy when new bonds are made that is required to break the bonds in the reactants.

Apparatus



- 250ml beaker
- Polystyrene cup
- Thermometer
- 10ml measuring cylinder
- Spatula
- Copper sulphate solution
- Dilute hydrochloric acid
- Sodium hydrogencarbonate solution
- Sodium hydroxide solution
- Magnesium ribbon
- Magnesium powder
- Citric acid

Risk assessment

Wear eye protection. Take care when handling acids and alkalis and magnesium powder. Wear gloves if skin easily irritated.

Method

Sodium hydroxide and hydrochloric acid reaction

1. Place the polystyrene cup in the beaker.
2. Measure out 10 cm^3 of sodium hydroxide. Pour it into the polystyrene cup.
3. Measure and record the temperature of the sodium hydroxide solution.
4. Measure out 10 cm^3 of hydrochloric acid. Add this to the polystyrene cup and stir.

5. Record the maximum or minimum temperature reached.
6. Work out the temperature change.
7. Repeat, work out average temperature change and if the reaction is endothermic or exothermic.

Sodium hydrogencarbonate solution and citric acid

1. Place the polystyrene cup in the beaker.
2. Measure out 10 cm^3 of sodium hydrogencarbonate. Pour it into the polystyrene cup.
3. Measure and record the temperature of the sodium hydrogencarbonate solution.
4. Add 4 small (not heaped) spatula measures of citric acid and stir.
5. Record the maximum or minimum temperature reached.
6. Work out the temperature change.
7. Repeat, work out average temperature change and if the reaction is endothermic or exothermic.

Copper sulfate solution and magnesium powder

1. Place the polystyrene cup in the beaker.
2. Measure out 10 cm^3 of copper sulfate solution. Pour it into the polystyrene cup.
3. Measure and record the temperature of the copper sulfate solution.
4. Add 1 **small** (not heaped) spatula measure of magnesium powder and stir
5. Record the maximum or minimum temperature reached.
6. Work out the temperature change.
7. Repeat, work out average temperature change and if the reaction is endothermic or exothermic.

Hydrochloric acid and magnesium ribbon

1. Place the polystyrene cup in the beaker.
2. Measure out 10 cm^3 of hydrochloric acid. Pour it into the polystyrene cup.
3. Measure and record the temperature of the hydrochloric acid.
4. Add one 2-3 cm piece of magnesium ribbon and stir.
5. Record the maximum or minimum temperature reached.
6. Work out the temperature change.
7. Repeat, work out average temperature change and if the reaction is endothermic or exothermic.

Results

Reaction	Start temperature (^oC)	End temperature (^oC)	Change in temperature (^oC)	Average temperature change (^oC)
Sodium hydroxide and hydrochloric acid				
Sodium hydrogencarbonate and citric acid				
Copper sulfate and magnesium powder				
Hydrochloric acid and magnesium ribbon				

Which reactions were endothermic and which were exothermic?

How did you make your results reliable?

Were there any anomalous results? What could you do if there were?

Additional information

Teacher's note: All experiments above are exothermic except sodium hydrogencarbonate + citric acid which is, obviously, endothermic.

Further information about endo and exothermic reactions:

http://www.bbc.co.uk/schools/gcsebitesize/science/add_ocr_21c/chemical_synthesis/whychemicalsrev8.shtml

Endothermic and exothermic reaction demonstrated: https://www.youtube.com/watch?v=BTDRtSGNMtM

2.9 Rates of reaction

2.9.3 Suggest appropriate practical methods to measure the rate of a reaction and collect reliable data (methods limited to measuring a change in mass or gas volume against time); the reactions studied should include:

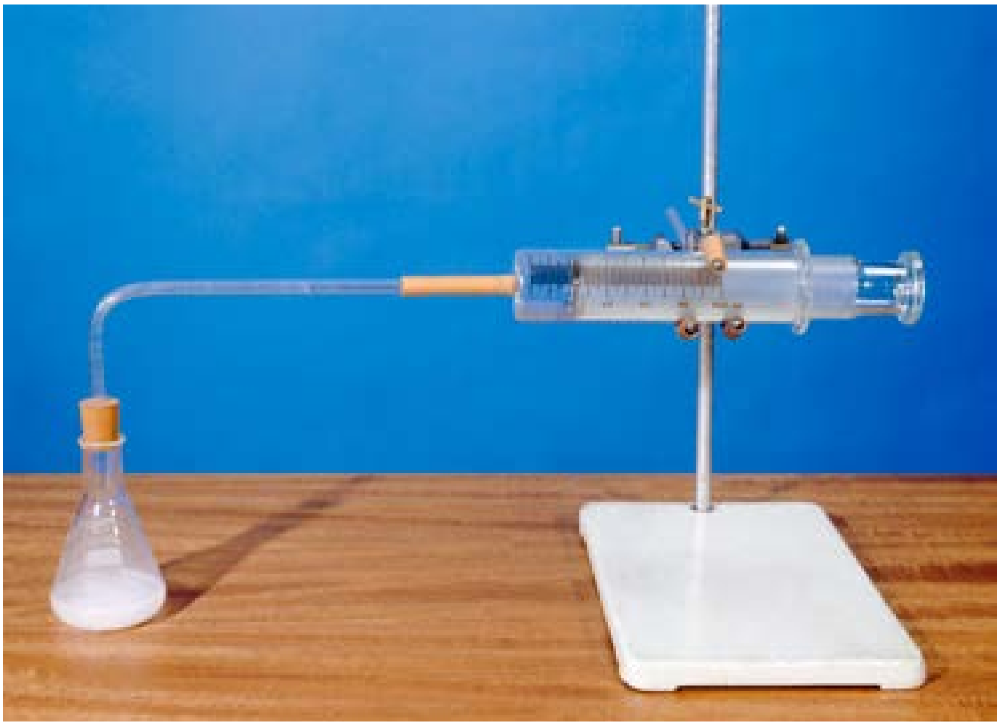
- metals with dilute acid; and
- metal carbonates with dilute acid;

2.9.7 Carry out practical work to investigate how changing a variable changes the rate of reaction (other practical).

Introduction

The rate of reaction is a way of working out how fast or slowly a reaction is progressing. This can be done in several different ways, such as measuring mass loss during a reaction or in measuring the volume of gas produced over the course of a reaction (see picture below). In order for a reaction to take place, reacting particles must collide with sufficient activation energy to react. Changing conditions in the reaction can result in more collisions between particles, which in turn would lead to more successful collisions and consequently a faster rate of reaction. The variables which can affect the rate of reaction include:

- concentration
- temperature
- surface area
- catalyst
- light



These experiments primarily look at the reaction between calcium carbonate and excess hydrochloric acid.

Calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide

$$CaCO_{3\,(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

Experiment 1: Surface area

Apparatus

- Conical flask
- Marble chips (Small, large and powdered calcium carbonate)
- 2M Hydrochloric acid
- Stopclock
- Goggles
- Measuring cylinder
- Balance
- Weighing boat

Risk assessment

Wear eye protection. Take care when handling acids and calcium carbonate powder. Wear gloves if skin easily irritated.

Method

1. Measure out 25ml of 2M hydrochloric acid and pour into conical flask.
2. Weigh out 2g of powdered calcium carbonate.
3. Add to flask and start the stopclock.
4. Record how long it takes for powder to dissolve.
5. Repeat steps 1-4 using 2g of small marble chips.
6. Repeat steps 1-4 using 2g of large marble chips.
7. For reliability, each experiment may be replaced 3 times.

Results

The smaller the particle size, the faster the rate of reaction and consequently the less time it will take to dissolve.

How did we make this a fair test?

Experiment 2: Concentration

Apparatus

- Conical flask
- marble chips (small)
- 0.5M Hydrochloric acid
- 1M Hydrochloric acid
- 2M Hydrochloric acid
- Stopclock
- Goggles
- Measuring cylinder
- Balance
- Weighing boat

Risk assessment

Wear eye protection. Take care when handling acids. Wear gloves if skin easily irritated.

Method

1. Measure out 25ml of 2M hydrochloric acid and pour into conical flask.
2. Weigh out 1g of small marble chips.
3. Add to flask and start the stopclock.
4. Record how long it takes for marble chips to dissolve.
5. Repeat steps 1-4 using 25ml of 1M hydrochloric acid.
6. Repeat steps 1-4 using 25ml of 0.5M hydrochloric acid.
7. For reliability, each experiment may be replaced 3 times if desired.

Results

The more concentrated the acid, the more particles there will be in solution. This will lead to more successful collisions, a faster rate of reaction and therefore, the marble chips will dissolve faster.

Experiment 3: Temperature

Apparatus

- Boiling tube
- Small marble chips
- 2M Hydrochloric acid
- Stopclock
- Goggles
- Measuring cylinder
- Balance
- Weighing boat
- Ice
- Water bath
- Thermometer

Risk assessment

Wear eye protection. Take care when handling acids. Wear gloves if skin easily irritated.

Method

1. Measure out 25ml of 2M hydrochloric acid and pour into boiling tube.
2. Fill a beaker with ice and put boiling tube in the beaker.
3. Record the temperature of the acid.
4. Weigh out 2g of small marble chips.
5. Add to boiling tube and start the stopclock.
6. Record how long it takes for marble chips to dissolve.
7. Measure out 25ml of 2M hydrochloric acid and pour into boiling tube.
8. Record the temperature of the acid.
9. Weigh out 2g of small marble chips.
10. Add to boiling tube and start the stopclock.
11. Record how long it takes for marble chips to dissolve.
12. Set a water bath to approximately 50oC.
13. Measure out 25ml of 2M hydrochloric acid and pour into boiling tube.
14. Record the temperature of the acid.
15. Weigh out 2g of small marble chips.
16. Add to boiling tube and start the stopclock.
17. Record how long it takes for marble chips to dissolve.
18. For reliability, each experiment may be replaced 3 times.

Results

The warmer the acid is, the more kinetic energy the acid particles will have. They will move around faster, leading to more successful collisions and a faster reaction rate. Therefore the acid in the water bath will dissolve the marble chips faster than the others.

Experiment 4: Mass loss

Apparatus

- Conical flask
- Marble chips (small)
- 2M Hydrochloric acid
- Stopclock
- Goggles
- Measuring cylinder
- Balance
- Weighing boat

Risk assessment

Wear eye protection. Take care when handling acids. Wear gloves if skin easily irritated.

Method

1. Measure out 25ml of 2M hydrochloric acid and pour into conical flask.
2. Weigh out 2g of small marble chips.
3. Place the flask on the balance and record the mass of the flask (remember to add 2g to the total to account for the marble chips)
4. Add the marble chips to flask and start the stopclock.
5. Record the mass of the flask every minute until the mass remains constant for a few minutes.

Results

Time (min)	Mass of flask (g)	Mass loss of flask (g)
0		0
1		
2		
3		
4		
5		
etc.		

Plot a graph of time (min) versus mass loss of flask (g)

When is the reaction at its fastest?

How can you tell from the graph?

Why is excess hydrochloric acid used?

How can you tell from the graph the reaction is over?

When did the reaction finish?

Additional information

Overview of rate of reaction: http://www.bbc.co.uk/schools/gcsebitesize/science/add_aqa/reaction/ratesrev1.shtml

An interesting analogy to help explain factors influencing rate of reaction: https://www.youtube.com/watch?v=OttRV5ykP7A

Questions on rates graphs: www.lesmahagow.s-lanark.sch.uk/wp-content/uploads/2013/06/RATES-HWK.pdf

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